Std-XI science

Unit I:

Some Basic Concepts of Chemistry

Vijaykumar N. Nazare

Grade I Teacher in Chemistry (Senior Scale)

vnn001@ chowgules.ac.in

1.1 IMPORTANCE OF CHEMISTRY

- Chemistry is the branch of science
- Chemistry plays a central role in science
- -intertwined with other branches of science etc.
- -plays an important role in daily life.

Chemistry in everyday life

Source:-Google images-Chemistry in everyday life

1.2 NATURE OF MATTER

- Matter
- Solids Liquids & Gases

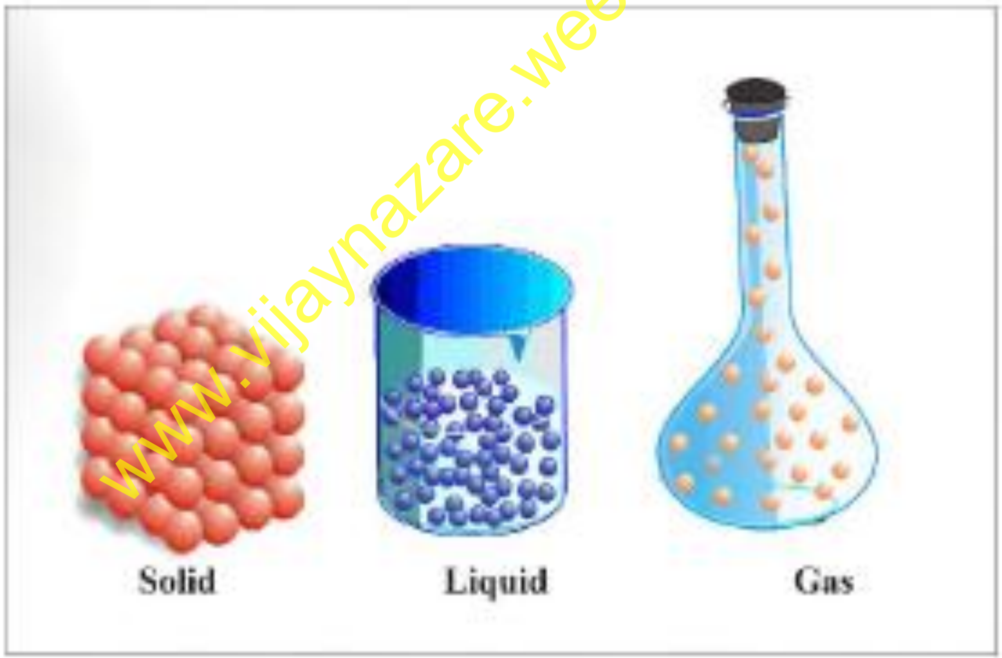


Fig. 1.1 Arrangement of particles in solid, liquid and gaseous state

Classification of matter

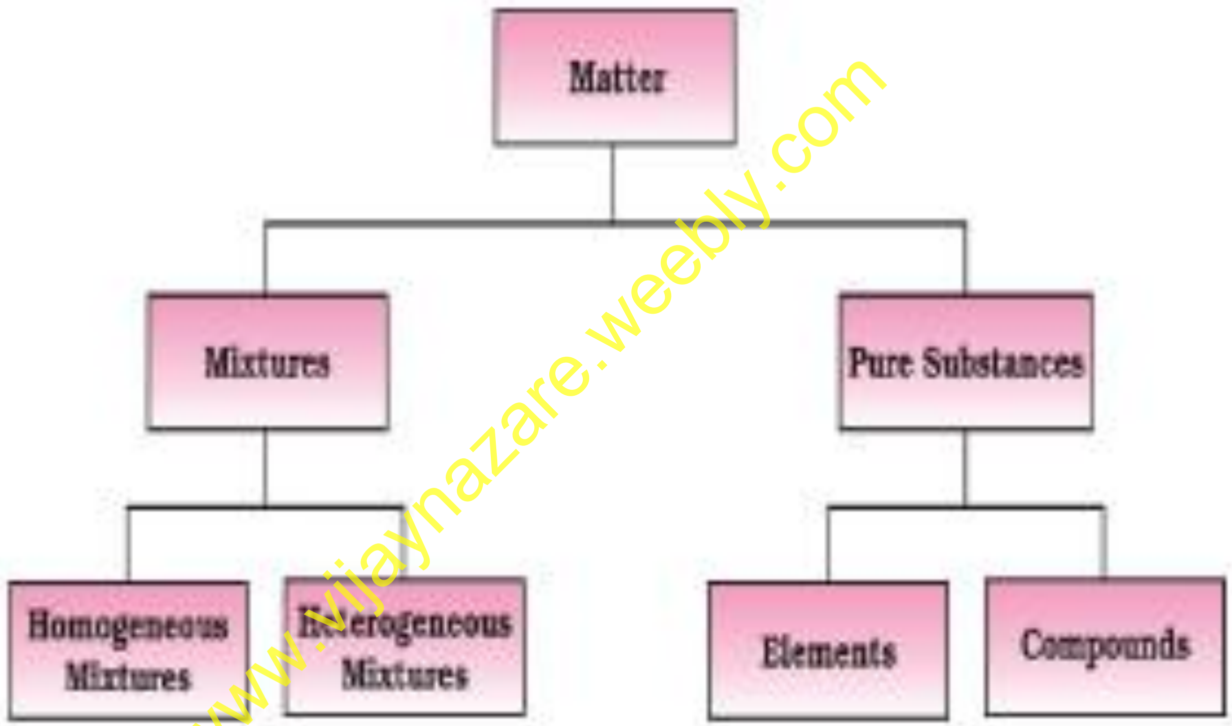


Fig. 1.2 Classification of matter

Law of Conservation of Mass

It states that *matter can neither be created nor destroyed*.

Law of Definite Proportions

- This law was given by, a French chemist, Joseph Proust. He stated *that a given compound always contains exactly the same proportion of elements by weight*.

Joseph Proust
(1754—1826)

Law of Multiple Proportions

- This law was proposed by Dalton in 1803. 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 small whole numbers.*

Gay Lussac's Law of Gaseous Volumes

- This law was given by Gay Lussac in 1808. 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.*

