01

Some Basic Concepts of Chemistry

IN THIS CHAPTER ....

Matter and Its Nature

Matter is anything which occupies space and has mass. All the things around us, e.g. water, air, book, table etc., are matter. There are five states of matter namely solid, liquid, gases, plasma and Bose-Einstein condensate. Out of these, three states, i.e. solid, liquid and gas are most common states and provide a basis for the physical classification of matter.

Solid, Liquid and Gas

- **Solids** have a definite volume and shape.
- **Liquids** have a definite volume but not definite shape.
- **Gases** have neither a definite volume nor a definite shape.

Thus, in these states a competition between intermolecular interactions, i.e. attractive force between molecules and thermal energy responsible for repulsion between molecules is visible.

Plasma

Among the rest 2 states, plasma is seen as a state containing gaseous ions and free electrons. It is formed when gaseous state is taken to very high temperatures (say 1000 to 1,000,000,000°C). In short we can say plasmas as low density ionised gases at very high temperatures.

Plasmas can be seen in northern lights or ball lightenings, neon lights, stars (in particular sun), clouds of gas and dust around stars.

Bose Einstein Condensate

BE condensate was predicted in 1924 by **Satyendra Nath Bose** and **Albert Einstein**. However, due to lack of equipments, it was only created in 1935 by **Cornell, Ketterle** and **Weimann.** Its concept and existence is totally opposite to plasmas.

The state is conceptualised by cooling a gas of extremely low density to supercold conditions. **Cornell** and **Weimann** developed BEC at such temperature with **rubidium.**

Dalton's Atomic Theory

J. Dalton in 1803, proposed the atomic theory of matter on the basis of laws of chemical combinations.

According to which

- matter is made up of indivisible and indestructible particles, called **atoms.**
- all atoms of an element have identical mass and similar chemical properties.
- atoms combine in the ratio of small whole numbers to form **compounds** or **molecules**.
- compounds formed by such combinations are alike in every respect.
- chemical reactions involve only combination, separation or rearrangement of atoms.
- atoms are neither created nor destroyed in the course of an ordinary chemical reaction.

Limitations of Dalton's Atomic Theory

It was failed to explain

- how atoms of different elements differ from each other.
- the nature of forces that bind different atoms together in a molecule.
- the difference between atoms and molecules.

The hypothesis of Dalton is even accepted today by the scientific community with two modifications only.

(i) Atom is divisible and destructible.

(ii) All atoms of an element are not identical in mass.

Atoms and Molecules

An **atom** is defined as "the smallest particle of matter which may or may not exist independently but can take part in a chemical reaction."

A **molecule** is defined as "the smallest particle of matter which can exist independently but cannot take part in a chemical reaction."

Both atoms and molecules are basic constituents of matter with the condition that usually atoms combine to form the molecules.

The molecules may be

- monoatomic, i.e. contain 1 atom only, e.g. Na, K etc.
- diatomic, i.e. contain 2 atoms, e.g. N_2, O_2 etc.
- triatomic, i.e. contain 3 atoms, e.g. O_3 etc.
- polyatomic, i.e. contain more than 3 atoms, e.g. P_4, S_8 etc.

Chemical Classification of Matter

On the basis of chemical composition and properties, matter can be classified as,

Matter

- Mixtures
  - Homogeneous mixture
  - Heterogeneous mixture
- Pure substances
  - Compounds
  - Elements

Mixtures

These have variable composition and variable properties due to the fact that their components retain their characteristic properties. These may be separated into pure components by applying physical methods.

These can be of following two types

(i) The **homogeneous mixtures**, have same composition throughout and their components are indistinguishable, e.g. a liquid solution of sugar and water etc.

(ii) A **heterogeneous mixture,** on the other hand, do not have the same composition throughout and the components here are distinguishable, i.e. each component maintain their own identity, e.g. a mixture of grains of sand and salt.

Pure Substances

These have fixed composition and non-variable properties. These cannot be separated into simpler substances by physical methods.

An **element** is a substance that contains only one type of atoms whereas a **compound** is formed when atoms of different elements combine in a fixed ratio.

Decomposition into simpler forms is achieved by physical methods in **elements** and by chemical methods in compounds.

Scientific Notation

In scientific notation, all numbers (however large or small) are expressed as a number multiplied or divided by 10^n. This simply means here a number is generally expressed in the form

$$N \times 10^n$$

Here, N is called **digit term.** Its value lies between 1.000 and 9.999. n is called an **exponent.**

If decimal is shifted towards left, the value of n is positive and if decimal is shifted towards right, the value of n is negative. e.g. 138.42 can be written as 1.3842×10^2 and 0.013842 can be written as 1.3842×10^{-2}

Physical Quantities and their Measurements

The description, interpretation and prediction of the behaviour of substances can be done on the basis of the knowledge of their physical and chemical properties. These properties are determined by careful experimental measurements.

Out of these the properties like mass, length, time, temperature etc., are **physical quantities** and their measurement does not involves any chemical reaction. These properties are expressed in numerals with suitable units.

Various Systems of Measurement

Units with universal acceptance came in 1960s. Till then following systems with different units were used for measurement.

(i) **CGS system** It is also called **Gaussian system** and is based on centimetre (cm), gram (g) and second (s) as the units of length, mass and time respectively.

(ii) **FPS system** It is a British system which used foot (ft), pound (lb) and second (s) as the fundamental units of length, mass and time.

(iii) **MKS system** It is called MKSA system later on. It is the system, which uses metre (m), kilogram (kg) and second (s) respectively for length, mass and time; Ampere (A) was added later on for electric current.

(iv) **SI system** It is internationally accepted system of 1960s. It is also called International system of units and contains following 7 basic and 2 supplementary units.

(a) **Basic units** includes metre (m) for length, kilogram (kg) for mass, second (s) for time, ampere (A) for electric current, kelvin (K) for thermodynamic temperature, *mole* (mol) for amount of substance and candela (Cd) for luminous intensity.

(b) **Supplementary units** includes radian (rad) for angle and steradian (sr) for solid angle.

All these units are called **fundamental units**.

Precision and Accuracy

Precision is the measure of reproducibility of an experiment while accuracy is the measurement of closeness of a result to its true value. Good accuracy means good precision but reverse is not always true, i.e.

Precision = individual value − arithmetic mean value

Accuracy = mean value − true value

Derived Units

The units of all other quantities, which are derived from the above mentioned units, are called the derived units. e.g.

Units of volume = length × breadth × height = m^3

Some Derived Properties and their Units

Quantity	Technical definition	Expression in terms of SI base units
Area	Length squared	m^2
Volume	Length cubed	m^3
Density	Mass per unit volume	kg/m^3 or $kg\ m^{-3}$
Velocity	Distance travelled per unit time	m/s or ms^{-1}
Acceleration	Velocity changed per unit time	m/s^2 or ms^{-2}
Force	Mass, times, acceleration of object	$kg\ m/s^2$ or $kg\ ms^{-2}$ (newton, N)
Pressure	Force per unit area	$kg/(ms^2)$ or $kg\ m^{-1}s^{-2}$ (pascal, Pa)
Energy (work, heat)	Force times distance travelled	$kg\ m^2/s^2$ or $kg\ m^2s^{-2}$ (joule, J)
Electric charge	Ampere, times	A-s (Coulomb, C)
Electric potential	Energy per unit charge	J/(A-s) potential difference (volt, V)

Multipliers and Submultipliers

Sometimes submultiples and multiples are used to reduce or enlarge the size of the different units. The names and symbols of some importance sub-multipliers and multipliers are listed in the table given below.

Sub-multipliers and Multipliers

Multiple	SI Prefix	Symbol	Multiple	SI Prefix	Symbol
10^{24}	yotta	Y	10^{-1}	deci	d
10^{21}	zetta	Z	10^{-2}	centi	c
10^{18}	exa	E	10^{-3}	milli	m
10^{15}	peta	P	10^{-6}	micro	μ
10^{12}	tera	T	10^{-9}	nano	n
10^{9}	giga	G	10^{-12}	pico	p
10^{6}	mega	M	10^{-15}	femto	f
10^{3}	kilo	k	10^{-18}	atto	a
10^{2}	hecto	h	10^{-21}	zepto	z
10	deca	da	10^{-24}	yocto	y

Example 1. *Two students performed the same experiment separately and each one of them recorded two readings, 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.* ***(NCERT Exemplar)***

Student	Readings (i)	Readings (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.

Sol. (*b*) Results of student *A* are close to true value as well as to each other, that's why these are precise as well as accurate.

Significant Figures

The meaningful digits in a properly recorded measurement are known as significant figures.

These figures include all those digits that are known with certainty along with one more which is uncertain or estimated. Always remember that greater the number of significant figures in a reported result, smaller will be the uncertainty.

Rules

The rules for determining the number of significant figures are as follows

- All non-zero digits are significant. e.g. In 852 cm, there are three significant figures and in 0.25 L there are two significant figures.
- Zeros preceding to first non-zero digit are not significant. Such zero indicates the position of decimal point. e.g. 0.03 has one significant figure and 0.0052 has two significant figures.
- Zeros between two non-zero digits are significant. e.g. 3.007 has four significant figures.
- Zeros at the end or right of a number are significant provided they are on the right side of the decimal point. e.g. 0.200 g has three significant figures. But, if otherwise, the terminal zeros are not significant if there is no decimal point.
  e.g. 100 has only one significant figure, but 10.0 has three significant figures and 100.0 has four significant figures.
- Such numbers are better represented in scientific notation. We can express the number 100 as 1×10^2 for one significant figure, 1.0×10^2 for two significant figures and 1.00×10^2 for three significant figures.
- Counting numbers of objects, for example, 2 balls or 20 eggs, have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal, i.e. $2 = 2.000000$ or $20 = 20.000000$.
- In numbers written in scientific notation, all digits are significant, e.g. 4.01×10^2 has three significant figures.

Significant Figures in Calculations

(i) **When adding or subtracting** The number of decimal places in the answer should not exceed the number of decimal places in either of the numbers. e.g.

0.13	2 significant figures
1.5	2 significant figures
20.911	5 significant figures
22.541	

Here, 1.5 has only one digit after the decimal point and the result should be reported only up to one digit after the decimal point which is 22.5.

(ii) **In multiplication and division** The significant figures in the answer should be the same as that in the quantity with the **least number** of significant figures.

e.g. $$\frac{0.01208}{0.0236} = 0.512$$

The number 0.0236 has only three significant figures that's why the answer must also be limited to three significant figures.

Similarly, for the product

$$132.07\times0.12 = 15.8484$$

The answer 15.8484 should be reported as 15 because 0.12 has only two significant figures.

(iii) **When a number is rounded off** The number of significant figures is reduced. The last digit retained is increased by 1 only if the following digit is > 5 and is left as such if the following digit is ≤ 4.

e.g. 12.696 can be written as 12.7
13.93 can be written as 13.9

If the following digit is 5, left the last digit retained as such if it is even or add 1 if it is odd.

e.g. 18.35 can be written as 18.4
18.45 can be written as 18.4

While calculating the significant figures of numbers, it is better to convert them into scientific notation because exponential term does not contribute to the significant figures.

Example 2. *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 ***(NCERT Exemplar)***
(b) 4680×10^{-3} g
(c) 4.680 g
(d) 46.80 g

Sol. (a) Mass = volume × density = 1.5 mL × 3.12 g mL^{-1}
= 4.68 g

The digit 1.5 has only two significant figures, so the answer must also be limited to two significant figures. Hence, it is rounded off to give 4.7g as the answer.

Dimensional Analysis

In calculations, generally there is a need to convert units from one system to other. This can be done with the help of conversion factor, so the method used to accomplish this is called **factor label method** or **unit factor method** or dimensional analysis.

The basic formulation used here is

information sought = information given × conversion factor

Some commonly conversion factors are given below

$$1\text{m} = 39.37\text{ inch}$$
$$1\text{ inch} = 2.54\text{ cm}$$
$$1\text{L} = 1000\text{ mL} = 1000\text{ cm}^3$$
$$= 10^{-3}\text{ m}^3 = 1\text{dm}^3$$
$$1\text{ lb} = 453.59237\text{g}$$
$$1\text{J} = 1\text{N} - \text{m} = 1\text{ kgm}^2\text{s}^{-2}$$
$$1\text{cal} = 4.184\text{ J} = 2.613 \times 10^{19}\text{eV}$$
$$1\text{eV} = 1.602 \times 10^{-19}\text{J}$$
$$1\text{eV / atom} = 96.485\text{ kJ mol}^{-1}$$
$$1\text{ u /amu} = 931.5\text{ MeV}$$
$$1\text{ J} = 10^7\text{ erg}$$
$$1\text{ dyne} = 10^{-5}\text{ N}$$
$$1\text{ atm} = 101325\text{Nm}^{-2} = 101325\text{ Pa}$$
$$1\text{ bar} = 1 \times 10^5\text{ Nm}^{-2} = 1 \times 10^5\text{ Pa}$$
$$1\text{ L-atm} = 101.3\text{ J} = 24.21\text{ cal}$$
$$1\text{ mol (gas)} = 22 \cdot 4\text{ L at STP}$$
$$1\text{ mol (substance)} = N_A\text{ molecules}$$
$$1\text{ g-atom} = N_A\text{ atoms}$$
$$1\text{ K} = t^\circ\text{C} + 273.15$$
$$t(^\circ\text{F}) = \frac{9}{5}t^\circ\text{C} + 32$$
$$1\text{ D (Debye)} = 1 \times 10^{-18}\text{ esu-cm}$$
$$1\text{ g-cm}^{-3} = 1000\text{ kg cm}^{-3}$$

Laws of Chemical Combinations

The combination of elements in atomic state to form compounds in molecular state is governed by following generalised basic laws.

(i) **Law of Conservation of Mass** (by Lavoisier) According to this law the matter can neither be created nor destroyed in a chemical reaction.

(ii) **Law of Definite Proportions** (by J. Proust) This law can be interpreted as "a sample of a pure compound always consists of the same elements combined in same proportions by mass, regardless of its source."

e.g. Ammonia always has the formula NH_3 i.e. one molecule of NH_3 always contains one atom of nitrogen and three atoms of hydrogen or 17.0 g of NH_3 always contains 14 g of nitrogen and 3 g of hydrogen. These findings always remain the same for NH_3.

(iii) **Law of Multiple Proportions** (by John Dalton) An element may form more than one compound with another element.

In such a case for a given mass of first element, the masses of other elements in its two or more compounds come in the ratio of small integers." This is called law of multiple proportions.

e.g. In NH_3, 14 g of nitrogen requires 3 g of hydrogen and in hydrazine (N_2H_4) 14 g of nitrogen requires 2 g of hydrogen. Hence, fixed mass of nitrogen requires hydrogen in the ratio 3 : 2 in two different compounds. Here, 3 : 2 is a simple ratio. Thus, this is in agreement with "law of multiple proportions".

Remember Law of definite proportions and law of multiple proportions do not hold good when same compound is prepared by different isotopes of the same element.
e.g. H_2O and D_2O or H_2O^{16} and H_2O^{18}. Moreover, law of conservation of mass does not hold good for nuclear reactions.

Gay Lussac's Law of Gaseous Volumes

According to this law, "the volume of reactants and products in a large number of chemical reactions are related to each other by small integers, provided the volumes are measured at same temperature and pressure".

e.g. In the reaction of hydrogen with oxygen to produce water, it was found that 2 vol of H_2 combines with 1 vol of O_2 to form 2 vol of H_2O (steam).

This simply means that 100 mL of H_2 gas combines with 50 mL of O_2 to produce exactly 100 mL of steam if volume of all the gases are measured at same temperature and pressure.

Avogadro's Law

According to this law (at constant pressure and temperature). "The volume of a gas is proportional to the number of moles (or molecules) of the gas present".

i.e. $$V \propto n$$

where, n = number of moles of gas

In simpler words, the law can also be stated as "equal volumes of all gases, under the same conditions of temperature and pressure contain equal number of molecules" which is infact 6.023×10^{23} or in multiples of it. This number is commonly called as **Avogadro's number**, i.e. N_0.

Atomic Mass

The atomic mass is measured by comparing mass of an atom with the mass of a particular atom chosen as standard. On the present atomic mass scale, ^{12}C is chosen as standard and is arbitrarily assigned the mass of 12 atomic mass unit (amu).

$\therefore$ Atomic mass

$$= \frac{\text{mass of one atom of the element}}{\frac{1}{12}\text{th part of the mass of one atom of C-12}}$$

Therefore, one amu or u (unified mass) is equal to exactly the 1/12th of the mass of ^{12}C atom.

$$1\ u = \frac{1}{12} \times \frac{12\ g}{6.022 \times 10^{23}} = 1.66 \times 10^{-24}\ g$$

Average Atomic Mass

Since, most of the elements have isotopes, the atomic mass of an element is, infact, the average of masses of its all the naturally occurring isotopes.

e.g. If an element exists in three isotopic forms having atomic masses, m_1, m_2 and m_3 in the ratio, *x, y* and *z,* the average atomic mass,

$$M_{av} = \frac{m_1 \times x + m_2 \times y + m_3 \times z}{x + y + z}$$

If in place of ratio, relative abundance is given, then

$$M_{av} = \frac{m_1 \times r_1 + m_2 \times r_2 + m_3 \times r_3}{r_1 + r_2 + r_3}$$

Here, r_1, r_2 and r_3 are relative abundances of the isotopes.

Example 3. *The average molar mass of chlorine is* $35.5\ g\ mol^{-1}$. *The ratio of* ^{35}Cl *to* ^{37}Cl *in naturally occurring chlorine is close to* ***(JEE Main 2020)***

(a) 4 : 1 (b) 3 : 1

(c) 2 : 1 (d) 1 : 1

Sol. (*b*) Given,

Average molar mass of chlorine is $35.5\ g\ mol^{-1}$.

Naturally occurring chlorine are

	^{35}Cl	^{37}Cl
Molar ratio	x	$1-x$

Total molar ratio $= x + 1 - x = 1$

$$\text{Mass average} = 35 \times x + 37 \times (1-x)$$

$$35.5 = 35x + 37 - 37x$$

$$2x = 37 - 35.5$$

$$x = \frac{1.5}{2} = 0.75 = \frac{3}{4}$$

For ^{35}Cl, $x = \frac{3}{4}$

For ^{37}Cl; $1 - \frac{3}{4}$

$$^{35}Cl : {^{37}Cl}$$

$$\frac{3}{4} : \left(1 - \frac{3}{4}\right) \Rightarrow \frac{3}{4} : \frac{1}{4}$$

On simplify this ratio, we get 3 : 1.

Dulong and Petit's Method to Calculate Average Atomic Mass

According to this method, the approximate atomic mass of solid elements except Be, B, C and Si, is related to their specific heats as

$$\text{Average atomic mass} = \frac{6.4}{\text{specific heat}}$$

Mass spectrometer is used to determine the atomic mass experimentally.

