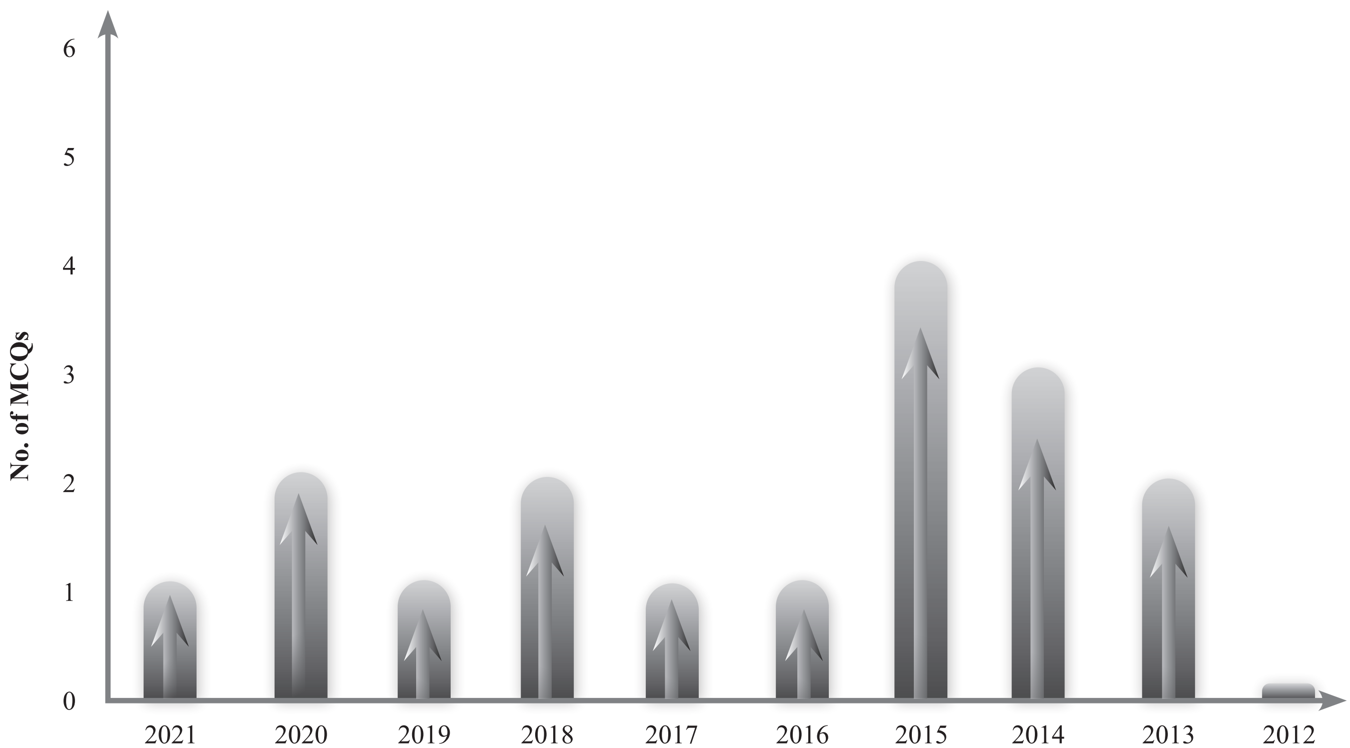
Some Basic Concepts of Chemistry

Past Years NEET Trend



Investigation Report

TARGET EXAM	PREDICTED NO. OF MCQs	CRITICAL CONCEPTS
NEET	1-2	• Mole concept • Stoichiometry and stoichiometric calculations

Perfect Practice Plan

Topicwise Questions	Learning Plus	Multiconcept MCQs	NEET Past 10 Years Questions	Total MCQs
105	39	22	17	183

INTRODUCTION

Chemistry is the branch of science that studies the composition, properties and interaction of matter.

Anything that has mass and occupies space is called **matter**.

For example: book, pen, pencil, water, air, all living beings etc. are composed of matter. You know that they have mass and they occupy space.

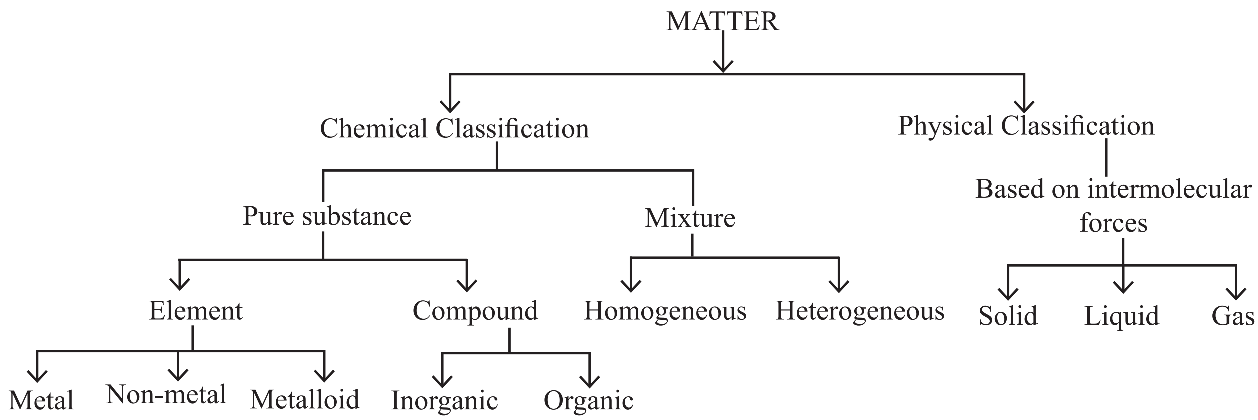
Material is another very common term used in chemistry. However, the term **material** has a limited meaning, which corresponds to matter having specific uses.

Classification of Matter

Matter can be classified in two ways:

(*i*) Physical classification of matter

(*ii*) Chemical classification of matter



Physical Classification of Matter

Depending upon the physical state of matter, it can be classified into three states, namely, **solid, liquid** and **gaseous state**.

Properties	Solid	Liquid	Gas
Shape	Definite	Indefinite	Indefinite
Volume	Definite	Definite	Indefinite
Attraction Force	Strongest	Moderate	Weakest
Examples	Sugar, Iron, Gold, Wood etc.	Water, Milk, oil, Mercury	Dihydrogen, Oxygen, carbon dioxide, etc

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

These three states of matter are interconvertible by changing temperature and pressure.

Chemical Classification of Matter

The chemical classification of matter is based upon its composition. At the macroscopic or bulk level, matter can be classified as **mixture or pure substances**.

Mixture : Mixtures are defined as the substances which are made up of two or more pure substances. They can possess variable composition and can be separated into constituent components by some suitable physical means/methods.

For Example : Alloys (Brass, Bronze) (Brass = Copper + Zinc) (Bronze = Copper + Tin) Water + alcohol, Water + Salt, Water + mustard Oil, Water + Sugar, Water + Kerosene

A mixture may be homogeneous or heterogeneous.

In a **homogeneous mixture**, the components completely mix with each other and its composition is uniform throughout. The components of such a mixture cannot be seen even under a microscope. Some examples of homogeneous mixtures are air, gasoline, sea water, brass, coloured glass, Alloys, Water + alcohol, Water + Salt, 22 carat Gold, Water + Sugar, etc.

Phase : A distinct portion of matter that is uniform in composition and property is called a phase.

In **heterogeneous mixtures**, the composition is not uniform throughout. These consist of two or more parts (called phases) which have different compositions.

For Example : Water + Sand, Water + Mustard oil, Milk, Blood Air, plastic, smoke, petrol etc.

Pure Substances Consist of Single Type of Particles

Pure substances can be further classified into elements and compounds.

1. **Element**: An element is the simplest form of a pure substance. It is defined as:

   The simplest form of a pure substance that can neither be decomposed into nor built from simpler substances by ordinary physical or chemical methods. For example Zn, B, Si.
2. **Compound:** A compound is defined as a pure substance that contains two or more than two elements combined together in a fixed proportion by mass and that can be break down into its constituent elements by suitable chemical methods.

Compounds are further classified into two categories.

1. **Organic Compound**

   For Example : Sources, Oils, fats, derivative of hydrocarbon.
2. **Inorganic Compound**

   For Example : HCl, H_2O, H_2SO_4, $HClO_4$, HNO_3 etc.

TRAIN YOUR BRAIN

Q. Which of the following mixture(s) are homogeneous?
Tap water, Air, Soil, Smoke

Ans. Tap water, Air.

Q. Classify the following as pure substances or mixtures. Also separate the pure substances into elements and compounds and divide mixture, into homogeneous and heterogeneous categories :

(*i*) Graphite (*ii*) Milk
(*iii*) Air (*iv*) Oxygen
(*v*) 22 carat gold (*vi*) Iodized table salt
(*vii*) Wood (*viii*) Cloud

Ans. Element: (*i*), (*iv*) Homogeneous Mixture : (*iii*), (*v*) Heterogeneous Mixture : (*ii*), (*vi*), (*vii*), (*viii*)

PROPERTIES OF MATTER AND THEIR MEASUREMENT

Physical properties are those properties that can be measured or observed without changing the identity or composition of the substance. Example: Colour, Odour etc.

Chemical properties are those in which a chemical change in the substance takes place. Example: pH, Heat of combustion etc.

Expressing a Physical Quantity

The value of a physical quantity is always expressed in two parts:

(*i*) Numerical value and (*ii*) Unit

The International System of Units (SI Units)

The scientists have generally agreed to use the International System of Units abbreviated as SI units.

The SI system has *seven base units* and are listed in table:

Physical Quantity	Symbol for quantity	Name of Unit	Symbol
Length	l	Metre	m
Mass	m	Kilogram	kg
Time	t	second	s
Thermodynamic Temperature	T	kelvin	K
Electric current	I	ampere	A
Amount of Substance	n	mole	mol
Luminous Intensity	I_v	candela	cd

The two temperature units are related as:

- Kelvin temperature (K) = °C + 273.15
- $1\overset{\circ}{A} = 10^{-10}\,m$

  $1nm = 10^{-9}\,m$

  $1pm = 10^{-12}\,m$

Some Commonly used Quantities

1. **Mass and Weight:** Mass of a substance is the amount of matter present in it. The SI unit of mass is kilogram.

   Weight is the force exerted on an object by the pull of gravity.
2. **Volume:** Volume is the amount of space occupied by an object. So in SI system, volume has units cubic meter, m^3.
3. **Density:** Density of a substance is its amount of mass per unit volume. SI unit of density is kg/m^3

Density

It is of two type

- **Absolute density** = $\dfrac{mass}{volume}$
- Relative density or specific gravity

$$= \frac{\text{density of the substance}}{\text{density of water at } 4°C}$$

We know that density of water at 4ºC = 1 g/ml.

For Gases

Absolute density (mass/volume) = $\dfrac{Molar\ mass}{Molar\ volume}$

Relative Density or Vapour Density

Vapour density is defined as the density of the gas with respect to that of hydrogen gas at the same temperature and pressure.

$$\text{Vapour density} = \frac{d_{gas}}{d_{H_2}} = \frac{PM_{gas}/RT}{PM_{H_2}/RT}$$

Where P is pressure of gas, M = mol. wt. of gas, R is the gas constant, T is the temperature.

$$\text{V.D.} = \frac{M_{gas}}{M_{H_2}} = \frac{M_{gas}}{2}$$

$$\mathbf{M_{gas} = 2\ V.D.}$$

Relative density can be calculated w.r.t. to other gases also.

4. **Temperature:** There are three common scales to measure temperature:

   1. The SI scale or Kelvin scale measured in Kelvin (K)
   2. Celsius scale measured in degree Celsius (°C).
   3. Fahrenheit scale measured in degrees Fahrenheit (°F)

(i) Conversion of celsius to Fahrenheit is

$$°F = \frac{9}{5}(°C) + 32°$$

(ii) Conversion of Fahrenheit to celsius

$$°C = \frac{5}{9}[°F - 32°]$$

UNCERTAINTY IN MEASUREMENT

Significant Figures: The uncertainty in the experimental or the calculated values is indicated by mentioning the number of significant figures.

Significant figures are those meaningful digits which are known with certainty. The uncertainty is indicated by writing the certain digits and the last uncertain digit.

Accuracy and Precision

Accuracy is a measure of the difference between the experimental value or the average value of a set of measurements and the true value.

Precision refers to closeness of two or more measurements of the same quantity that agree with one another.

Rules for Determining the Number of Significant Figures:

1. All *non-zero* digits are significant. For example, 3.132 has four significant figures.
2. Zeros between two non zero digits are significant. For example, 3.01 has three significant figures.
3. The zeros preceding to the first non-zero number (i.e. to the left of the first non-zero number) are not significant. Such zeros indicate the position of decimal point. For example, 0.324 has three significant figures.
4. All zeros at the end or to the right of a number are significant provided they are on the right side of the decimal point. For example, 0.0200 has three significant figures.
5. Exponential form: $N \times 10^n$. Where N shows the significant figure.
   E.g., 1.86×10^4 has three significant figure.
6. Rounding off the uncertain digit:
   (*i*) If the left most digit to be rounded off is more than 5, the preceding number is increased by one.
   E.g., 2.16 is rounded to 2.2
   (*ii*) If the left most digit to be rounded off is less than 5, the preceding number is retained.
   E.g., 2.14 is rounded off to 2.1
   (*iii*) If the left most digit to be rounded off is equal to 5, the preceding number is not changed if it is even and increased by one if it is odd.
   E.g., 3.25 is rounded off to 3.2
   2.35 is rounded off to 2.4

TRAIN YOUR BRAIN

Q. How many significant figure are there in each of the following numbers:
(*i*) 1.00×10^6 (*ii*) 0.00010 (*iii*) π

Ans. (*i*) Three (*ii*) Two (*iii*) An infinite number

LAWS OF CHEMICAL COMBINATIONS

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

1. Law of Conservation of Mass/Law of Indestructibility of Matter

Given by – Lavoisier

Tested by – Landolt

According to law of conservation of mass in all physical & chemical changes total mass of the system remains constant.
In a physical or chemical change mass is neither be created nor destroyed.
i.e. Total mass of the reactant = Total mass of the product.
This relationship holds good when reactants are completely converted into products.
If reactants are not completely consume then the relationship will be:

Total mass of reactant

= Total mass of product + Mass of unreacted reaction

> ✍ **KEY NOTE**
> - Nuclear reactions are exception of law of conservation of mass. In nuclear reaction mass + energy is conserved.
> - According to the modern views, the law of conservation of mass is not always valid. The law hold good only in case of such chemical reactions where **there is no evolution of heat or light**.
> - During **chemical processes**, the loss of mass is negligible. But in **nuclear reactions**, tremendous amount of energy is evolved. Consequently, the change in mass is quite significant. Thus, it is clear that the law of conservation of mass and law of conservation of energy are two ways of looking at the same law.
> - Thus, combining the two we get general law known as **law of conservation of mass energy**. It states that, Mass and energy are inter convertible. But the total sum of mass and energy of the system remains constant.

TRAIN YOUR BRAIN

Q. 10 g of $CaCO_3$ on heating gives 4.4 g of CO_2 then determine weight of produced CaO in quintal.

Ans. Total mass of reactant = 10 g
Mass of CO_2 = 4.4 g
Mass of produced CaO = x
According to law of conservation of mass
$10 = 4.4 + x$
$10 - 4.4 = x$
$x = 5.6$ g
$\therefore$ 1 quintal = 100 kg
$\therefore$ 1 Kg = 1000 g

$$= 5.6 \times \frac{Kg}{1000} = 5.6 \times 10^{-3} \times \text{Kg}$$

$$= 5.6 \times 10^{-3} \times \frac{1}{100} \text{ quintal} = 5.6 \times 10^{-5} \text{ quintal}$$

2. Law of Definite Proportions

Given by → Joseph Proust: A chemical compound always contains same elements combined together in same proportion by mass. i.e, chemical compound has a fixed composition & it does not depends on the method of its preparation or the source from which it has been obtained.

Example: Carbon dioxide can be produced by different methods such as burning of carbon, heating lime stone etc. It has been observed that each sample of CO_2 contains carbon and oxygen combined in the ratio 3:8 by mass. This means that the composition of a compound always remain the same irrespective of the method by which it is prepared.

3. Law of Multiple Proportions

Given by → John Dalton

According to this law, if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of smallest whole numbers.

Example: Carbon (C) can combine with oxygen (O) to form more than one compound, namely CO, CO_2. Here ratio of masses of O that combine with fixed mass of C is 16:32 or 1:2.

TRAIN YOUR BRAIN

Q. On analysis it was found that the black oxide of copper and the red oxide of copper contain 79.9% and 88.8% metal respectively. Establish the law of multiple proportions with the help of this data.

Ans. In the black oxide, 79.9 g copper combines with (100 – 79.9), i.e. 20.1 g oxygen

$\therefore$ In red oxide 88.8 g copper will combine with 100 – 88.8 = 11.2 g

$\therefore$ According to red oxide 79.9 copper will combine with $\frac{11.2\times 79.9}{88.8}$ = = 10.08 g oxygen

Thus the weights of oxygen that combine with the same 79.9 g copper are 20.1 g and 10.08 g respectively. These are in the ratio 20.1 : 10.08 = 2 : 1

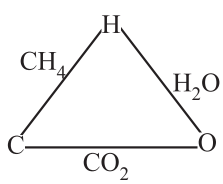
It is a simple whole number ratio. Hence, the law of multiple proportions is established.

4. Law of Reciprocal Proportion

Given by → Richter: The ratio of the weights of two elements A and B that combine separately with fixed weight of the third element C is either the same or some simple multiple of this ratio of the weights in which A and B combine directly with each other.

Example: The elements C and O combine separately with the third element H to form CH_4 and H_2O and they combine directly with each other to form CO_2 as shown in the below figure.

In CH_4, 12 parts by weight of carbon combine with 4 parts by weight of hydrogen. In H_2O, 2 parts by weight of hydrogen combine with 16 parts by weights of oxygen. Thus the weights of C and O which combine with fixed weight of hydrogen (say 4 parts of weight) are 12 and 32, i.e. they are in the ratio 12 : 32 or 3 : 8.