Gay Lussac's Law of Gaseous Volumes

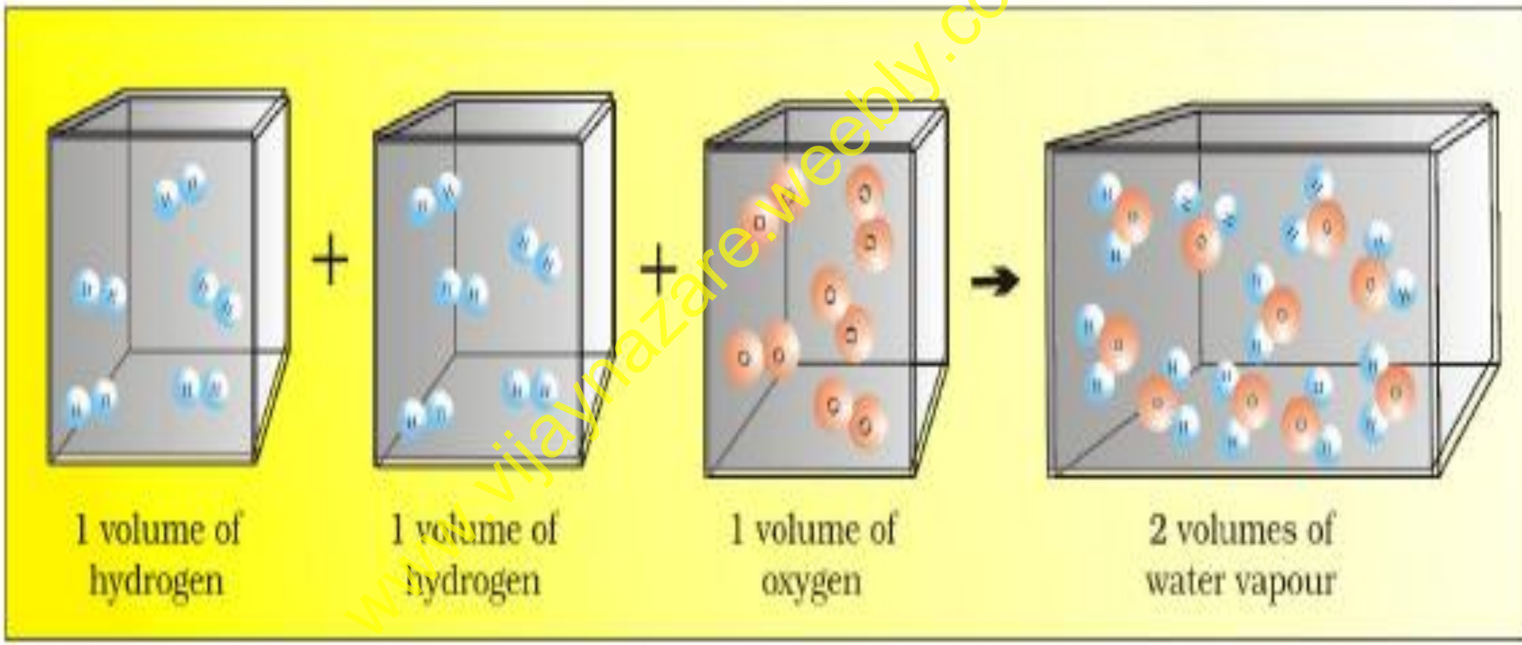


Fig. 1.9 Two volumes of hydrogen react with One volume of oxygen to give Two volumes of water vapour

Avogadro Law

In 1811, Avogadro proposed that *equal volumes of gases at the same temperature and pressure should contain equal number of molecules*

DALTON'S ATOMIC THEORY

John Dalton
(1776—1884)

Dalton's Atomic Theory (postulates)

1) **All elements are composed of tiny indivisible particles called atoms**
2) **Atoms of the same element are identical. Atoms of any one element are different from those of any other element.**
3) **Atoms of different elements combine in simple whole-number ratios to form chemical compounds**
4) **In chemical reactions, atoms are combined, separated, or rearranged – but never changed into atoms of another element.**

Atomic Mass

1 amu = 1.66056×10^{-24} g

Mass of an atom of hydrogen

$= 1.673610^{-24}$ g

Thus, in terms of amu, the mass

of hydrogen atom $= \dfrac{1.6736 \times 10^{-24}\ \text{g}}{1.66056 \times 10^{-24}\ \text{g}}$