Molecular Mass

It is the sum of atomic masses of the elements present in a molecule. *It is normally calculated for covalent compounds as they have a definite molecular structure.*

It is obtained by multiplying the atomic mass of each element by the number of its atoms and adding them together.

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

$$= 6(12.011u) + 12(1.008u) + 6(16.00u)$$

$$= (72.066u) + (12.096u) + (96.00u)$$

$$= 180.162u \approx 180$$

Formula Mass

The formula mass of a substance is the sum of the atomic masses of all atoms in the formula unit of the compound. *It is normally calculated for ionic compounds as they don't have any definite molecule.*

e.g. Formula mass of NaCl is $23 + 35.5 = 58.5$ amu or 58.5u.

Equivalent Mass or Equivalent Weight

Besides the molecular and formula mass, the masses of compounds are also calculated on comperative basis.

The number of parts of a substance that combines with or displaces, directly or indirectly, 1.008 parts by mass of hydrogen or 35.5 parts by mass of chlorine or 8 parts by mass of oxygen is called the equivalent mass of that substance.

Mathematically,

$$\frac{\text{Equivalent weight of substance } A}{\text{Equivalent weight of substance } B} = \frac{\text{Weight of substance } A}{\text{Weight of substance } B}$$

If A is metal and B is H (or O or Cl) then

Equivalent weight of metal

$$= \frac{\text{mass of metal}}{\text{mass of hydrogen displaced}} \times 1.008$$

or $$= \frac{\text{mass of metal}}{\text{mass of oxygen combined}} \times 8.0$$

or $= \dfrac{\text{mass of metal}}{\text{mass of chlorine combined}} \times 35.5$

Equivalent weight of acid

$$= \frac{\text{molecular weight of acid}}{\text{basicity (number of replaceable } H^+)}$$

e.g. Equivalent weight of $H_2SO_4 = \dfrac{98}{2} = 49$

Equivalent weight of base

$$= \frac{\text{molecular weight of base}}{\text{acidity (number of replaceable } OH^-)}$$

e.g. Equivalent weight of NaOH $= \dfrac{40}{1} = 40$

Equivalent weight of salt

$$= \frac{\text{molecular weight of salt}}{\text{total positive valency of metal atoms}}$$

e.g. Equivalent weight of $Na_2CO_3 = \dfrac{106}{2} = 53$

Equivalent weight of a substance that undergoes

$$\text{oxidation/ reduction} = \frac{\text{molecular weight}}{\text{change in oxidation number}}$$

e.g. When $KMnO_4$ reacts under acidic conditions, change in oxidation number (from +7 to +2) is 5. Hence, equivalent weight of $KMnO_4$ in acidic medium

$$= \frac{158}{5} = 31.6$$

For volatile metal chlorides, eq. wt. and atomic weight are related as

$$\text{Atomic wt.} = \text{Eq. wt.} \times \text{valency}$$

Remember

- Atomic and molecular masses of elements and compounds respectively are always constant but equivalent mass may vary with change of valency.
- The valencies of elements forming isomorphous compounds, i.e. the compounds that have similar constitution and chemical formulae are same, e.g. valencies of Cr, Se and S in K_2CrO_4, K_2SeO_4 and K_2SO_4 are same.

Percentage Composition of Compounds

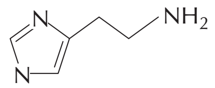
For compounds this is the relative mass of the each of the constituent elements in 100 parts of it. Percentage composition is readily calculated from the formula of the compound as,

Mass percentage of an element

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

Example 4. *The mass percentage of nitrogen in histamine is ....* ***(JEE Main 2020)***

Sol. (*37.84*)



Histamine

Molecular formula $= C_5H_9N_3$. Molecular mass $= 111$

Number of nitrogen atoms per molecule $= 3$

Mass due to N $= 3 \times 14 = 42$

$$\% \text{ of N} = \frac{42}{111} \times 100 = 37.84\ \%$$

The value have a range from 37.80 to 38.20.

Example 5. *The ratio of the mass percentages of 'C' and H' and 'C and O' of a saturated acyclic organic compound 'X' are* 4 : 1 *and* 3 : 4 *respectively. Then, the moles of oxygen gas required for complete combustion of two moles of organic compound 'X' is ....* ***(JEE Main 2020)***

Sol. (*5*) For the molecular formula, $C_nH_{2n+2}O_2$ (M = Molar mass)

$$C\% = \frac{12n}{M} \times 100, \ H\% = \frac{2n+2}{M} \times 100, \ O\% = \frac{32}{M} \times 100$$

To be saturated and alicyclic, $n = 2$, is,

$$C\% = \frac{24}{M} \times 100, \ H\% = \frac{6}{M} \times 100 \text{ and } O\% = \frac{32}{M} \times 100$$

$\Rightarrow$ $C\% : H\% = 24 : 6 = 4 : 1$

$C\% : O\% = 24 : 32 = 3 : 4$

So, X is $C_2H_6O_2$ and its combination equation is

$$C_2H_6O_2 + \left(2 + \frac{6}{4} - \frac{1}{2}\right) O_2 \longrightarrow 2CO_2 + \frac{6}{2} H_2O$$

$$\Rightarrow \quad \underset{1 \text{ mol}}{C_2H_6O_2} + \underset{\frac{5}{2}\text{ mol}}{\frac{5}{2} O_2} \longrightarrow 2CO_2 + 3H_2O$$

2 mol of compound X required $\dfrac{5}{2} \times 2$ mol of O_2

$= 5$ mol of O_2 will be required.

Empirical and Molecular Formulae

The chemical formula of a compound which shows relative number of atoms in their simplest ratio is called **empirical formula.**

On the other hand, the chemical formula of a compound which gives the actual number of atoms of each element in a molecule is called its **molecular formula.**

It is important to note that the molecular formula is generally an integral multiple of the empirical formula, i.e.

$$\text{Molecular formula} = (\text{empirical formula}) \times n$$

where, $n = 1, 2, 3, \ldots$ etc.

The molecular formula conveys following two informations

(i) The relative number of each type of atoms in a molecule.

(ii) The total of atoms of each element in the molecule.

Deriving Empirical Formula

Use the following steps to determine the empirical formula of the compound.

I. Calculate the amount of elements and their percentage composition.

II. Divide the percentage of each element by its atomic mass. It gives atomic ratio of the elements present in the compound.

III. Divide the atomic ratio of each element by the minimum value of atomic ratio. As the result you will get simplest ratio of the atoms of elements present in the compound.

IV. If the simplest ratio is fractional, then multiply the values of simplest ratio of each element by a smallest integer to get a simplest whole number for each of the element.

V. To get the empirical formula, write symbols of various elements present side by side with their respective whole number ratio as a subscript to the lower right hand corner of the symbol.

Deriving Molecular Formula

The molecular formula of a substance may be determined from the empirical formula if the molecular mass of the substance is known.

The molecular formula is always a simple multiple of empirical formula and the value of simple multiple is obtained by dividing molecular mass with empirical formula mass.

Example 6. *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. *The molecular formula of the compound is* ***(NCERT)***

(a) CH (b) CH_2 (c) C_2H_2 (d) CH_4

Sol. (c) (i) 44 g CO_2 contains 12 g carbon

$$3.38 \text{ g } CO_2 \text{ contains } \frac{12}{44} \times 3.38 \text{ g} = 0.9218 \text{ g carbon}$$

$$18 \text{ g } H_2O \text{ contains } 2 \text{ g hydrogen}$$

$$0.690 \text{ g } H_2O \text{ contains } \frac{2}{18} \times 0.690 \text{ g} = 0.0767 \text{ g hydrogen}$$

As the compound contains only carbon and hydrogen so the total mass of compound = 0.9218 + 0.0767 = 0.9985g

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

$$\text{Percentange of H in the compound} = \frac{0.0767}{0.9985} \times 100 = 7.68$$

Calculation for Empirical Formula

Element	Percent by mass	Atomic mass	Relative number of moles of elements	Simplest molar ratio
C	92.32	12	$\frac{92.32}{12} = 7.69$	$\frac{7.69}{7.68} \approx 1$
H	7.68	1	$\frac{7.68}{1} = 7.68$	$\frac{7.68}{7.68} = 1$

Hence, empirical formula = CH

(ii) **Calculation for molar masss of the gas**

10.0 L of the given gas at STP weigh = 11.6 g

$\therefore$ 22.4 L of the given gas at STP will weigh

$$= \frac{11.6 \times 22.4}{10} = 25.984 \text{ g}$$

Molar mass = $25.984 \approx 26$ g mol^{-1}

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

$$\therefore \quad n = \frac{\text{molecular mass}}{\text{empirical formula mass}} = \frac{26}{13} = 2$$

Hence, molecular formula = $n \times CH = 2 \times CH = C_2H_2$

Example 7. *A compound having empirical formula* $(C_3H_4O)_n$ *has vapour density* 84. *The molecular formula of this compound is*

(a) C_3H_4O (b) $C_6H_8O_2$

(c) $C_6H_{12}O_3$ (d) $C_9H_{12}O_3$

Sol. (d) Molecular mass = $2 \times VD = 2 \times 84 = 168$

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

$$= \frac{168}{12 \times 3 + 4 + 16} = \frac{168}{56} = 3$$

Thus, molecular formula = $(C_3H_4O) \times 3$

$= C_9H_{12}O_3$

Mole Concept

The word mole was introduced around 1896 by W. Ostwald who derived it from Latin word moles means a heap or pile. It is represented by the symbol mol.

One mole of a substance always means 6.023×10^{23} entities of that substance. It is called **Avogadro's number/constant** represented by N_A.

Technically, one mole of any substance is defined as,

- The amount which weighs exactly same as its **formula weight in gram** or **atomic mass in gram** or **molecular/formula mass in gram**.
- The amount which has same number of entities, i.e. atoms, molecules or other particles as there are atoms in exactly 0.012 kg (or 12 g) of carbon-12 isotope. This number is equal to Avogadro's number, i.e. 6.023×10^{23} entities.

- The amount which occupies 22.4 L at STP if it is taken for a gas.

Thus, the formulae used to convert amount of substance into moles are

- Number of moles of molecule $= \dfrac{\text{weight in gram}}{\text{molecular/formula weight}}$
- Number of moles of atoms $= \dfrac{\text{weight in gram}}{\text{atomic weight}}$
- Number of moles of particles $= \dfrac{\text{number of particles}}{\text{Avogadro's number}}$
- Number of moles of gases $= \dfrac{\text{volume of gas at STP (in L)}}{22.4}$

For example, one mole of NaCl $= 6.023 \times 10^{23}$ NaCl units

$= 6.023 \times 10^{23}$ units of Na^+

$+ \; 6.023 \times 10^{23}$ units of Cl^-

Mass of one mole of NaCl = 23.0 g + 35.5 g = 58.5 g NaCl

Thus, we can say 58.5 g of NaCl contains these units.

While solving problems related to mole concept, always keep in mind the following points

- Convert all in same unit, i.e. mol/atoms/mass and then compare.
- Moles $\propto$ molecules

Example 8. *The number of atoms present in one mole of an element is equal to Avogadro number. Which of the following element contains the greatest number of atoms?* ***(NCERT Exemplar)***

(a) 4 g He (b) 46 g Na (c) 0.4 g Ca (d) 12 g He

Sol. (*d*) For comparing number of atoms, first we calculate the moles as all are monoatomic and hence,
moles $\times N_A =$ number of atoms.

$$\text{Moles of 4g He} = \frac{4}{4} = 1 \text{ mol}$$

$$\text{46g Na} = \frac{46}{23} = 2 \text{ mol}$$

$$\text{0.40 g Ca} = \frac{0.40}{40} = 0.1 \text{ mol}$$

$$\text{12g He} = \frac{12}{4} = 3 \text{ mol}$$

Therefore, number of atoms are highest in 12g He as it has the maximum number of moles.

Molar Mass and Molar Volume

Molar mass of an element is defined as mass of 1 mole of that element, i.e. mass of 6.023×10^{23} entities or particles of that element.

e.g. molar mass of oxygen = 32 g/mol. It means that 6.023×10^{23} molecules of oxygen weigh 32 g, similarly, molar mass of Na = 23 g/mol, that means 6.023×10^{23} monoatomic molecules of Na weigh 23 g.

When molar mass is divided by density, **molar volume** is obtained. It is the volume of one mole of a substance.

Since molar volumes of solids and liquids do not vary much with temperature and pressure, these can be calculated easily by the following relation

$$\text{Molar volume} = \frac{\text{Molar mass}}{\text{Density}}$$

The molar volumes of gases, change considerably with temperature and pressure. For an ideal gas, the molar volume at 0°C and 1 atm pressure is 22.4 L.

Example 9. *A* 10 mg *effervescent tablet containing sodium bicarbonate and oxalic acid releases* 0.25 mL *of* CO_2 *at* $T = 298.15$ K *and* $p = 1$ bar. *If molar volume of* CO_2 *is* 25.0 L *under such condition, what is the percentage of sodium bicarbonate in each tablet ?*

[Molar mass of $NaHCO_3 = 84 \text{ g mol}^{-1}$*]*

(JEE Main 2019)

(a) 8.4 (b) 0.84 (c) 16.8 (d) 33.6

Sol. (a) $\underset{2\text{ mol}}{2NaHCO_3} + \underset{1\text{ mol}}{H_2C_2O_4} \longrightarrow \underset{2\text{ mol}}{2CO_2} + Na_2C_4O_4 + 2H_2O$

$\Rightarrow$ In the reaction, number of mole of CO_2 produced.

$$n = \frac{pV}{RT} = \frac{1 \text{ bar} \times 0.25 \times 10^{-3} \text{ L}}{0.082 \text{ L atm K}^{-1}\text{mol}^{-1} \times 298.15 \text{ K}}$$
$$= 1.02 \times 10^{-5} \text{ mol}$$

Now, one mole of CO_2 is produced by one mole of $NaHCO_3$.

$\therefore$ The number of moles of $NaHCO_3$ in the given reaction

= number of moles of CO_2

$$\text{Number of mole of } NaHCO_3 = \frac{\text{Weight of } NaHCO_3}{\text{Molecular mass of } NaHCO_3}$$

$\therefore$ $w_{NaHCO_3} = 1.02 \times 10^{-5} \times 84 \times 10^3$ mg

$= 0.856$ mg

$\Rightarrow$ $NaHCO_3 \% = \dfrac{0.856}{10} \times 100$

$= 8.56\% \approx 8.4$

Mole in Relation to Concentration of a Solution

Concentration of a solution tells us how much solute is dissolved in a given quantity of solution. Terms for the measurement of concentration of various components of a solution are given below

1. Mole Fraction (χ)

It is defined as "the ratio of the number of moles of a particular component to the number of moles of the solution". It is denoted by symbol χ (chi).

Mole fraction of solute (χ_{solute})

$$= \frac{\text{moles of solute}}{\text{total moles of solution, i.e. [solute + solvent]}}$$

$$\chi_{\text{solute}} = \frac{n_{\text{solute}}}{n_{\text{solute}} + n_{\text{solvent}}}$$

Similarly, mole fraction of solvent (χ_{solvent})

$$= \frac{\text{moles of solvent}}{\text{total moles of solution, i.e. [solute + solvent]}}$$

$$\chi_{\text{solvent}} = \frac{n_{\text{solvent}}}{n_{\text{solute}} + n_{\text{solvent}}}$$

But, we all know that being fraction

$$\chi_{\text{solute}} + \chi_{\text{solvent}} = 1, \text{ so, } \frac{\chi_{\text{solute}}}{\chi_{\text{solvent}}} = \frac{n_{\text{solute}}}{n_{\text{solvent}}}$$

2. Molarity (*M*)

It is defined as the number of moles of solute per litre of solution or the same numerically, as the number of milligram molecules per millilitre of solution. The molarity is usually designated by *M*,

Thus, molarity is given as

$$\text{Molarity, } M = \frac{\text{moles of solute}}{\text{volume of solution (in L)}}$$

e.g. If the molarity of H_3PO_4 is 0.18, it means a concentration corresponding to 0.18 mol of H_3PO_4 per litre of solution.

If specific gravity is given,

$$\text{Molarity} = \frac{\text{specific gravity} \times \text{\% strength} \times 10}{\text{molecular weight}}$$

3. Molality (*m*)

It is defined as the number of moles of solute dissolved in 1000 g of the solvent. It is designated by *m*.

$$\text{Molality, } m = \frac{\text{moles of solute}}{\text{weight of solvent (in g)}} \times 1000$$

Molality is independent of temperature, as it depends only upon the mass which does not vary with temperature.

4. Formality (*F*)

It is practically same as molarity. The difference lies in the fact that formality is used for ionic solutes.

$$\text{Thus, formality, } F = \frac{\text{gram formula weight}}{\text{volume in litre}}$$

5. Normality (*N*)

It is defined as the number of gram-equivalents of solute per litre of solution or as the number of mg-equivalents of a substance per millilitre of solution. e.g. 0.12 N H_2SO_4 means a solution which contains 0.12 gram-equivalent of H_2SO_4 per litre of solution.

This also means that each millilitre of this solution can react, with for example, 0.12 mg eq. of CaO or with 0.12 mg-eq. of Na_2CO_3.

$$\text{Thus, normality, } N = \frac{\text{g -equivalent of solute}}{\text{volume of solution (in L)}}$$

or

$$= \frac{\text{g-equivalent of solute}}{\text{volume of solution (in mL)}} \times 1000$$

If specific gravity is known, normality is calculated as

$$\text{Normality} = \frac{\text{specific gravity} \times \text{\% strength} \times 10}{\text{equivalent weight}}$$

Useful Formulae Related to Moles and Concentration Terms

- Number of millimoles = molarity $\times$ volume in mL

  or $\text{Moles} = \dfrac{\text{millimoles}}{1000}$
- Number of equivalents of a substance

  $$= \frac{\text{weight (in g)}}{\text{equivalent weight}}$$
- Number of milli-equivalent (M. eq.)

  $$= \frac{1000 \times \text{weight (in g)}}{\text{equivalent weight}}$$

  $$= \text{normality} \times \text{volume (mL)}$$
- Normality = molarity $\times$ balance factor (y).

  where, y = acidity/basicity/number of replaceable H-atoms/change in oxidation number.
- Number of equivalents = $y \times$ number of moles
- Number of milli-equivalents = $y \times$ number of millimoles
- Mole fraction of solute in the solution = $\dfrac{n}{n + N}$

  where, n = moles of solute, N = moles of solvent
- For very dilute solutions,

  $$\text{mole fraction of solute in solution} = \frac{Mm}{1000}$$

  where, M = molecular weight of solvent, m = molality
- Density of solution

  $$= \text{molarity}\left(\frac{1}{\text{molality}} + \frac{\text{molecular weight of solute}}{1000}\right)$$
- Percent by weight of solute in solution

  $$= \frac{\text{weight of solute (in g)} \times 100}{\text{weight of solution (in g)}}$$
- Percent by volume of solute in solution

  $$= \frac{\text{weight of solute (in g)} \times 100}{\text{volume of solution (mL)}}$$
- 1 mL, 1 N $KMnO_4 \equiv$ 1 mL, 1NH_2O_2
- 1 mL, 1 N $Na_2S_2O_3 \equiv$ 1 mL, 1N I_2 solution
- 1 volume H_2O_2 = 0.1785 NH_2O_2
- If two compounds *A* and *B* neutralise completely each other then milli-equivalents of *A* = milli-equivalents of *B*
- Molecular weight = 2 $\times$ vapour density (for gaseous phase).

