CHEMISTRY

CLASS-XI NEET

MODULE-01

Some Basic Concept of Chemistry

| Structure of Atom | Classification of Elements
| Chemical Bonding & Molecular Structure

Physics Wallah

NEET

Full Course

CHEMISTRY

Module-1

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Some Basic Concepts of Chemistry

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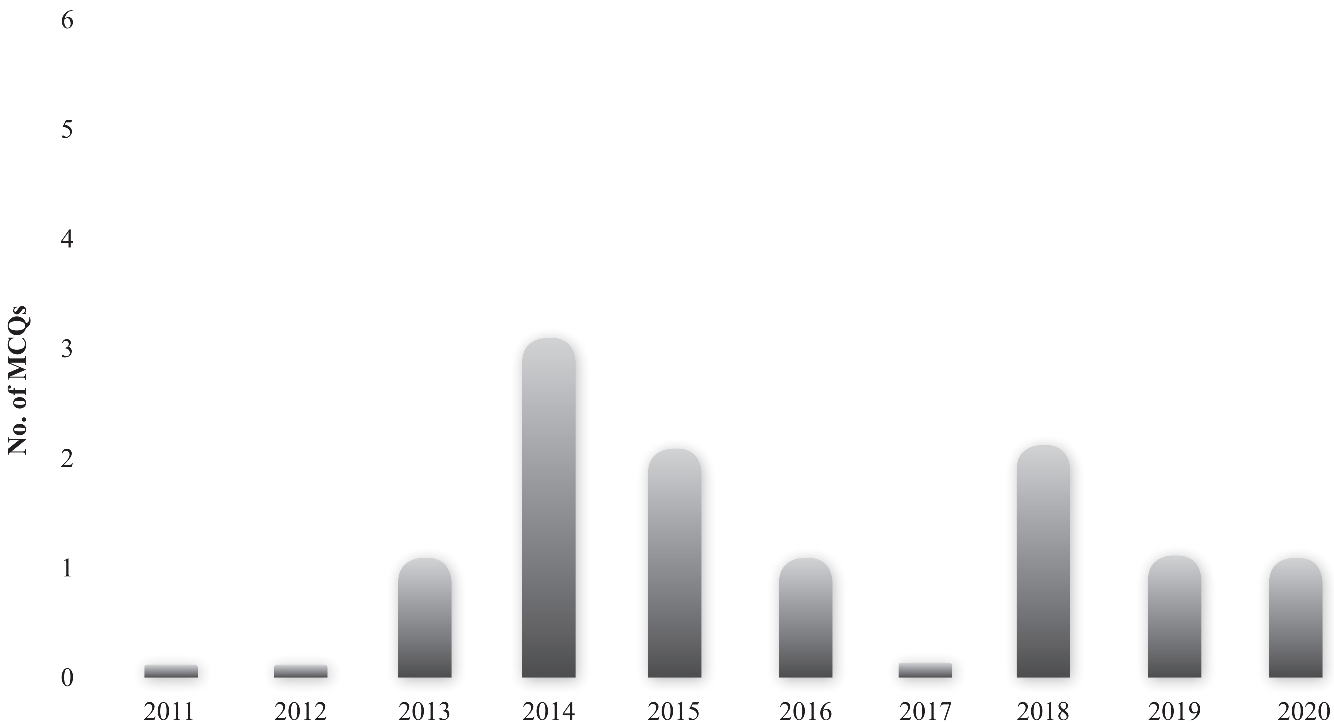
Past Year NEET Trend



Investigation Report

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

Perfect Practice Plan

TOPIC-WISE MCQs	NCERT BASED QUESTIONS	MULTI-CONCEPT QUESTIONS	NEET PAST 10 YEAR QUESTIONS	TOTAL MCQs
110	40	24	18	192

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

- MATTER
  - Chemical Classification
    - Pure substance
      - Element
        - Metal
        - Non-metal
        - Metalloid
      - Compound
        - Inorganic
        - Organic
    - Mixture
      - Homogeneous
      - Heterogeneous
  - Physical Classification
    - Based on intermolecular forces
      - Solid
      - Liquid
      - Gas

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

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

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 posses 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 hetergeneous 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) Compound : (vii), (viii)
Homogeneous Mixture : (iii), (v) Hetergeneous Mixture : (ii), (vi)

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.

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

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 Sub-stance	n	mole	mol
Luminous Intensity	I_v	candela	cd

The two temperature units are related as:

☺ Kelvin temperature (K) = °C + 273.15

☺ $1\overset{\circ}{A} = 10^{-10}\,\text{m}$

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

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

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

☺ Relative density

☺ **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) = $\dfrac{\text{Molar mass}}{\text{Molar volume}}$

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.

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

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

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. **Exponetial 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:**
   1. 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
   2. 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
   3. 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

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

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

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 $\dfrac{11.2 \times 79.9}{88.8} =$ = 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 respectively. These are in the ratio 20.1 : 10.08 = 2 : 1

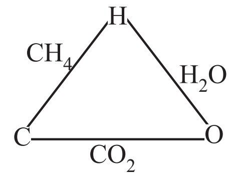
It is a simple whole number ratio. Hence, the law of multiple proportions is established.