$= 1.0078$ amu

$= 1.0080$ amu

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

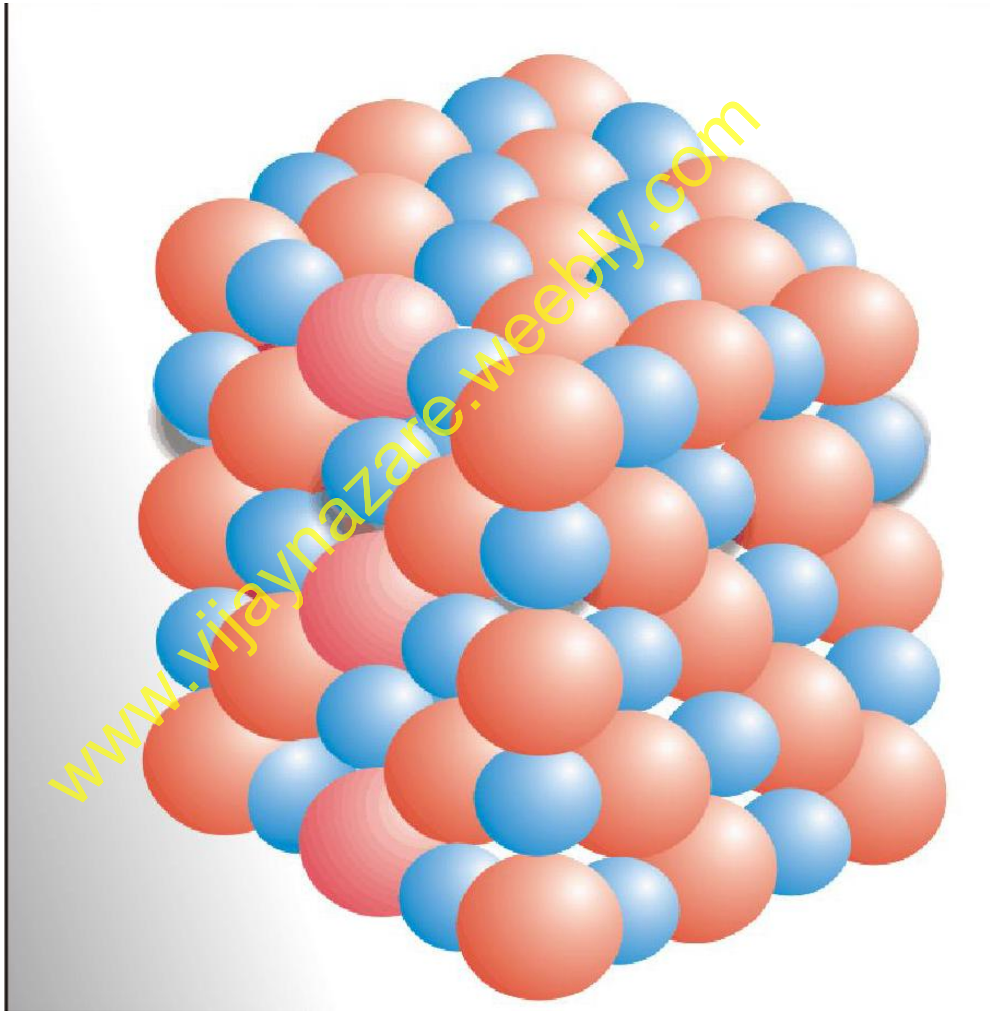
Average Atomic Mass

Many naturally occurring elements exist as more than one isotope. When we take into account the existence of these isotopes and their relative abundance (per cent occurrence), the average atomic mass of that element can be computed

Molecular Mass

- Molecular mass 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



formula mass of NaCl

- It may be noted that in sodium chloride, one Na+ is surrounded by six Cl–and *vice-versa.*
- The formula such as NaCl is used to calculate the **formula mass instead of** molecular mass as in the solid state sodium chloride does not exist as a single entity.
- Thus, formula mass of sodium chloride =
- atomic mass of sodium + atomic mass of Chlorine = 23.0 u + 35.5 u = 58.5 u

One mole

One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the 12C isotope.

Knowing that one mole of carbon weighs 12 g, the number of atoms in it is equal to :

One mole

.

$$\frac{12\,g/mol\ {}^{12}C}{1.992648 \times 10^{-23}\,g/{}^{12}C\ atom}$$

$=6.0221367 \times 10^{23}$ atoms/mol

One mole =6.022 x 10^{23} Particles

- This number of entities in 1 mol is so important that it is given a separate name and symbol.
- It is known as '**Avogadro constant**',
- denoted by N_A in honour of Amedeo Avogadro.
- To really appreciate largeness of this number, let us write it with all the zeroes without using any powers of ten.
  - 602213670000000000000000

Examples

- 1 mol of hydrogen atoms = 6.022×10^{23} atoms
- 1 mol of Nitrogen atoms = 6.022×10^{23} atoms
- 1 mol of Diydrogen Molecules = 6.022×10^{23} Diydrogen Molecules