Example 10. *Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is* 0.040. **(NCERT)**

(a) 2.314 (b) 1.123 (c) 3.139 (d) 1.762

Sol. (*a*) Molarity is defined as the moles of solute (ethanol) in 1 L of the solution.

1 L of ethanol solution (as it is diluted) = 1 L of water

$$\text{Number of moles of } H_2O \text{ in 1 L water} = \frac{1000 \text{ g}}{18} = 55.55 \text{ moles}$$

For a binary solution (binary solution contains two components)

$$X_1 + X_2 = 1$$

Hence,
$$X_{H_2O} = 1 - X_{C_2H_5OH}$$
$$X_{H_2O} = 1 - 0.040 = 0.96$$
$$X_{H_2O} = \frac{n_{H_2O}}{n_{H_2O} + n_{C_2H_5OH}}$$
$$0.96 = \frac{55.55}{55.55 + n_{C_2H_5OH}}$$
$$53.328 + 0.96 n_{C_2H_5OH} = 55.55$$
$$0.96 n_{C_2H_5OH} = 55.55 - 53.328 = 2.222$$
$$n_{C_2H_5OH} = \frac{2.222}{0.96} = 2.3145 \text{ mol}$$

Example 11. *Mole fraction of* I_2 *in benzene is* 0.1. *The molality of* I_2 *in* C_6H_6 *is*

(a) 1.42 (b) 3.205

(c) 2.06 (d) 1.86

Sol. (*a*) $\text{Mole fraction of } I_2 = \left[\frac{n}{n+N}\right] = 0.1$ ...(i)

$\text{Mole fraction of benzene} = \left[\frac{N}{n+N}\right] = 0.9$ ...(ii)

where, n and N are moles of I_2 and C_6H_6 in solution respectively.

So, molality,
$$m = \frac{1000 \times \text{mole fraction of } I_2}{(1 - \text{mole fraction of } I_2) \times M}$$
$$= \frac{1000 \times 0.1}{0.9 \times 78} = 1.42 \text{ m}$$

Example 12. *One mole of any substance contains* 6.022×10^{23} atoms/molecules. *Number of molecules of* H_2SO_4 *present in* 100 mL 0.02 M H_2SO_4 *solution is* **(NCERT Exemplar)**

(a) 12.044×10^{20} molecules (b) 6.022×10^{23} molecules

(c) 1×10^{23} molecules (d) 12.044×10^{23} molecules

Sol. (*a*) Number of millimoles of H_2SO_4
$$= \text{molarity} \times \text{volume in mL}$$
$$= 0.02 \times 100 = 2 \text{ millimol}$$
$$= 2 \times 10^{-3} \text{ mol}$$

$$\text{Number of molecules} = \text{Number of moles} \times N_A$$
$$= 2 \times 10^{-3} \times 6.022 \times 10^{23}$$
$$= 12.044 \times 10^{20} \text{ molecules}$$

Example 13. *Sulphuric acid reacts with sodium hydroxide as follows*

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

When 1 L *of* 0.1 M *sulphuric acid solution is allowed to react with* 1 L *of* 0.1 M *sodium hydroxide solution, the amount of sodium sulphate formed and its molarity in the solution obtained are* **(NCERT Exemplar)**

(a) 0.1 mol L^{-1}, 3.55 g (b) 0.025 mol L^{-1}, 7.10 g

(c) 0.01 mol L^{-1}, 5.33 g (d) 0.25 mol L^{-1}, 5.33 g

Sol. (*b*) Moles of $H_2SO_4 = 1\text{L} \times 0.1\text{M} = 0.1$ mol

$$\text{Moles of NaOH} = 1\text{L} \times 0.1\text{M} = 0.1 \text{ mol}$$

$$\begin{array}{cccc} H_2SO_4 + & 2NaOH & \longrightarrow Na_2SO_4 + & 2H_2O \\ 1\text{mol} & 2 \text{ mol} & 1 \text{ mol} & \\ 0.1\text{mol} & 0.1\text{mol} & 0.05 \text{ mol} & \end{array}$$

$\therefore$ Molarity mass of $Na_2SO_4 = 0.05 \times 142$

$= 7.1\text{g}$ ($\therefore$ Molar mass of $Na_2SO_4 = 142$ mol g mol^{-1})

$$\text{Molarity} = \frac{\text{moles of } Na_2SO_4}{\text{volume of solution}} = \frac{0.05}{1+1} = 0.025 \text{ mol L}^{-1}$$

Example 14. *If* 500 mL *of a* 5 M *solution is diluted to* 1500 mL, *what will be the molarity of the solution obtained?* **(NCERT Exemplar)**

(a) 1.5 M (b) 1.66 M (c) 0.017 M (d) 1.59 M

Sol. (*b*) For dilution, $\underset{\text{(Before dilution)}}{M_1V_1} = \underset{\text{(After dilution)}}{M_2V_2}$

$$500 \times 5\text{M} = 1500 \times M$$
$$M = \frac{5}{3} = 1.66 \text{ M}$$

Example 15. 6.023×10^{22} *molecules are present in* 10 g *of a substance 'x'. The molarity of a solution containing* 5 g *of substance 'x' in 2 L solution is* ... $\times 10^{-3}$. **(JEE Main 2020)**

Sol. (25) Given, 6.023×10^{22} molecules are present in 10 g substance. So,

$$\text{Number of moles} = \frac{\text{Number of molecules}}{6.023 \times 10^{23}} = \frac{\text{Mass (given)}}{\text{Molar mass}}$$

$$\text{Molar mass} = \frac{10 \times 6.023 \times 10^{23}}{6.023 \times 10^{22}} = 100 \text{ g/mol}$$

We know that,

$$\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution } (l)}$$

[Given, molarity of solution containing = 5 g of subtance and volume = 2L]

$$= \frac{(5/100)}{2} = \frac{5}{100 \times 2} = 0.025 = 25 \times 10^{-3}$$

Example 16. A 100 mL *solution was made by adding* 1.43 g *of* $Na_2CO_3 \cdot xH_2O$. *The normality of the solution is* 0.1 N. *The value of x is* .... **(JEE Main 2020)**

Sol. (*10*) (Atomic mass of Na = 23, C = 12, O = 16)

Molar mass of $Na_2CO_3 \cdot xH_2O = 23 \times 2 + 12 + 48 + 18x$

$$= 46 + 12 + 48 + 18x = (106 + 18x)$$

Equivalent weight of $Na_2CO_3 \cdot xH_2O$

$$= \frac{\text{Molar mass}}{n \text{ factor}}$$

$$= \frac{M}{2} = (53 + 9x)$$

(here, M = molar mass and n factor = 2)

$$\text{Gram equivalent} = \frac{\text{Weight}}{\text{Equivalent weight}}$$

(Given, weight of $Na_2CO_3 \cdot xH_2O = 1.43$ g)

Hence, gram equivalent of

$$Na_2CO_3 \cdot xH_2O = \frac{1.43}{53 + 9x}$$

$$\text{Normality} = \frac{\text{Gram equivalent weight}}{V_{\text{litre}}}$$

$$0.1 = \frac{\frac{1.43}{53 + 9x}}{0.1}$$

As, volume $= 100 \text{ mL} = 0.1 \text{ L}$

So, $$10^{-2} = \frac{1.43}{53 + 9x} \Rightarrow 53 + 9x = 143$$

$$9x = 90 \quad \Rightarrow \quad x = 10.00$$

Chemical Equations and Stoichiometry

Calculations based on the quantitative relationship between the reactants and products are referred as **stoichiometry**. The goal of these calculations is to predict the relationship between the amounts of the reactants and products of chemical reaction.

Word stoichiometry is taken from Greek words *stoichion* = element; *metron* = measure. Solving of stoichiometric problems require

- a firm grasp of mole concept,
- balancing chemical equations and
- care in consistent use of units.

The numerals used to balance a chemical equation are called **stoichiometric coefficients.**

Always remember, never lose your goal while solving such problems. Use mole concept carefully and then also find type of stoichiometric problem before solving.

To solve out stoichiometric problems, follow a right approach that is based on a few relatively easy ground rules, which can be understood by following solved examples.

Example 17. *Predict the amount of oxygen that must be inhaled to digest (burn)* 100 g *of sugar. The sugar burns according to the following equation.*

$$C_{12}H_{22}O_{11}(s) + 12\,O_2(g) \longrightarrow 12\,CO_2(g) + 11H_2O(l)$$

(a) 95.28 g
(b) 112. 28 g
(c) 125 g
(d) 35.12 g

Sol. (*b*) Follow the following steps to solve such type of problems

Step I	Identify the goal of the problem. e.g. in the given problem.	Find out how many grams of O_2 are consumed when 100 g of sugar is burned
Step II	Write down the key elements of the problem, or draw a simple picture that summarizes the key information in the problem.	We have 100 g of sugar. Sugar has formula, $C_{12}H_{22}O_{11}$. We know molar equation.
Step III	Try to do what can be done.	We can convert the mass (100 g) in moles, as we know molar mass of $C_{12}H_{22}O_{11}$, i.e. $\frac{100}{342} = 0.29$ mol
Step IV	Never lose sight of what you have accomplished.	What else we can do? We know the molar relationship of O_2 and sugar in balanced chemical equation. We can calculate the number of moles of oxygen needed to burn 0.29 mole of sugar as $0.29 \times \frac{12}{1} = 3.48$ mole of O_2 by unitary method.
Step V	Don't try to work the problem in your head. Write down all the intermediate steps.	Now, we can calculate the weight of O_2 as 3.48×32 $= 112.28$ g O_2

Limiting Reactant or Limiting Reagent

If a reaction involves two or more reactants, the reactant that is consumed first is called limiting reagent or limiting reactant as it limits the amount of product formed. The other reactant present in the larger amounts is called **excess reagent**.

Remember Limiting reagent can be easily identified from the molar ratios of reactants. Infact, the reactant with lower molar ratio limits the amount of product.

Calculation of Limiting Reagent

This can be best understood by considering the following sample problem.

Example 18. *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 the amount of C formed is* ***(NCERT Exemplar)***

(a) 3 mol (b) 4 mol (c) 5.5 mol (d) 4.5 mol

Sol. (*d*) $2A + 4B \rightarrow 3C + 4D$

According to the above equation, 2 moles of 'A' require 4 moles of 'B' for the reaction.

Hence, for 5 moles of 'A' the moles of 'B' required

$$= 5\text{moles of } A \times \frac{4 \text{ moles of } B}{2 \text{ moles of } A} = 10 \text{ mol } B$$

But we have only 6 moles of '*B*' hence, '*B*' is the limiting reagent. So amount of '*C*' formed is determined by amount of '*B*'.
Since 4 moles of '*B*' give 3 moles of '*C*'.
Hence, 6 moles of '*B*' will give

$$6 \text{ moles of } B \times \frac{3 \text{ moles of } C}{4 \text{ moles of } B} = 4.5 \text{ moles of } C$$

Example 19. *The mass of ammonia in grams produced when* 2.8 kg *of dinitrogen quantitatively reacts with* 1 kg *of dihydrogen is* .... ***(JEE Main 2020)***

Sol. (*3400*) To find the limiting reagent, we have to follow the following sequence of steps

***Step* I** Consider the equation,

$$\underset{\substack{1 \text{ mol} \\ 28 \text{ g}}}{N_2(g)} + \underset{\substack{3 \text{ mol} \\ 6 \text{ g}}}{3H_2(g)} \longrightarrow \underset{\substack{2 \text{ mol} \\ 34 \text{ g}}}{2NH_3(g)}$$

From the above data, it is clear that N_2 and H_2 react in the simplest ratio 7 : 1.5 (28 : 6) to form NH_3.

***Step* II** Now, compare the ratio of their amounts given, i.e. $2.8 \times 10^3 : 1.00 \times 10^3$ or 2.8 : 1. As both the ratios differ, this proves the presence of limiting reagent.

You can easily identify that If N_2 and H_2 are required in the ratio = 7:1.5 and their availability ratio is only 2.8 : 1 then N_2 is the limiting reagent in the reaction.

Now, $\quad 2800 \text{ g } N_2$ will reacts with $\dfrac{2800 \times 6}{28}$

$$= 600 \text{ g } H_2$$

This confirns that, N_2 is the limiting reagent and H_2 is in excess. N_2 limits the amount of ammonia produced.

Amount of H_2 remain unreacted

$$= 1000 - 600$$
$$= 400 \text{ g}$$

***Step* III** Once you have identified the limiting reagent, calculate the amount of the product formed as,

$$28 \text{ g } N_2 \text{ produces } 34 \text{ g } NH_3$$
$$1 \text{ g } N_2 \text{ produces } \frac{34}{28} \text{ g } NH_3$$
$$\therefore \; 2800 \text{ g } N_2 \text{ will produce } \frac{34}{28} \times 2800 = 3399.99 \text{ g } NH_3$$
$$\approx 3400 \text{ g } NH_3$$

Thus, the amount of NH_3 produced by the reaction of 2.80 kg N_2 and 1 kg H_2 = 3400 g

Stoichiometric Problems of Different Kinds

Depending on the units, stoichiometric problems may be classified into the following relationship

(i) mass-mass relationship, i.e. gravimetric analysis
(ii) mass-volume relationship
(iii) volume-volume relationship

I. Calculations Based on Mass-Mass Relationship

For mass-mass related calculations, following methods can be used

1. Mole Method (Based on Mole Concept)

It involves the following steps

***Step* I** Write the complete and balanced chemical reaction as per the information given in the problem.

***Step* II** The stoichiometric coefficients in the balanced chemical reactions represent the relative number of moles of the different reaction components. Relate the amounts of the reaction components concerned, with the help of mole.

***Step* III** Calculate the unknown amount of substance using unitary method.

Example 20. *Calculate the weight of carbon dioxide formed by complete combustion of* 1.5 g *ethane.*

(a) 22 g (b) 2.2 g (c) 44 g (d) 4.4 g

Sol. (*d*) $\underset{\substack{1 \text{ mol} \\ = 30 \text{ g}}}{C_2H_6} + \frac{7}{2} O_2 \longrightarrow \underset{\substack{2 \text{ mol} \\ = 2 \times 44 \text{ g}}}{2CO_2} + 3H_2O$

$\because$ 30 g C_2H_6 produces 2×44 g CO_2 on complete combustion

$\therefore$ 1.5 g C_2H_6 will produce $\dfrac{2 \times 44}{30} \times 1.5 = 4.4 \text{ g } CO_2$

2. Principle of Atomic Conservation (POAC) Method

This is another approach of solving problems by mole method, but without balancing the reaction. Here, the condition must be, *as the reactions are balanced by conserving the atoms of each element, the mole of atoms of each element in the reactant side should be equal to that in the product side*.

On applying the conservation of atoms of each element with the help of mole, we may get relations needed to solve the problem.

e.g. The sample problem 21 given above may be solved by POAC method as,

Moles of C-atoms in C_2H_6 = moles of C-atoms in CO_2

or $\quad 2 \times \text{mole of } C_2H_6 = 1 \times \text{mole of } CO_2$

or $\quad 2 \times \dfrac{1.5}{30} = 1 \times \dfrac{w}{44}$

$\therefore \quad w = 4.4 \text{ g}$

3. Factor-Label Method

In this method, the required amount is determined directly by

- converting the given amount of substance into its moles,
- then relating the moles of given substance with the moles of required substance as per balanced chemical equation or atomic conservation and then finally,

- converting the moles into the amount of the required substance.

For example, the given solved example 20 may be solved by factor label method as,

Procedure

$$\text{Mass of } C_2H_6 \text{ (given)} \xrightarrow[\text{into}]{\text{Convert}} \text{Moles of } C_2H_6$$

$$\xrightarrow[\text{moles of } CO_2]{\text{Relate with}} \text{moles of } CO_2$$

$$\xrightarrow[\text{into}]{\text{Convert}} \text{Mass of } CO_2 \text{ (required)}$$

Application

$$\text{Mass of } CO_2 = (1.5\text{g } C_2H_6) \times \left(\frac{1 \text{ mol } C_2H_6}{30 \text{ g } C_2H_6}\right) \times \left(\frac{2 \text{ mol } CO_2}{1 \text{ mol } C_2H_6}\right) \times \left(\frac{44\text{g } CO_2}{1 \text{ mol } CO_2}\right)$$

$$= 4.4 \text{ g}$$

4. Equivalent Method (Based on Equivalent Weight Concept)

In a reaction, the number of gram-equivalents of each reactants reacted results to same number of gram-equivalents of each products. e.g. The example 20 may be solved by equivalent concept as,

$$\text{no. of g-equivalents of } C_2H_6 = \text{no. of g-equivalents of } CO_2$$

or
$$\frac{1.5}{\left(\frac{30}{14}\right)} = \frac{w}{\left(\frac{44}{7}\right)}$$

$$w = 4.4 \text{ g}$$

$\therefore$ The factors 14 and 7 are the changes in oxidation states of carbon for C_2H_6 and CO_2 respectively.

Example 21. *How many grams of oxygen are required to burn* 570 g *of octane completely?*

(a) 570 g (b) 5 mol
(c) 25×32 g (d) 2000 g

Sol. (*d*) Balanced equation,

$$\underset{\substack{2 \text{ mol} \\ = 2 \times 114}}{2C_8H_{18}} + \underset{\substack{25 \text{ mol} \\ = 25 \times 32}}{25O_2} \longrightarrow 16\,CO_2 + 18H_2O$$

By direct method For burning 2×114 g of octane, oxygen required $= 25 \times 32$ g

For burning 1g octane, oxygen required $= \frac{25 \times 32}{2 \times 114}$ g

Thus, for burning 570 g of octane, oxygen required

$$= \frac{25 \times 32}{2 \times 114} \times 570 \text{ g} = 2000 \text{ g}$$

By mole method Number of moles of octane in 570 g

$$= \frac{570}{114} = 5.0$$

For burning 2.0 moles of octane, oxygen required

$$= 25 \text{ mol} = 25 \times 32 \text{ g}$$

For burning 5 moles octane, oxygen required

$$= \frac{25 \times 32}{2.0} \times 5.0\text{g} = 2000 \text{ g}$$

Calculation Involving Per cent Yield

In general, when a reaction is carried out in the laboratory we do not obtain actually the theoretical amount of the product. The amount of the product that is actually obtained is called the **actual yield**.