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.

$= \dfrac{66.6 \times 40}{33.4} = 79.8$ 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

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

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

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

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

$$\text{Mass of 1 amu} = \frac{12}{6.022 \times 10^{23}} \times \frac{1}{12}$$

$$= \mathbf{1.}67 \times 10^{-24} \text{ g}$$

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 $= \dfrac{\text{Atomic mass}}{\text{Valency}}$
- E.M. of an acid $= \dfrac{\text{Molecular mass}}{\text{Basicity}}$
- E.M. of a base $= \dfrac{\text{Molecular mass}}{\text{Acidity}}$

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

moles of $H_2SO_4 = \dfrac{294}{98} = 3$ 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. $\text{Mass \% of N in } NH_3 = \dfrac{\textit{Mass of N in 1 mol } NH_3}{\textit{Mass of 1 mol of } NH_3} \times 100$

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

$\text{Mass \% of H in } NH_3 = \dfrac{\textit{Mass of H in 1 mol } NH_3}{\textit{Mass of 1 mole of } NH_3} \times 100$

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

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) × n;

Where n is a simple whole number.

Molecular formula = n × Empirical formula

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

— KEY NOTE —

$$n = \frac{Molecular\ weight}{Emprical\ weight}$$

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 – 1.1615 = 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 = $\frac{126}{18} = 7$

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 either 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 $\frac{Mass\ of\ KClO_3}{Mass\ of\ KCl} = \frac{2\times122.5}{2\times74.5}$

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

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

$$\frac{Mass\ of\ KClO_3}{volume\ of\ O_2\ at\ STP} = \frac{2\times 122.5}{3\times 22.4\ lt} \quad ...(i)$$

$$\frac{Mass\ of\ KCl}{volume\ of\ O_2\ at\ STP} = \frac{2\times 74.5}{3\times 22.4\ lt} \quad ...(ii)$$

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

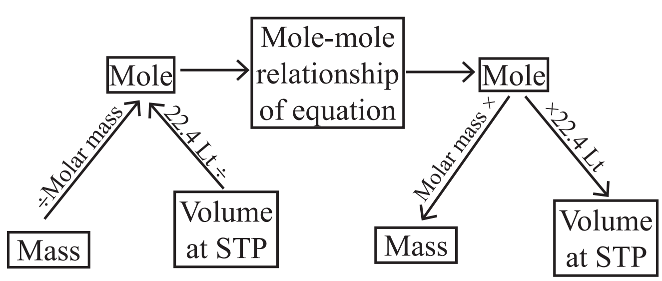
$a\,A + b\,B \rightarrow c\,C + d\,D$

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

$$\frac{mass\ of\ KClO_3}{mass\ of\ KCl} = \frac{2\times 122.5}{2\times 74.5} \Rightarrow \frac{367.5}{W} = \frac{122.5}{74.5}$$

W = 3 × 74.5 = 223.5 g

$$\frac{Mass\ of\ KClO_3}{Mass\ of\ O_2} = \frac{2\times 122.5}{3\times 32} \Rightarrow \frac{367.5}{W} = \frac{2\times 122.5}{3\times 32}$$

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. $\frac{mass\ of\ KClO_3}{volume\ of\ O_2\ at\ STP} = \frac{2\times 122.5}{3\times 22.4\ lt} \Rightarrow \frac{367.5}{V} = \frac{2\times 122.5}{3\times 22.4\ lt}$

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.

Step: III Now if we find limiting reagent then our focus should be on limiting reagent.

Train Your Brain

Q. 6 moles of Na_2O_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

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

$\therefore$ **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.

Q. If 0.5 mol of $BaCl_2$ is mixed with 0.2 mole of Na_3PO_4, the maximum amount of $Ba_3(PO_4)_2$ that can be formed is :

Ans. 0.10 mol

	$3BaCl_2$ +	$2Na_3PO_4$ $\rightarrow$	$6NaCl$ +	$Ba_3(PO_4)_2$
molar ratio	3	2	6	1
initial moles	0.5	0.2	0	0

Limiting reagent is Na_3PO_4 hence it would be consumed, and the yield would be decided by it initial moles.

2 moles of Na_3PO_4 give 1 mole of $Ba_3(PO_4)_2$,

$\therefore$ 0.2 moles of Na_3PO_4 would give 0.1 mole of of $Ba_3(PO_4)_2$

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

1. Mass percent

It is obtained by using the following relation:

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

2. Mole Fraction

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

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.

3. 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 = $\frac{2}{10} = 0.2\ M$

4. 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

∴ 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})

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

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

ABOUT PHYSICS WALLAH

Alakh Pandey is one of the most renowned faculty in NEET & JEE domain's Physics. On his YouTube channel, Physics Wallah, he teaches the Science courses of 11th and 12th standard to the students aiming to appear for the engineering and medical entrance exams.

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