Now in CO_2, 12 parts by weight of carbon combine directly with 32 parts by weight of oxygen, i.e. they combine directly in the ratio 12 : 32 or 3 : 8 that is the same as the first ratio.

TRAIN YOUR BRAIN

Q. Copper sulphide contains 66.6% Cu, copper oxide contains 79.9% copper and sulphur trioxide contains 40% sulphur. Show that these data illustrate law of reciprocal proportions.

Ans. In copper sulphide,

Cu : S mass ratio is 66.6 : 33.4

In sulphur trioxide, O : S mass ratio is 60 : 40

Now in copper sulphide

33.4 parts of sulphur combines with Cu = 66.6 parts

40.0 parts of sulphur combines with Cu.

$$= \frac{66.6\times 40}{33.4} = 79.8 \text{ parts}$$

Now ratio of the masses of Cu and O which combines with same mass (40 parts) of sulphur separately is 79.8 : 60

Cu : O ratio by mass in CuO is 79.9 : 20.1

$$\text{Ratio I : Ratio II} = \frac{79.8}{60}\times\frac{20.1}{79.9} = 3:1$$

Which is simple whole number ratio.

Hence, law of reciprocal proportion is proved.

5. Gay Lussac's Law of Gaseous Volumes

Given by → Gay Lussac: He observed 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.

Example: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

$2H_2(g)$	+	$O_2(g)$	$\rightarrow$	$2H_2O(g)$
100 ml		50 ml		100 ml
2 volumes		1 volume		2 volumes
vol of H_2	:	vol of O_2	:	vol of steam
2	:	1	:	2

TRAIN YOUR BRAIN

Q. For the gaseous reaction $H_2 + Cl_2 \rightarrow 2HCl$

If 40 ml of hydrogen completely reacts with chlorine then find out the required volume of chlorine and volume of produced HCl ?

Ans. According to Gay Lussac's Law :

$H_2 + Cl_2 \rightarrow 2HCl$

1 ml of H_2 will react will 1 ml of Cl_2 and 2 ml of HCl will be produced.

$\therefore$ 40 ml of H_2 will react with 40 ml of Cl_2 and 80 ml of HCl will produce.

Required vol. of Cl_2 = 40 ml

Produced vol. of HCl = 80 ml.

6. Avogadro Law

Given by → Amedeo Avogadro

Avogadro proposed that equal volumes of gases at the same temperature and pressure should contain equal number of molecules.

Example: 22.4 L of every gas at STP (Standard temperature and Pressure, ie. T = 273 K, P = 1 atm) contains equal number of molecules, which is equal to 6.022×10^{23}

TRAIN YOUR BRAIN

Q. Which of the following contains the largest number of oxygen atoms? 1.0 g of O atoms, 1.0 g of O_2, 1.0 g of ozone O_3.

Ans. All have the same number of oxygen atoms.

DALTON'S ATOMIC THEORY

The assumption of Dalton's Atomic theory are:

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

   Dalton's theory could explain the laws of chemical combination.

The main failures of Dalton's atomic theory are:

1. It failed to explain how atoms of different elements differ from each other i.e., did not tell anything about structure of the atom.

2. It does not explain how and why atoms of different element combine with each other to form compound.
3. It failed to explain the nature of forces present between different atoms in a molecule.
4. It fails to explain Gay Lussac's law of combining volumes.
5. It did not make any difference between ultimate particle of an element that takes part in reaction (atoms) and ultimate particle that has independent existence (molecules).

ATOMIC MASS AND MOLECULAR MASS

Atomic Mass Unit

It is defined as exactly $\frac{1}{12}$th of the mass of a carbon-12 atom. It is represented as amu. [Now a new symbol 'u' called unified mass is used].

$$\text{Mass of 1 amu} = \frac{12}{6.022\times10^{23}}\times\frac{1}{12} = 1.67 \times 10^{-24}\ \text{g}$$

Today, 'amu' has been replaced by 'u' which is known as unified mass.

Average Atomic Mass

When we take into account the existence of the isotopes and their relative abundance (Percent occurrence), the average atomic mass of that element is calculated.

Average atomic mass of an element is the sum of the masses of its isotopes each multiplied by its natural abundance.

Mathematically, average atomic mass of X (A_x)

$$= \frac{a_1x_1 + a_2x_2 + ... + a_nx_n}{100}$$

a_1 = atomic mass ; x_1 % occurence in nature

> ✍ **KEY NOTE**
>
> Relative atomic mass is nothing but the number of nucleons present in the atom.

TRAIN YOUR BRAIN

Q. Naturally occuring chlorine is 75% Cl^{35} which has an atomic mass of 35 amu and 25% Cl^{37} which has a mass of 37 amu. Calculate the average atomic mass of chlorine -

(*a*) 35.5 amu (*b*) 36.5 amu
(*c*) 71 amu (*d*) 72 amu

Ans. (a) Average atomic mass =

$$\frac{\%\ of\ \text{I}\ isotope \times its\ atoms\ mass + \%\ of\ \text{II}\ isotope \times its\ atomic\ mass}{100}$$

$$\frac{75 \times 35 + 25\times 37}{100} = 35.5\ \text{amu}$$

Molecular Mass

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

Formula Mass

In ionic compounds we use formula mass instead of molecular mass. **Formula mass** of an ionic compound is the sum of the atomic masses of all atoms in a formula unit of compound.

> ✍ **KEY NOTE**
>
> **Equivalent Mass (E.M)**
>
> - E.M. of an element $= \frac{\text{Atomic mass}}{\text{Valency}}$
> - E.M. of an acid $= \frac{\text{Molecular mass}}{\text{Basicity}}$
> - E.M. of a base $= \frac{\text{Molecular mass}}{\text{Acidity}}$

MOLE CONCEPT AND MOLAR MASSES

'Mole' was introduced as the seventh base quantity for the amount of substance in SI system.

One mole of a substance contains as many particles and their number is equal to the number of particles in 12 g of the ^{12}C isotope. This number is known as avogardo constant ($N_A = 6.022 \times 10^{23}$).

Mole Concept in Gaseous Reaction

Molar volume is the mole related to volume of gaseous substance. The volume occupied by 1 mol of a gaseous substance is called molar volume. 1 mole occupies 22.414 L or 22414 ml at STP ie. 273 K and 1 atm.

$$\text{Number of moles} = \frac{\text{Volume}}{\text{Molar volume}}$$

Molar Mass

The mass of 1 mol of a substance in grams is called its molar mass.

Mass-Mole-Number Relationship

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molar mass in g mol}^{-1}}$$

TRAIN YOUR BRAIN

Q. The molecular mass of H_2SO_4 is 98 amu. Calculate the number of moles of each element in 294 g of H_2SO_4.

Ans. Gram molecular mass of H_2SO_4 = 98 gm

$$\text{moles of } H_2SO_4 = \frac{294}{98} = 3 \text{ moles}$$

H_2SO_4	H	S	O
One molecule	2 atom	one atom	4 atom
$1 \times N_A$	$2 \times N_A$ atoms	$1 \times N_A$ atoms	$4 \times N_A$ atoms
∴ One mole	2 mole	one mole	4 mole
∴ 3 mole	6 mole	3 mole	12 mole

PERCENTAGE COMPOSITION

We know that according to law of definite proportions any sample of a pure compound always possess constant ratio with their combining elements.

Mass percentage of an element

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

TRAIN YOUR BRAIN

Q. Every molecule of ammonia always has formula NH_3 irrespective of method of preparation or sources. i.e. 1 mole of ammonia always contains 1 mole of N and 3 mole of H. In other words 17 gm of NH_3 always contains 14 gm of N and 3 gm of H. Now find out % of each element in the compound.

Ans. Mass % of N in NH_3 $= \dfrac{Mass\ of\ \ N\ in\ 1\ mol\ \ NH_3}{Mass\ of\ \ 1\ mol\ \ of\ \ NH_3} \times 100$

$= \dfrac{14\ gm}{17\ gm} \times 100 = 82.35\ \%$

Mass % of H in NH_3 $= \dfrac{Mass\ of\ \ H\ in\ 1\ mol\ \ NH_3}{Mass\ of\ \ 1\ mol\,e\ \ of\ \ NH_3} \times 100$

$= \dfrac{3}{17} \times 100 = 17.65\ \%$

Chemical Formula

It is of two types:

(a) Empirical formula: It represent the simplest whole number ratio of various atoms present in a compound. eg. EF of benzene (C_6H_6) is CH.

(b) Molecular formula: It shows the exact number of different types of atoms present in a molecule of a compound.
eg., MF of benzene is C_6H_6

Determination of Chemical Formula

(a) Determination of empirical formula:

Step (I): Determination of percentage of each element

Step (II): Determination of mole ratio

Step (III): Making it whole number ratio

Step (IV) : Simplest whole number ratio

(b) Determination of Molecular Formula

$$MF = (EF) \times n;$$

Where n is a simple whole number.
Molecular formula = n × Empirical formula

$= 2 \times (C_5H_4) = C_{10}H_8$

> ✍ **KEY NOTE**
>
> $$n = \frac{\text{Molecular weight}}{\text{Empirical weight}}$$

TRAIN YOUR BRAIN

Q. Phosgene, a poisonous gas used during World war-I, contains 12.1% C, 16.2% O and 71.7% Cl by mass. What is the empirical formula of phosgene.

(*a*) $COCl_2$ (*b*) COCl (*c*) $CHCl_3$ (*d*) $C_2O_2Cl_4$

Ans. (a)

Element	Symbol	% of element	A.mu of element	Relative no. of atoms	Simplest ratio	Simple whole no. atomic ratio
Carbon	C	12.1	12	$\frac{12.1}{12} = 1.01$	$\frac{1.01}{1.01} = 1$	1
Oxygen	O	16.2	16	$\frac{16.2}{16} = 1.01$	$\frac{1.01}{1.01} = 1$	1
Chlorine	Cl	71.7	35.5	$\frac{71.7}{35.5} = 2.02$	$\frac{2.02}{1.01} = 2$	2

Then empirical formula = $COCl_2$

Q. 1.615 g of anhydrous $ZnSO_4$ was left in moist air. After a few days its weight was found to be 2.875 g. What is the molecular formula of hydrated salt?

(AT. masses: Zn = 65.5, S = 32, O = 16, H = 1)

Ans. Molecular mass of anhydrous $ZnSO_4$
= 65.5 + 32 + 4 × 16 = 161.5 g
So, 1.615 g of anhydrous $ZnSO_4$ combine with water
= 2.875 – 1.615 = 1.260 g
1.615 g of anhydrous $ZnSO_4$ combine with water = 1.260 g
161.5 g of anhydrous $ZnSO_4$ combine with

$$= \frac{1.260}{1.615} \times 161.5 = 126\,g$$

No. of moles of water $= \dfrac{126}{18} = 7$

Hence, Formula is $ZnSO_4 . 7H_2O$.

STOICHIOMETRY AND STOICHIOMETRIC CALCULATIONS

Chemical Equation and Balanced Chemical Equation

Chemical Reaction: It is a process in which two or more than two substances interact with each other where old bonds are broken and new bonds are formed.

Chemical equation is a scientific method of representing a chemical change in terms of symbols and formula of reactants and products involved in it.

e.g., $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

However, a balanced chemical equation tells us a lot of quantitative information. Mainly the molar ratio in which reactants combine and the molar ratio in which products are formed.

Features of a Balanced Chemical equation

(*a*) It contains a same number of atoms of each element on both sides of equation.(POAC)
(*b*) It should follow law of charge conservation on both side.
(*c*) Physical states of all the reagents/reactants should be included in brackets.
(*d*) All reagents/reactants should be written in thier standard forms (Molecular, Atomic, Solid etc.)
(*e*) The coefficients give the relative molar ratios of each reagent/reactant.

Stoichiometry deals with the calculation of masses and sometimes volumes also of the reactants and products in a reaction. The coefficients of reactants and products in a balanced chemical equation is called the **stoichiometric coefficients**.

Steps:

1. Write the balanced chemical equation.
2. See the number of moles of various reactants that take part in the reaction and products formed.
3. Calculate the number of moles or amount of substance formed.

Interpretation of balanced chemical equations:

Once we get a balanced chemical equation then we can interpret a chemical equation by following ways:

- Mass - mass analysis
- Mass - volume analysis
- Mole - mole analysis

Mass-mass Analysis

In the following reaction

$$Mass - mass\ ratio: \underset{2\times122.5}{2KClO_3} \rightarrow \underset{2\times74.5}{2KCl} + \underset{3\times32}{3O_2} \quad \left(\begin{array}{l}\text{According to stoichiometry}\\ \text{of the reaction}\end{array}\right)$$

or $$\frac{Mass\ of\ KClO_3}{Mass\ of\ KCl} = \frac{2\times122.5}{2\times74.5} \qquad \frac{Mass\ of\ KClO_3}{Mass\ of\ O_2} = \frac{2\times122.5}{3\times32}$$

Mass–Volume Analysis

Considering decomposition of $KClO_3$

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

mass volume ratio : 2 × 122.5 g : 2 × 74.5 g : 3 × 22.4 litre at STP

We can use two relation for volume of oxygen

$$\frac{Mass\ of\ KClO_3}{volume\ of\ O_2\ at\ STP} = \frac{2\times122.5}{3\times22.4\ lt} \qquad ...(i)$$

$$\frac{Mass\ of\ KCl}{volume\ of\ O_2\ at\ STP} = \frac{2\times74.5}{3\times22.4\ lt} \qquad ...(ii)$$

Mole-mole Analysis

This analysis is very much important for quantative analysis point of view. Consider again the decomposition of $KClO_3$.

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

In very first step of mole-mole analysis you should read the balanced chemical equation like **2 *moles* $KClO_3$ *on decomposition gives us* 2 *moles KCl and* 3 *moles* O_2.** and from the stoichiometry of reaction we can write

$$\frac{Moles\ of\ KClO_3}{2} = \frac{Moles\ of\ KCl}{2} = \frac{Moles\ of\ O_2}{3}$$

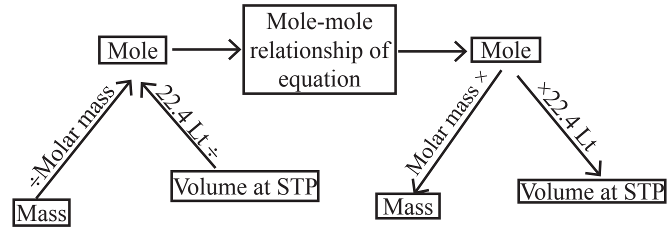
Now for any general balance chemical equation like

$$aA + bB \rightarrow cC + dD$$

You can write.

$$\frac{Moles\ of\ A\ reacted}{a} = \frac{moles\ of\ B\ reacted}{b} = \frac{moles\ of\ C\ formed}{c} = \frac{moles\ of\ D\ formed}{d}$$

> ✍ **KEY NOTE**
>
> In fact mass-mass and mass-vol analysis are also interpreted in terms of mole-mole analysis you can see in the following chart also.



TRAIN YOUR BRAIN

Q. 367.5 gram $KClO_3$ (M = 122.5) when heated how many gram of KCl and oxygen is produced.

Ans. Balance chemical equation for heating of $KClO_3$ is

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

mass-mass ratio : 2 × 122.5 g : 2 × 74.5 g : 3 × 32 g

$$\frac{mass\ of\ KClO_3}{mass\ of\ KCl} = \frac{2\times122.5}{2\times74.5} \Rightarrow \frac{367.5}{W} = \frac{122.5}{74.5}$$

W = 3 × 74.5 = 223.5 g

$$\frac{Mass\ of\ KClO_3}{Mass\ of\ O_2} = \frac{2\times122.5}{3\times32} \Rightarrow \frac{367.5}{W} = \frac{2\times122.5}{3\times32}$$

W = 144 g

Q. 367.5 g $KClO_3$ (M = 122.5) when heated, how many litre of oxygen gas is produced at STP

Ans. $$\frac{mass\ of\ KClO_3}{volume\ of\ O_2\ at\ STP} = \frac{2\times122.5}{3\times22.4\ lt} \Rightarrow \frac{367.5}{V} = \frac{2\times122.5}{3\times22.4\ lt}$$

V = 3 × 3 × 11.2 ⇒ V = 100.8 lt

Limiting Reagent

- The reactant that gets consumed during the reaction & limits the amount of product formed is known as the limiting reagent.
- Limiting reagent is present in least stoichiometric amount and therefore, controls amount of product.
- The remaining or leftout reactant is called the excess reagent.
- If we are dealing with balance chemical equation then if number of moles of reactants are not in the ratio of stoichiometric coefficient of balanced chemical equation, then there should be one reactant which should be limiting reactant.

TRAIN YOUR BRAIN

Q. Three moles of Na_2CO_3 are reacted with 6 moles of HCl solution. Find the volume of CO_2 gas produced at STP. The reaction is $Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$

Ans. From the reaction : $Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$

given moles	3 mol		6 mol
given mole ratio	1	:	2
Stoichiometric coefficient ratio	1	:	2

Given moles of reactant are in stoichiometric coefficient ratio therefore no reactant is left over.

Mole-mole analysis to calculate V of CO_2 produced at STP

$$\frac{Moles\ of\ Na_2CO_3}{1} = \frac{Mole\ of\ CO_2\ Produced}{1}$$

Moles of CO_2 produced = 3

Volume of CO_2 produced at STP = 3 × 22.4 L = 67.2 Lt

How to Find Limiting Reagent

Step: I Divide the given moles of reactant by the respective stoichiometric coefficient of that reactant.

Step: II See that for which reactant this division come out to be minimum. The reactant having minimum value is limiting reagent for you.

TRAIN YOUR BRAIN

Q. 6 moles of Na_2CO_3 and 4 moles of HCl are made to react. Find the volume of CO_2 gas produced at STP. The reaction is

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 + H_2O$

Ans. From Step I & II Na_2CO_3 HCl

$$\frac{6}{1} = 6 \qquad \frac{4}{2} = 2 \ (\textit{division is minimum})$$

$\therefore$ **HCl is limiting reagent**

From Step III

$$\frac{\textit{Mole of HCl}}{2} = \frac{\textit{Moles of } CO_2 \textit{ produced}}{1}$$

$\therefore$ mole of CO_2 produced $= \frac{4}{2} = 2$

$\therefore$ volume of CO_2 produced at S.T.P. $= 2 \times 22.4 = 44.8$ lt.