Knowing the actual yield and theoretical yield the per cent yield can be calculated by using the formula,

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Example 22. *For the reaction,*

$$CaO + 2HCl \longrightarrow CaCl_2 + H_2O$$

1.23 g *of* CaO *is reacted with excess of hydrochloric acid and* 1.85 g *of* $CaCl_2$ *is formed. What is the per cent yield ?*

(a) 76.1 (b) 86.3 (c) 95.1 (d) None of these

Sol. (*a*) The balanced equation is

$$\underset{\substack{1 \text{ mol} \\ 56 \text{ g}}}{CaO} + 2HCl \longrightarrow \underset{\substack{1 \text{ mol} \\ 111 \text{ g}}}{CaCl_2} + H_2O$$

56 g of CaO produces $CaCl_2 = 111$ g

1.23 g of CaO will produce $CaCl_2 = \frac{111}{56} \times 1.23 = 2.43$ g

Thus, theoretical yield = 2.43 g

Actual yield = 1.85 g

$$\text{Percent yield} = \frac{1.85}{2.43} \times 100 = 76.1$$

Calculations Involving Per cent Purity

In this type of calculations depending upon the mass of the product, the equivalent amount of the reactant present can be determined with the help of a chemical equation.

Knowing the actual amount of the reactant taken and the amount calculated with the help of a chemical equation, the percentage purity can be determined as

$$\% \text{ purity} = \frac{\text{Weight of reactant required}}{\text{Weight of reactant taken}} \times 100$$

Example 23. 0.2415 g *sample of a mixture of* $MgSO_4 \cdot 7H_2O$ *and* $MgCl_2 \cdot 6H_2O$ *containing inert impurities was subjected to suitable treatment, as a result of which there were obtained* 0.1362 g *of* $BaSO_4$ *and* 0.1129 g *of* $Mg_2P_2O_7$. *Calculate the percentage impurity in the original mixture.*

(a) 5% (b) 4%
(c) 3% (d) 8%

Sol. (*b*) Let the original sample contains *a* mol of $MgSO_4 \cdot 7H_2O$ and *b* mol of $MgCl_2 \cdot 6H_2O$. As $BaSO_4$ will be formed only from $MgSO_4 \cdot 7H_2O$ by the conservation of SO_4^{2-} ion

So, mole of $BaSO_4$ will formed as

Now from the question, $a \times 233 = 0.1362$ ... (i)

Now, as $Mg_2P_2O_7$ will be formed from both $MgSO_4 \cdot 7H_2O$ and $MgCl_2 \cdot 6H_2O$, by the conservation of Mg atom.

So, mole of $Mg_2P_2O_7$ formed as $a/2 + b/2$.

Now from question, $\left(\frac{a}{2}+\frac{b}{2}\right) \times 222 = 0.1129$ ...(ii)

From Eq. (i) and (ii), $a = 0.000585$ and $b = 0.000432$

Hence, the composition of original mixture

$MgSO_4 \cdot 7H_2O = 0.000585$ mol $= 0.000585 \times 246 = 0.1439$ g

$$\text{Percentage purity} = \frac{0.1439}{0.2415} \times 100 = 59.59\%$$

$MgCl_2 \cdot 6H_2O = 0.000432$ mol $= 0.000432 \times 203 = 0.0877$ g

$$\text{Percentage purity} = \frac{0.0877}{0.2415} \times 100 = 36.31\%$$

$$\text{Impurity} = 100 - (59.59 + 36.31) = 4.1\%$$

Analysis of Mixtures

In problems related to analysis of mixtures one of the components is supposed to be x g and the other will be the difference from the total.

Balanced chemical equations for the reactions of both the components are now written and the total amount of the common product produced by the components of the mixture is calculated. It is equated with the data given and the unknown factors are, thus, worked out.

Example 24. *A solid mixture* (5.0 g) *consisting of lead nitrate and sodium nitrate was heated below* $600°$ C *until the mass of the residue was constant. If the loss in mass is* 28.0 *percent, find the amount of lead nitrate in the mixture.*

(a) 1.67 (b) 3.32 (c) 1.40 (d) 5.00

Sol. (*b*) Let the amount of $NaNO_3$ in the mixture be x g.

$\therefore$ The amount of $Pb(NO_3)_2$ in the mixture $= (5.0 - x)$ g

$$\underset{(2 \times 85)\text{ g}}{2NaNO_3} \xrightarrow{\text{Heat}} 2NaNO_2 + \underset{32\text{ g}}{O_2}$$

$$\underset{\substack{(2\times 331)\text{ g} \\ 662\text{ g}}}{2Pb(NO_3)_2} \xrightarrow{\text{Heat}} 2PbO + \underbrace{\underset{(4\times 46)\text{ g}}{4NO_2} + \underset{32\text{ g}}{O_2}}_{216\text{ g}}$$

170 g of $NaNO_3$ evolve oxygen $= 32$ g

x g of $NaNO_3$ evolve oxygen $= \frac{32}{170}\, x$ g

662 g of $Pb(NO_3)_2$ evolve gases $= 216$ g

$(5.0 - x)$ g of $Pb(NO_3)_2$ evolve gases $= \frac{216}{662} \times (5.0 - x)$ g

$$\therefore \quad \text{Total loss} = \frac{32}{170}(x) + \frac{216}{662}(5.0 - x)$$

$$\text{Loss given in the problem} = \frac{28}{100} \times 5 = 1.4\text{ g}$$

$$\therefore \quad \text{Total loss} = \frac{32}{170}(x) + \frac{216}{662}(5.0 - x) = 1.4$$

On solving, $x = 1.676$ g

Thus, mass of $NaNO_3 = 1.676$ g

Mass of $Pb(NO_3)_2 = (5.0 - 1.676)\text{g} = 3.324$ g

Example 25. *In order to oxidise a mixture of one mole of each of* FeC_2O_4, $Fe_2(C_2O_4)_3$, $FeSO_4$ *and* $Fe_2(SO_4)_3$ *in acidic medium, the number of moles of* $KMnO_4$ *required is*

(JEE Main 2019)

(a) 2 (b) 1

(c) 3 (d) 1.5

Sol. (*a*) The oxidation of a mixture of one mole of each of FeC_2O_4, $Fe_2(C_2O_4)_3$, $FeSO_4$ and $Fe_2(SO_4)_3$ in acidic medium with $KMnO_4$ is as follows

$$FeC_2O_4 + KMnO_4 \longrightarrow Fe^{3+} + CO_2 + Mn^{2+} \quad \text{...(i)}$$

$$Fe_2(C_2O_4)_3 + KMnO_4 \longrightarrow Fe^{3+} + CO_2 + Mn^{2+} \quad \text{...(ii)}$$

$$FeSO_4 + KMnO_4 \longrightarrow Fe^{3+} + SO_4^{2-} + Mn^{2+} \quad \text{...(iii)}$$

Change in oxidation number of Mn is 5. Change in oxidation number of Fe in (i), (ii) and (iii) are $+3, +6, +1$, respectively.

$$n_{eq}KMnO_4 = n_{eq}[FeC_2O_4 + Fe_2(C_2O_4)_3 + FeSO_4]$$

$$n \times 5 = 1\times 3 + 1 \times 6 + 1 \times 1 \Rightarrow n = 2$$

Reactions in Succession

In such problems, the amount of any one of the reaction component belonging to a reaction is to be determined. It is done with the help of the given amount of some other component belonging to some other reaction. Thus, here calculation are done with the help of common components belonging to different reactions.

II. Calculations Involving Mass-Volume Relationship

In such calculations masses of reactants are given and volume of the product is required and *vice-versa*. As 1 mole of a gas occupies 22.4 L volume at STP thus, mass of a gas can be related to volume according to the following gas equation

$$pV = nRT \quad \text{or} \quad pV = \frac{w}{m}RT$$

Example 26. *A vessel contains* 1.6 g *of dioxygen at STP* (273.15 K, 1 atm *pressure). The gas is now transferred to another vessel at constant temperature, where pressure becomes half of the original pressure. The volume of the new vessel, and number of molecules of dioxygen, are respectively.* ***(NCERT Exemplar)***

(a) $2.24, 6.02 \times 10^{23}$ (b) $1.12, 3.01 \times 10^{22}$

(c) $2.24, 3.01 \times 10^{21}$ (d) $2.24, 3.01 \times 10^{22}$

Sol. (*d*) Given, $p_1 = 1$atm $T_1 = 273.15$ K, $V_1 = ?$

32 g of oxygen occupies 22.4 L of volume at STP.

Hence, 1.6 g of oxygen will occupy,

$$1.6\text{ g oxygen} \times \frac{22.4\text{ L}}{32\text{ g oxygen}} = 1.12\text{ L volume}$$

$$V_1 = 1.12\text{L}$$

$$p_2 = \frac{p_1}{2} = \frac{1}{2} = 0.5\text{ atm.}$$

$$V_2 = ?$$

According to Boyle's law,

$$p_1V_1 = p_2V_2$$

$$\therefore \quad V_2 = \frac{p_1 \times V_1}{p_2} = \frac{1 \text{ atm} \times 1.12 \text{ L}}{0.5 \text{ atm}} = 2.24 \text{ L}$$

Now, number of molecules of oxygen in the vessel

$$= \frac{6.022 \times 10^{23} \times 1.6}{32} = 3.011 \times 10^{22}$$

III. Calculations Based on Volume-Volume Relationship

Sub calculations further be classified are as follows

For Gases (Gas Analysis or Eudiometry)

These calculations are based on various gas laws given in initial topics of this chapter. Let us go through some examples.

Example 27. *What volume of air having 21% oxygen by volume is required to completely burn 1 kg of carbon containing 100% combustible substances?*

(a) 8.9985 L (b) 8.8885 L
(c) 8.7785 L (d) 8.9898 L

Sol. (*b*) Combustion of carbon may be given as,

$$\underset{\substack{1 \text{ mol} \\ 12 \text{ g}}}{C(s)} + \underset{\substack{1 \text{ mol} \\ 1 \text{ mol}}}{O_2(g)} \longrightarrow CO_2(g)$$

$\because$ 12 g carbon requires 1 mole of O_2 for complete combustion.

$\therefore$ 1000 g carbon will require $\frac{1}{12} \times 1000$ mol O_2 for combustion, i.e. 83.33 mol O_2

Volume of O_2 at NTP $= 83.33 \times 22.4$ L $= 1866.592$ L

$\because$ 21 L O_2 is present in 100 L air

$\therefore$ 1866.592 L O_2 will be present in $\frac{100}{21} \times 1866.592$ L air

$$= 8888.5 \text{ L} = 8.8885 \times 10^3 \text{ L}$$

Example 28. *At 300 K and 1 atmospheric pressure, 10 mL of a hydrocarbon required 55 mL of O_2 for complete combustion and 40 mL of CO_2 is formed. The formula of the hydrocarbon is* ***(JEE Main 2019)***

(a) C_4H_7Cl (b) C_4H_6 (c) C_4H_{10} (d) C_4H_8

Sol. (*b*) In eudiometry,

$$C_xH_y + \left(x + \frac{y}{4}\right) O_2 \xrightarrow[1 \text{ atm}]{300 \text{ K}} x\, CO_2 + \frac{y}{2} H_2O$$

1 mol	$\left(x + \frac{y}{4}\right)$ mol	x mol
1 mL	$\left(x + \frac{y}{4}\right)$ mL	x mL
10 mL	$\left(x + \frac{y}{4}\right) \times 10$ mL	$10x$ mL

Given, (i) $V_{CO_2} = 10x = 40$ mL $\Rightarrow x = 4$

(ii) $V_{O_2} = 10\left(x + \frac{y}{4}\right)$ mL $= 55$ mL

$$\Rightarrow 10\left(4 + \frac{y}{4}\right) = 55 \qquad [\because x = 4]$$

$$\Rightarrow 40 + \frac{y \times 10}{4} = 55$$

$$\Rightarrow y \times \frac{10}{4} = 15 \Rightarrow y = 15 \times \frac{4}{10} = 6$$

So, the hydrocarbon (C_xH_y) is C_4H_6.

For Solutions (Volumetric Analysis or Titrations)

The calculations based on volume relationships for solutions usually involve one of the following titrations

1. Simple Titrations

In this, a solution of substance A of unknown concentration is made to react with a solution of B with known concentration.

The two are reacted in such a way that the volume of B required to completely react with A can be found out by using some indicators. Thus, the concentration of A may be calculated.

If the normality of B is N_1 and its volume used is V_1 then the gram-equivalents of B reacted $= N_1V_1$.

According to the law of equivalents, the gram-equivalents of A would be equal to that of B.

$\therefore$ Gram equivalents of $A = N \times V$

If the volume of A is V then the normality of A would be

$$\frac{N_1V_1}{V}$$

Example 29. *The ammonia (NH_3) released on quantitative reaction of 0.6 g urea (NH_2CONH_2) with sodium hydroxide (NaOH) can be neutralised by* ***(JEE Main 2020)***

(a) 100 mL of 0.2 N HCl (b) 200 mL of 0.4 N HCl
(c) 200 mL of 0.2 N HCl (d) 100 mL of 0.1 N HCl

Sol. (a) $\underset{(\text{Molar mass}=60)}{NH_2CONH_2} + 2\,NaOH \longrightarrow 2\,NH_3 + Na_2CO_3$

No. of moles of urea = Given mass/molar mass

$$= \frac{0.6}{60} = 0.01 \text{ mol } NH_2CONH_2$$

$\therefore$ It gives 0.02 mol NH_3 as per the equation.

To neutralise it, 0.02 equivalents of HCl needed.

As we know, no. of equivalents of HCl = Normality $\times$ Volume

Therefore,

(a) $100 \times 10^{-3} \times 0.2 = 0.02$

(b) $200 \times 10^{-3} \times 0.4 = 0.08$

(c) $200 \times 10^{-3} \times 0.2 = 0.04$

(d) $100 \times 10^{-3} \times 0.1 = 0.01$

Thus, option (a) is correct.

2. Back Titration

This is the titration method in which the concentration of one analyte is determined by reacting it with a known amount of excess reagent. The remaining excess reagent is then titrated with another (second) reagent.

The results of second titration show that how much of the excess reagent was used in the first titration. Now, from this data the original analyte concentration is calculated.

Example 30. 1.6 g *of pyrolusite ore was treated with* 50 cc *of* 1.0 N *oxalic acid and some sulphuric acid. The oxalic acid left undecomposed was raised to* 250 cc *in a flask.* 25 cc *of this solution when titrated with* 0.1 N $KMnO_4$ *requires* 32 cc *of the solution. Find out the percentage of pure* MnO_2 *in the sample.*

(a) 49% (b) 50%
(c) 81% (d) 62%

Sol. (a) M.eq of excess dilute oxalic acid in 25 cc = M.eq. of $KMnO_4$ $= 0.1 \times 32$

$\therefore$ M.eq of oxalic acid in 250 cc dilute solution = 32

M.eq of MnO_2 = Meq. of oxalic acid added − M.eq. of excess oxalic acid $= 1 \times 50 - 32 = 18$

$\therefore$ Let the weight of MnO_2 is Wg.

$$\therefore \quad \frac{W}{E} \times 1000 = 18$$

$$\therefore \quad \frac{W}{87/2} \times 1000 = 18$$

$$\text{or} \quad W = \frac{18 \times 87}{2000} = 0.783 \text{ g}$$

$$\% \text{ of } MnO_2 = \frac{0.783}{1.6} \times 100 = 48.93$$

3. Double titration

This is a titration of specific compound using different indicators.

- This method involves two indicators namely phenolphthalein and methyl orange.
- These indicators are weak organic acids or bases which shows completion of reaction by changing its colour.
- **Phenolphthale**in acts as an indicator in weakly basic medium, but not in acidic medium. On the other hand, **methyl orange** acts as an indicator both in acidic medium as well as in basic medium.

Extent of different bases with acid (HCl) using these two indicators is summarised below.

Indicators Used with Different Bases

	Indicators	
Bases	**Phenolphthalein**	**Methyl orange**
NaOH	100% reaction is indicated	100% reaction is indicated
Na_2CO_3	50% of reaction up to $NaHCO_3$ stage is indicated	100% reaction is indicated
$NaHCO_3$	No reaction is indicated	100% reaction is indicated

Suppose, the volume of given standard acid solution (HCl) is required,

for complete reaction of NaOH $= V_a$ mL
for complete reaction of $Na_2CO_3 = V_b$ mL
for complete reaction of $NaHCO_3 = V_c$ mL

There may be different combinations of mixture of bases are possible. We will opt the following two methods,

I. **Method I** We carry two titration separately with two different indicators.
II. **Method II** We carry single titration but adding second indicator after first end point is reached.

Results with Two Indicators in Two Different Methods

Mixture	Method I		Method II	
	Volume of HCl used with indicator		Volume of HCl used with indicator	
	Phenolphthalein	**Methyl orange**	**Phenolphthalein**	**Methyl orange added after the first end point is reached**
$NaOH + Na_2CO_3$	$V_a + \frac{V_b}{2}$	$V_a + V_b$	$V_a + \frac{V_b}{2}$	$\frac{V_b}{2}$
$NaOH + NaHCO_3$	V_a	$V_a + V_c$	V_a	V_c
$Na_2CO_3 + NaHCO_3$	$\frac{V_b}{2}$	$V_a + V_c$	$\frac{V_b}{2}$	$\frac{V_b}{2} + V_a$

If mixture contains NaOH, Na_2CO_3 and $NaHCO_3$, then following are the cases that occurs

Case 1 The titre readings of HCl, using phenolphthalein from beginning $= V_a + \frac{V_b}{2}$

Case 2 If methyl orange is added after the first end point, then the titre readings of HCl $= \frac{V_b}{2} + V_c$

Case 3 If methyl orange is added from the very beginning, the titre readings of HCl $= V_a + V_b + V_c$

Example 31. *A solution contains a mixture of* Na_2CO_3 *and* NaOH. *Using phenolphthalein (tiPh) as indicator,* 25 mL *of mixture required* 19.5 mL of 0.995 N HCl *for the end point. With methyl orange* (MeOH), 25mL *of solution required* 25 mL *of the same* HCl *for the end point. The grams per litre of* Na_2CO_3 *in the mixture is*

(a) 23.2 (b) 18.5 (c) 19.9 (d) 12.8

Sol. (a) Let the moles of Na_2CO_3 and NaOH in 25 mL mixture be x and y respectively.