- 1 mol of water molecules = **6.022×10^{23}** water molecules
- 1 mol of ammonia molecules **= 6.022×10^{23}** ammonia molecules

- 1 mol of sodium chloride **= 6.022×10^{23}** formula units of sodium chloride

Molar Mass=Mass of 1 mole particles

- 6.022×10^{23} Oxygen atoms=16 grams
- 1 mole of Oxygen atoms= 16 grams
- 6.022×10^{23} Oxygen Molecules
- =32grams
- 6.022×10^{23} Water Molecules
- =36grams
- 6.022×10^{23} Ammonia Molecules
- =17grams

One mole of various substances

PERCENTAGE COMPOSITION

Mass % of an element =

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

- Examples
- CO_2
- HNO_3
- CH_4
- C_2H_6

Empirical Formula for Molecular Formula

- An **empirical formula represents the simplest** whole number ratio of various atoms present in a compound

- whereas the **molecular formula shows the exact number of different** types of atoms present in a molecule of a compound.

Empirical Formula for Molecular Formula

1. A compound contains 43.4 % Sodium, 11.3 % carbon and 45.3% Oxygen. What is its empirical formula ?

Element	symbol	%	At mass	Relative No of moles	Simple Ratio	Simple Whole No Ratio
Sodium	Na	43.4	23			
Carbon	C	11.3	12			
Oxygen	O	45.3	16			

Empirical Formula for Molecular Formula

1. A compound contains 43.4 % Sodium, 11.3 % carbon and 45.3% Oxygen. What is its empirical formula ?

Element	symbol	%	At mass	Relative No of moles	Simple Ratio	Simple Whole No Ratio
Sodium	Na	43.4	23	43.4/23=1.88	1.88/0.94	2
Carbon	C	11.3	12	11.3/12=0.94	0.94/0.94	1
Oxygen	O	45.3	16	45.3/16=2.83	2.83 /0.94	3

Empirical Formula for Molecular Formula

- A compound contains 80 % Carbon and 20% Hydrogen. If its Molecular mass is 30. calculate its empirical formula & Molecular formula ?