Reactions in Solutions

The concentration of a solution or the amount of substance present in its given volume can be expressed in any of the following ways.

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

Mass Percent

It is obtained by using the following relation:

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

Mole Fraction

It is no. of moles of a certain component to the total no. of moles of the solution.

$$\text{Mole fraction of A} = \frac{\text{No.of moles of A}}{\text{No.of moles of solution}} = \frac{n_A}{n_A + n_B}$$

Mole fraction is a pure number. It will remain independent of temperature changes.

Molarity

It is defined as the number of moles of the solute in 1 liter of the solution. It is denoted by M

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

Molarity is an unit that depends upon temperature. It varies inversely with temperature.

Mathematically : Molarity decreases as temperature increases.

$$\textit{Molarity} \propto \frac{1}{\textit{temperature}} \propto \frac{1}{\textit{volume}}$$

If a particular solution having volume V_1 and molarity M_1 is diluted to V_2 mL then

$M_1V_1 = M_2V_2$

M_2 : Resultant molarity

If a solution having volume V_1 and molarity M_1 is mixed with another solution of same solute having volume V_2 & molarity M_2

then $M_1V_1 + M_2V_2 = M_R (V_1 + V_2)$

$$M_R = \text{Resultant molarity} = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

TRAIN YOUR BRAIN

Q. 149 gm of potassium chloride (KCl) is dissolved in 10 Lt of an aqueous solution. Determine the molarity of the solution (K = 39, Cl = 35.5)

Ans. Molecular mass of KCl = 39 + 35.5 = 74.5 gm

$\therefore$ Moles of KCl $= \frac{149 \text{ gm}}{74.5 \text{ gm}} = 2$

$\therefore$ Molarity of the solution $= \frac{2}{10} = 0.2\ M$

Molality

It is defined as the number of moles of solute present in 1 kg of solvent. it is denoted by m.

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

Molality is independent of temperature changes.

There are other terms also used to express concentration of solution

Normality (N)

It is the number of gram equivalent of a solute dissolved per liter of the solution.

$$\text{Normality}(N) = \frac{\text{No. of gram equivalents of solute}}{\text{Vol. of solution in litres}}$$

$$= \frac{\text{Mass of solute in gram}}{\text{Equivalent weight in gram} \times \text{vol.of soltution in litres}}$$

Normality equation: $N_1V_1 = N_2V_2$

$$\text{Formality} = \frac{\text{Wt. of ionic solute}}{\text{Formula Wt. of solute} \times \text{Vol. in lit.}}$$

TRAIN YOUR BRAIN

Q. 255 g of an aqueous solution contains 5 g of urea. What is the concentration of the solution in terms of molality. (Mol. wt. of urea = 60)

Ans. Mass of urea = 5 g

Molecular mass of urea = 60 g

Number of moles of urea = 0.083

Mass of solvent = (255 – 5) = 250 g

$\therefore$ Molality of the solution

$$= \frac{\text{Number of moles of solute}}{\text{Mass of solvent in gram}} \times 1000$$

$$= \frac{0.083}{250} \times 1000 = 0.332 \text{ m}$$

TRAIN YOUR BRAIN

Q. 0.5 g of a substance is dissolved in 25 g of a solvent. Calculate the percentage amount of the substance in the solution.

Ans. Mass of substance = 0.5 g

Mass of solvent = 25 g

$$\therefore \text{ Percentage of the substance (w/w)} = \frac{0.5}{0.5+25} \times 100$$

$$= 1.96$$

Q. 20 cm^3 of an alcohol is dissolved in 80 cm^3 of water. Calculate the percentage of alcohol in solution.

Ans. Volume of alcohol = 20 cm^3

Volume of water = 80 cm^3

$$\therefore \text{ percentage of alcohol} = \frac{20}{20+80} \times 100 = 20.$$

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

Ans. Molarity of solution (mol L^{-1})

$$= \frac{\text{mass of solute (g)}}{\text{M. Mass}} \times \frac{1000}{\text{V in mL}}$$

conc. of sugar $= \frac{20}{342} \times \frac{1000}{2000} = 0.0292 \; mol \; L^{-1}$

Topicwise Questions

UNCERTAINTY IN MEASUREMENT

1. The number of significant figures in 0.0045 are

(*a*) Two (*b*) Three (*c*) Four (*d*) Five

2. Light travels with a speed of 3×10^8 m/sec. The distance travelled by light in 1 Femto sec is:

(*a*) 0.03 mm

(*b*) 0.003 mm

(*c*) 3 mm

(*d*) 0.0003 mm

3. Area of nuclear cross-section is measured in "Barn". It is equal to:

(*a*) 10^{-20} m^2 (*b*) 10^{-30} m^2

(*c*) 10^{-28} m^2 (*d*) 10^{-14} m^2

4. Two students X and Y report the mass of the same substance as 7.0 g and 7.00 g respectively, which of the following statement is correct?

(*a*) Both are equally accurate

(*b*) X is more accurate than Y

(*c*) Y is more accurate than X

(*d*) Both are inaccurate scientifically

5. The number of significant figures in value of π are:

(*a*) 1 (*b*) 2 (*c*) 3 (*d*) ∞

6. 5.041 has how many significant figures.

(*a*) 1 (*b*) 2 (*c*) 3 (*d*) 4

7. The correctly reported answer of the addition of 29.4406, 3.2 and 2.25 will have significant figures:

(*a*) 3 (*b*) 4 (*c*) 2 (*d*) 5

8. What is the area of rectangle which is 12.34 cm wide and 1.23 cm long?

(*a*) 15.2 m^2 (*b*) 15.18 cm^2

(*c*) 16.2 cm^2 (*d*) 16.2 m^2

9. If an object has a mass of 0.2876 g, then find the mass of nine such objects:

(*a*) 2.5884 g (*b*) 2.5886 g (*c*) 2.588 g (*d*) 2.5 g

10. The value of Plank's constant is 6.62618×10^{-34} Js. The number of significant figures in it is

(*a*) Six (*b*) Five

(*c*) Three (*d*) Thirty four

LAW OF CHEMICAL COMBINATIONS

11. In Habers process, the volume at S.T.P of ammonia relative to the total volume of reactants at STP is :

(*a*) One fourth (*b*) One half

(*c*) Same (*d*) Three fourth

12. 6 g of carbon combines with 32 g of sulphur to form CS_2, 12 g of C also combine with 32 g oxygen to form CO_2. 10 g of sulphur combines with 10 g of oxygen to form Sulphur dioxide. Which law is illustrated by this?

(*a*) Law of multiple proportions

(*b*) Law of constant composition

(*c*) Law of reciprocal proportions

(*d*) Gay Lussac's law

13. Which of the following data illustrates the law of conservation of mass?

(*a*) 56 g of C reacts with 32 g of Oxygen to produce 44 g of CO_2

(*b*) 1.70 g of $AgNO_3$ reacts with 100 ml of 0.1M HCl to produce 1.435 g of AgCl and 0.63 g of HNO_3

(*c*) 12 g of C is heated in vacuum and on cooling, there is no change in mass

(*d*) 36 g of S reacts with 16 g of O_2 to produce 48 g of SO_2

14. One part of an element A combines with two parts of another element B, 6 parts of element C combines with 4 parts of B. If A and C combine together the ratio of their weights, will be governed by

(*a*) law of definite proportion

(*b*) law of multiple proportion

(*c*) law of reciprocal proportion

(*d*) law of conservation of mass

15. The law of conservation of mass holds good for all of the following except.

(*a*) All chemical reactions (*b*) Nuclear reaction

(*c*) Endothermic reactions (*d*) Exothermic reactions

16. The % of copper and oxygen in samples of CuO obtained by different methods were found to be the same. This proves the law of:

(*a*) Constant Proportion (*b*) Reciprocal Proportion

(*c*) Multiple Proportion (*d*) Conservation of mass.

17. Two elements X and Y combine in gaseous state to form XY in the ratio 1:35.5 by mass. The mass of Y that will be required to react with 2 g of X is:

(*a*) 7.1 g (*b*) 3.55 g

(*c*) 71 g (*d*) 35.5 g

18. 4.4 g of an oxide of nitrogen gives 2.24 L of nitrogen and 60 g of another oxide of nitrogen gives 22.4 L of nitrogen at S.T.P. The data illustrates:

(*a*) Law of conservation of mass

(*b*) Law of constant proportions

(*c*) Law of multiple proportions

(*d*) Law of reciprocal proportions

19. The law of multiple proportions is illustrated by the two compounds

(*a*) Sodium chloride and sodium bromide

(*b*) Ordinary water and heavy water

(*c*) Caustic soda and caustic potash

(*d*) Sulphur dioxide and sulphur trioxide.

20. How many grams of H_3PO_4 is required to completely neutralize 120g of NaOH

(*a*) 49 (*b*) 98 (*c*) 196 (*d*) 9.8

21. The weight of oxygen required to completely react with 27 gm of 'Al' is

(*a*) 8 gm (*b*) 16 gm (*c*) 32 m (*d*) 24 gm

22. If law of conservation of mass was to hold true, then 20.8 g of $BaCl_2$ on reaction with 9.8 g of H_2SO_4 will produce 7.3 g of HCl and $BaSO_4$ equal to:

(*a*) 11.65 g (*b*) 23.3 g (*c*) 25.5 g (*d*) 30.6 g

23. One of the following combinations which illustrates the law of reciprocal proportions is:

(*a*) N_2O_3, N_2O_4, N_2O_5 (*b*) NaCl, NaBr, NaI

(*c*) CS_2, CO_2, SO_2 (*d*) PH_3, P_2O_3, P_2O_5

24. 23g of sodium will react with ethyl alcohol to give

(*a*) 1 mole of H_2 (*b*) 1/2 mole of H_2

(*c*) 1 mole of O (*d*) 1 mole of NaOH

25. Hydrogen and oxygen combine to form H_2O_2 and H_2O containing 5.93% and 11.2% hydrogen respectively, the data illustrates:

(*a*) Law of conservation of mass

(*b*) Law of Constant proportions

(*c*) Law of reciprocal proportions

(*d*) Law of multiple proportions

26. Two elements X (of mass 16) and Y (of mass 14) combine to form compounds A, B and C. The ratio of different masses of Y which combine with a fixed mass of X in A, B and C is 1 : 3 : 5, if 32 parts by mass of X combines with 84 parts by mass of Y in B, then in C, 16 parts by mass of X will combine with;

(*a*) 14 parts by mass of Y (*b*) 42 parts by mass of Y

(*c*) 70 parts by mass of Y (*d*) 84 parts by mass of Y

ATOMIC AND MOLECULAR MASSES

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

(*a*) 94.117 u (*b*) 1884 u

(*c*) 941 u (*d*) 976 u

28. If we assume 1/24 th part of mass of carbon instead of 1/12 th part of it as 1 amu., mass of 1 mole of a substance will

(*a*) Remain unchanged (*b*) get doubled

(*c*) Get halved (*d*) can't be predicted

29. Boron has two isotopes 10B and 11B whose relative abundances are 20% and 80% respectively. Atomic weight of Boron is

(*a*) 10 (*b*) 11

(*c*) 10.5 (*d*) 10.8

30. 10 grams of each O_2, N_2 and Cl_2 are kept in three bottles. The correct order of arrangment of bottles containing decreasing number of Molecules.

(*a*) O_2, N_2, Cl_2 (*b*) Cl_2, N_2, O_2

(*c*) Cl_2, O_2, N_2 (*d*) N_2, O_2, Cl_2

31. Avogadro's number is the number of molecules present in

(*a*) 1 g of molecule (*b*) 1atom of molecule

(*c*) gram molecular mass (*d*) 1 litre of molecule

32. Maximum number of atoms are present in

(*a*) 14 gms. of carbon monoxide

(*b*) 2 gms. of hydrogen

(*c*) 11.2 lit. of nitrogen at STP

(*d*) 1.5 gm atoms of helium

33. One amu is equal to

(*a*) 1.66×10^{-8} g (*b*) 1.66×10^{-4} g

(*c*) 1.66×10^{-16} g (*d*) 1.66×10^{-24} g

34. The number of molecules present in one milli litre of a gas at STP is known as

(*a*) Avogadro number

(*b*) Boltzman number

(*c*) Loschmidt number

(*d*) Universal gas constant

35. Which of the following gases contain the same number of molecules as that of 16 grams of oxygen?

(*a*) 16gm of O_3 (*b*) 32 grams of SO_2

(*c*) 16gm of SO_2 (*d*) All

36. If the atomic mass unit 'u' were defined to be $\frac{1}{5}^{th}$ of the mass of an atom of C-12, what would be the atomic weight of nitrogen in amu or 'u' in this state? Atomic weight of N on conventional scale is 14:

(*a*) 6.77 u (*b*) 5.834 u

(*c*) 14 u (*d*) 23 u

37. A 100 g sample of Haemoglobin on analysis was found to contain 0.34% Fe by mass. If each haemoglobin molecule has four Fe^{2+} ions, the molecular mass of haemoglobin is (Fe = 56 amu)

(*a*) 77099.9 g (*b*) 12735 g

(*c*) 65882 g (*d*) 96359.9 g

MOLE CONCEPT AND MOLAR MASSES

38. 1 g-atom of nitrogen represents:

(*a*) 6.02×10^{23} N_2 molecules

(*b*) 22.4 L of N_2 at S.T.P

(*c*) 11.2 L of N_2 at S.T.P

(*d*) 28 g of nitrogen

39. Which is correct for 10 g of $CaCO_3$?

(*a*) It contains 1 g atom of carbon

(*b*) It contains 0.3 g atoms of oxygen

(*c*) It contains 12 g of calcium

(*d*) It refers to 0.1 g equivalent of $CaCO_3$

40. The number of oxygen atoms present in 14.6 g of magnesium bicarbonate is:

(*a*) 6 N_A (*b*) 0.6 N_A (*c*) N_A (*d*) $\frac{N_A}{2}$

41. Which of the following has the highest mass?

(*a*) 20 g of sulphur

(*b*) 4 mol of carbon dioxide

(*c*) 12×10^{24} atoms of hydrogen

(*d*) 11.2 L of helium at N.T.P.

42. If isotopic distribution of C-12 ad 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*) 6.02×10^{23}

43. 5.6 L of a gas at S.T.P. weights equal to 8 g. The vapour density of gas is:

(*a*) 32 (*b*) 16 (*c*) 8 (*d*) 40

44. One atom of an element weighs 1.8×10^{-22} g, its atomic mass is:

(*a*) 29.9 g (*b*) 18 g

(*c*) 108.36 g (*d*) 154 g

45. If H_2SO_4 ionises as $H_2SO_4 + 2H_2O \rightarrow 2H_3O^+ + SO_4^{2-}$. Then total number of ions produced by 0.1 mol 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}

46. Which of the following will not have a mass of 10 g?

(*a*) 0.1 mol $CaCO_3$. (*b*) 1.51×10^{23} Ca^{2+} ions

(*c*) 0.16 mol of CO_3^{2-} ions (*d*) 7.525×10^{22} Br atom

47. xL of N_2 at S.T.P. contains 3×10^{22} molecules. The number of molecules in x/2 L of ozone at S.T.P. will be:

(*a*) 3×10^{22} (*b*) 1.5×10^{22}

(*c*) 1.5×10^{21} (*d*) 1.5×10^{11}

48. A person adds 1.71 gram of sugar ($C_{12}H_{22}O_{11}$) in order to sweeten his tea. The number of carbon atoms added are: (mol mass of sugar = 342)

(*a*) 3.6×10^{22} (*b*) 7.2×10^{21}

(*c*) 0.05 (*d*) 6.6×10^{22}

49. The number of atoms present in 0.1 mole of P_4 (At. mass = 31) are:

(*a*) 2.4×10^{24} atoms

(*b*) Same as in 0.05 mol of S_8

(*c*) 6.02×10^{22} atoms

(*d*) Same as in 3.1g of phosphorus

50. Which one contains maximum number of molecules?

(*a*) 2.5 g molecule of N_2

(*b*) 4 g atom of nitrogen

(*c*) 3.01×10^{24} atoms

(*d*) 82 g of dinitrogen

51. Out of 1.0 g dioxygen, 1.0 g (atomic) oxygen and 1.0 g ozone, the maximum number of oxygen atoms 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

52. The maximum volume at S.T.P. is occupied by:

(*a*) 12.8 g of SO_2 (*b*) 6.02×10^{22} molecules of CH_4

(*c*) 0.5 mL of NO_2 (*d*) 1g molecule of CO_2

53. If N_A is Avogadro's number, then the number of oxygen atoms in one g-equivalent of oxygen is:

(*a*) N_A (*b*) $N_A/2$

(*c*) $N_A/4$ (*d*) $2N_A$

54. If 224 ml. of a triatomic gas has a mass of 1g at 273 K and 1 atm pressure, then the mass of one atom is:

(*a*) 8.30×10^{-23} g (*b*) 6.24×10^{-23}

(*c*) 2.08×10^{-23} g (*d*) 5.54×10^{-23} g

55. The rest mass of an electron is 9.11×10^{-31} kg. Molar mass of the electron is:

(*a*) 1.5×10^{-31} kg mol^{-1} (*b*) 9.11×10^{-31} kg mol^{-1}

(*c*) 5.5×10^{-7} kg mol^{-1} (*d*) 6.02×10^{23} kg mol^{-1}

56. A sample of ammonium phosphate, $(NH_4)_3\ PO_4$, contains 3.18 moles of hydrogen atoms. The number of moles of oxygen atoms in the sample is:

(*a*) 0.265 (*b*) 0.795

(*c*) 1.06 (*d*) 3.18

57. What is the total number of atoms present in 25.0 mg of camphor, $C_{10}H_{16}O$?

(*a*) 9.89×10^{19} (*b*) 6.02×10^{20}

(*c*) 9.89×10^{20} (*d*) 2.67×10^{21}

58. 4.0 g of caustic soda (NaOH) (mol mass 40) contains same number of sodium ions as are present in-

(*a*) 10.6 g of Na_2CO_3 (mol. mass 106)

(*b*) 58.5 g of NaCl (Formula mass 58.5)

(*c*) 100 ml of 0.5 M Na_2SO_4 (Formula mass 142)

(*d*) 1mol of $NaNO_3$ (mol. mass 85)

59. Total number of atoms present in 64 gm of SO_2 is -

(*a*) $2 \times 6.02 \times 10^{23}$ (*b*) 6.02×10^{23}

(*c*) $4 \times 6.02 \times 10^{23}$ (*d*) $3 \times 6.02 \times 10^{23}$

60. The total number of protons, electrons and neutrons in 12gm of ${}_6C^{12}$ is -

(*a*) 1.084×10^{25} (*b*) 6.022×10^{23}

(*c*) 6.022×10^{22} (*d*) 18

61. Number of Ca^{+2} and Cl^- ion in 111 g of anhydrous $CaCl_2$ are -

(*a*) N_A, $2N_A$ (*b*) $2N_A$, N_A

(*c*) N_A, N_A (*d*) None

62. The maximum volume at N.T.P. is occupied by
(*a*) 12.8 gm of SO_2
(*b*) 6.02×10^{22} molecules of CH_4
(*c*) 0.5 mol of NO_2
(*d*) 1 gm-molecule of CO_2

63. Number of moles of water in 488 g of $BaCl_2.2H_2O$ are - (Ba = 137)
(*a*) 2 moles (*b*) 4 moles (*c*) 3 moles (*d*) 5 moles

64. 4.4 g of CO_2 and 2.24 litre of H_2 at STP are mixed in a container. The total number of molecules present in the container will be -
(*a*) 6.022×10^{23} (*b*) 1.2044×10^{23}
(*c*) 2 moles (*d*) 6.023×10^{24}

65. One mole of nitrogen gas has volume equal to
(*a*) 1 litre of nitrogen at S.T.P.
(*b*) 32 litre of nitrogen at S.T.P.
(*c*) 22.4 litre of nitrogen at S.T.P.
(*d*) 11.2 litre of nitrogen at S.T.P.