Case 1 when HPh is used as indicator.

$$\underset{x}{NaOH} + \underset{x}{HCl} \longrightarrow NaCl + H_2O$$

$$\underset{y}{Na_2CO_3} + \underset{y}{HCl} \longrightarrow NaHCO_3 + NaCl$$

So, $$x + y = 19.5 \times 10^{-3} \times (0.995 \times 1) \qquad \ldots(i)$$

Case 2 when MeOH is used as indicator.

$$\underset{x}{NaOH} + \underset{x}{HCl} \longrightarrow NaCl + H_2O$$

$$\underset{y}{Na_2CO_3} + \underset{2y}{2HCl} \longrightarrow 2NaCl + H_2O + CO_2$$

$$x + 2y = 25 \times 10^{-3} \times 0.995 \times 1 \qquad \ldots(ii)$$

On solving Eqs. (i) and (ii), we get,

$$x = 13.93 \times 10^{-3} \text{ mol and } y = 5.4725 \times 10^{-3} \text{ mol}$$

Now, weight of NaOH in 25 mL mixture

$$= 13.93 \times 10^{-3} \times 40 = 557.2 \times 10^{-3} \text{ g}$$

Weight of Na_2CO_3 in 25 mL mixture $= 5.4725 \times 10^{-3} \times 106$

$$= 580.085 \times 10^{-3} \text{ g}$$

$\therefore$ Weight of NaOH per litre $= \dfrac{557.2 \times 10^{-3} \times 1000}{25}$

$$= 22.288 \text{ g/L}$$

Weight of Na_2CO_3 per litre $= \dfrac{580.085 \times 10^{-3} \times 1000}{25}$

$$= 23.2034 \text{ g/L}$$

4. Iodimetry

This is a simple redox titration between a solution of iodine and $Na_2S_2O_3$ solution.

$$\underset{n\,=\,2}{I_2} + \underset{n\,=1}{2S_2O_3^{2-}} \longrightarrow 2I^- + S_4O_6^{2-}$$

5. Iodometry

This is an indirect way of doing iodimetry. Here, an oxidising agent A is made to react with excess of solid KI. The oxidising agent oxidises I^- to I_2. This iodine is then made to react with $Na_2S_2O_3$ solution.

As can be seen the gram-equivalents of $Na_2S_2O_3$ would be equal to I_2 which in turn will be equal to that of reacted KI. This all would be equal to the number of gram equivalents of A.

Example 32. 5.7g *of bleaching powder was suspended in* 500*mL* of water. 25mL of this solution *on treatment* with KI *in the presence of* HCl *liberated iodine which reacted with* 24.35 mL *of* $N/10$ $Na_2S_2O_3$. *Calculate the percentage of available chlorine in the bleaching powder.*

(a) 30% (b) 25%
(c) 80% (d) 70%

Sol. (a) Millieq. of $Na_2S_2O_3 = 24.35 \times 1/10 = 2.435$

This would be the milliequivalent of I_2 and therefore that of Cl_2 (which liberates I_2 from KI).

Milliequivalents of Cl_2 in 500mL $= 2.435 \times 20 = 48.7$

M.eq. of Cl_2 = M.eq. of bleaching powder
= M.eq. of available Cl_2 in the bleaching powder.

Percentage of chlorine $= \dfrac{48.7}{1000} \times \dfrac{35.5}{5.7} \times 100$

$$= 30.33\%$$

Practice Exercise

ROUND I Topically Divided Problems

Matter and its Nature

1. The smallest matter particle that can take part in chemical reaction is
(a) atom (b) molecule
(c) Both (a) and (b) (d) None of these

2. Matter is anything which occupies ... A... and has... B.... Here, A and B are
(a) density and mass (b) volume and mass
(c) space and mass (d) None of these

3. The solid like conducting state of gases with free electrons is called
(a) sol state (b) gel state
(c) plasma state (d) All of these

4. Which of the following statements about a compound is incorrect? *(NCERT Exemplar)*
(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. Which of the following is not a mixture?
(a) Gasoline (b) Distilled alcohol
(c) LPG (d) Iodised table salt

Physical Quantities, their Units and Significant Figures

6. The units J Pa^{-1} is equivalent to
(a) m^3 (b) cm^3
(c) dm^3 (d) None of these

7. In which of the following numbers all zeros are significant?
(a) 0.500 (b) 30.000
(c) 0.00030 (d) 0.0050

8. A measured temperature on Fahrenheit scale is 200 F. What will this reading be on celsius scale? *(NCERT Exemplar)*
(a) 40°C (b) 94°C
(c) 93.3°C (d) 30°C

9. Convert the following into basic units.
(i) 28.7 pm
(ii) 15.15 μs
(iii) 25365 mg
The correct answer is *(NCERT)*
(a) 28.7×10^{-11}, 1.515×10^{-6}, 2.5365×10^{-3}
(b) 2.87×10^{-11}, 1.515×10^{-5}, 2.5365×10^{-2}
(c) 2.87×10^{-10}, 1.515×10^{-5}, 2.5365×10^{-3}
(d) 2.87×10^{-10}, 1.515×10^{-6}, 2.5365×10^{-2}

10. The result of which of the following has least significant figures? *[NCERT]*
(a) $\dfrac{0.02856 \times 298.15 \times 0.112}{0.5785}$
(b) 5×5.364
(c) $0.0125 + 0.7864 + 0.0215$
(d) All have same number of significant figures

Laws of Chemical Combinations

11. Which of the following statements is correct about the reaction given below? *(NCERT Exemplar)*

$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(g)$$

(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 taking any one of the reactants (iron or oxygen) in excess
(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

12. 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?

(a) 4.5 g (b) 3.3 g
(c) 2.6 g (d) 2.8 g

13. 45.4L of dinitrogen reacted with 22.7 L of dioxygen and 45.4L 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 ? *(NCERT Exemplar)*

(a) Law of definite proportion
(b) Law of conservation of mass
(c) Law of multiple proportion
(d) None of the above

14. The mass of nitrogen per gram hydrogen in the compound hydrazine is exactly one and half times the mass of nitrogen in the compound ammonia. The fact illustrates the

(a) law of conservation of mass
(b) multiple valency of nitrogen
(c) law of multiple proportions
(d) law of definite proportions

15. Zinc sulphate contains 22.65% of zinc and 43.9% of water of crystallisation. If the law of constant proportions is true then the weight of zinc required to produce 20 g of the crystals will be

(a) 45.3 g (b) 4.53 g
(c) 0.453 g (d) 453 g

16. 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, each weighing 5 g. In the combinations AB, AB_2, A_2B and A_2B_3 which law is applicable. *(NCERT Exemplar)*

(a) Law of definite proportion
(b) Law of multiple proportion
(c) Law of conservation of mass
(d) None of the above

Atomic Masses, Molecular Mass, Equivalent Mass and Empirical Formula

17. If w_1 g of a metal X displaces w_2 g of another metal Y from its salt solution and if the equivalent weights are E_1 and E_2 respectively, the correct expression for the equivalent weight of X is

(a) $E_1 = \frac{w_1}{w_2} \times E_2$ (b) $E_1 = \frac{w_2 \times E_2}{w_1}$

(c) $E_1 = \frac{w_1 \times w_2}{E_2}$ (d) $E_1 = \sqrt{\frac{w_1}{w_2} \times E_2}$

18. The empirical formula and molecular mass of a compound are CH_2O and 180 g mol^{-1} respectively. What will be the molecular formula of the compound ? *(NCERT Exemplar)*

(a) $C_9H_{18}O_9$ (b) CH_2O
(c) $C_6H_{12}O_6$ (d) $C_2H_4O_2$

19. The strengths of 5.6 volume hydrogen peroxide (of density 1 g/mL) in terms of mass percentage and molarity (M), respectively are
(Take molar mass of hydrogen peroxide as 34 g/mol) *(JEE Main 2020)*

(a) 0.85 and 0.25 (b) 0.85 and 0.5
(c) 1.7 and 0.5 (d) 1.7 and 0.25

20. A 400 mg iron capsule contains 100 mg of ferrous fumarate, $(CHCOO)_2Fe$. The percentage of iron present in it is approximately

(a) 33% (b) 25% (c) 14% (d) 8%

21. The mass of potassium dichromate crystals required to oxidise 750 cm^3 of 0.6 M Mohr's salt solution is (Given, molar mass : Potassium dichromate = 294, Mohr's salt = 392) *[AIEEE 2011]*

(a) 0.49 g (b) 0.45 g (c) 22.05 g (d) 2.2 g

22. The density (in g mL^{-1}) of a 3.60 M sulphuric acid solution that is 29% H_2SO_4 (molar mass = 98 g mol^{-1}) by mass will be *[AIEEE 2007]*

(a) 1.64 (b) 1.88 (c) 1.22 (d) 1.45

23. 5.6 L of a gas at NTP are found to have a mass of 11 g. The molecular mass of the gas is

(a) 36 (b) 48 (c) 40 (d) 44

24. An organic compound contains 20.0% C, 6.66% H, 47.33% N and the rest was oxygen. Its molar mass is 60 g mol^{-1}. The molecular formula of the compound is

(a) CH_4N_2O (b) $C_2H_4NO_2$
(c) CH_3N_2O (d) $CH_4N_2O_2$

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

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

The molar mass of argon is *(NCERT)*

(a) 39.948 (b) 39.665 (c) 38.678 (d) 38.442

26. A certain amount of a metal whose equivalent mass is 28 displaces 0.7 L of H_2 at STP from an acid. Hence, mass of the element is

(a) 1.75 g (b) 0.875 g
(c) 3.50 g (d) 7.00 g

27. There are two isotopes of an element with atomic mass z. Heavier one has atomic mass $z + 2$ and lighter one has $z - 1$, then abundance of lighter one is

(a) 66.6% (b) 96.7% (c) 6.67% (d) 33.3%

28. One atom of an element weighs 1.8×10^{-22} g. Its atomic mass is

(a) 29.9 (b) 18 (c) 108.36 (d) 154

29. 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. Given that molar mass of the oxide is 159.89 g mol^{-1}. *(NCERT)*

(a) FeO (b) FeO_2 (c) Fe_2O_3 (d) Fe_3O_2

30. The equivalent mass of chlorine is 35.5 and the atomic mass of copper is 63.5. The equivalent mass of copper chloride is 99.0. Hence, formula of copper chloride is

(a) CuCl (b) Cu_2Cl
(c) $CuCl_2$ (d) None of these

31. Insulin contains 3.4% sulphur. What will be the minimum molecular weight of insulin?

(a) 94.117 (b) 1884 (c) 941.176 (d) 976

32. Two oxides of a metal contain 50% and 40% metal (M) respectively. If the formula of first oxide is MO_2, the formula of second oxide will be

(a) MO_2 (b) MO_3 (c) M_2O (d) M_2O_5

33. What quantity of ammonium sulphate is necessary for the production of NH_3 gas sufficient to neutralize a solution containing 292 g of HCl? [HCl = 36.5, $(NH_4)_2SO_4$ = 132, NH_3 = 17]

(a) 272 g (b) 403 g
(c) 528 g (d) 1056 g

34. On analysis a certain compound was found to contain iodine and oxygen in the ratio of 254 g of iodine (at. mass 127) and 80 g oxygen (at. mass 16). What is the formula of the compound?

(a) IO (b) I_2O
(c) I_5O_3 (d) I_2O_5

35. The hardness of a water sample (in terms of equivalents of $CaCO_3$) containing 10^{-3} M $CaSO_4$ is (Molar mass of $CaSO_4$ = 136 g mol^{-1}) *(JEE Main 2019)*

(a) 100 ppm (b) 10 ppm
(c) 50 ppm (d) 90 ppm

36. The most abundant elements by mass in the body of a healthy human adult are oxygen (61.4%); carbon (22.9%), hydrogen (10.0 %); and nitrogen (2.6%). The weight which a 75 kg person would gain if all ^1H atoms are replaced by ^2H atoms is *(JEE Main 2017)*

(a) 15 kg (b) 37.5 kg (c) 7.5 kg (d) 10 kg

37. The ratio of mass per cent of C and H of an organic compound $(C_xH_yO_z)$ is 6 : 1. If one molecule of the above compound $(C_xH_yO_z)$ contains half as much oxygen as required to burn one molecule of compound C_xH_y completely to CO_2 and H_2O. The empirical formula of compound $C_xH_yO_z$ is *(JEE Main 2018)*

(a) $C_3H_6O_3$ (b) C_2H_4O (c) $C_3H_4O_2$ (d) $C_2H_4O_3$

38. 1.020 g of metallic oxide contains 0.540 g of the metal. If the specific heat of the metal, M is 0.216 cal $deg^{-1}g^{-1}$, the molecular formula of its oxide is

(a) MO (b) M_2O_3 (c) M_2O_4 (d) M_2O

Mole Concept and Chemical Stoichiometry

39. What will be the molarity of a solution, which contains 5.85 g of NaCl(s) per 500 mL? *(NCERT Exemplar)*

(a) 4 mol L^{-1} (b) 20 mol L^{-1}
(c) 0.2 mol L^{-1} (d) 2 mol L^{-1}

40. The mass of 112 cm^3 of CH_4 gas at STP is

(a) 0.16 g (b) 0.8 g
(c) 0.08 g (d) 1.6 g

41. Out of 1.0 g dioxygen, 1.0 g (atomic) oxygen and 1.0 g ozone, the maximum number of molecules are contained in

(a) 1.0 g of atomic oxygen
(b) 1.0 g of ozone
(c) 1.0 g of oxygen gas
(d) All contain same number of atoms

42. The number of water molecules present in a drop of water (volume 0.0018 mL) at room temperature is

(a) 6.023×10^{19} (b) 1.084×10^{18}
(c) 4.84×10^{17} (d) 6.023×10^{23}

43. A signature written with carbon pencil weighs 1 mg. What is the number of carbon atoms present in the signature ?

(a) 6.02×10^{20} (b) 0.502×10^{20}
(c) 5.02×10^{23} (d) 5.02×10^{20}

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

(a) 6.02×10^{23} (b) 1.6×10^{-19}
(c) 1.99×10^{-23} (d) 1.67×10^{-23}

45. A sample of AlF_3 contains 3.0×10^{24} F^- ions. The number of formula units of this sample are
(a) 9.0×10^{24} (b) 3.0×10^{24}
(c) 0.75×10^{24} (d) 1.0×10^{24}

46. 19.7 kg of gold was recovered from a smuggler. How many atoms of gold were recovered (Au = 197)?
(a) 100 (b) 6.02×10^{23}
(c) 6.02×10^{24} (d) 6.02×10^{25}

47. How many moles of magnesium phosphate, $Mg_3(PO_4)_2$ will contain 0.25 moles of oxygen atoms?
(AIEEE 2006)
(a) 0.02 (b) 3.125×10^{-2}
(c) 1.25×10^{-2} (d) 2.5×10^{-2}

48. Which of the following pairs contains equal number of atoms ?
(a) 11.2 cc (STP) of nitrogen and 0.015 g of nitric oxide
(b) 22.4 L (STP) of nitrous oxide and 22.4 L of nitric oxide
(c) 1 millimole of HCl and 0.5 millimole of H_2S
(d) 1 mole of H_2O_2 and 1 mole of N_2O_4

49. Calculate the number of moles left after removing 10^{21} molecules from 200 mg of CO_2.
(a) 0.00454 (b) 0.00166
(c) 2.88×10^{-3} (d) None of these

50. If H_2SO_4 ionises as,
$$H_2SO_4 + 2H_2O \longrightarrow 2H_3O^+ + SO_4^{2-}$$
then total number of ions produced by 0.1 M H_2SO_4 will be
(a) 9.03×10^{21} (b) 3.01×10^{22}
(c) 6.02×10^{22} (d) 1.8×10^{23}

51. The density of 3 molal solution of NaOH is 1.110 g mL^{-1} Calculate the molarity of the solution.
(NCERT Exemplar)
(a) 2.97 M (b) 1.67 M
(c) 3.64 M (d) 2.32 M

52. If the density of methanol is 0.793 kg L^{-1}, what is its volume needed for making 2.5 L of its 0.25M solution? *(NCERT)*
(a) 5 mL (b) 25.2 mL
(c) 50 mL (d) 2.525 mL

53. 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) Determine the molarity of chloroform in the water sample.
(NCERT)
(a) 1.45×10^{-4} (b) 1.26×10^{-4}
(c) 1.89×10^{-6} (d) 1.34×10^{-5}

54. x g of Ag was dissolved in HNO_3 and the solution was treated with excess of NaCl when 2.87 g of AgCl was precipitated. The value of x is
(a) 1.08 g (b) 2.16 g
(c) 2.70 g (d) 1.62 g

55. 1 g of a carbonate (M_2CO_3) on treatment with excess HCl produces 0.01186 mole of CO_2. The molar mass of M_2CO_3 in g mol^{-1} is *(JEE Main 2017)*
(a) 1186 (b) 84.3 (c) 118.6 (d) 11.86

56. 100 mL of a water sample contains 0.81 g of calcium bicarbonate and 0.73 g of magnesium bicarbonate. The hardness of this water sample expressed in terms of equivalents of $CaCO_3$ is (molar mass of calcium bicarbonate is 162 g mol^{-1} and magnesium bicarbonate is 146 g mol^{-1})
(JEE Main 2019)
(a) 5,000 ppm (b) 1,000 ppm
(c) 100 ppm (d) 10,000 ppm

57. A sample of ammonium phosphate, $(NH_4)_3PO_4$ contains 6.36 moles of hydrogen atoms. The number of moles of oxygen atom in the sample is
(atomic mass of $N = 14.04, H = 1, P = 31, O = 16$)
(a) 0.265 mol (b) 0.795 mol
(c) 2.12 mol (d) 4.14 mol

58. A sample of copper sulphate pentahydrate contains 8.64 g of oxygen. How many gram of Cu is present in this sample ? (Atomic mass of Cu = 63.6, S = 32.06, O = 16)
(a) 0.952 g (b) 3.816 g
(c) 3.782 g (d) 8.64 g

59. Cortisone is a molecular substance containing 21 atoms of carbon per molecule. The molecular weight of cortisone is 360.4. What is the percentage of carbon in cortisone?
(a) 59.9% (b) 75%
(c) 69.98% (d) None of these

60. How many moles of $MgIn_2S_4$ can be made from 1 g each of Mg, In and S? (Atomic mass : Mg = 24, In = 114.8, S = 32)
(a) 6.47×10^{-4} (b) 3.0×10^{-1}
(c) 9.17×10^{-2} (d) 8.7×10^{-3}

61. If an iodised salt contains 1% KI and a person takes 2 g of the salt every day, the iodide ions going into his body every day would be approximately
(a) 7.2×10^{21} (b) 7.2×10^{19}
(c) 3.6×10^{21} (d) 9.5×10^{19}

62. If the density of water is 1 g cm^{-3} then the volume occupied by one molecule of water is approximately

(a) 18 cm^{-3} (b) 22400 cm^{-3}

(c) 6.02×10^{-23} cm^3 (d) 3.0×10^{-23} cm^3

63. 5 moles of AB_2 weight 125×10^{-3} kg and 10 moles of A_2B_2 weight 300×10^{-3} kg. The molar mass of $A(M_A)$ and molar mass of $B\ (M_B)$ in kg mol^{-1} are *(JEE Main 2019)*

(a) $M_A = 10 \times 10^{-3}$ and $M_B = 5 \times 10^{-3}$

(b) $M_A = 50 \times 10^{-3}$ and $M_B = 25 \times 10^{-3}$

(c) $M_A = 25 \times 10^{-3}$ and $M_B = 50 \times 10^{-3}$

(d) $M_A = 5 \times 10^{-3}$ and $M_B = 10 \times 10^{-3}$

64. For the reaction ,

$$X + 2Y \longrightarrow Z$$

5 moles of X and 9 moles of Y will produce

(a) 5 moles of Z (b) 8 moles of Z

(c) 14 moles of Z (d) 4 moles of Z

65. If 0.5 mole of $BaCl_2$ is mixed with 0.2 mole of Na_3PO_4, the maximum number of moles of $Ba_3(PO_4)_2$ that can be formed is