Element	symbol	%	At mass	Relative No of moles	Simple Ratio	Simple Whole No Ratio
Carbon	C	80	12			
Hydrogen	H	20	1			

empirical formula =

Empirical Formula for Molecular Formula

- A compound contains 80 % Carbon and 20% Hydrogen. If its Molecular mass is 30. calculate its empirical formula & Molecular formula ?

Element	symbol	%	At mass	Relative No of moles	Simple Ratio	Simple Whole No Ratio
Carbon	C	80	12	80/12=6.66	6.66/6.66	1
Hydrogen	H	20	1	20/1=20	20 /6.66	3

empirical formula $=CH_3$

Calculate Molecular Formula

Molecular formula = empirical formula x n

Where n= 1,2,3……..

n= Molecular formula / empirical formula

n= Molecular formula mass / empirical formula mass

Element	symbol	%	At mass	Relative No of moles	Simple Ratio	Simple Whole No Ratio
Carbon	C	80	12	80/12=6.66	6.66/6.66	1
Hydrogen	H	20	16	20/1=20	20 /6.66	3

n= Molecular formula mass / empirical formula mass

- n= Molecular formula mass / empirical formula mass
- Molecular formula =n x empirical formula

Stoichiometry

- The word 'stoichiometry' is derived from two Greek words –
- *stoicheion* *(meaning element)* and *metron* *(meaning measure).*
- Stoichiometry, thus, deals with the calculation of masses (sometimes volumes also) of the reactants and the products involved in a chemical reaction.

Information available from the *balanced chemical equation*

- Let us consider the combustion of methane.
- A balanced equation for this reaction is as given below

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

Information available from the *balanced chemical equation*

- Methane and dioxygen are called *reactants and carbon dioxide and water are* called *products.*
- *All the reactants and* the products are gases in the above reaction and this has been indicated by letter (g) in the brackets next to its formula.
- Similarly, in the case of solids and liquids, (s) and (l) are written respectively

Information available from the *balanced chemical equation*

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2\,H_2O\,(g)$$

- The coefficients 2 for O_2 and H_2O are called stoichiometric coefficients.
- Similarly the coefficient for CH_4 and CO_2 is one in each case.
- They represent the number of molecules (and moles as well) taking part in the reaction or formed in the reaction

Thus, according to the above chemical reaction,

- One **mole of CH_4(g) reacts with two moles** of O_2(g) to give one **mole of CO_2(g) and** two **moles of H_2O(g)**

- One **molecule of CH_4(g) reacts with** 2 **molecules of O_2(g) to give one molecule** of CO_2(g) and 2 molecules of H_2O(g)

Thus, according to the above chemical reaction,

- 22.4 L of CH_4(g) reacts with 44.8 L of O_2 (g) to give 22.4 L of CO_2 (g) and 44.8 L of H_2O(g)
- 16 g of CH_4 (g) reacts with 2×32 g of O_2 (g) to give 44 g of CO_2 (g) and 2×18 g of H_2O (g).

Problem

$$CH_4\ (g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$$

- Calculate the amount of water (g) produced by the combustion of 16 g of methane.
- Ans:- 36 g.

Limiting Reagent

- The reactant which is present in the lesser amount gets consumed after sometime and after that no further reaction takes place whatever be the amount of the other reactant present.

- Hence, the reactant which gets consumed, limits the amount of product formed and is, therefore, called the **limiting reagent.**

Reactions in Solutions

- A majority of reactions in the laboratories are carried out 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 per cent or weight per cent (w/w %)
  - 2. Mole fraction
  - 3. Molarity
  - 4. Molality

1. Mass per cent

- It is obtained by using the following relation:

$$\boxed{\text{Mass per cent} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100}$$

1. Mass per cent

Problem

A solution is prepared by adding 4 g of a Sugar to 36 g of water. Calculate the mass per cent of the Sugar.

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

Ans:- Mass % of sugar **=10%**

2. *Mole Fraction*

- It is the ratio of number of moles of a particular component to the total number of moles of the solution.
- If a substance 'A' dissolves in substance 'B' and their number of moles are
- *nA and nB respectively; then the mole fractions* of A and B are given as

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

$$= \frac{n_A}{n_A + n_B}$$

Mole fraction of B

$$= \frac{\text{No. of moles of B}}{\text{No. of moles of solution}}$$

$$= \frac{n_B}{n_A + n_B}$$

3. Molarity

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

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

4. Molality

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

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

- *Molarity of a solution depends upon temperature because volume of a solution is temperature dependent.*

Thank you

- Vijaykumar N. Nazare.
- Grade I Teacher Senior Scale
- vnn001@ chowgules.ac.in