PERCENTAGE COMPOSITION

66. The percentage of C, H and N in an organic compound are 40 %, 13.3% and 46.7% respectively then empirical formula is
(*a*) $C_3H_{13}N_3$ (*b*) CH_2N
(*c*) CH_4N (*d*) CH_6N

67. B_1 g of an element gives B_2 g of its chloride, the equivalent mass of the element is:
(*a*) $\frac{B_1}{B_2 - B_1} \times 35.5$ (*b*) $\frac{B_2}{B_2 - B_1} \times 35.5$
(*c*) $\frac{B_2 - B_1}{B_1} \times 35.5$ (*d*) $\frac{B_2 - B_1}{B_2} \times 35.5$

68. 60 g of a compound on analysis gave 24 g C, 4 g H and 32 g O. The empirical formula of the compound is:
(*a*) $C_2H_4O_2$ (*b*) $C_2H_2O_2$
(*c*) CH_2O_2 (*d*) CH_2O

69. A compound made of two elements A and B are found to contain 25% A (at mass 12.5) and 75% B (at mass 37.5). The simplest formula of the compound is:
(*a*) AB (*b*) AB_2
(*c*) AB_3 (*d*) A_3B

70. 400 mg of capsule contains 100 mg of ferrous fumarate. The percentage of Fe present in the capsule is approximately: (formula of ferrous fumarate is $(CHCOO)_2$ Fe).
(*a*) 8.2% (*b*) 25%
(*c*) 16% (*d*) Unpredictable

71. Simplest formula of compound containing 50% of element X (at mass 10) and 50% of element Y (at mass 20) is:
(*a*) XY (*b*) X_2Y
(*c*) XY_2 (*d*) X_2Y_3

72. A compound having the empirical formula (C_3H_4O) has a molecular mass of 170 ± 5. The molecular formula of it compound is:
(*a*) C_3H_4O (*b*) $C_6H_8O_2$
(*c*) $C_6H_{12}O_3$ (*d*) $C_9H_{12}O_3$

73. Two oxides of a metal contains 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

74. The vapour density of gas A is four times that of B. If molecular mass of B is M, then molecular mass of A is:
(*a*) M (*b*) 4 M
(*c*) $\frac{M}{4}$ (*d*) 2 M

75. A metal nitride M_3N_2 contains 28% of nitrogen. The atomic mass of metal M is:
(*a*) 24 (*b*) 54
(*c*) 9 (*d*) 87.62

76. A container of volume V, contains 0.28 g of N_2 gas. If same volume of an unknown gas under similar conditions of temperature and pressure weights 0.44 g, the molecular mass of gas is:
(*a*) 22 (*b*) 44
(*c*) 66 (*d*) 88

77. A gaseous hydrocarbon on complete combustion gives 3.38 g of CO_2 and 0.690 g of H_2O and no other products. The empirical formula of hydrocarbon is:
(*a*) CH (*b*) CH_2
(*c*) CH_3 (*d*) The data is not complete

78. The percentage of Carbon in CO_2 is
(*a*) 27.27% (*b*) 29.27%
(*c*) 30.27% (*d*) 26.97%

79. The haemoglobin form red blood corpuscles of most mammals contain approximately 0.33% of iron by mass. The molecular mass of haemoglobin is 67200. The number of iron atoms in each molecule of haemoglobin is:
(*a*) 3 (*b*) 4
(*c*) 2 (*d*) 6

80. 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 compound?
(*a*) IO (*b*) I_2O
(*c*) I_5O_3 (*d*) I_2O_5

81. 0.5 mol of potassium ferrocyanide contains carbon equal to: (Formula of potassium ferrocyanide is $K_4[Fe(CN)_6]$.
(*a*) 1.5 mol (*b*) 36 g
(*c*) 18 g (*d*) 3.6 g

82. 14 g of element X combine with 16g of oxygen. On the basis of this information, which of the following is a correct statement :

(*a*) The element X could have an atomic weight of 7 and its oxide formula XO

(*b*) The element X could have an atomic weight of 14 and its oxide formula X_2O

(*c*) The element X could have an atomic weight of 7 and its oxide is X_2O

(*d*) The element X could have an atomic weight of 14 and its oxide is XO_2

83. A compound has 20% of nitrogen by weight. If one molecule of the compound contains two nitrogen atoms, the molecular weight of the compound is

(*a*) 35 (*b*) 70 (*c*) 140 (*d*) 280

STOICHIOMETRY & STOICHIOMETRIC CALCULATION

84. 'X' litres of carbon monoxide is present at STP. It is completely oxidized to CO_2. The volume of CO_2 formed is 11.207 litres at STP. What is the value of 'X' in litres?

(*a*) 22.414

(*b*) 11.207

(*c*) 5.6035

(*d*) 44.828

85. The moles of O_2 required for reacting with 6.8 g ammonia. (.....NH_3+....O_2→.....NO+.....H_2O) is:

(*a*) 5 (*b*) 2.5

(*c*) 1 (*d*) 0.5

86. What mass of $CaCl_2$ in grams would be enough to produce 14.35 gm of AgCl?

(*a*) 5.55 g (*b*) 8.29 g

(*c*) 16.59 g (*d*) 10 g

87. What weight of sodium hydroxide is required to neutralise 100 ml of 0.1N HCl ?

(*a*) 4 g (*b*) 0.4 g

(*c*) 0.04 g (*d*) 40 g

88. H_2O_2 is sold as a solution of approximately 5.0 g H_2O_2 per 100 ml of the solution. The molarity of this solution is approximately

(*a*) 0.15 M

(*b*) 1.5 M

(*c*) 3.0 M

(*d*) 3.4 M

89. The amount of oxalic acid (eq.wt.63) required to prepare 500 ml of its 0.10 N solution is

(*a*) 0.315 g

(*b*) 3.150 g

(*c*) 6.300 g

(*d*) 63.00 g

90. The molarity of pure water is

(*a*) 100 M (*b*) 55.6M

(*c*) 50 M (*d*) 18 M

91. The mass of 70% H_2SO_4 by mass is required for neutralisation of 1 mole of NaOH is:

(*a*) 65 (*b*) 98 (*c*) 70 (*d*) 54

92. If potassium chlorate is 80% pure then 48 g of oxygen would be produced from:

(*a*) 153.12 g of $KClO_3$

(*b*) 120 g of $KClO_3$

(*c*) 20 g of $KClO_3$

(*d*) 90 g of $KClO_3$

93. Density of a solution containing x% by mass of H_2SO_4 is y. The normality is

(*a*) $\frac{xy \times 10}{98}$ (*b*) $\frac{xy \times 10}{98y} \times 2$

(*c*) $\frac{xy \times 10}{98} \times 2$ (*d*) $\frac{x \times 10}{98y}$

94. Mass percentage (w/w) of ethylene glycol ($HOCH_2$ - CH_2OH) in a aqueous solution is 20, then mole fraction of solute is

(*a*) 0.5 (*b*) 0.067

(*c*) 0.1 (*d*) 0.4

95. Number of gram equivalents of solute in 100 ml of 5 N HCl solution is

(*a*) 50 (*b*) 500 (*c*) 5 (*d*) 0.5

96. If 1.26 grams of oxalic acid is dissolved in 250 ml of solution then its normality is

(*a*) 0.05 (*b*) 0.04 (*c*) 0.02 (*d*) 0.08

97. 100ml of ethyl alcohol is made upto a litre with distilled water. If the density of C_2H_5OH is 0.46 gm/ml. Then its molality is

(*a*) 0.55 m (*b*) 1.11m

(*c*) 2.22 m (*d*) 3.33m

98. A solution of 0.1 mole of a metal chloride MCl_x required 500 mL of 0.6 molal $AgNO_3$ solution for complete ppt. The value of x is:

(*a*) 5 (*b*) 4

(*c*) 3 (*d*) 1

99. If 20 g of $CaCO_3$ is treated with 100 ml 20% HCl solution. The amount of CO_2 produced is:

(*a*) 22.4l g (*b*) 8.8 g

(*c*) 2.2 g (*d*) 8 l

100. The mass of $CaCO_3$ required to react with 25 mL of 0.75 molar HCl is:

(*a*) 0.94 g (*b*) 0.68 g

(*c*) 0.76 g (*d*) 0.52 g

101. 2 moles of H_2S and 11.2 L of SO_2 at N.T.P. reacts to form x moles of sulphur. The value of x is:

(*a*) 1.5 (*b*) 3.5

(*c*) 7.8 (*d*) 12.7

102. Sulphuryl chloride (SO_2Cl_2) reacts with H_2O to give a mixture of H_2SO_4 & HCl. Aqueous solution of 1 mole SO_2Cl_2 will be neutralised by:

(*a*) 3 moles of NaOH

(*b*) 2 moles of $Ca(OH)_2$

(*c*) Both (*a*) & (*b*)

(*d*) None of these

103. A sample of pure compound contains 1.15 g of sodium, 3.01×10^{22} atoms of carbon and 0.1 mol of oxygen atom. Its empirical formula is:

(*a*) Na_2CO_3 (*b*) $NaCO_2$ (*c*) Na_2CO (*d*) Na_2CO_2

104. If 0.30 mol of zinc are added to 0.52 mol of HCl, the moles of H_2 formed are:

(*a*) 0.52 (*b*) 0.30 (*c*) 0.26 (*d*) 0.60

105. The specific gravity of 98% H_2SO_4 is 1.8 g/cc. 50 ml of this solution is mixed with 1750 ml of pure water. Molarity of resulting solution is

(*a*) 0.2 M (*b*) 0.5 M (*c*) 0.1 M (*d*) 1 M

Learning Plus

1. Which of the following is/are not affected by temperature?

(*a*) Molarity

(*b*) Molality

(*c*) Normality

(*d*) None of these

2. Ferric sulphate on heating gives sulphur trioxide. The ratio between the weights of oxygen and sulphur present in SO_3 obtained by heating 1 kg of ferric sulphate is

(*a*) 2 : 3 (*b*) 1 : 3

(*c*) 3 : 1 (*d*) 3 : 2

3. The number of atoms present in 4.25 grams of NH_3 is approximately

(*a*) 1×10^{23}

(*b*) 8×10^{20}

(*c*) 2×10^{23}

(*d*) 6.02×10^{23}

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

Students	Readings	
	(i)	(ii)
A	3.01	2.99
B	3.05	2.95

(*a*) Results of both the students are neither accurate nor precise

(*b*) Results of student A are both precise and accurate

(*c*) Results of student B are neither precise nor accurate

(*d*) Results of student B are both precise and accurate

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

(*a*) 4 mol L^{-1}

(*b*) 20 mol L^{-1}

(*c*) 0.2 mol L^{-1}

(*d*) 2 mol L^{-1}

6. Number of atoms in 55.85 gram Fe (at. wt. of Fe = 55.85 g mol^{-1}) is

(*a*) Twice that 60 g carbon

(*b*) 6.023×10^{22}

(*c*) Half that in 8g He

(*d*) $5558.5 \times 6.023 \times 10^{23}$

7. Neon has two isotopes Ne^{20} and Ne^{22}. If atomic weight of Neon is 20.2, the ratio of the relative abundances of the isotopes is

(*a*) 1 : 9 (*b*) 9 : 1

(*c*) 70 % (*d*) 80 %

8. The total weight of 10^{22} molecular units of $CuSO_4 . 5H_2O$ is nearly

(*a*) 4.144 g (*b*) 5.5 g

(*c*) 24.95 g (*d*) 41.45 g

9. The number of Cl^- and Ca^{+2} ions in 222g. of $CaCl_2$ are

(*a*) 4N, 2N (*b*) 2N, 4N

(*c*) 1N, 2N (*d*) 2N, 1N

10. The empirical formula of a gaseous compound is 'CH_2'. The density of the compound is 1.25 gm/lit. at S.T.P. The molecular formula of the compound is 'X'

(*a*) C_2H_4 (*b*) C_3H_6

(*c*) C_6H_{12} (*d*) C_4H_8

11. If 500 mL of a 5 M solution is diluted to 1500 mL, what will be the molarity of the solution obtained?

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

12. 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?

(*a*) 4 g He (*b*) 46 g Na
(*c*) 0.40 g Ca (*d*) 12 g He

13. The empirical formula of an organic compound is CH_2O. Its vapour density is 45. The molecular formula of the compound is

(*a*) CH_2O (*b*) $C_2H_4O_2$
(*c*) $C_3H_6O_3$ (*d*) $C_6H_{12}O_6$

14. 0.132 g of an organic compound gave 50 ml of N_2 at NTP. The weight percentage of nitrogen in the compound is close to

(*a*) 15 (*b*) 20
(*c*) 48.9 (*d*) 47.34

15. 0.7 moles of potassium sulphate is allowed to react with 0.9 moles of barium chloride in aqueous solutions. The number of moles of the substance precipitated in the reaction is

(*a*) 1.4 moles of potassium chloride
(*b*) 0.7 moles of barium sulphate
(*c*) 1.6 moles of potassium chloride
(*d*) 1.6 moles of barium sulphate

16. The number of moles of Fe_2O_3 formed when 0.5 moles of O_2 and 0.5 moles of Fe are allowed to react are

(*a*) 0.25 (*b*) 0.5 (*c*) 1/3 (*d*) 0.125

17. Amount of oxalic acid required to prepare 250ml of N/10 solution (MW of oxalic acid = 126) is

(*a*) 1.5759 g (*b*) 3.15 g (*c*) 15.75 g (*d*) 63.0 g

18. If the concentration of glucose ($C_6H_{12}O_6$) in blood is 0.9 g L^{-1}, what will be the molarity of glucose in blood?

(*a*) 5 M (*b*) 50 M
(*c*) 0.005 M (*d*) 0.5 M

19. What will be the molality of the solution containing 18.25 g of HCl gas in 500 g of water?

(*a*) 0.1 m (*b*) 10 m (*c*) 0.5 m (*d*) 1 m

20. Increasing order of number of moles of the species

(*i*) 3 grams of NO
(*ii*) 8.5 grams of PH_3 and
(*iii*) 8 grams of methane is

(*a*) (*i*) < (*ii*) < (*iii*) (*b*) (*iii*) < (*ii*) < (*i*)
(*c*) (*i*) < (*iii*) < (*ii*) (*d*) (*ii*) < (*iii*) < (*i*)

21. The number of molecules present in 1.12×10^{-7} cc of a gas at STP is

(*a*) 6.02×10^{23} (*b*) 3.01×10^{12}
(*c*) 6.02×10^{12} (*d*) 3.01×10^{23}

22. From 320 mg. of O_2, 6.023×10^{20} molecules are removed, the no. of moles remained are

(*a*) 9×10^{-3} moles (*b*) 9×10^{-2} moles
(*c*) Zero (*d*) 3×10^{-3} moles

23. An oxide of nitrogen has a molecular weight 92. Find the total number of electrons in one gram mole of that oxide.

(*a*) 4.6 N (*b*) 46 N
(*c*) 23 N (*d*) 2.3 N

24. No. of moles of water in 488.6 gms of $BaCl_2.2H_2O$ are (molecular weight of $BaCl_2.2H_2O$=244.33)

(*a*) 2 moles (*b*) 4 moles (*c*) 3 moles (*d*) 5 moles

25. One mole of any substance contains 6.022×10^{23} atoms/molecules. Number of molecules of H_2SO_4 present in 100 mL of 0.02 M H_2SO_4 solution is:

(*a*) 12.044×10^{20} molecules (*b*) 6.022×10^{23} molecules
(*c*) 1×10^{23} molecules (*d*) 12.044×10^{23} molecules

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

(*a*) 3, 4 and 5, respectively (*b*) 3, 4 and 4, respectively
(*c*) 3, 3 and 4, respectively (*d*) 3, 3 and 3, respectively

27. A certain compound contains magnesium, carbon and Nitrogen in the mass ratio 12 : 12 : 14. The formula of the compound is

(*a*) MgCN (*b*) Mg_2CN
(*c*) $MgCN_2$ (*d*) $Mg(CN)_2$

28. An oxide of nitrogen contains 36.8% by weight of nitrogen. The formula of the compound is

(*a*) N_2O (*b*) N_2O_3 (*c*) NO (*d*) NO_2

29. 40 ml. of a hydrocarbon undergoes combustion in 260 ml of oxygen and gives 160 ml of carbon dioxide. If all gases are measured under similar conditions of temperature and pressure, the formula of hydrocarbon is

(*a*) C_3H_8 (*b*) C_4H_8 (*c*) C_6H_{14} (*d*) C_4H_{10}

30. The mass of Hydrogen at S.T.P. that is present in a vessel which can hold 4 grams of oxygen under similar conditions is

(*a*) 1 gram (*b*) 0.5 grams
(*c*) 0.25 gms. (*d*) 0.125 gm

31. Which of the following solutions has the highest normality?

(*a*) 172 milli equivalents in 200 ml
(*b*) 84 milli equivalents in 100 ml
(*c*) 275 milli equivalents in 250 ml
(*d*) 43 milli equivalents in 60 ml

32. What volume of 75 % H_2SO_4 by mass is required to prepare 1.5 litres of 0.2 M H_2SO_4? (Density of the sample is 1.8 g/cc)

(*a*) 14.2cc (*b*) 28.4cc
(*c*) 21.7cc (*d*) 7.1 cc

33. The empirical formula and molecular mass of a compound are CH_2O and 180 g respectively. What will be the molecular formula of the compound?