(a) 0.7 (b) 0.5

(c) 0.03 (d) 0.10

66. In the following reaction,

$$MnO_2 + 4HCl \longrightarrow MnCl_2 + 2H_2O + Cl_2$$

2 moles of MnO_2 react with 4 moles of HCl to form 11.2 L Cl_2 at STP. Thus, percent yield of Cl_2 is

(a) 25% (b) 50%

(c) 100% (d) 75%

67. In the reaction,

$$I_2 + 2S_2O_3^{2-} \longrightarrow 2I^- + S_4O_6^{2-}$$

equivalent weight of iodine will be equal to

(a) molecular weight

(b) 1/2 of molecular weight

(c) 1/4 of molecular weight

(d) twice of molecular weight

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

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

How many grams of HCl reacts with 5.0 g of manganese dioxide? *(NCERT)*

(a) 36.5 g (b) 4.12 g

(c) 8.39 g (d) 13.25 g

69. If ten volumes of dihydrogen gas react with five volumes of dioxygen gas, how many volumes of water vapour would be produced? *(NCERT)*

(a) 5 (b) 10 (c) 2 (d) 20

70. Consider the reaction, $A + B_2 \longrightarrow AB_2$, A acts as the limiting reagent, in which of the following reactions are mixtures ? *(NCERT)*

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

(b) 2 moles of A + 3 moles of B

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

(d) 5 moles of A + 2.5 moles of B

71. What is the number of moles of $Fe(OH)_3(s)$ that can be produced by allowing 1 mole of Fe_2S_3, 2 moles of H_2O and 3 moles of O_2 to react as,

$$2Fe_2S_3 + 6H_2O + 3O_2 \longrightarrow 4Fe(OH)_3 + 6S\ ?$$

(a) 1 mol (b) 1.84 mol (c) 1.34 mol (d) 1.29 mol

72. 25 g of an unknown hydrocarbon upon burning produces 88 g of CO_2 and 9 g of H_2O. This unknown hydrocarbon contains *(JEE Main 2019)*

(a) 20 g of carbon and 5 g of hydrogen

(b) 22 g of carbon and 3 g of hydrogen

(c) 24 g of carbon and 1 g of hydrogen

(d) 18 g of carbon and 7 g of hydrogen

73. For the following reaction, the mass of water produced from 445 g of $C_{57}H_{110}O_6$ is

$$2C_{57}H_{110}O_6(s) + 163O_2(g) \longrightarrow 114CO_2(g) + 110\ H_2O\ (l)$$

(JEE Main 2019)

(a) 490 g (b) 495 g (c) 445 g (d) 890 g

74. In the reaction of oxalate with permanganate in acidic medium, the number of electrons involved in producing one molecule of CO_2 is *(JEE Main 2019)*

(a) 2 (b) 5 (c) 1 (d) 10

75. For a reaction, $N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$, identify dihydrogen ($H_2$) as a limiting reagent in the following reaction mixtures. *(JEE Main 2019)*

(a) 56 g of N_2 + 10 g of H_2 (b) 35 g of N_2 + 8 g of H_2

(c) 14 g of N_2 + 4 g of H_2 (d) 28 g of N_2 + 6 g of H_2

76. The minimum amount of $O_2(g)$ consumed per gram of reactant is for the reaction (Given atomic mass : Fe = 56, O = 16, Mg = 24, P = 31, C = 12, H = 1) *(JEE Main 2019)*

(a) $C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$

(b) $P_4(s) + 5O_2(g) \longrightarrow P_4O_{10}(s)$

(c) $4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$

(d) $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

ROUND II Mixed Bag

1. The law of definite proportions is not applicable to nitrogen oxide because
(a) nitrogen atomic weight is not constant
(b) nitrogen molecular weight is variable
(c) nitrogen equivalent weight is variable
(d) oxygen atomic weight is variable

2. The number of water molecules in 1 L of water is
(a) 18 (b) 18×1000 (c) N_A (d) $55.55\, N_A$

3. Haemoglobin contains 0.33% of iron by weight. The molecular weight of haemoglobin is approximately 67200. The number of iron atoms (at. wt. of Fe = 56) present in one molecule of haemoglobin is
(a) 6 (b) 1 (c) 4 (d) 2

4. If isotopic distribution of C-12 and C-14 is 98% and 2% respectively then the number of C-14 atoms in 12 g of carbon is
(a) 1.032×10^{22} (b) 3.01×10^{22}
(c) 5.88×10^{23} (d) 6.023×10^{23}

5. The number of moles of oxygen in 1 L of air containing 21% oxygen by volume, in standard conditions, is
(a) 0.186 mol (b) 0.21 mol
(c) 2.10 mol (d) 0.0093 mol

6. Which of the following statements indicates that law of multiple proportion is being followed? *(NCERT Exemplar)*
(a) Sample of carbon dioxide taken from any source will always have carbon and oxygen in the ratio 1:2.
(b) Carbon forms two oxides namely CO_2 and CO, where masses of oxygen which combine with fixed mass of carbon are in the simple ratio 2:1
(c) When magnesium burns in oxygen, the amount of magnesium taken for the reaction is equal to the amount of magnesium in magnesium oxide formed
(d) At constant temperature and pressure 200 mL of hydrogen will combine with 100 mL oxygen to produce 200 mL of water vapour

7. 1.520 g of the hydroxide of a metal on ignition gave 0.995 g of oxide. The equivalent weight of metal is
(a) 1.520 (b) 0.995 (c) 19.00 (d) 9.00

8. 100 mL of 20.8% $BaCl_2$ solution and 50 mL of 9.8% H_2SO_4 solution will form $BaSO_4$
(Ba = 137, Cl = 35.5, S = 32 , H = 1, O = 16)

$$BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$$

(a) 23.3 g (b) 11.65 g
(c) 30.6 g (d) None of these

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

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

In this reaction, 250 mL of 0.76 M HCl reacts with 1000 g of $CaCO_3$. Calculate the mass of $CaCl_2$ formed in the reaction. *(NCERT Exemplar)*
(a) 11.1 g (b) 10.54 g
(c) 5.25 g (d) 2.45 L

10. One gram of hydrogen is found to combine with 80 g of bromine. One gram of calcium (valency=2) combines with 4 g of bromine. The equivalent weight of calcium is
(a) 10 g (b) 20 g
(c) 40 g (d) 80 g

11. The percentage of P_2O_5 in diammonium hydrogen phosphate, $(NH_4)_2HPO_4$ is
(a) 23.48 % (b) 46.96 %
(c) 53.78 % (d) 71.00 %

12. A mixture of CH_4, N_2 and O_2 is enclosed in a vessel of one litre capacity at 0°C. The ratio of partial pressures of gases is 1 : 4 : 2. Total pressure of the gaseous mixture is 2660 mm. The number of molecules of oxygen present in the vessel is
(a) $\dfrac{6.02 \times 10^{23}}{22.4}$ (b) 6.02×10^{23}
(c) 22.4×10^{22} (d) 1000

13. An ore contains 1.34% of the mineral argentite, Ag_2S, by mass. How many gram of this ore would have to be processed in order to obtain 1.00 g of pure solid silver, Ag ?
(a) 74.6 g (b) 85.7 g
(c) 107.9 g (d) 134.0 g

14. What weight of silver chloride will be precipitated when a solution containing 4.77 g of NaCl is added to a solution of 5.77 g of $AgNO_3$? (Na = 23, Cl = 35.5, Ag = 108, N = 14 and O = 16)
(a) 4.37 g (b) 4.87 g (c) 5.97 g (d) 3.87 g

15. 0.75 mole of a solid A_4 and 2 moles of $O_2(g)$ are heated in a sealed vessel, completely using up the reactants and produces only one compound. It is found that when the temperature is used to initial temperature, the contents of the vessel exhibit a pressure equal to 1/2 of the original pressure. The formula of the product will be
(a) A_2O_3 (b) A_3O_8 (c) A_3O_4 (d) AO_2

16. A sample of a mixture of $CaCl_2$ and NaCl weighing 4.22 g was treated to precipitate all the Ca as $CaCO_3$. This $CaCO_3$ is then heated and quantitatively converted into 0.959g of CaO. Calculate the percentage of $CaCl_2$ in the mixture.
(Atomic mass of Ca = 40, O = 16, C = 12, and Cl = 35.5)
(a) 31.5% (b) 21.5% (c) 45.04% (d) 68.48%

17. Which one of the following will have the largest number of atoms? *(NCERT)*
(a) 1 g Au (s) (b) 1 g Na (s) (c) 1 g Li (s) (d) 1 g Cl_2 (g)

Numeric Value Questions

18. The molarity of HNO_3 in a sample which has density 1.4 g/mL and mass percentage of 63% is …… M. (Molecular weight of $HNO_3 = 63$) *(JEE Main 2020)*

19. $NaClO_3$ is used, even in spacecrafts, to produce O_2. The daily consumption of pure O_2 by a person is 492 L at 1 atm 300 K. How much amount of $NaClO_3$, in grams, is required to produce O_2 for the daily consumption of a person at 1 atm, 300 K .... ?

$$NaClO_3(s) + Fe(s) \longrightarrow O_2(g) + NaCl(s) + FeO(s)$$

$$R = 0.082 \text{ L atm mol}^{-1}\text{K}^{-1}$$

(JEE Main 2020)

20. The volume (in mL) of 0.125 M $AgNO_3$ required to quantitatively precipitate chloride ions in 0.3 g of $[Co(NH_3)_6]Cl_3$ is ……… .

$M_{[Co(NH_3)_6]Cl_3} = 267.46$ g/mol

$M_{AgNO_3} = 169.87$ g/mol *(JEE Main 2020)*

21. In an estimation of bromine by Carius method, 1.6 g of an organic compound gave 1.88 g of AgBr. The mass percentage of bromine in the compound is .... (Atomic mass, Ag = 108, Br = 80 g mol^{-1}) *(JEE Main 2020)*

22. A solution of phenol in chloroform when treated with aqueous NaOH gives compound P as a major product. The mass percentage of carbon in P is ……… . (to the nearest integer) (Atomic mass: C = 12; H = 1 and O = 16) *(JEE Main 2020)*

23. The minimum number of moles of O_2 required for complete combustion of 1 mole of propane and 2 moles of butane is …… . *(JEE Main 2020)*

24. The volume, in mL, of 0.02 M $K_2Cr_2O_7$ solution required to react with 0.288 g of ferrous oxalate in acidic medium is …… .
(Molar mass of Fe = 56 g mol^{-1}) *(JEE Main 2020)*

25. The volume (in mL) of 0.1 N NaOH required to neutralise 10 mL of 0.1 N phosphonic acid is .... *(JEE Main 2020)*

26. The mole fraction of glucose ($C_6H_{12}O_6$) in an aqueous binary solution is 0.1. The mass percentage of water in it, to the nearest integer, is .... *(JEE Main 2020)*

27. Arsenic forms two oxides, one of which contains 65.2% and the other 75.7% of the element. Hence, equivalent masses of arsenic are in the ratio 3 : x. The value of x will be ……… .

28. 2.76 g of silver carbonate on being strongly heated yields a residue weighing …… g.

29. One atom of an element X weigh 6.643×10^{-23} g. Number of moles of atom in 20 kg is …… .

30. 100 tons of Fe_2O_3 containing 20% impurities will give iron by reduction with H_2 equal to …… tons.

31. The molality of the solution containing 18.25 g of HCl gas in 500 g of water will be ……… . *(NCERT Exemplar)*

32. The number of sodium atoms in 2 moles of sodium ferrocyanide is … $a \times 10^{23}$. The value of a will be ……… .

33. The mole fraction of a solute in a 100 molal aqueous solution ……… $\times 10^{-2}$.

[Given : Atomic masses, H : 1.0 u, O : 16.0 u] *(JEE Main 2021)*

Answers

Round I

1. (a)	**2.** (c)	**3.** (c)	**4.** (c)	**5.** (b)	**6.** (a)	**7.** (b)	**8.** (c)	**9.** (b)	**10.** (a)
11. (a)	**12.** (b)	**13.** (d)	**14.** (c)	**15.** (b)	**16.** (b)	**17.** (a)	**18.** (c)	**19.** (c)	**20.** (d)
21. (c)	**22.** (c)	**23.** (d)	**24.** (a)	**25.** (a)	**26.** (a)	**27.** (a)	**28.** (c)	**29.** (c)	**30.** (a)
31. (c)	**32.** (b)	**33.** (c)	**34.** (d)	**35.** (a)	**36.** (c)	**37.** (d)	**38.** (b)	**39.** (c)	**40.** (c)
41. (a)	**42.** (a)	**43.** (b)	**44.** (c)	**45.** (d)	**46.** (d)	**47.** (b)	**48.** (a)	**49.** (c)	**50.** (d)
51. (a)	**52.** (b)	**53.** (b)	**54.** (b)	**55.** (b)	**56.** (d)	**57.** (c)	**58.** (b)	**59.** (c)	**60.** (c)
61. (b)	**62.** (d)	**63.** (d)	**64.** (d)	**65.** (d)	**66.** (b)	**67.** (b)	**68.** (c)	**69.** (b)	**70.** (b)
71. (c)	**72.** (c)	**73.** (b)	**74.** (c)	**75.** (d)	**76.** (c)				

Round II

1. (c)	**2.** (d)	**3.** (c)	**4.** (a)	**5.** (d)	**6.** (b)	**7.** (d)	**8.** (b)	**9.** (b)	**10.** (b)
11. (c)	**12.** (a)	**13.** (b)	**14.** (b)	**15.** (c)	**16.** (c)	**17.** (c)	**18.** (14.00)	**19.** (2130)	**20.** (26.92)
21. (50)	**22.** (69)	**23.** (18)	**24.** (50)	**25.** (10)	**26.** (47.00)	**27.** (5)	**28.** (2.16)	**29.** (500)	**30.** (56)
31. (1)	**32.** (48)	**33.** (64.28)							

Solutions

Round I

4. The properties of a compound are quite different from the properties of constituent elements.

6. $JPa^{-1} = \frac{J}{Pa} = \frac{\text{work}}{\text{pressure}} = \frac{\text{N-m}}{\text{N/m}^2} = \text{m}^3$

8.
$$°F = \frac{9}{5} t° C + 32$$
$$200 - 32 = \frac{9}{5} t° C$$
$$\Rightarrow \quad \frac{9}{5} t° C = 168$$
$$t° C = \frac{168 \times 5}{9} = 93.3°C$$

9. (a) $28.7 \text{ pm} \times \frac{10^{-12}\text{m}}{1\text{pm}} = 2.87 \times 10^{-11}\text{m}$

(b) $15.15 \text{ ms} \times \frac{10^{-6}\text{s}}{1 \text{ ms}} = 1.515 \times 10^{-5}\text{s}$

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

10. (a) In multiplication and division, the least precise term 0.112 has 3 significant figures. Hence, the answer should not have more than three significant figures.

(b) In multiplication, 5 is the exact number and the other number has 4 significant figures. Hence, the answer should have 4 significant figures.

(c) In addition (or in subtraction), the answer cannot have more digits to the decimal point than either of the original members. Hence, the answer should have 4 significant figures.

11. According to the law of conservation of mass,
Total mass of reactants = total mass of products
Amount of Fe_2O_3 is decided by limiting reagent.

12. According to law of conservation of mass,
mass of reactants = mass of products

$\therefore \quad 6.3 + 15.0 = 18.0 + x$

or $\quad x = 21.3 - 18.0 = 3.3$ g

13.
$$\underset{2V}{2N_2(g)} + \underset{1V}{O_2(g)} \longrightarrow \underset{2V}{2N_2O(g)}$$
$$\frac{45.4}{22.7} = 2 \quad \frac{22.7}{22.7} = 1 \quad \frac{45.4}{22.7} = 2$$

Hence, the ratio between the volumes of the reactants and the product in the given question is simple, i.e. 2 : 1 : 2. It proves the **Gay- Lussac's law of gaseous volume**s, which states that when gases combine or are produced in a chemical reaction, they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

14. As ratio of masses of nitrogen per gram of hydrogen in hydrazine and $NH_3 = 1\frac{1}{2} : 1 = \frac{3}{2} : 1 \quad \text{or} \quad 3 : 2$

i.e. the law of multiple proportions.

15. To prepare 20 g zinc sulphate crystals, zinc required

$= \frac{22.65}{100} \times 20 = 4.53$ g

17.
$$\frac{\text{Weight of metal } X}{\text{Weight of metal } Y} = \frac{\text{Equivalent weight of metal } X}{\text{Equivalent weight of metal } Y}$$
$$\frac{w_1}{w_2} = \frac{E_1}{E_2} \text{ or } E_1 = \frac{w_1}{w_2} \times E_2$$

18. Empirical formula mass = CH_2O

$$= 12 + 2 \times 1 + 16 = 30$$

Molecular mass = 180

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

$$= \frac{180}{30} = 6$$

$\therefore$ Molecular formula = $n \times$ empirical formula

$$= CH_2O \times 6 = C_6H_{12}O_6$$

19. Volume of hydrogen = 5.6

Volume strength = 11.2 × molarity

$$\Rightarrow \quad \text{Molarity} = \frac{\text{Volume strength}}{11.2} = \frac{5.6}{11.2} = 0.5$$

Suppose, that the solution taken is = 1 L

Mass of solution = 1000 mL × 1 g/mL = 1000 g

Mass of solute = Moles × Molar mass = 0.5 × 34

[$\because$ Molar mass of H_2O_2 = 34]

= 17 g

$$\Rightarrow \quad \text{Mass \%} = \frac{17}{1000} \times 100 = 1.7\%$$

20. Molecular mass of $(CHCOO)_2Fe$ = 170

$\therefore$ In 100 g $(CHCOO)_2Fe$, iron present $= \frac{56}{170} \times 100$ mg

= 32.9 mg

Since, this quantity of Fe is present in 400 mg of capsule,

$$\therefore \quad \text{\% of Fe in capsule} = \frac{32.9}{400} \times 100 = 8.2\%$$

24. Ratio of atoms

$$C : H : N : O :: \frac{20.0}{12} : \frac{6.66}{1} : \frac{47.33}{14} : \frac{26.01}{16}$$

$$= 1.67 : 6.66 : 3.38 : 1.63$$

$$= 1 : 4 : 2 : 1$$

Empirical formula = CH_4N_2O

Empirical formula mass = 60 g

Molar mass = 60 g mol^{-1}

$\therefore$ Molecular formula = CH_4N_2O

25. Average molar mass of

$$Ar = \sum f_i \times A_i = (0.00337 \times 35.96755) + (0.00063 \times 37.96272) + (0.99600 \times 39.9624)$$

$$= 0.121 + 0.024 + 39.803 = 39.948 \text{ g mol}^{-1}$$

28. Mass of 1 atom = 1.8×10^{-22} g

Mass of 6.02×10^{23} atoms

$$= 6.02 \times 10^{23} \times 1.8 \times 10^{-22} \text{ g}$$

$$= 6.02 \times 1.8 \times 10 \text{ g}$$

$$= 108.36 \text{ g}$$

$\therefore$ Atomic mass of element = 108.36

29.

Element	Symbol	%by mass	Atomic mass	Relative number of moles of element	Simple molar ratio	Simple whole number molar ratio
Iron	Fe	69.9	55.85	$\frac{69.9}{55.85} = 1.25$	$\frac{1.25}{1.25} = 1$	1×2 = 2
Oxygen	O	30.1	16.00	$\frac{30.1}{16.00} = 1.88$	$\frac{1.88}{1.25} = 1.5$	1.5 × 2 = 3

$\therefore$ Empirical formula = Fe_2O_3

Empirical formula mass of Fe_2O_3

$$= (2 \times 55.85) + (3 \times 16.00)$$

$$= 159.7 \text{ g mol}^{-1}$$

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

$$= \frac{159.8}{159.7} = 1$$

Hence, molecular formula is same as empirical formula, i.e. Fe_2O_3.

30. Equivalent mass of copper chloride = 99

Equivalent mass of chlorine = 35.5

$\therefore$ Equivalent mass of copper = 99 − 35.5 = 63.5

$$\therefore \text{Valency of copper} = \frac{\text{Atomic mass of copper}}{\text{Equivalent mass of copper}} = 1$$

$\therefore$ Formula of copper chloride is CuCl.

31. For minimum molecular mass, there must be one S atom per insulin molecule.

If 3.4 g of S is present, then the molecular mass = 100

$\therefore$ If 32 g of S is present, then the molecular mass

$$= \frac{100 \times 32}{3.4} = 941.176$$

32.

	Oxide I	Oxide II
Metal, M	50%	40%
Oxygen, O	50%	60%

As first oxide is MO_2

Let atomic mass of $M = x$

$$\therefore \quad \% \text{ O} = \frac{32}{x + 32} \times 100$$

$$\text{or} \quad \frac{50}{100} = \frac{32}{x + 32} \text{ or } 0.5 = \frac{32}{x + 32}$$

$$\text{or} \quad 0.5x + 16 = 32$$

$$0.5x = 16$$

$$x = 32$$

$\therefore$ At. mass of metal M = 32

Let formula of second oxide is M_2O_n

$$\% M = \frac{2x}{2x + 16n} \times 100 = \frac{64}{64 + 16n} \times 100$$

$$\frac{40}{100} = \frac{64}{64 + 16n}$$

or $\dfrac{100}{40} = \dfrac{64+16n}{64}$

$2.5 = 1 + 0.25n$

$n = \dfrac{1.5}{0.25} = 6$

Therefore, formula of second oxide $= M_2O_6$

or $= MO_3$

33. $(NH_4)_2SO_4 \longrightarrow 2NH_3 + H_2O + SO_3$

$2NH_3 + 2HCl \longrightarrow 2NH_4Cl$

$\underset{132\text{ g}}{(NH_4)_2SO_4} \equiv 2NH_3 \equiv \underset{73\text{ g}}{2HCl}$

73 g HCl $\equiv$ 132 g of $(NH_4)_2SO_4$

$$292 \text{ g of HCl} \equiv \frac{132 \times 292}{73} \text{ g of } (NH_4)_2SO_4$$

$$= 528 \text{ g of } (NH_4)_2SO_4$$

35. Hardness of water sample can be calculated in terms of ppm concentration of $CaCO_3$.

Given, molarity $= 10^{-3}$ M

i.e. 1000 mL of solution contains 10^{-3} mole of $CaCO_3$.

$\therefore$ Hardness of water = ppm of $CaCO_3$

$$= \frac{10^{-3} \times 100}{1000} \times 10^6 = 100 \text{ ppm}$$

36. Given, abundance of elements by mass oxygen = 61.4%, carbon = 22.9%, hydrogen = 10% and nitrogen = 2.6%

Total weight of person = 75 kg

$$\text{Mass due to } {}^1H = \frac{75 \times 10}{100} = 7.5 \text{ kg}$$

${}^1H$ atoms are replaced by ${}^2H$ atoms,

Mass due to ${}^2H = (7.5 \times 2)$ kg

$\therefore$ Mass gain by person = 7.5 kg

37. We can calculate the simplest whole number ratio of C and H from the data given, as

Element	Relative mass	Molar mass	Relative mole	Simplest whole number ratio
C	6	12	$\frac{6}{12} = 0.5$	$\frac{0.5}{0.5} = 1$
H	1	1	$\frac{1}{1} = 1$	$\frac{1}{0.5} = 2$

Alternatively this ratio can also be calculated directly in the terms of x and y as

$\dfrac{12x}{y} = \dfrac{6}{1}$ (given and molar mass of C = 12, H = 1)

Now, after calculating this ratio look for condition 2 given in the question, i.e. quantity of oxygen is half of the quantity required to burn one molecule of compound C_xH_y completely to CO_2 and H_2O. We can calculate number of oxygen atoms from this as consider the equation.

$$C_xH_y + \left[x + \frac{y}{4}\right]O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O$$

Number of oxygen atoms required

$$= 2 \times \left[x + \frac{y}{4}\right] = \left[2x + \frac{y}{2}\right]$$

Now given, $z = \dfrac{1}{2}\left[2x + \dfrac{y}{2}\right] = \left[x + \dfrac{y}{4}\right]$

Here we consider x and y as simplest ratios for C and H so now putting the values of x and y in the above equation.

$$z = \left[x + \frac{y}{4}\right] = \left[1 + \frac{2}{4}\right] = 1.5$$

Thus, the simplest ratio figures for x, y and z are $x = 1$, $y = 2$ and $z = 1.5$

Now, put these values in the formula given, i.e. $C_xH_yO_z = C_1H_2O_{1.5}$

So, empirical formula will be $[C_1H_2O_{1.5}] \times 2 = C_2H_4O_3$

38. Mass of oxygen in the oxide

$$= (1.020 - 0.540) = 0.480 \text{ g}$$

$$\text{Equivalent mass of the metal} = \frac{0.540}{0.480} \times 8 = 9.0$$

According to Dulong and Petit's law,

$$\text{Approx, atomic mass} = \frac{6.4}{\text{sp. heat}} = \frac{6.4}{0.216} = 29.63$$

$$\text{Valency of the metal} = \frac{\text{at. mass}}{\text{eq. mass}} = \frac{29.63}{9.0} = 3$$

Hence, the formula of the oxide is M_2O_3.

39. $\text{Molarity} = \dfrac{\text{wt.} \times 1000}{\text{mol. wt.} \times \text{volume (mL)}} = \dfrac{5.85 \times 1000}{58.5 \times 500}$

$= 0.2 \text{ mol L}^{-1}$

40. $\because$ Mass of 22400 cm^3 CH_4 = 16 g

$\therefore$ Mass of 112 cm^3 $CH_4 = \dfrac{16 \times 112}{22400} = 0.08$ g

41. Number of molecules $= \dfrac{\text{mass} \times N_A}{\text{molar mass}}$

42. Number of moles $= \dfrac{\text{weight}}{\text{molecular wt.}} = \dfrac{0.0018}{18} = 1 \times 10^{-4}$

[$\because$ 0.0018 mL = 0.0018 g]

$\therefore$ Number of water molecules $= 1 \times 10^{-4} \times 6.02 \times 10^{23}$

$= 6.023 \times 10^{19}$

43. Number of C atoms $= \dfrac{1 \times 10^{-3}}{12} \times N_A = 0.50 \times 10^{20}$

44. Mass of 1 atom of ${}^{12}C = \dfrac{\text{Atomic mass of C}}{\text{Avogadro's number}}$

$$= \frac{12 \text{ g}}{6.022 \times 10^{23}} = 1.9927 \times 10^{-23} \text{ g}$$

45. $3F^- \equiv 1$ Formula unit (AlF_3)

3.0×10^{24} $F^- = 1 \times 10^{24}$ Formula units (AlF_3)

48. Number of atoms in $N_2 = \dfrac{11.2 \times 10^{-3} \times 6.023 \times 10^{23} \times 2}{22.4}$

$= 6.023 \times 10^{20}$

Number of atoms in NO $= \dfrac{0.015 \times 2 \times 6.023 \times 10^{23}}{30}$

$= 6.023 \times 10^{20}$

49. Molecular mass of $CO_2 = 12 + 32 = 44$

44 g of CO_2 has 6.023×10^{23} molecule

0.2 g of CO_2 has $\dfrac{6.023 \times 10^{23}}{44} \times 0.2$

$= 0.0273 \times 10^{23}$ molecules

If 10^{21} molecules are removed then number of molecules

$= 1.73 \times 10^{21}$

$\because$ 6.023×10^{23} molecules $= 1$ mol

$\therefore$ 1.73×10^{21} molecules $= \dfrac{1}{6.023 \times 10^{23}} \times 1.73 \times 10^{21}$

$= 0.0028$ mol or 2.88×10^{-3} mol

50. 1 mole of H_2SO_4 gives 3 moles of ions or $3 \times 6.023 \times 10^{23}$ ions

$\therefore$ 0.1 mole of H_2SO_4 will give $0.1 \times 3 \times 6.023 \times 10^{23}$ ions

$= 1.8 \times 10^{23}$ ions

51. 3 molal solution of NaOH means 3 moles of NaOH are dissolved in 1 kg solvent. So, the mass of solution

= 1000 g solvent + 120 g NaOH = 1120 g solution

(Molar mass of NaOH = 23 + 16 + 1 = 40 g;

3 moles of NaOH = 3 × 40 = 120 g)

Volume of solution $= \dfrac{\text{Mass of solution}}{\text{Density of solution}}$

$$V = \frac{1120 \text{ g}}{1.110 \text{ g mL}^{-1}} = 1009 \text{ mL}$$

$$\text{Molarity} = \frac{\text{moles of solute} \times 1000}{\text{volume of solution (mL)}} = \frac{3 \times 1000}{1009} = 2.973 \text{ M}$$

52. Given, $d = 0.793$ kg L^{-1} = 0.793×10^3 g L^{-1}

Molar mass of methanol,

$$CH_3OH = (1 \times 12.01) + (4 \times 1.0079) + (1 \times 16.00)$$

$$= 32.0416 \approx 32$$

$$\text{Molarity} = \frac{0.793 \times 10^3 \text{ g L}^{-1}}{32 \text{ g mol}^{-1}} = 24.781 \text{ mol}^{-1}$$

$$M_1 V_1 = M_2 V_2$$

$$24.781 \times V_1 = 0.25 \times 2.5$$

$$V_1 = \frac{0.25 \times 2.5}{24.781}$$

$$= 0.02522 \text{ L} = 25.22 \text{ mL}$$

53. 15 ppm means 15 parts in one million (10^6) parts.

Therefore, % by mass $= \dfrac{15 \times 100}{10^6} = 1.5 \times 10^{-3}\%$

Molar mass of $CHCl_3 = 12.01 + 1.0079 + (3 \times 35.45)$

$$M_{CHCl_3} = 119.367 \approx 119 \text{ g mol}^{-1}$$

$1.5 \times 10^{-3}\%$ means 1.5×10^{-3} g chloroform is present in 100 g sample.

Molarity, $M = \dfrac{w \times 1000}{m \times \text{volume of sample}}$

(For water, density = 1g cm^{-3}, so mass = volume)

$$M = \frac{1.5 \times 10^{-3} \times 1000}{119 \times 100}$$

$$= 0.000126 = 1.26 \times 10^{-4} \text{ M}$$

54. $2Ag + 2HNO_3 \longrightarrow 2AgNO_3 + H_2$

$2AgNO_3 + 2NaCl \longrightarrow 2AgCl + NaNO_3$

$$\underset{143.5 \text{ g}}{AgCl} \equiv \underset{170 \text{ g}}{AgNO_3} \equiv \underset{108 \text{ g}}{Ag}$$

$\because$ 143.5g AgCl is obtained from Ag = 108 g

$\therefore$ 2.87 g AgCl is obtained from Ag $= \dfrac{108 \times 2.87}{143.5} = 2.16$ g

55. $$\underset{1 \text{ g}}{M_2CO_3} + 2HCl \longrightarrow 2MCl + H_2O + \underset{0.01186 \text{ mol}}{CO_2}$$

Number of moles of M_2CO_3 reacted

= Number of moles of CO_2 evolved

$\dfrac{1}{M} = 0.01186$ [M = molar mass of M_2CO_3]

$M = \dfrac{1}{0.01186} = 84.3$ g mol^{-1}

56. Given, $W_{Ca(HCO_3)_2} = 0.81$ g

$W_{Mg(HCO_3)_2} = 0.73$ g

$M_{Ca(HCO_3)_2} = 162$ g mol^{-1},

$M_{Mg(HCO_3)_2} = 146$ mol^{-1}

$V_{H_2O} = 100$ mL

Now, $n_{eq}(CaCO_3) = n_{eq}[Ca(HCO_3)_2] + n_{eq}[Mg(HCO_3)_2]$

$$\frac{W}{100} \times 2 = \frac{0.81}{162} \times 2 + \frac{0.73}{146} \times 2$$

$\therefore$ $$\frac{W}{100} = 0.005 + 0.005$$

$$W = 0.01 \times 100 = 1$$

Thus, hardness of water sample $= \dfrac{1}{100} \times 10^6 = 10{,}000$ ppm

57. 1 mole of $(NH_4)_3PO_4$ contains 12 moles of hydrogen atoms.

$\therefore$ 12 moles of hydrogen atoms $\equiv$ 1 mole of $(NH_4)_3PO_4$.

$\therefore$ 1 mole of hydrogen atom $= \dfrac{1}{12}$ mole of $(NH_4)_3PO_4$

$\therefore$ 6.36 moles of hydrogen atom $= \dfrac{1}{12} \times 6.36$

$= \dfrac{6.36}{12}$ mole of $(NH_4)_3PO_4$

1 mole of $(NH_4)_3PO_4$ = 4 moles of oxygen

So, $\dfrac{6.36}{12}$ mole of $(NH_4)_3PO_4$ $= \dfrac{4 \times 6.36}{12}$

$= 2.12$ mol of oxygen

58. $CuSO_4 \cdot 5H_2O$ has 1 mole of copper and 9 moles of oxygen atoms.

63.5 g of Cu $\simeq 9 \times 16$ g of oxygen

8.64 g of oxygen $= \frac{63.5 \times 8.64}{9 \times 16} = 3.81$ g of Cu

59. Molecular weight of cortisone $= 360.4$

Molecular weight of 21 carbon atom $= 21 \times 12 = 252$

Percentage of carbon in cortisone $= \frac{252 \times 100}{360.4} = 69.9\%$

61. The mass of KI in 2g salt $= \frac{2 \times 1}{100} = 0.02$ g

$$= \frac{0.02}{39 + 127} \text{ mol}$$

$$= \frac{0.02}{166} \times 6.02 \times 10^{23} \text{ ions}$$

$$= 7.2 \times 10^{19} \text{ ions}$$

62. Mass of one molecule of water

$$= \frac{\text{mol. mass}}{\text{Avagadro number } (N_0)} = \frac{18}{6.02 \times 10^{23}} \text{ g}$$

$\therefore$ Volume of 1 molecule of water $= \frac{\text{mass}}{\text{density}}$

$$= \frac{18 \times 10^{-23}}{6.02 \times 1}$$

$$= 3 \times 10^{-23} \text{ cm}^3$$

63. To find the mass of A and B in the given question, mole concept is used.

$$\text{Number of moles } (n) = \frac{\text{given mass } (w)}{\text{molecular mass } (M)}$$

Compound	Mass of A (g)	Mass of B (g)
AB_2	M_A	$2M_B$
A_2B_2	$2M_A$	$2M_B$

We know that,

$$\text{Number of moles } (n) = \frac{\text{given mass } (w)}{\text{molecular mass } (M)}$$

$$n \times M = w \qquad \ldots(A)$$

Using equation (A), it can be concluded that

$$5(M_A + 2M_B) = 125 \times 10^{-3} \text{ kg} \qquad \ldots\text{(i)}$$

$$10(2M_A + 2M_B) = 300 \times 10^{-3} \text{ kg} \qquad \ldots\text{(ii)}$$

From equation (i) and (ii)

$$\frac{1}{2}\frac{(M_A + 2M_B)}{(2M_A + 2M_B)} = \left(\frac{125}{300}\right)$$

On solving the equation, we obtain

$$M_A = 5 \times 10^{-3}$$

and $M_B = 10 \times 10^{-3}$

So, the molar mass of $A(M_A)$ is 5×10^{-3} kg mol^{-1} and $B(M_B)$ is 10×10^{-3} kg mol^{-1}.

64. In a chemical reaction, coefficient represents mole of that substance.

$$X + 2Y \longrightarrow Z$$

This indicates 1 mole of X reacts with 2 moles of Y to form 1 mole of Z.

So, 5 moles of X will require 10 moles of Y. But we have taken only 9 moles of Y.

Hence, Y is in limiting quantity. Hence, we determine product from Y.

Thus, 5 moles of X react with 9 moles of Y to form 4 moles of Z.

66. $\underset{2\text{ mol}}{MnO_2} + \underset{4\text{ mol}}{4HCl} \longrightarrow MnCl_2 + 2H_2O + \underset{\substack{1\text{ mol}\\22.4\text{ L}}}{Cl_2}$

But the yield is 11.2.

$\therefore$ Percentage yield $= \frac{11.2}{22.4} \times 100 = 50\%$

68. $\underset{4 \times 36.5}{4HCl(aq)} + \underset{87\text{ g}}{MnO_2(s)} \longrightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$

According to the balanced chemical equation,

87 g of MnO_2 react with 4× 36.5 g of HCl

5 g of MnO_2 will react with $\frac{4 \times 36.5 \times 5}{87} = 8.39$ g of HCl

Amounts of one reactant required to react a particular amount of another reactant can be determined by using stoichiometric calculations.

69. $\underset{2V}{2H_2(g)} + \underset{1V}{O_2} \longrightarrow \underset{2V}{2H_2O(g)}$

According to Gay-Lussac's law of gaseous volume, 2 volumes of dihydrogen react with 1 volume of O_2 to produce 2 volumes of water vapour. Therefore, 10 volumes of dihydrogen on reaction with 5 volumes of dioxygen will produce 10 volumes of water vapour.