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

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

(*a*) 4.7 g (*b*) 4680×10^{-3} g

(*c*) 4.680 g (*d*) 46.80 g

35. 4.9 grams of H_2SO_4 is present is 100 ml of the solution, then its molarity and normality are

(*a*) 1, 0.5 (*b*) 1, 1

(*c*) 0.5, 1 (*d*) 0.5, 2

36. In order to prepare one litre normal solution of $KMnO_4$, how many grams of $KMnO_4$ required if the solution is to be used in acidic medium for oxidation

(*a*) 158 (*b*) 79 (*c*) 31.6 (*d*) 790

37. 50 gm of sample of sodium hydroxide required for complete neutralisation, 1L 1N HCl. What is the percentage purity of NaOH is

(*a*) 50 (*b*) 60 (*c*) 70 (*d*) 80

38. Which of the following statements is correct about the reaction given below?

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

(*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

39. Which of the following statements indicates that law of multiple proportion is being followed?

(*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

Multiconcept MCQs

1. In HABER's process, 30 L of H_2 and 30 L of N_2 were taken for a reaction which yielded only 50% of the expected product. What will be the composition of the gaseous mixture under these conditions?

(*a*) 10 L NH_3, 25 L N_2, 15 L H_2

(*b*) 20 L NH_3, 10 L N_2, 3 L H_2

(*c*) 20 L NH_3, 25 L N_2, 15 L H_2

(*d*) None of these

2. Diborane (B_2H_6) can be prepared by the following reaction-

$$3NaBH_4 + 4BF_3 \longrightarrow 3NaBF_4 + 2B_2H_6$$

If the reaction has a 70% yield, how many moles of $NaBH_4$ should be used with excess BF_3 in order to obtain 0.200 mol of B_2H_6?

(*a*) 0.21 moles (*b*) 0.429 moles

(*c*) 0.300 mol (*d*) 0.175 moles

3. An ore contains 1.24% of mineral argentate, Ag_2S by mass. How many grams of this ore would have to be processed in order to obtain 1g of pure solid silver?

(*a*) 92.6 g

(*b*) 88.1 g

(*c*) 101.11 g

(*d*) 107.25 g

4. In an organic compound of molar mass 108 g mol^{-1} C, H and N atoms are present in 9 : 1 : 3.5 by weight. Molecular formula can be –

(*a*) $C_6H_8N_2$ (*b*) $C_7H_{10}N$

(*c*) $C_5H_6N_3$ (*d*) $C_4H_{18}N_3$

5. 6.02×10^{20} molecules of urea are present in 100 mL of its solution. The concentration of urea solution is -

(*a*) 0.001 M (*b*) 0.01 M

(*c*) 0.02 M (*d*) 0.1 M

6. What is the empirical formula of a compound composed of O & Mn in equal weight ratio? (At. wt of Mn = 55)

(*a*) MnO (*b*) MnO_2

(*c*) Mn_2O_3 (*d*) Mn_2O_7

7. 1 g sample of alkaline earth metal react completely with 4.08 g H_2SO_4 and yields an ionic product MSO_4. Then find out the atomic mass of alkaline earth metal (M)?

(*a*) 9 (*b*) 24

(*c*) 40 (*d*) 87

8. A compound contains 36% C by mass. If each molecule contains two C atoms the number of moles of the compound in its 10 g is/are-

(*a*) 0.15 (*b*) 1.5

(*c*) 150 (*d*) 1500

9. When a certain amount of octane is burnt completely, 7.04 g of CO_2 is formed. What mass of H_2O is formed simultaneously?

(*a*) 3.24 g (*b*) 6.68 g

(*c*) 6.48 g (*d*) 6.16 g

10. A 1.50 g sample of $KHCO_3$ having 80% purity is strongly heated. Assuming the impurity to be thermally stable, the loss in weight of the sample, on heating is:

(*a*) 26.4 g (*b*) 2.64 g

(*c*) 0.264 g (*d*) 0.0264 g

11. Chlorine can be prepared by reacting HCl with MnO_2. The reaction is represented by the equation

$$MnO_2(g) + 4HCl(aq) \rightarrow MnCl_2(aq) + Cl_2(g) + 2H_2O(\ell)$$

Assuming that the reaction goes to completion. What mass of conc. HCl solution (36% by mass) is needed to produce 2.5g Cl_2?

(*a*) 14.27 g (*b*) 25 g

(*c*) 25.21 g (*d*) 30 g

12. 29.2% (w/w) HCl stock solution has a density of 1.25 g mL^{-1}. The molecular weight of HCl is 36.5 g mol^{-1}. The volume (in mL) of stock solution required to prepare a 200 mL solution of 0.4 M HCl is-

(*a*) 16 mL (*b*) 61 mL

(*c*) 80 mL (*d*) 8 mL

13. The composition of residual mixture will be, if 30 g of Mg combines with 30 g of O_2 -

(*a*) 40 g MgO + 20 g O_2

(*b*) 45 g MgO + 15 g O_2

(*c*) 50 g MgO + 10 g O_2

(*d*) 60 g MgO only

14. If the yield of given reaction is 33.33% what volume of O_2 gas will be produced if 4 moles of $K_2Cr_2O_7$ are taken initially;

$$K_2Cr_2O_7 \xrightarrow{\Delta} K_2CrO_4 + Cr_2O_3 + O_2 \quad (\text{unbalanced})$$

(*a*) 11.2 L (*b*) 22.4 L

(*c*) 33.6 L (*d*) 67.2 L

15. 0.8 mole of a mixture of CO and CO_2 requires exactly 40 gram of NaOH in solution for complete conversion of all the CO_2 into Na_2CO_3, if the mixture (0.8 mole) is completely oxidised to CO_2, find further required moles of NaOH.

(*a*) 0.2 (*b*) 0.6

(*c*) 1 (*d*) 1.5

16. 3.68 g of mixture of $CaCO_3$ and $MgCO_3$ is heated to liberate 0.04 mole of CO_2. The mole % of $CaCO_3$ and $MgCO_3$ in the mixture is respectively:

(*a*) 50%, 50% (*b*) 60%, 40%

(*c*) 40%, 60% (*d*) 30%, 70%

17. An element is found in nature in two isotopic forms with mass numbers (A–1) and (A + 3). If the average atomic mass of the element is found to be A, then the relative abundance of the heavier isotope in the nature will be.

(*a*) 66.6% (*b*) 75%

(*c*) 25% (*d*) 33.3%

18. Which of the following molarity values of ions in a aqueous solution of 5.85% w/v NaCl, 5.55% w/v $CaCl_2$ and 6% w/v NaOH are correct [Na = 23, Cl = 35.5, Ca = 40, O = 16]

(*a*) $[Cl^-] = 2M$ (*b*) $[OH^-] = 1.5$ M

(*c*) $[Ca^{2+}] = 0.5$ M (*d*) All of these

19. How many moles of ferric alum, $(NH_4)_2SO_4Fe_2(SO_4)_3.24H_2O$ can be made from the sample of Fe containing 0.0056 g of it?

(*a*) 10^{-4} mol (*b*) 0.5×10^{-4} mol

(*c*) 0.33×10^{-4} mol (*d*) 2×10^{-4} mol

20. The following substances are present in different containers

(*i*) one gram atom of nitrogen

(*ii*) one mole of calcium

(*iii*) One atom of silver

(*iv*) One mole of oxygen molecules

(*v*) 10^{23} atoms of carbon

(*vi*) One gram of iron.

The correct order of increasing masses (in grams) is/are

(*a*) (*iii*) < (*iv*) < (*i*) < (*v*) (*b*) (*iii*) < (*vi*) < (*iv*) < (*ii*)

(*c*) (*vi*) < (*v*) < (*i*) < (*iv*) (*d*) (*b*) and (*c*) both are correct

21. How many molecules of HCl gas will be produced by reacting 112 L of H_2 (0 °C, 1 atm) with 213 g of Cl_2?

(*a*) 3.61×10^{24} (*b*) 6.13×10^{23}

(*c*) 6.13×10^{24} (*d*) 1.63×10^{24}

22. 85 g $CaCO_3$ (limestone sample), on heating produces exactly the same amount of CO_2 which converts 30 g of MgO to $MgCO_3$. The percentage purity of limestone sample is

(*a*) 80% (*b*) 82.4%

(*c*) 88.24% (*d*) 84.8%

NEET Past 10 Years Questions

1. An organic compound contains 78% (by wt.) carbon and remaining percentage of hydrogen. The right option for the empirical formula of this compound is: [Atomic wt. of C is 12, H is 1] ***(2021)***

(*a*) CH_2 (*b*) CH_3 (*c*) CH_4 (*d*) CH

2. Which one of the followings has maximum number of atoms? ***(2020)***

(*a*) 1 g of Mg(s) [Atomic mass of Mg = 24]

(*b*) 1 g of O_2(g) [Atomic mass of O = 16]

(*c*) 1 g of Li(s) [Atomic mass of Li = 7]

(*d*) 1 g of Ag(s) [Atomic mass of Ag = 108]

3. One mole of carbon atom weighs 12g, the number of atoms in it is equal to. ***(2020 Covid Re-NEET)***

(Mass of carbon- 12 is 1.9926×10^{-23} g)

(*a*) 6.022×10^{22} (*b*) 12×10^{22}

(*c*) 6.022×10^{23} (*d*) 12×10^{23}

4. The number of moles of hydrogen molecules required to produce 20 moles of ammonia through Haber's process is : ***(2019)***

(*a*) 10 (*b*) 20 (*c*) 30 (*d*) 40

5. A mixture of 2.3 g formic acid and 4.5 g oxalic acid is treated with conc. H_2SO_4. The evolved gaseous mixture is passed through KOH pellets. Weight (in g) of the remaining product at STP will be: ***(2018)***

(*a*) 1.4 (*b*) 3.0 (*c*) 4.4 (*d*) 2.8

6. In which case is number of molecules of water maximum? ***(2018)***

(*a*) 18 mL of water (*b*) 0.18 g of water

(*c*) 10^{-3} mol of water

(*d*) 0.00224 L of water vapours at 1 atm and 273 K

7. A hydrocarbon contains 85.7% of Carbon and 14.3% of Hydrogen. If 42 mg of the compound contains 3.01×10^{20} molecules, the molecular formula of the compound will be: ***(2017-Gujarat)***

(*a*) C_2H_4 (*b*) C_3H_6 (*c*) C_6H_{12} (*d*) $C_{12}H_{24}$

8. Suppose the elements X and Y combine to form two compounds XY_2 and X_3Y_2. When 0.1 mole of XY_2 weighs 10 g and 0.05 mole of X_3Y_2 weighs 9 g, the atomic weights of X and Y are: ***(2016–II)***

(*a*) 20, 30 (*b*) 30, 20 (*c*) 40, 30 (*d*) 60, 40

9. The number of water molecules is maximum in: ***(2015 Re)***

(*a*) 18 moles of water (*b*) 18 molecules of water

(*c*) 1.8 gram of water (*d*) 18 gram of water

10. If Avogadro number N_A, is changed from 6.022×10^{23} mol^{-1} to 6.022×10^{20} mol^{-1}, this would change: ***(2015 Re)***

(*a*) The ratio of elements to each other in a compound

(*b*) The definition of mass in units of grams

(*c*) The mass of one mole of carbon

(*d*) The ratio of chemical species to each other in a balanced equation

11. What is the mass of the precipitate formed when 50 mL of 16.9% solution of $AgNO_3$ is mixed with 50 mL of 5.8% NaCl solution?

(Ag = 107.8, N = 14, O = 16, Na = 23, Cl = 35.5) ***(2015 Re)***

(*a*) 3.5 g (*b*) 7 g (*c*) 14 g (*d*) 28 g

12. 20.0 g of a magnesium carbonate sample decomposes on heating to give carbon dioxide and 8.0 g magnesium oxide. What will be the percentage purity of magnesium carbonate in the sample?

(Atomic weight of Mg = 24) ***(2015 Re)***

(*a*) 96 (*b*) 60 (*c*) 84 (*d*) 75

13. When 22.4 litres of H_2(g) is mixed with 11.2 litres of Cl_2(g), each at STP, the moles of HCl(g) formed is equal to: ***(2014)***

(*a*) 2 mol of HCl(g) (*b*) 0.5 mol of HCl(g)

(*c*) 1.5 mol of HCl(g) (*d*) 1 mol of HCl(g)

14. 1.0 g of magnesium is burnt with 0.56 g O_2 in a closed vessel. Which reactant is left in excess and how much? ***(2014)***

(Atomic weight Mg = 24; O = 16)

(*a*) O_2, 0.16 g (*b*) Mg, 0.44 g

(*c*) O_2, 0.28 g (*d*) Mg, 0.16 g

15. Equal masses of H_2, O_2 and methane have been taken in a container of volume *V* at temperature 27°C in identical conditions. The ratio of the volumes of gases $H_2 : O_2$: methane would be: ***(2014)***

(*a*) 8 : 16 : 1 (*b*) 16 : 8 : 1

(*c*) 16 : 1 : 2 (*d*) 8 : 1 : 2

16. 6.02×10^{20} molecules of urea are present in 100 mL of its solution. The concentration of solution is: ***(2013)***

(*a*) 0.02 M (*b*) 0.01 M

(*c*) 0.001 M (*d*) 0.1 M

17. An excess of $AgNO_3$ is added to 100 mL of a 0.01 M solution of dichlorotetraaquachromium(III) chloride. The number of moles of AgCl precipitated would be: ***(2013)***

(*a*) 0.001 (*b*) 0.002

(*c*) 0.003 (*d*) 0.01

ANSWER KEY

Topicwise Questions

1. (*a*)	**2.** (*d*)	**3.** (*c*)	**4.** (*c*)	**5.** (*d*)	**6.** (*d*)	**7.** (*a*)	**8.** (*b*)	**9.** (*c*)	**10.** (*a*)
11. (*b*)	**12.** (*c*)	**13.** (*b*)	**14.** (*c*)	**15.** (*b*)	**16.** (*a*)	**17.** (*c*)	**18.** (*c*)	**19.** (*d*)	**20.** (*b*)
21. (*d*)	**22.** (*b*)	**23.** (*c*)	**24.** (*b*)	**25.** (*d*)	**26.** (*c*)	**27.** (*c*)	**28.** (*a*)	**29.** (*d*)	**30.** (*d*)
31. (*c*)	**32.** (*b*)	**33.** (*d*)	**34.** (*c*)	**35.** (*b*)	**36.** (*b*)	**37.** (*c*)	**38.** (*c*)	**39.** (*b*)	**40.** (*b*)
41. (*b*)	**42.** (*a*)	**43.** (*b*)	**44.** (*c*)	**45.** (*d*)	**46.** (*c*)	**47.** (*b*)	**48.** (*a*)	**49.** (*b*)	**50.** (*d*)
51. (*d*)	**52.** (*d*)	**53.** (*c*)	**54.** (*d*)	**55.** (*c*)	**56.** (*c*)	**57.** (*d*)	**58.** (*c*)	**59.** (*d*)	**60.** (*a*)
61. (*a*)	**62.** (*d*)	**63.** (*b*)	**64.** (*b*)	**65.** (*c*)	**66.** (*c*)	**67.** (*a*)	**68.** (*d*)	**69.** (*a*)	**70.** (*a*)
71. (*b*)	**72.** (*d*)	**73.** (*b*)	**74.** (*b*)	**75.** (*a*)	**76.** (*b*)	**77.** (*a*)	**78.** (*a*)	**79.** (*b*)	**80.** (*d*)
81. (*b*)	**82.** (*c*)	**83.** (*c*)	**84.** (*b*)	**85.** (*d*)	**86.** (*a*)	**87.** (*b*)	**88.** (*b*)	**89.** (*b*)	**90.** (*b*)
91. (*c*)	**92.** (*a*)	**93.** (*c*)	**94.** (*b*)	**95.** (*d*)	**96.** (*d*)	**97.** (*b*)	**98.** (*c*)	**99.** (*b*)	**100.** (*a*)
101. (*a*)	**102.** (*b*)	**103.** (*b*)	**104.** (*c*)	**105.** (*b*)					

Learning Plus

1. (*b*)	**2.** (*d*)	**3.** (*d*)	**4.** (*b*)	**5.** (*c*)	**6.** (*a*)	**7.** (*b*)	**8.** (*a*)	**9.** (*a*)	**10.** (*a*)
11. (*b*)	**12.** (*d*)	**13.** (*c*)	**14.** (*d*)	**15.** (*b*)	**16.** (*a*)	**17.** (*a*)	**18.** (*c*)	**19.** (*d*)	**20.** (*a*)
21. (*b*)	**22.** (*a*)	**23.** (*b*)	**24.** (*b*)	**25.** (*a*)	**26.** (*d*)	**27.** (*d*)	**28.** (*b*)	**29.** (*d*)	**30.** (*c*)
31. (*c*)	**32.** (*c*)	**33.** (*c*)	**34.** (*a*)	**35.** (*c*)	**36.** (*c*)	**37.** (*d*)	**38.** (*a*)	**39.** (*b*)	

Multiconcept MCQs

1. (*a*)	**2.** (*b*)	**3.** (*a*)	**4.** (*a*)	**5.** (*b*)	**6.** (*d*)	**7.** (*b*)	**8.** (*a*)	**9.** (*a*)	**10.** (*c*)
11. (*a*)	**12.** (*d*)	**13.** (*c*)	**14.** (*b*)	**15.** (*b*)	**16.** (*a*)	**17.** (*c*)	**18.** (*d*)	**19.** (*b*)	**20.** (*d*)
21. (*a*)	**22.** (*c*)								

NEET Past 10 Years Questions

1. (*b*)	**2.** (*c*)	**3.** (*c*)	**4.** (*c*)	**5.** (*d*)	**6.** (*a*)	**7.** (*c*)	**8.** (*c*)	**9.** (*a*)	**10.** (*c*)
11. (*b*)	**12.** (*c*)	**13.** (*d*)	**14.** (*d*)	**15.** (*c*)	**16.** (*b*)	**17.** (*a*)			

Solution

Topicwise Questions

1. (*a*) The number of significant figures in 0.0045 are two because zeros to the left of the first non-zero digit are not significant.

2. (*d*) $d = s \times t$

1 femto = 10^{-15} sec = $(3 \times 10^{8}$ m/s$)(10^{-15}$ s$)$
$= 3 \times 10^{-7}$ m
$= 3 \times 10^{-4}$ mm = 0.0003 mm

3. (*c*) 1 Barn = 10^{-28} m^2

4. (*c*) 7.00 is more accurate than 7.0 because the former has three significant figures while the later has two.

5. (*d*) π is irrational number (please note that $\pi \neq 22/7$ as 22/7 is a rational number) and it means it has infinite number of significant figures.

6. (*d*) According to rule (2) zero between '5' and '4' is also significant

7. (*a*) 29.4406 + 3.2 + 2.25 = 34.8906. As 3.2 has least number of decimal places, i.e., one, therefore sum should contain one decimal place only. After round off, reported sum is 34.9, which has three significant figures.

8. (*b*) Area = length × width
$= (12.34\text{cm}) \times (1.23\text{cm}) = 15.1782\ \text{cm}^2$
$\simeq 15.18\ \text{cm}^2$

9. (*c*) 0.2876 g × 9 = 2.5884 g
= 2.588 g

10. (*a*) According to rule (1) all non zero digits are significant

11. (*b*) Haber's process:

$$\underset{1L}{N_2} + \underset{3L}{3H_2} \rightarrow \underset{2L}{2NH_3}$$

Total volume of reactant at STP = 4L

Volume of NH_3 = 2L

Volume of NH_3 = 1/2 volume of Reactant

Hence (*b*) one half.

12. (*c*) 6g of carbon combines with 32g sulphur to give CS_2.

12g of carbon combine with 32g of O to give CO_2.

10g of S combine with 10g of O to give SO_2.

In CS_2, C : S = 6 : 32 = 3 : 16

In CO_2, C : O = 12 : 32 = 3 : 8

From CS_2 and CO_2, S : O = 16 : 8 = 2 : 1

Also in SO_2, S : O = 10 : 10 = 1 : 1

2 : 1 is multiple of 1 : 1

So law of reciprocal proportion is satisfied.

13. (*b*) Hint: Total mass of reactants = Mass of products.

14. (*c*) Law of reciprocal proportion

15. (*b*) Law of conservation of mass not applicable to nuclear reactions

16. (*a*) According to law of constant proportion

17. (*c*) 1g of X combines with Y = 35.5 g
$\therefore$ 2 g of X combines with Y = 2 × 35.5 g = 71 g

18. (*c*) **In First Oxide N_2O:** Mass of 2.24 L of nitrogen at STP = 2.8g
$\therefore$ Mass of Oxygen = 4.4 – 2.8 = 1.6 g

In Second Oxide N_2O_2: Mass of 22.4 L of nitrogen at STP = 28 g
$\therefore$ Mass of oxygen = 60 – 28 = 32 g
$\therefore$ In second oxide 2.8 g of nitrogen combines with 3.2 g of oxygen.

Keeping the mass of nitrogen same in both the oxides, the different masses of oxygen which combines with 2.8 g of nitrogen are 1.6 g : 3.2 g or 1 : 2, This is a simple whole number ratio. This illustrates the law of multiple proportions.

19. (*d*) SO_2 , SO_3

20. (*b*) $H_3PO_4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$
As 1 mole H_3PO_4 neutralize 3 mole of NaOH. 120 g of NaOH is 3 moles. So, 98 g of H_3PO_4 will neutralize 120 g of NaOH.

21. (*d*) $4Al + 3O_2 \rightarrow 2Al_2O_3$

22. (*b*) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + 2HCl$

20.8 g   9.8 g   x   7.3 g

mass of $BaSO_4$ produced

x = 20.8 + 9.8 – 7.3 = 23.3 g

23. (*c*) In CO_2, 12 parts by mass of C combine with 32 parts by mass of oxygen while in SO_2, 32 parts by mass of S combine with 32 parts by mass of oxygen. The ratio of masses of carbon and sulphur which combine with a fixed mass of oxygen is 12 : 32 or 3 : 8. In CS_2, 12 parts of carbon combines with 64 parts by mass of sulphur therefore 12 : 64, i.e., 3 : 16.

$\therefore$ The ratios are $\frac{3}{8}:\frac{3}{16}$ or $2:1$

24. (*b*) $C_2H_5OH + Na \rightarrow C_2H_5ONa + \frac{1}{2}H_2$

25. (*d*)

H_2O	H_2O_2
H : O	H : O
5.93 : 94.07	11.2 : 88.8 or
	5.93 : 47.0

Ratio of different masses of O which combines with fixed mass of H is 94.07 : 47.0 or 2 : 1

26. (*c*) In B, 32 parts of X combines with Y = 84 parts

$\therefore$ 16 parts of X combine with Y = 42 parts

Now, number of parts of X in both B and C is equal.

Different masses of Y which combine with same mass of X in B and C are in the ratio 3:5

$$\therefore \frac{\text{Mass of Y in B}}{\text{Mass of Y in C}} = \frac{3}{5}$$

$$\therefore \text{Mass of Y in C} = \frac{5}{3} \times 42$$

= 70 parts

27. (*c*) 3.4 amu S $\rightarrow$ 100 amu insulin

$$32\,\text{amu S} \rightarrow \frac{100}{3.4} \times 32 = 941 \text{ amu}$$

28. (*a*) Mass of one mole of a substance remains same

29. (*d*) Average relative at Atm.wt

$$A.W = \frac{20 \times 10 + 80 \times 11}{100}$$

$$= \frac{200 + 880}{100} = \frac{1080}{100} = 10.8g$$

30. (*d*) No. of molecules = $\frac{W}{M.w} \times N$

31. (*c*) Gram molecular weight of any substance contains Avagadro number of molecules.

32. (*b*) No. of atoms = $\frac{Wt}{MW} \times N \times atomicity$

33. (*d*) One amu = 1.66×10^{-24}g

34. (*c*) The number of molecules present in 1 ml of a gas at STP is known as loschmidt number.

35. (*b*) Verify options by calculating no. of moles

36. (*b*) Apply,

$$\text{weight} = \frac{\text{absolute at. wt.}}{1\,\text{amu}}$$

$$14 = \frac{\text{absolute at. wt.}}{\frac{1}{12} \times \frac{12}{N_A} g} \quad \ldots(1)$$

$$x = \frac{\text{absolute at. wt.}}{\frac{1}{5} \times \frac{12}{N_A} g} \quad \ldots(2)$$

x = 5.83 amu

37. (*c*) 100 g Haemoglobin contains Fe = 0.34 g

1 mole haemoglobin contains Fe = 4 mole Fe^{2+}

= 4 × 56 = 224 g

0.34 g Fe is present in Haemoglobin = 100 g

224 g Fe is present in Haemoglobin $= \frac{100}{0.34} \times 224$

= 65882 g

38. (*c*) 1 g – atom of nitrogen = 6.02×10^{23} N atoms

1 mol of N = 1/2 mole of N_2 = 11.2 L at S.T.P

39. (*b*) 10 g of $CaCO_3$ = 0.1 mol of $CaCO_3$

= 0.1 × 3 g – atom of Oxygen

= 0.3 g atoms of oxygen

40. (*b*) Magnesium bicarbonate is $Mg(HCO_3)_2$. So, 146 g of $Mg(HCO_3)_2$ contains O atom = 0.6 N_A

41. (*b*) Molecular mass of CO_2 = 44g mol^{-1}

$\therefore$ 4 mol of CO_2 = 44 × 4 = 176 g

(*c*) 6.02×10^{23} atoms of hydrogen has mass = 1.008 g

= 12×10^{24} atoms of hydrogen has mass = 20.1 g

(*d*) 22.4 L of helium at N.T.P. has mass = 4 g 11.2 L of helium at N.T.P. has mass = 2 g

Thus; 4 moles of CO_2 has maximum mass

42. (*a*) In 12 g of carbon, the amount of C – 14

$$= \frac{12 \times 2}{100} = 0.24$$

$\therefore$ C-14 atoms in 0.24 g $= \frac{0.24 \times 6.02 \times 10^{23}}{14}$

= 1.03×10^{22} atoms

43. (*b*) Mass of 22.4 L gas at S.T.P. = $\frac{8 \times 22.4}{5.6}$ = 32 g (it is also equal to mol. mass)

Mol. mass = 32

$\therefore$ V.D. = $\frac{32 \times 1}{2}$ = 16

44. (*c*) Mass of 1 atom = 1.8×10^{-22} g

Mass of 6.022×10^{23} atoms

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

= 108.36 g

$\therefore$ Atomic mass of element = 108.36 g

45. (*d*) 1 M produces ions = 3 moles
$\therefore$ 0.1 M H_2SO_4 produce ions = 0.3 mol
Number of ions = $0.3 \times 6.02 \times 10^{23}$
$= 1.8 \times 10^{23}$

46. (*c*) 0.1 Mol $CaCO_3$ = 0.1 × 100 = 10 g

$Ca^{2+} = \dfrac{1.51 \times 10^{23}}{6.023 \times 10^{23}} \times 40 = 10$ g

$CO_3^{2-} = 0.16 \times 60 = 9.6$ g

$Br = \dfrac{7.525 \times 10^{22}}{6.02 \times 10^{23}} \times 80 = 10$ g

47. (*b*) Equal volume under similar conditions of temperature and pressure have fixed number of molecules.

48. (*a*) Moles of sugar added = $\dfrac{1.71}{342} = 5 \times 10^{-3}$

Carbon atoms added = $12 \times 5 \times 10^{-3} \times 6.02 \times 10^{23}$
$= 3.61 \times 10^{22}$

49. (*b*) 0.1 mole P_4 contains = $4 \times 0.1 \times 6.02 \times 10^{23}$
0.05 mole of S_8 contains = $8 \times 0.05 \times 6.02 \times 10^{23}$ atoms

50. (*d*) 82 g of $N_2 = \dfrac{82}{28}$ g-molecule = 2.92 moles

51. (*d*) Since; mass of oxygen is fixed. Therefore, the number of g-atom and also the number of atoms will be fixed.

i.e, Atoms of Oxygen $= \dfrac{1 \times 2}{32}$ in O_2

$= \dfrac{1}{16}$ in atomic Oxygen

$= \dfrac{1 \times 3}{48}$ in O_3

52. (*d*) 12.8 g of SO_2 = 4.48 L
6.02×10^{22} molecules of CH_4= 2.24 L
0.5 mol of NO_2 = 11.2 L
1 g molecule of CO_2 = 22.4 L

53. (*c*) G.M.M of O_2 = 32 g

$$\text{eq mass} = \frac{\text{At.mass of O}}{\text{valency}} = \frac{16}{2} = 8 = \frac{1}{4} = \frac{N_A}{4}$$

54. (*d*) No of mol of triatomic gas

$= \dfrac{224\,\text{ml}}{22400\,\text{ml mole}^{-1}} = 10^{-2}$ mol

No. of molecules = 10^{-2} mol × (6.02×10^{23} molecules mol^{-1})
no. of atoms of gas = 6.02×10^{21}
$= 3 \times 6.02 \times 10^{21} = 18.06 \times 10^{21}$
1.806×10^{22} atoms has mass = 1 g

1 atom has mass $= \dfrac{1}{1.806 \times 10^{22}\,\text{g}}$

$= \dfrac{10^{-22}\,\text{g}}{1.806}$

$= 0.554 \times 10^{-22}$ g

$= 5.54 \times 10^{-23}$ g

55. (*c*) Mass of 1 mole of electrons
$= 9.11 \times 10^{-31}$ kg × 6.02×10^{23} = 5.5×10^{-7} kg mol^{-1}

56. (*c*) If 3.18 mol of hydrogen atom are there, then number of moles of oxygen atoms

$= \dfrac{3.18}{3}$ mol = 1.06 mol

57. (*d*) Molar mass of $C_{10}H_{16}O$ = 120 + 16 +16
= 152 g mol^{-1}

$$25.0\,\text{mg} = 25.0\,\text{mg} \times \left(\frac{1\,\text{g}}{10^3\,\text{mg}}\right) \times \left(\frac{1\,\text{mol}}{152\,\text{g}}\right) \times \left(\frac{27 \times 6.02 \times 10^{23}\ \text{atoms}}{1\,\text{mol}}\right)$$

$$= \frac{25 \times 27 \times 6.02}{152} \times 10^{20}\ \text{atom}$$

$= 2.67 \times 10^{21}$ atoms

58. (*c*) $\dfrac{4}{40} \times N_A = \dfrac{N_A}{10}$ atom

0.5 M $Na_2SO_4 = \dfrac{0.5 \times 2}{100} \times N_A = \dfrac{N_A}{10}$

59. (*d*) $SO_2 \rightarrow$ Contain 3 atom

$n = \dfrac{\text{wt.}}{\text{M.wt.}} = \dfrac{64}{64} \times 3N_A$

$= 3 \times 6.02 \times 10^{23}$

60. (*a*) Total $e^- + P + n^o = 18$
$18 \times N_A = 1.084 \times 10^{25}$

61. (*a*) 111 g anhydrous $CaCl_2$ = 1 mole $CaCl_2$
1 mole $CaCl_2$ contain 1 mole Ca^{+2} and 2 mole Cl^-

62. (*d*) 1 g molecule of CO_2 = 1 mole CO_2
1 mole CO_2 occupy $\rightarrow$ 22.4 L

63. (*b*) 224 g contain 36 g of water
488 g contain 2 × 36 g of water
18 g of water = 1 mole
72 g of water = 4 mole

64. (*b*) 4.4 g CO_2 = 0.1 mole CO_2 molecule
2.24 L H_2 at STP = 0.1 mole H_2 molecule
= 0.2 mole molecules
$= 0.2 \times 6.023 \times 10^{23} = 1.20 \times 10^{23}$

65. (*c*) 22.4 L Vol. of gas at STP = 1 mole

66. (*c*) $C\% : H\% : N\% = \dfrac{40}{12} : \dfrac{13.3}{1} : \dfrac{46.7}{14}$

67. (*a*) Mass of chloride = $(B_2 - B_1)$ g
$(B_2 - B_1)$g combines with metal = B_1 g
$\therefore$ 35.5 g combines with metal = $\dfrac{35.5 \times B_1}{B_2 - B_1}$ g

68. (*d*)

	Mass in "g"	% age	H. Ratio	S.R
C	24	40	40/12 = 3.33	1
H	4	6.66	6.66/1=6.66	2
O	32	53.33	53.33/16=3.33	1

Empirical formula : CH_2O

69. (*a*)

Element	%	At. mass	$\frac{\%}{\text{At.mass}}$	Simplest ratio
A	25	12.5	25/12.5 = 2	1
B	75	37.5	75/37.5 = 2	1

The formula of the compound is AB

70. (*a*) Mol. mass of Fe $(CHCOO)_2$ = 170 g

Fe in 100 gms of Fe $(CHCOO)_2 = \frac{56\times100}{170}$
$= 32.9$ mg

Total Fe in 400mg of capsule = 32.9 mg

percentage of Fe $= \frac{32.9\times100}{400} = 8.2\%$

71. (*b*)

Element	%	At. mass	Mass Ratio	S.R
X	50	10	5	2
Y	50	20	2.5	1

E.F. = X_2Y

72. (*d*) E.F. = C_3H_4O

E.F. Mass = 12 × 3+ 4 × 1 + 16 = 56

$n = \frac{170\pm5}{56} = 3$

$\therefore$ M.F. = $(C_3H_4O)_3 = C_9H_{12}O_3$

73. (*b*)

		Oxide (I)	Oxide (II)
Metal;	M	50%	40%
Oxygen;	O	50%	60%

As first Oxide is MO_2

let atomic mass of M = x

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

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

x = 32

At. mass of metal M, x = 32

let formula of second Oxide M_2O_n

% of M $= \frac{2x}{2x+16n}\times100 = \frac{64}{64+16n}\times100$

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

0.25 n = 2.5 – 1 = 1.5

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

Now, formula of second oxide

$= M_2O_6$ or MO_3

74. (*b*) Molecular mass of B = M

V.D. of B = M/2

V.D. of A = 4 × V.D. of B

$4 \times \frac{M}{2} = 2$ M

Molecular mass of A = 2 × 2 M = 4M

75. (*a*) Moles of N $= \frac{28}{14} = 2$

Moles of metal, M $= \frac{100-28}{a}=\frac{72}{a}$

Mole ratio, M:N $= \frac{72}{a}:2 = 3:2$

$\frac{72}{a} = 3$

$a = \frac{72}{3} = 24$

76. (*b*) $\frac{M(G)}{M(N_2)}=\frac{0.44}{0.28}$ (Given)