71. H_2O is the limiting reagent for the given equation.

$\because$ 6 moles of H_2O produces $= 4$ mol of $Fe(OH)_3$

$\therefore$ 2 moles of H_2O will produce $= \frac{4 \times 2}{6} = 1.34$ mol of $Fe(OH)_3$

72. Hydrocarbon (containing C and H) upon burning produces CO_2 and water vapour. The equation is represented as,

$$C_xH_y + (x + y/4)O_2 \longrightarrow xCO_2 + (y/2)H_2O$$

$$\text{Mass of carbon} = \frac{12}{44} \times \text{mass of } CO_2$$

$$= \frac{12}{44} \times 88 \text{ g} = 24 \text{ g}$$

$$\text{Mass of hydrogen} = \frac{2}{18} \times \text{mass of } H_2O$$

$$= \frac{2}{18} \times 9 = 1 \text{ g}$$

So, the unknown hydrocarbon contains 24 g of carbon and 1g of hydrogen.

73. $2\,C_{57}H_{110}O_6(s) + 163\,O_2(g) \longrightarrow 110\,H_2O\,(l) + 114\,CO_2(g)$

Molecular mass of $C_{57}H_{110}O_6$

$$= 2 \times (12 \times 57 + 1 \times 110 + 16 \times 6) \text{ g} = 1780 \text{ g}$$

Molecular mass of 110 $H_2O = 110\,(2+16) = 1980$ g

1780 g of $C_{57}H_{110}O_6$ produced = 1980 g of H_2O.

445 g of $C_{57}H_{110}O_6$ produced $= \frac{1980}{1780} \times 445$ g of H_2O

$= 495$ of H_2O

75. The reactant which is present in the lesser amount, i.e. which limits the amount of product formed is called limiting reagent. When 56 g of N_2 + 10 g of H_2 is taken as a combination then dihydrogen (H_2) act as a limiting reagent in the reaction.

$$\underset{\substack{2\times 14\text{g}\\ 28\text{g}}}{N_2(g)} + \underset{\substack{3\times 2\text{g}\\ 6\text{g}}}{3H_2(g)} \longrightarrow \underset{\substack{2(14+3)\\ 34\text{g}}}{2NH_3(g)} \quad \ldots\text{(i)}$$

28 g of N_2 requires 6 g of H_2 gas.

56 g of N_2 requires $\frac{6\text{ g}}{28\text{ g}} \times 56\text{ g} = 12$ g of H_2

12 g of H_2 gas is required for 56 g of N_2 gas but only 10 g of H_2 gas is present in option (a).

Hence, H_2 gas is the limiting reagent.

In option (b), i.e. 35 g of N_2 + 8 g of H_2.

As 28 g N_2 requires 6 g of H_2.

35g N_2 requires $\frac{6\text{ g}}{28\text{ g}} \times 35\text{ g } H_2 \Rightarrow 7.5$ g of H_2.

Here, H_2 gas does not act as limiting reagent since 7.5 g of H_2 gas is required for 35 g of N_2 and 8 g of H_2 is present in reaction mixture.

Mass of H_2 left unreacted = 8 – 7.5 g of H_2 = 0.5 g of H_2.

Similarly, in option (c) and (d), H_2 does not act as limiting reagent.

For 14 g of N_2 + 4 g of H_2.

As we know 28 g of N_2 reacts with 6 g of H_2.

14 g of N_2 reacts with $\frac{6}{28} \times 14$ g of $H_2 \Rightarrow$ 3 g of H_2.

For 28 g of N_2 + 6 g of H_2, i.e. 28 g of N_2 reacts with 6 g of H_2 (by equation i).

76. (a) $\underset{44\text{ g}}{C_3H_8(g)} + \underset{160\text{ g}}{5O_2(g)} \longrightarrow 3CO_2(g) + 4H_2O(l)$

$\Rightarrow$ 1g of reactant $= \frac{160}{44}$ g of O_2 consumed= 3.64 g

(b) $\underset{124\text{ g}}{P_4(s)} + \underset{160\text{ g}}{5O_2(g)} \longrightarrow P_4O_{10}(s)$

$\Rightarrow$1 g of reactant $\frac{160}{124}$ g of O_2 consumed = 1.29 g

(c) $\underset{244\text{ g}}{4Fe(s)} + \underset{96\text{ g}}{3O_2(g)} \longrightarrow 2Fe_2O_3(s)$

$\Rightarrow$ 1 g of reactant $\frac{96}{224}$ g of O_2 consumed $= 0.43$ g

(d) $\underset{48\text{ g}}{2Mg(s)} + \underset{32\text{ g}}{O_2(g)} \longrightarrow 2MgO(s)$

$\Rightarrow$ 1 g of reactant $\frac{32}{48}$ g of O_2 consumed = 0.67 g

So, minimum amount of O_2 is consumed per gram of reactant (Fe) in reaction (c).

Round II

1. Nitrogen shows variable valency and thus, have variable equivalent weight.

2. For water, 1 g = 1 mL $\qquad (\because d$ for water = 1)

$\therefore \quad$ 18 g = 18 mL

18 mL water $= 6.02 \times 10^{23}$ molecules $= N_A$ molecules

$\because$ In 1000 mL, number of water molecules

$$= \frac{N_A \times 1000}{18} = 55.55\, N_A$$

3. In 100 g haemoglobin, mass of iron = 0.33 g

$\therefore$ In 67200 g haemoglobin, mass of iron $= \frac{67200 \times 0.33}{100}$

$= 672 \times 0.33$ g

$\therefore$ The number of Fe atoms in one Hb molecule

$$= \frac{672 \times 0.33}{56} = 4$$

4. In 12 g carbon, mass of C-14 isotope $= 12 \times \frac{2}{100} = 0.24$ g

$\therefore$ Number of C-14 atoms in 12 g of C $= \frac{0.24}{14} \times 6.02 \times 10^{23}$

$= 1.032 \times 10^{22}$

5. In 1 L air, volume of O_2 = 210 cc

$\because \quad 22400\text{ cm}^3 = 1$ mol

$\therefore \quad 210\text{ cm}^3 = \frac{210}{22400} = 0.0093$ mol

6. [CO_2 and CO] This is an example of law of multiple proportion.

7.
$$\frac{\text{Wt. of metal hydroxide}}{\text{Wt. of metal oxide}} = \frac{\text{Eq. wt. of metal + Eq. wt. of OH}^-}{\text{Eq. wt. of metal + Eq. wt. of O}_2^{2-}}$$

$$\Rightarrow \quad \frac{1.520}{0.995} = \frac{E+17}{E+8}$$

On solving, $\quad E = 9.0$

8. $\underset{208\text{ g}}{BaCl_2} + \underset{98\text{ g}}{H_2SO_4} \longrightarrow \underset{233\text{ g}}{BaSO_4} + 2HCl$

100 mL of 20.8% $BaCl_2$ solution contains 20.8 g $BaCl_2$

50 mL of 9.8% H_2SO_4 solution contains 4.9 g H_2SO_4

Here, H_2SO_4 is the limiting reactant.

$\because$ 98 g of H_2SO_4 gives $BaSO_4$ = 233 g

$\therefore$ 4.9 g of H_2SO_4 will give $BaSO_4 = \frac{233 \times 4.9}{98} = 11.65$ g

9. Molar mass of $CaCO_3 = 40 + 12 + 3 \times 16 = 100\text{ g mol}^{-1}$

Moles of $CaCO_3$ in 1000 g,

$$n_{CaCO_3} = \frac{\text{Mass (g)}}{\text{Molar mass}}$$

$$n_{CaCO_3} = \frac{1000\text{ g}}{100\text{ g mol}^{-1}} = 10\text{ mol}$$

$$\text{Molarity} = \frac{\text{Moles of solute (HCl)} \times 1000}{\text{Volume of solution}}$$

$$0.76 = \frac{n_{HCl} \times 1000}{250}$$

$$n_{HCl} = \frac{0.76 \times 250}{1000} \; 0.19 \text{ mol}$$

$$\underset{1 \text{ mol}}{CaCO_3 (s)} + \underset{2 \text{ mol}}{2HCl (aq)} \longrightarrow CaCl_2(aq) + CO_2 (g) + H_2O (l)$$

According to the equation,

1 mole of $CaCO_3$ reacts with 2 moles of HCl

$\therefore$ 10 moles of $CaCO_3$ will react with $\frac{10 \times 2}{1}$

$= 20$ moles of HCl.

But we have only 0.19 moles HCl, so HCl is the limiting reagent and it limits the yield of $CaCl_2$.

2 moles of HCl, produces 1 mole of $CaCl_2$.

0.19 mole of HCl will produce $\frac{1 \times 0.19}{2} = 0.095$ mol of $CaCl_2$.

Molar mass of $CaCl_2 = 40 + (2 \times 35.5) = 111 \text{ g mol}^{-1}$

$\therefore$ 0.095 mole of $CaCl_2 = 0.095 \times 111 = 10.54$ g

10. Since, 1 g hydrogen combines with 80 g bromine, the equivalent weight of bromine = 80

$\because$ 4 g bromine combines with Ca = 1 g

$\therefore$ 80 g bromine will combine with Ca $= \frac{1 \times 80}{4} = 20$ g

$\therefore$ Equivalent weight of Ca is 20 g.

11. $\underset{264 \text{ g}}{2(NH_4)_2HPO_4} \equiv \underset{142 \text{ g}}{P_2O_5}$

$$\text{Percentage of } P_2O_5 = \frac{\text{wt. of } P_2O_5}{\text{wt. of salt}} \times 100$$

$$= \frac{142}{264} \times 100 = 53.78\%$$

12. Partial pressure of oxygen $= \frac{2}{1 + 4 + 2} \times 2660$

$= 760$ mm

Thus, 1 L oxygen gas is present at 0°C and 760 mm pressure.

$\therefore$ Number of oxygen molecules $= \frac{6.023 \times 10^{23}}{22.4}$

13. $\underset{248 \text{ g}}{Ag_2S} \equiv \underset{2 \times 108 \text{ g}}{2Ag}$

2×108 g Ag is obtained from Ag_2S = 248 g

1 g Ag will be obtained from $Ag_2S = \frac{248 \times 1}{2 \times 108}$

$= \frac{248}{216}$ g

But, the ore contains only 1.34% Ag_2S.

Thus, 1 g Ag is obtained from ore $= \frac{248}{216} \times \frac{100}{1.34}$ g

$= 85.68 \text{ g} \approx 85.7$ g

15. As both the reactants are consumed completely, thus the ratio of stoichiometric coefficients would be 0.75 : 2 or 3 : 8.

So, $3A_4 + 8O_2 \longrightarrow$ Product

Now, as final pressure is half of oxygen initially, thus the molecular formula will be A_3O_4 to balance the equation correctly, i.e.

$$3A_4 + 8O_2 \longrightarrow 4A_3O_4$$

16. $\underset{111 \text{ g}}{CaCl_2} + CO_3^{2-} \longrightarrow \underset{100 \text{ g}}{CaCO_3} + 2Cl^-$

$$\underset{100 \text{ g}}{CaCO_3} \xrightarrow{\Delta} \underset{56 \text{ g}}{CaO} + CO_2$$

$\because$ 56 g CaO is obtained by the decomposition of

$$CaCO_3 = 100 \text{ g}$$

$\therefore$ 0.959 g CaO will be obtained by the decomposition of

$$CaCO_3 = \frac{100 \times 0.959}{56} = 1.71 \text{ g}$$

Further,

100 g of $CaCO_3 \equiv$ 111 g of $CaCl_2$

1.71 g of $CaCO_3 = \frac{111 \times 1.71}{100} = 1.89$ g of $CaCl_2$

Percentage of $CaCl_2$ in the mixture $= \frac{1.89}{4.22} \times 100$

$= 44.78$

$\approx 45\%$

18. For 100 g sample of the solution,

mass of HNO_3 = 63 g, $\therefore n_{HNO_3} = \frac{63}{63} = 1$ mol

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

$$= \frac{100 \text{ g}}{1.4 \text{ g/mL}} = \frac{100}{1.4} \text{ mL} = \frac{1}{14} \text{ L}$$

$$\text{Molarity} = \frac{n}{V(\text{L})} = \frac{1}{(1/14)} = 14 \text{ M}$$

19. Volume of O_2 = 492 L

$$\therefore n = \frac{pV}{RT} = \frac{1 \text{ atm} \times 492 \text{ L}}{0.082 \text{ L atm K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}$$

$$= \frac{492}{8.2 \times 3} \text{ mol}$$

From the reaction,

$$NaClO_3 (s) + Fe(s) \longrightarrow O_2(g) + NaCl(s) + FeO(s)$$

$$n_{NaClO_3} = n_{O_2} = \frac{492}{8.2 \times 3} = 20 \text{ mol}$$

$$m_{NaClO_3} = n \times \text{molar mass of } NaClO_3 \text{ (i.e. 106.5)}$$

$$= 20 \times 106.5 = 2130 \text{ g}$$

The value in different situations may range from 2120 to 2140.

20. Given, Mass of $[Co(NH_3)_6]Cl_3 = 0.3$ g

Molar mass $= 267.46 \text{ g mol}^{-1}$

$$\therefore \quad n = \frac{0.3}{267.46} \text{ mol}$$

From the formula, 1 mol complex will give 3 mol Cl^-.

$\therefore \qquad n_{Cl^-} = \frac{0.3}{267.46} \times 3 = \frac{0.9}{267.46}$

The precipitation reaction is takes place as follows :

$$\underset{(0.3g)}{[Co(NH_3)_6]Cl_3} + \underset{(V\ mL, 0.125g)}{3AgNO_3} \longrightarrow 2AgCl$$

$\therefore \qquad \frac{0.3 \times 3}{267.46} = 0.125 \times V \times 10^{-3}$

$\therefore$ Volume of $AgNO_3$ (in mL) $= \frac{0.9 \times 1000}{0.125 \times 267.46} = 26.92$ mL.

This value may vary from 26.60 to 27.00.

21. Molar mass of AgBr = 108 + 80 $\Rightarrow$ 188 g/mol

Weight of organic compound = 1.6 g

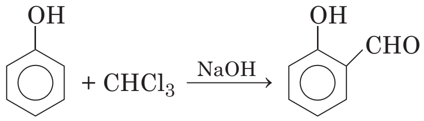
Weight of AgBr = 1.88 g

$$\% \text{ of Br} = \frac{\text{Weight of AgBr} \times 80}{\text{Weight of organic compound} \times 188} \times 100$$

$$= \frac{1.88 \times 80}{1.6 \times 188} \times 100 = 50\%$$

Hence, the correct answer is 50.

22. A solution of phenol in chloroform react with aqueous NaOH gives compound P (salicylaldehyde).



(P)
(Salicylaldehyde)

$\therefore$ Mass % of C in P (Compound)

$$= \frac{\text{Mass of C}}{\text{Mass of compound}} \times 100 \left\{ \begin{array}{l} \text{Here, mass of C} = 12 \times 7 \\ \text{mass of compound } (P) = \\ \quad \text{C} \qquad \text{H} \qquad \text{O} \\ (12 \times 7 + 6 \times 1 + 16 + 2) \end{array} \right\}$$

$$= \frac{12 \times 7}{84 + 6 + 32} \times 100$$

$$= 68.85\% = 69\%$$

23. Combustion of propane

$$\underset{\text{Propane}}{C_3H_8(g)} + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(l)$$

For complete combustion of 1 mole of propane 5 mole of O_2 is required.

Combustion of butane

$$\underset{\substack{\text{Butane} \\ 2\text{ mol}}}{C_4H_{10}(g)} + \frac{13}{2} O_2(g) \longrightarrow 4CO_2(g) + 5H_2O(l)$$

For combustion of 2 moles of butane 13 moles of O_2 is required.

Hence, minimum number of moles of O_2 required to oxidise 1 mole of propane and 2 moles of butane

$$= 5 + 2 \times \frac{13}{2} = 18$$

24. m. eq. of $K_2Cr_2O_7$ = m eq. of FeC_2O_4

$$K_2Cr_2O_7 + 7H_2SO_4 + FeCrO_4 \longrightarrow Fe_2(SO_4)_3 + K_2SO_4 + Cr_2(SO_4)_3 + 4CO_2 + 7H_2O$$

$$FeC_2O_4 + Cr_2O_7^{2-} \longrightarrow Fe^{3+} + CO_2 + Cr^{3+}$$

$$V \times 0.02 \times 6 = \frac{0.288 \times 3 \times 1000}{144} \Rightarrow V = 50 \text{ mL}$$

25. Base + Acid $\longrightarrow$ Salt + H_2O

$$\underset{\text{Base}}{NaOH} + \underset{\text{Phosphonic acid}}{H_3PO_3} \longrightarrow \underset{\text{Salt}}{NaH_2PO_3} + \underset{\text{Water}}{H_2O}$$

$$\text{Given} \left\{ \begin{array}{ll} 0.1\text{ N} & 0.1\text{ N} \\ V = ? & 10\text{ mL} \end{array} \right.$$

On dilution, $(N_1V_1)_{NaOH} = (N_2V_2)_{H_3PO_3}$

$$0.1\, V_1 = 0.1 \times 10 \text{ mL}$$

$$\Rightarrow \qquad V_1 = \frac{0.1 \times 10 \text{ mL}}{0.1} = 10 \text{ mL}$$

Volume of NaOH = 10 mL

26. Let, for glucose

Mass = w_B g, mole = n_B and mole fraction = χ_B

(Molar mass, $M_B = 180$ g mol^{-1})

For water : mass = w_A g,

mole = n_A and mole fraction = χ_A

(Molar mass, $M_A = 18$ g mol^{-1})

where, $\qquad \chi_A + \chi_B = 1$

$$\Rightarrow \qquad \frac{\chi_B}{\chi_A} = \frac{\chi_B}{1 - \chi_B} = \frac{0.1}{0.9} = \frac{1}{9}$$

$$\Rightarrow \qquad \frac{\frac{n_B}{n_A + n_B}}{\frac{n_A}{n_A + n_B}} = \frac{1}{9} \quad \Rightarrow \quad \frac{n_B}{n_A} = \frac{1}{9}$$

$$\Rightarrow \qquad \frac{\frac{w_B}{180}}{\frac{w_A}{18}} = \frac{1}{9} \quad \Rightarrow \quad \frac{w_B}{180} \times \frac{18}{w_A} = \frac{1}{9}$$

$$\Rightarrow \qquad \frac{w_B}{w_A} = \frac{10}{9} \Rightarrow \quad \frac{w_B}{w_A} + 1 = \frac{10}{9} + 1$$

$$\Rightarrow \qquad \frac{w_A}{w_A + w_B} = \frac{9}{19}$$

$$\Rightarrow \qquad \frac{w_A}{w_A + w_B} \% = \frac{9}{19} \times 100 = 47.368 \simeq 47.00$$

So, mass percentage of water (A) in the solution is 47.00.

33. 100 molal aqueous solution means there is 100 mole solute in 1 kg = 1000 gm water. Now,

$$\text{mole fraction of solute} = \frac{n_{\text{solute}}}{n_{\text{solute}} + n_{\text{solvent}}}$$

$$= \frac{100}{100 + \frac{1000}{18}} = \frac{1800}{2800} \simeq 0.6428$$

$$= 64.28 \times 10^{-2}$$