$\therefore M_{(G)} = \frac{0.44}{0.28}\times 28 = 44$ g

77. (*a*) 44 CO_2 = 1 mol carbon

3.38 g CO_2 = 0.0768 moles carbon

18 g H_2O = 2 mol of hydrogen atoms

0.690 g H_2O = 0.0767 mol hydrogen atoms

Molar ratio of C : H = 1:1

$\therefore$ Empirical formula of hydrocarbon is CH

78. (*a*) % of $C = \frac{Wt}{M.Wt(CO_2)}\times100 = \frac{12}{44}\times100 = 27.27$

79. (*b*) 100 g haemoglobin has = 0.33 g Fe

67200 g haemoglobin has Fe

$= \frac{0.33}{100}\times 67200 = 221.76$ g

1 mole of haemoglobin

$= \frac{221.76}{56}$ g atom of Fe

= 3.96 g atom of Fe ≈ 4.0 g atoms

80. (*d*) Moles of Iodine $= \frac{254}{127} = 2$

Moles of Oxygen $= \frac{80}{16} = 5$

formula : I_2O_5

81. (*b*) 0.5 mol of $K_4[Fe(CN)_6]$

has C $= \frac{1}{2}\times 6$ mol = 3 mol or 3 ×12 = 36 g

82. (*c*) X_2O

X = 7

X = 7 × 2 = 14 g

83. (*c*) $100 \rightarrow 20$

$? \rightarrow 28$

84. (*b*) $CO + \frac{1}{2}O_2 \rightarrow CO_2$

85. (*d*) The balanced equation is:

$4\,NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$ ...(1)

4×17 g of NH_3 (eq^{-1})

Required O_2 = 5 mol

6.8 of NH_3 require $O_2 = \frac{5 \times 6.8}{4 \times 17} = 0.5$ mol

86. (*a*) $CaCl_2 + 2AgNO_3 \rightarrow Ca(NO_3)_2 + 2\,AgCl$

111 g 2×143.5 g

$CaCl_2$ require to produce 2×143.5 of AgCl = 111 g

$CaCl_2$ required to produce 14.35 g of AgCl

$= \frac{111 \times 14.35}{2 \times 143.5} = 5.55$ g

87. (*b*) $NaOH + HCl \rightarrow NaCl + H_2O$

88. (*b*) $M = \frac{w}{GM.w} \times \frac{1000}{V\ in\ ml}$

89. (*b*) $w = \frac{N \times E.wt \times V(\text{in ml})}{1000}$

90. (*b*) $M = \frac{w}{Molar\ mass} \times \frac{1}{V\ in\ litres}$

91. (*c*) $2\,NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

Pure H_2SO_4 is required for 1 mole of NaOH = 1/2 mole = 49 g

70% H_2SO_4 required for 1 mole of NaOH

$= \frac{49 \times 100}{70} = 70$

92. (*a*) $2KClO_3 \rightarrow 2\,KCl + 3O_2$

2×122.5 g 3×32 g

3×32 g of O_2 is produced from $KClO_3$ = 245 g

48g of O_2 is produced from $KClO_3$ = 245 g

$\therefore$ 80% $KClO_3$ needed $= \frac{122.5 \times 100}{80} = 153.12$ g

93. (*c*) $N = \frac{\text{percentage} \times d \times 10}{E.wt}$

94. (*b*) $Mass\,\%_{(s)} = \frac{n_s}{n_s + n_o}$

95. (*d*) $\text{No. gram equivalents} = \frac{N \times V(\text{in } ml)}{1000}$

96. (*d*) $N = \frac{w}{E.wt} \times \frac{1000}{V(\text{in ml})}$

97. (*b*) $m = \frac{w}{E.wt} \times \frac{1000}{b(\text{in gms})}$

98. (*c*) $MCl_x + x\,AgNO_3 \rightarrow M(NO_3)_x + xAgCl$

Molar ratio: 1   x   1   x

Molecules of $AgNO_3$ given $= \frac{0.6}{100} \times 500 = 3$

Moles of MCl_x = 0.1

$\therefore$ The value of x = 3

99. (*b*) $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

100 g   73 g   44 g

100 ml of 20% HCl solution = 20 g HCl

Here, $CaCO_3$ is a limiting reagent

Hence,

20 g of $CaCO_3 = \frac{44}{100} \times 20\,g$ of CO_2 = 8.8 g of CO_2

100. (*a*) $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

25 ml of 0.75 molar HCl $= \frac{25}{1000} \times 1 \times 0.75 = 0.01875$ moles

Moles of $CaCO_3$ required $= \frac{\text{moles of HCl}}{2} = \frac{0.01875}{2}$

$= 9.375 \times 10^{-3}$ moles

Mass of $CaCO_3$ required $= 9.375 \times 10^{-3}$ moles $\times$ 100 g/moles

= 0.94 g

101. (*a*) 22.4 L of SO_2 at N.T.P = 1 moles

11.2 L of SO_2 at N.T.P = 1/2 moles

$SO_2 + 2H_2S \rightarrow 3S + 2H_2O$

Here SO_2 is a limiting reactant

$\therefore$ moles of sulphur produced = $3 \times$ moles of SO_2

= 1.5 moles

102. (*b*) $SO_2Cl_2 + 2H_2O \rightarrow H_2SO_4 + 2HCl$

$H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4$

$\therefore$ 1 mol of SO_2Cl_2 in aqueous solution required is 2 moles $Ca(OH)_2$

103. (*b*) Moles of Na $= \frac{1.15\,g}{23\,g\,mol^{-1}} = 0.05$ mol

Moles of C $= \frac{3.01 \times 10^{22}}{6.02 \times 10^{23}} = 0.05$ mol

Moles of O = 0.1

Moles ratio, Na : C: O = 1 : 1 :2

$\therefore$ Empirical formula – $NaCO_2$

104. (*c*) $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

Here: HCl is a limiting reagent

$\therefore$ Moles of H_2 produced $= \frac{\text{moles of HCl}}{2}$

$= \frac{0.52}{2} = 0.26$

105. (*b*) $M = \frac{\text{Percentage} \times \text{specific gravity} \times 10}{\text{GMW}}$

$M_1V_1 = M_2V_2$

Learning Plus

1. (*b*) Molality is not effected by temperature as molality is not involved with volume (volume changes with temperature).

Molality = [(Moles of solute / Weight of solvent (kg)]

2. (*d*) We have

$Fe_2(SO_4)_3 \rightarrow Fe_2O_3 + 3SO_3$

So, 400 g of ferric sulphate is producing 240 g of SO_3.

So, 1 kg (1000 g) of ferric sulphate produce

SO_3 = (240/400) × 1000 = 600 g

In 80 g of SO_3, mass of oxygen = 48 g

Mass of sulphur = 32 g,

So, mass of oxygen in 600 g of SO_3 = (48/80) × 600 = 360g

Mass of sulphur in 600 g of SO_3 = (32/80) × 600 = 240 g

(Mass of oxygen/Mass of sulphur) = (360/240) = (3/2)

3. (*d*) Molar mass = 14 + 3 = 17

No. of moles = (4.25/17) = 0.25 moles

Total atoms = 4[1 + 3] NH_3

1 mole = $4 \times 6.022 \times 10^{27}$ atoms

0.25 mole = $0.25 \times 4 \times 6.022 \times 10^{23} = 6.022 \times 10^{23}$ atoms.

4. (*b*) Average of readings of student,

$$A = \frac{3.01 + 2.99}{2} = 3.00$$

Average of readings of student,

$$B = \frac{3.05 + 2.95}{2} = 3.00$$

Correct reading = 3.00

For both the students, average value is close to the correct value. Hence, readings of both are accurate.

But the readings recorded by student A are more precise as they differ only by ± 0.01, whereas readings recorded by the student B are differed by +0.05. Thus, the results of student A are both precise and accurate.

5. (*c*) Since, molarity (M) is calculated by following equation

$$\text{Molarity} = \frac{\text{weight} \times 1000}{\text{molecular weight} \times \text{volume (mL)}}$$

$$= \frac{5.85 \times 1000}{58.5 \times 500} = 0.2 \text{ molL}^{-1}$$

Note: Molarity of solution depends upon temperature because volume of a solution depends upon temperature.

6. (*a*) No of atoms = (558.5/55.85) = 10 mole

Total number of atom

$= 10 \times 6.022 \times 10^{23} = 6.022 \times 10^{24}$

12 g of carbon means Avogadro's number.

Therefore, 60 g of carbon means 5 moles of carbon.

Thus contains $5 \times N_A$ atoms, twice of this equals 10 N_A atoms of carbon.

Total number of atom

$= 10 \times 6.022 \times 10^{23} = 6.022 \times 10^{24}$

6.022×10^{24} no. of atoms found in twice that in 60 g carbon.

7. (*b*) Let W% of Ne^{20} = x

W % of Ne^{22} = 100 – x

$$20.2 = \frac{x \times 20 + (100 - x)22}{100}$$

w % of Ne^{20} = 90 %

w% of Ne^{22} = 100 – 90 = 10%

Ne^{20} : Ne^{22} = 90 : 10 = 9 : 1

8. (*a*) Molecular mass of $CuSO_4$, $5H_2O$ = 249.5

= 63.5 + 32 + 4 × 16 + 5 × 18 = 249.5

Thus 249.5 g of $CuSO_4.5H_2O$ contains Avogadro no. of molecules

So mass of 6.023×10^{23} molecules = 249.5

$$\text{So mass of } 10^{22} \text{ molecules} = \frac{249.5}{6.023 \times 10^{23}} \times 10^{22} = 4.144 \text{ g}$$

9. (*a*) 111 g $CaCl_2$ has = N ions of Ca^{2+}

$$\text{No. of molecules} = \frac{222}{111} \times N = 2N \text{ ions of } Ca^{2+}$$

111 g $CaCl_2$ = 2 N ions of Cl^-

$$222 \text{ g of } CaCl_2 = \frac{2N \times 222}{111} \text{ ions of } Cl^-$$

= 4 N ions of Cl^-

10. (*a*) E.F. Wt = 14 gm

MW = d × 22.4

$$\text{M.F} = \text{E.F} \times \frac{MW}{EF\ Wt.}$$

11. (*b*) Given that,

M_1 = 5 M

V_1 = 500 mL

V_2 =1500 mL

M_2 = M

For dilution, a general formula is

$M_1V_1 = M_2V_2$

(Before dilution) (After dilution)

500 × 5 M =1500 × M

$$M = \frac{5}{3} = 1.66 \text{ M}$$

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

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

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

$$0.40\text{ g Ca} = \frac{0.4}{40} = 0.01\,\text{mol}$$

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

Hence, 12 g He contains greatest number of atoms as it possesses maximum number of moles.

13. (*c*) E.F. Wt = 30 gm

MW = 2 × V.D

$$\text{M.F} = \text{E.F} \times \frac{MW}{EF\ Wt.}$$

14. (*d*) $\text{W \% of } N_2 = \frac{28}{22400} \times \frac{V\ of\ N_2\ at\ S.T.P}{Wt.\ of\ compound} \times 100$

15. (*b*) The balanced chemical reaction is,

$K_2SO_4 + BaCl_2 \rightarrow BaSO_4 + 2KCl$

1 mole of potassium sulphate reacts with 1 mole of barium chloride to precipitate 1 mole of barium sulphate and 2 moles of potassium chloride is left in the solution.

Hence, in the given situation, sulphate ions are the limiting agents.

So, 0.7 moles of potassium sulphate reacts with 0.7 moles of barium chloride to precipitate 0.7 moles of barium sulphate.

16. (*a*) The reaction is given below:

$4Fe + 3O_2 \rightarrow 2Fe_2O_3$

Fe is the limiting reagent in this reaction. Limiting reagent decides the amount of product formed.

Therefore, 4 moles of Fe produce 2 moles of Fe_2O_3.

Thus, 0.5 moles of Fe will produce 0.25 moles of Fe_2O_3.

17. (*a*) $\text{w} = \frac{\text{N} \times Eq.wt \times \text{V(in ml)}}{1000} = \frac{0.1 \times 63 \times 250}{1000} = 1.575\text{ g}$

18. (*c*) In the given question, 0.9 g L^{-1} means that 1000 mL (or 1 L) solution contains 0.9 g of glucose.

$$\text{M} = \frac{\text{conc in gL}^{-1}}{\text{Molar mass}} = \frac{0.90}{180} = 0.005\text{ M}$$

19. (*d*) $\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mass of solvent (in kg)}}$ …(*i*)

Given that, Mass of solvent (H_2O) = 500 g = 0.5 kg

Weight of HCl = 18.25 g

Molecular weight of HCl = 36.5 g

$$\text{Moles of HCl} = \frac{18.25}{36.5} = 0.5 = \frac{0.5}{0.5} = 1\text{ m}$$

20. (*a*) No. of moles = $\frac{W}{MW}$

21. (*b*) 22400 cc of a gas at STP = 6.023 × 10^{23} molecules

So, 1.12 × 10^{-7}cc of a gas at STP = ?

$$= \frac{1.12 \times 10^{-7} \times 6.023 \times 10^{23}}{22400} = 3.01 \times 10^{12}\text{ molecules}$$

22. (*a*) 32 g of O_2 = 1 mol. = 6.023 × 10^{23} molecules.

∴ 320 mg = 0.32 g = 0.01 mol = 6.3023 × 10^{21}

From this 6.023 × 10^{20} moles are removed = 0.001 moles

Remaining = 0.01 – 0.001 = 0.009 moles = 9 × 10^{-3} moles

23. (*b*) The at. no. of N and O are 1/2 of their Mass no. So, one molecule of 92 g molecular weight will have 46 electrons so 1 gm of molecule contain 46 N e^{-}s.

24. (*b*) No. of moles = $\frac{W}{MW}$

1 mole contain 2 mole of water

25. (*a*) One mole of any substance contains 6.022 × 10^{23} atoms/molecules.

Hence, Number of millimoles of H_2SO_4

= molarity × volume in mL

= 0.02 × 100 = 2 millimoles

= 2 × 10^{-3} mol

Number of molecules = number of moles × 6.022 × 10^{23}

= 2 × 10^{-3} × 6.022 × 10^{23}

= 12.044 × 10^{20} molecules

26. (*d*) 3, 3, 3

27. (*d*) $\frac{Wt}{At.wt} = no.of\ gram\ atoms$

$$Mg : C : N = \frac{12}{24} : \frac{12}{12} : \frac{14}{14}$$

28. (*b*) Let us assume the weight of the compound = 100 gm

Weight of nitrogen = 36.8 gm

Moles of nitrogen = (36.8/14)

∴ Weight of oxygen = 100 – 36.8 = 63.2 gm

Moles of oxygen = (63.2/16)

The ratio of number of moles of N and O is,

N : O = (36.8/14) : (63.2/16)

= 2.628 : 3.95 = 1 : 1.5 = 2 : 3

So, formula is N_2O_3.

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

$$x = \frac{160}{40} = 4 \quad and \quad x + y = \frac{26}{4}$$

$$\frac{Y}{4} = \frac{10}{4}$$

∴ Y = 10

30. (*c*) Equal volumes of different gases at STP contains an equal number of moles.

4 g oxygen corresponds to (4/32) = 0.125 mole. This is also equal to the number of moles of hydrogen. 1 mole of hydrogen corresponds to 2 gm.

Hence, 0.125 mole will correspond to 2 × 0.125 = 0.25 gm.

Hence, the correct option is C.

31. (*c*) $N = \frac{\text{No. of milli equivalents}}{\text{V (in m}l)}$

32. (*c*) $M = \frac{\text{percentage} \times 10 \times d}{\text{GMW}}$

33. (*c*) Empirical formula mass = CH_2O

= 12 + 2 × 1+ 16 = 30

Molecular mass = 180

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

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

∴ Molecular formula = n × empirical formula

= 6 × CH_2O

= $C_6H_{12}O_6$

34. (*a*) Given that density of solution = 3.12 g mL^{-1}

Volume of solution = 1.5 mL

For a solution,

Mass = volume × density

= 1.5 mL × 3.12 gm L^{-1} = 4.68 g

Hence, the answer is reported as 4.7 g.

2 significant figures.

35. (*c*) $M = \frac{w}{GMW} \times \frac{1000}{V\ in\ ml}$

N = M × Basicity of an acid

36. (*c*) $w = \frac{N \times GEW \times V\,(\text{in ml})}{1000}$

37. (*d*) W = N × G.EW × V (In L)

$$\%\ \text{Purity of NaOH} = \frac{40}{50} \times 100$$

= 80 %

38. (*a*) 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.

39. (*b*) The element, carbon, combines with oxygen to form two compounds, i.e., carbon dioxide and carbon monoxide in CO_2, 12 parts by mass of carbon combine with 32 parts by mass of oxygen while in CO, 12 parts by mass of carbon combine with 16 parts by mass of oxygen.

Therefore, the masses of oxygen combine with a fixed mass of carbon (12 parts) in CO_2 and CO are 32 and 16 respectively. These masses of oxygen have ratio of 32 : 16 or 2 : 1 to each other.

This is an example of law of multiple proportion.

Multiconcept MCQs

1. (*a*) $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

1 L of N_2 reacts with 3 L of H_2 to form 2 L of NH_3.

So, N_2 is the limiting reagent here because when 30L of H_2 will be consumed, the volume of N_2 consumed will be $\frac{1}{3}^{rd}$ i.e., $\frac{1}{3} \times 30 = 10L$.

Since actual yield is 50% of the expected value, NH_3 formed = 10 L , N_2 reacted = 5 L and H_2 reacted 15 L.

So, the final mixture contains 10 L NH_3, 25 L N_2 and 15 L H_2.

2. (*b*) $3NaBH_4 + 4BF_3 \longrightarrow 3NaBF_4 + 2B_2H_6$

Since BF_3 is in excess, the limiting reagent is $NaBH_4$.

To obtain 2 mole B_2H_6, $NaBH_4$ required = 3 mole

To obtain 0.200 mole B_2H_6, $NaBH_4$ required = $\frac{3}{2} \times 0.200$

= 0.300 mole

Because, the yield is 70%, Hence

$$x \times \frac{70}{100} = 0.300$$

$$x = \frac{0.300}{1000} \times \frac{100}{70} = \frac{3}{7} = 0.429 \text{ moles}$$

3. (*a*) $Ag_2S \rightarrow 2Ag + S$

To obtain 2 mole Ag, mole of Ag_2S required = 1

To obtain $\frac{1}{108}$ mole Ag, mole of Ag_2S required

$$= \frac{1}{2} \times \frac{1}{108} = \frac{1}{216} \text{mole}$$

grams of Ag_2S required $= \frac{1}{216} \times 248$ g

1.24 g Ag_2S is obtained from ore = 100 g

$\frac{248}{216}$ g Ag_2S is obtained from ore $= \frac{100}{1.24} \times \frac{248}{216}$

$= \frac{24800}{267.84} = 92.6$ g

4. (*a*) C : H : N = 9 : 1 : 3.5

$\therefore$ mole ratio $= C : H : N = \frac{9}{12} : \frac{1}{1} : \frac{3.5}{14} = 3 : 4 : 1$

$(C_3H_4N)_x \Rightarrow x = 2$ & $C_6H_8N_2$

5. (*b*) $[\text{urea}] = \frac{6.02 \times 10^{20}}{6.02 \times 10^{23}} = 10^{-3}\,\text{mol}$

$\therefore \quad [\text{urea}] = \frac{10^{-3}}{0.1} = 0.01$ M

6. (*d*) Mn = x gm $= \frac{x}{55} = 0.018x$

O = x gm $= \frac{x}{16} = 0.0625\,x$

$\frac{0.018\,x}{0.018\,x} = 1 \Rightarrow \frac{0.0625\,x}{0.018\,x} \approx 3.5$

$Mn_1O_{3.5} \Rightarrow Mn_2O_7$

7. (*b*) $M + H_2SO_4 \rightarrow MSO_4 + H_2$

$\frac{\text{Mole of M}}{1} = \frac{\text{Mole of } H_2SO_4}{1}$

$\frac{1}{a} = \frac{4.08}{98}$

$a = \frac{98}{4.08} = 24.01$

Thus atomic weight of M = 24 g.

8. (*a*) Let molar mass of compound = y g

% by wt of C = 36

So, $\frac{24}{y} \times 100 = 36$

$\frac{2400}{36} = y$

y = 66.6 g

So, 66.6 g has moles = 1

10 g has moles $= \frac{1}{66.6} \times 10 = 0.15$

9. (*a*) $C_8H_{18} + \frac{25}{2}O_2 \rightarrow 8CO_2 + 9H_2O$

Moles of CO_2 formed $= \frac{7.04}{44}$ moles $= 0.16$

now, from equation

Moles of H_2O formed when 8 moles of CO_2 react = 9

Moles of H_2O formed when $\frac{7.04}{44}$ moles of CO_2 react

$= \frac{9}{8} \times \frac{7.04}{44} = \frac{63.36}{8 \times 44} = 0.18$ moles

wt. of H_2O formed = 0.18 × 18 = 3.24 g

10. (*c*) $2KHCO_3 \xrightarrow{\Delta} K_2CO_3 + H_2O + CO_2 \uparrow$

Loss of wt. will be because of CO_2 escaped

Total $KHCO_3$ chosen = 1.50 g

percentage purity = 80%

$\therefore$ Pure $KHCO_3 = 1.50 \times \frac{80}{100} = 1.2$ g

M. mass of $KHCO_3$ = 39 + 1 + 12 + 3 × 16 = 100 g

moles of pure $KHCO_3 = \frac{1.2}{100} = 0.012$ moles

From balanced equation-

2 moles of $KHCO_3$ yield moles of CO_2 = 1

0.012 moles of $KHCO_3$ yield moles of $CO_2 = \frac{1}{2} \times 0.012$

= 0.006 moles

wt. of CO_2 formed = 0.264 g

11. (*a*) $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$

From balanced equation,

71 gm Cl_2 is produced from HCl = 4 × 36.5 g

2.5 gm Cl_2 is produced from HCl $= \frac{4 \times 36.4 \times 2.5g}{71}$

= 5.14 g

Now, HCl is 36% by mass, which means-

36 g HCl is obtained from HCl solution = 100 g

5.14 g HCl is obtained from HCl solution $= \frac{100 \times 5.14}{36}$

= 14.27 g

12. (*d*) Density of solution = 1.25 g/mL

29.2% (w/w) means that 29.2 g of HCl is present in 100 gms of solution

So, $\rho(\text{density of solution}) = \frac{\text{wt. of soln.}}{\text{vol. of soln.}}$

$1.25 \text{ g/mL} = \frac{100}{\text{vol. of soln.}}$

$v = \frac{100}{1.25}$ mL

$\text{molarity of solution} = \frac{\text{no. of moles of solute}}{\text{vol. of solution (in ltrs)}}$

$M = \frac{29.2/36.5}{100/1.25} \times 1000 = 10$ M

Apply, $M_1V_1 = M_2V_2$

0.4 × 200 = 10 × V

V = 8 mL

13. (*c*) $2\,Mg + O_2 \rightarrow 2\,MgO$

2×24 g Mg reacts with $O_2 = 32$ g

30 g Mg reacts with $O_2 = \frac{32}{2 \times 24} \times 30 = 20$ g

So, 10 g O_2 will be left unreacted

Mg is the limiting reagent

2×24 g Mg forms MgO $= 2 \times 40$ g

30 g Mg forms MgO $= \frac{2 \times 40}{2 \times 24} \times 30 = 50$ g

14. (*b*) $4K_2Cr_2O_7 \xrightarrow{\Delta} 4K_2CrO_4 + 2Cr_2O_3 + 3O_2$

mole 4 3 (100% yield)

= 1 (for 33.33%)

Volume of O_2 gas produced $= 1 \times 22.4 = 22.4$ L

15. (*b*) $CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$

$n_{NaOH} = 1$

$\therefore$ CO_2 present in mixture = 0.5 and CO present = 0.3 mole

When more CO_2 produced = 0.3, more NaOH required $= 0.3 \times 2 = 0.6$ mole

16. (*a*) $CaCO_3 + MgCO_3 \xrightarrow{\Delta} CaO + MgO + 2CO_2$

Let the mass of $CaCO_3$ = x g

Then, mass of $MgCO_2$ = (3.68 – x)g

moles of $CaCO_3 = \frac{x}{100}$

moles of $MgCO_3 = \frac{3.68 - x}{84}$

Applying POAC for C-atoms

$$\frac{x}{100} + \frac{3.68 - x}{84} = 0.04$$

x = 2

$$\therefore \quad n_{CaCO_3} = \frac{2}{100} = 0.02 \text{ and } n_{MgCO_3} = \frac{1.68}{84} = 0.02$$

$$\text{mole\% of } CaCO_3 = \frac{0.02}{0.04} \times 100 = 50\%$$

$$\text{mole\% of } MgCO_3 = \frac{0.02}{0.04} \times 100 = 50\%$$

17. (*c*) Let relative abundance of heavier isotope be x %

$$A = \frac{(100 - x)(A - 1) + x(A + 3)}{100}$$

100 A = 100A – xA – 100 + x + 3x + xA

4x = 100

$\therefore$ x = 25.

18. (*d*) Only single solution have all these

Means 100 ml solution have 5.85 gm NaCl = 0.1 mole and 5.55 gm $CaCl_2$ = 0.05 mole

$$[Cl^-] = \frac{(0.1 + 0.05 \times 2) \times 1000}{100} = 2\,M$$

$$\Rightarrow [Na^+] = \frac{(0.1 + 0.15) \times 1000}{100} = 2.5\,M$$

$$[Ca^{2+}] = \frac{0.5}{100} \times 1000 = 0.5\,M \quad \Rightarrow [OH^-] = 1.5\,M$$

19. (*b*) Moles of Fe $= \frac{0.0056}{56} = 10^{-4}$

1 mol of alum = 2 mol of Fe

2 mol of Fe = 1 mol of alum

10^{-4} mol of Fe $= \frac{1}{2} \times 10^{-4}$ mol of alum

$= 0.5 \times 10^{-4}$ mol of alum

20. (*d*) (*i*) 14 g (*ii*) 40 g

(*iii*) $\frac{108}{6.022 \times 10^{23}} = 1.79 \times 10^{-22}$ g

(*iv*) 32 g (v) 1.99 g (vi) 1 g

Hence, the correct order of increasing masses is

(*iii*) < (*vi*) < (*v*) < (*i*) < (*iv*) < (*ii*)

21. (*a*) $H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$

$112L = \frac{112}{22.4}$ moles $= 5$ moles $213g = \frac{213}{71} = 3$ moles

Here, $Cl_{2(g)}$ is limiting reagent,

1 Mole $Cl_{2(g)}$ produces $HCl_{(g)}$

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

3 Moles $Cl_{2(g)}$ produces $HCl_{(g)}$

$= 3 \times 2 \times 6.022 \times 10^{23}$ molecules

$= 3.61 \times 10^{24}$ molecules

22. (*c*) $MgO + CO_2 \rightarrow MgCO_3$

40 g 44 g

40g MgO needs CO_2 = 44 g

30g MgO needs $CO_2 = \frac{44 \times 30}{40} = 33g$

$CaCO_3 \rightarrow CaO + CO_2$

100 g 44 g

44 g CO_2 is obtained from $CaCO_3$ = 100 g

33 g CO_2 is obtained from $CaCO_3 = \frac{100 \times 33}{44} = 75\,g$

Percentage purity of $CaCO_3$ sample

$$= \frac{75}{85} \times 100 = 88.24\%$$

NEET Past 10 Years Questions

1. (*b*)

Element	%	At Weight	% / At weight	Simplest Ratio
C	78	12	6.5	1
H	22	1	22	3

Empirical formula of this compound is CH_3

2. (*c*)

(*a*) Number of Mg atoms $= \frac{1}{24} \times N_A$

$= \frac{1}{24} \times 6.022 \times 10^{23}$ atom

(*b*) Number of O atoms $= \frac{1}{32} \times N_A$

$= \frac{1}{32} \times 2 \times 6.022 \times 10^{23}$ atom

(*c*) Number of Li atoms $= \frac{1}{7} \times N_A$

$= \frac{1}{7} \times 6.022 \times 10^{23}$ atom

(*d*) Number of Ag atoms $= \frac{1}{108} \times N_A$

$= \frac{1}{108} \times 6.022 \times 10^{23}$ atom

Hence, 1g lithium has the largest number of atoms.

3. (*c*)

No of atom in 12 g carbon = $12 \div (1.9926 \times 10^{-23})$

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

Thus Number of atoms in 1 mole carbon

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

4. (*c*)

Haber's process

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

20 moles need to be produced

2 moles of $NH_3 \rightarrow$ 3 moles of H_2

Hence 20 moles of $NH_3 \rightarrow \frac{3 \times 20}{2}$ = 30 moles to H_2

5. (*d*)

$$\underset{2.3\,g\ \text{or}\left(\frac{1}{20}\text{mol}\right)}{HCOOH} \xrightarrow{\text{Conc}\,H_2SO_4} CO(g) + \underset{\frac{1}{20}\text{mol}}{H_2O(l)}$$

$$\underset{4.5\ \text{or}\left[\frac{1}{20}\text{mol}\right]}{\begin{matrix}COOH\\ |\\ COOH\end{matrix}} \xrightarrow{\text{Conc.}H_2SO_4} \underset{\frac{1}{20}\text{mol}}{CO\,(g)} + \underset{\frac{1}{20}\text{mol}}{CO_2\,(g)} + H_2O\,(l)$$

Gaseous mixture formed is CO and CO_2. When it is passed through KOH, only CO_2 is absorbed. So the remaining gas is CO. KOH pellets absorbs all CO_2, H_2O is absorbed by H_2SO_4 thus CO is remaining product.

So, weight of remaining gaseous product CO is

$\frac{2}{20} \times 28 = 2.8g$

So, the correct option is (*d*)

6. (*a*)

(*a*) Mass of water = 18 × 1 = 18 g

Molecules of water = mole × N_A = $\frac{18}{18} N_A$

$= 1\ N_A$

(*b*) Molecules of water = mole × N_A $= \frac{0.18}{18} N_A$

$= 10^{-2}\ N_A$

(*c*) Molecules of water = mole × N_A = $10^{-3}\ N_A$

(*d*) Moles of water $= \frac{0.00224}{22.4} = 10^{-4}$

Molecules of water = mole × N_A = $10^{-4}\ N_A$

7. (*c*)

	%	Moles	Relative moles
C	85.7	$\frac{85.7}{12} = 7.14$	1
H	14.3	$\frac{14.3}{1} = 14.3$	2

Hence, empirical formula = CH_2.

empirical weight = 14

$\frac{3.01 \times 10^{20}}{6.022 \times 10^{23}}$ = No. of moles $= \frac{42 \times 10^{-3}}{M}$

$\frac{1}{2} \times 10^{-3} = \frac{42 \times 10^{-3}}{M}$

M = 84

$\therefore$ Atomicity $= \frac{84}{14} = 6$

Molecular formula = C_6H_{12}.

8. (*c*)

For XY_2, let atomic weight of X = A_x

and of Y = A_y

So, $n_{XY_2} = 0.1 = \frac{10}{A_x + 2A_y}$

$A_x + 2A_y = 100$ ...(1)

Similarly for X_3Y_2,

$3A_x + 2A_y = 180$ ...(2)

On solving (1) and (2)

$A_x = 40$ and $A_y = 30$

9. (*a*)

(*a*) 18 moles of water will contain

$= 18 \times 6.022 \times 10^{23}$ molecules of H_2O

(*b*) 18 molecules

(*c*) $\frac{1.8}{18} = 0.1$ mole will contain

$= 0.1 \times 6.022 \times 10^{23}$ molecules of H_2O

(*d*) $\frac{18}{18}\text{g} = 1\,\text{mole} = 1 \times 6.022 \times 10^{23}$ molecules of H_2O

So, maximum number of molecules is present in 18 moles of H_2O.

10. (*c*) Avogadro's number 6.022×10^{23} is ideally the mass of number of atoms present in 1 mole that is 12 grams of C. If we change the Avogadro's number it will directly change the mass of 1 mole that is 12 g of C.

11. (*b*) Molecular weight of $AgNO_3 = 170$
Molecular weight of NaCl = 58.5

1. 16.9% solution of $AgNO_3$ means 16.9 g of $AgNO_3$ in 100 mL of solution
so, 8.45 g of $AgNO_3$ in 50 mL of solution.
2. 5.8% solution of NaCl means 5.8 g of NaCl is in 100 mL solution. So, in 50 mL = 2.9 g NaCl

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

Initial Mole: $\frac{8.45}{170}$ $\quad \frac{2.9}{58.5}$ $\quad 0 \quad 0$

$= 0.049 = 0.049 \quad 0 \quad 0$

Final mole: $0 \quad 0 \quad 0.049 \quad 0.049$

Mass of AgCl precipitated = 0.049 mole = 0.049×143.3

$= 7.02$ gm $\simeq 7$ gm

12. (*c*)

$$MgCO_3 \rightarrow MgO + CO_2(g)$$

Mw $\rightarrow$ 84 g $\quad$ 40 g

According to question

84 g $MgCO_3$ gives = 40 g MgO

1 g $MgCO_3$ gives = $\frac{40}{84}$

20 g $MgCO_3$ gives = $\frac{40}{84} \times 20 = 9.52$ g of MgO

But according to question
yield of MgO is = 8 g

$$\%\ \text{purity} = \frac{8}{9.52} \times 100 = 84\%$$

13. (*d*) 1 mole = 22.4 litres at S.T.P.

$n_{H_2} = \frac{22.4}{22.4} = 1\,\text{mol}\,;\ n_{Cl_2} = \frac{11.2}{22.4} = 0.5\,\text{mol}$

Reaction is as,

	$H_2(g)$	+ $Cl_2(g)$	$\rightarrow$ 2HCl (g)
Initial	1 mol	0.5 mol	0
Final	(1 – 0.5) = 0.5 mol	(0.5 – 0.5) = 0 mol	2 × 0.5 = 1 mol

Here, Cl_2 is limiting reagent. So, 1 mole of HCl (g) is formed.

14. (*d*)

$n_{Mg} = \frac{1}{24} = 0.0416$ moles

$n_{O_2} = \frac{0.56}{32} = 0.0175$ moles

The balanced chemical equation:

	Mg	+ $\frac{1}{2}O_2$	$\rightarrow$ MgO
Initial	0.0416 moles	0.0175 moles	0
Final	$(0.0416 - 2 \times 0.0175)$	0	2×0.0175

= 0.0066 moles (O_2 is limiting reagent)

$\therefore$ Mass of Mg left in excess $= 0.0066 \times 24$

$= 0.16$ g

15. (*c*)

According to Avogadro's principle, ratio of volume of gases will be equal to the ratio of their number of moles

$$\text{mole} = \frac{w}{M_w}$$

$n_{H_2} : n_{O_2} : n_{CH_4}$

$\frac{w}{2} : \frac{w}{32} : \frac{w}{16} \Rightarrow 16 : 1 : 2$

16. (*b*)

6.02×10^{23} number of molecules = 1 mole

$6.02 \times 10^{20} = 0.001$ mole

$$\text{Concentration} = \frac{\text{mole}}{V(\text{mL})} \times 1000 = \frac{0.001}{100} \times 1000$$

$\Rightarrow 0.01$ M

17. (*a*) Molarity of solution of dichlorotetraaquachromium(III) chloride = 0.01 M.

Volume of solution of dichlorotetraaquachromium(III) chloride = 100 ml.

The formula of dichlorotetraaquachromium(III) chloride is $[Cr(H_2O)Cl_2]Cl$.

On ionisation,

$$[Cr(H_2O)Cl_2]Cl \rightarrow [Cr(H_2O)Cl_2]^+ + Cl^-$$

Initial	100 × 0.01	0	0
Final	0	1 mol	1 mol

So 1 mol of Cl^- ions will react with 1 mol of $AgNO_3$ mole of $[Cr(H_2O)Cl_2]Cl$ 0.1 M 100 ml solution is,

No. of moles = Molarity × Volume

$= 0.01 \times 0.1 = 0.001$ mol

Hence, 0.001 mol of Cl^- ions will react with 0.001 mole of $AgNO_3$.

So number of moles of AgCl formed is 0.001 mol.