ATOMS & MOLECULES

CONTENTS

Introduction

Matter is called 'padarth' in Hindi. Kanad was one of the first persons to propose that matter (or padarth) is made up of very small particles called 'parmanu'. John Dalton called these particles by the name of atom. The word 'atom' means 'indivisible' something which cannot be divided further. The particles of matter (atoms or parmanu) normally exist in a combined form. This combined form of atoms is now called 'molecules'. All matter is made up of small particles called atoms and molecules. Different kinds of atoms and molecules have different properties due to which different kinds of matter also show different properties.

Laws of chemical combination

The laws of chemical combination played a significant role in the development of Dalton's atomic theory of matter.

There are two important laws of chemical combination.

(A) Law of conservation of mass

- Law of conservation of mass was given by Lavoisier in 1774. According to the law of conservation of mass : Matter is neither created nor destroyed in a chemical reaction. The substances which combine together in a chemical reaction are known as 'reactants' whereas the new substances formed as a result of chemical reaction are called 'products'. The law of conservation of mass means that in a chemical reaction, the total mass of products is equal to the total mass of reactants. There is no change in mass during a chemical reaction.

Ex. Lavoisier showed that when mercuric oxide was heated, it produced free mercury and oxygen. The sum of masses of mercury and oxygen was found to be equal to the mass of mercuric oxide

Mercuric oxide $\rightarrow$ Mercury + Oxygen

100 g 92.6 g 7.4 g

Sol. Mass of the reactant = 100 g

Mass of the products = 92.6 + 7.4 g = 100.0 g

= 100 g

Since the total mass of the products formed is equal to the total mass of the reactants undergoing reaction, so the data is in agreement with low of conservation of mass.

(B) Law of constant proportions

This law was discovered by the french chemists, A. Lavoisier and Joseph Proust. A pure chemical compound always contains same elements combined together in same proportion by mass

Ex. Pure water obtained from different sources such as river, well, spring, sea, etc., always contains hydrogen and oxygen combined together in the ratio 1 : 8 by mass.

Limiting reagent : - In a reaction having more than one reactant we must identify the limiting reagent **"Reagent which is finished early is known as limiting reagent "**.

Example : - In Haber's process to manufacture NH_3,

$N_2 + 3H_2 \rightarrow 2NH_3$

If we take 2 moles of N_2 (56 grams) along with 3 moles of H_2 (6 grams) we can see that only 1 moles of N_2 (28 grams) is sufficient to react with 3 moles of H_2. It means 1 mole N_2 (28 grams) is in excess & H_2 will be finished when reaction will be completed or H_2 is limiting reagent.

➢ Dalton's atomic theory

On the basis of laws of chemical combination John Dalton, an English school teacher in Manchester, proposed that behaviour of matter could be explained using an atomic theory. He published his work about atomic theory in 1808. The main points of Dalton's atomic theory are:

- All the matter is made up of very small particles called "atoms".
- Atoms cannot be divided.
- Atoms can neither be created nor destroyed.
- Atoms are of various kinds. There are as many kinds of atoms as are elements
- All the atoms of a given element are identical in every respect, having the same mass, size and chemical properties.
- Atoms of different elements differ in mass, size and chemical properties.
- Chemical combination between two (or more) elements consists in the joining together of atoms of these elements to form molecules of compounds.
- The "number" and "kind" of atoms in a given compound is fixed.
- During chemical combination, atoms of different elements combine in small whole numbers to form compounds.

Atoms of the same elements can combine in more than one ratio to form more than one compounds.

❖ Drawbacks of Dalton's atomic theory

Some of the drawbacks of the Dalton's atomic theory of matter are given below :

- One of the major drawbacks of Dalton's atomic theory of matter is that atoms were thought to be indivisible (which cannot be divided). We now know that under special circumstances, atoms can be further divided into still smaller particles called electrons, protons and neutrons. So, atoms are themselves made up of three particles : electrons, protons and neutrons.
- Dalton's atomic theory says that all the atoms of an element have exactly the same mass. it is, however, now known that atoms of the same element can have slightly different masses.
- Dalton's atomic theory said that atoms of different elements have different masses. it is, however, now known that even atoms of different elements can have the same mass.
- It failed to explain how atoms of different elements differ from each other, i.e., it did not tell anything about internal structure of the atom.
- It could not explain how and why atoms of different elements combine with each other to form compound atoms or molecules.
- It failed to explain the nature of forces that hold together different atoms in a molecule.
- It did not make any distinction between ultimate particle of an element that takes part in reactions (atom) and ultimate particle that has independent existence (molecule).

➢ Atoms

- All the matter is made up of atoms. An atom is the smallest particle of an element that can take part in a chemical reaction. Atoms of most of the elements are very reactive and do not exist in the free state. They exist in combination with the atoms of the same element of another element.
- Atoms are very, very small in size. The size of an atom is indicated by its radius which is called 'atomic radius'. Atomic radius is measured in 'nanometres'. The symbol of a nanometre is nm.

$$1 \text{ nanometre} = \frac{1}{10^9} \text{ metre}$$

or $$1 \text{ nm} = \frac{1}{10^9} \text{ m}$$

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

Hydrogen atom is the smallest atom of all. They cannot be viewed by simple optical microscopes. However, through modern techniques such as scanning tunneling microscope it is possible to produce magnified images of surfaces of elements showing atoms.

Symbols of elements

In order to represent the elements, instead of using full lengthy names, scientists use abbreviated names. These abbreviated names of the elements are known as symbols. Thus, symbol may be defined as the abbreviation used for the name of an elements.

Dalton's symbols of elements

Dalton was the first scientist to use the symbols to represent elements in a short way.

Element	Dalton's symbol
Hydrogen	⊙
Carbon	●
Oxygen	○
Phosphorus	
Sulphur	⊕
Platinum	Ⓟ
Iron	Ⓘ
Copper	Ⓒ
Silver	Ⓢ
Gold	Ⓖ
Lead	Ⓛ
Mercury	

Dalton's symbols for elements were difficult to draw and inconvenient to use. So, Dalton's symbols are only of historical importance. They are not used at all.

Modern symbols of elements

IUPAC (International Union of Pure and Applied Chemistry) approves names of elements. The symbols of elements are generally either the first letter or the first two letters or the first and the third letters of the name of the elements. **The symbols of the following elements are the first letter of the name of that elements.**

Element	Symbol
Hydrogen	H
Carbon	C
Nitrogen	N
Oxygen	O
Fluorine	F
Phosphorus	P
Sulphur	S
Iodine	I

Some symbols derived from the first two letters of the names of the elements.

Element	Symbol
Aluminium	Al
Barium	Ba
Lithium	Li
Beryllium	Be
Neon	Ne
Silicon	Si
Argon	Ar
Calcium	Ca
Nickel	Ni

Some symbols derived from the first and the third letters of the names of the elements

Element	Symbol
Arsenic	As
Magnesium	Mg
Chlorine	Cl
Chromium	Cr
Manganese	Mn
Zinc	Zn
Rubidium	Rb

Some symbols derived from the latin names of the elements are given below

Element	Latin name	Symbol
Iron	Ferrum	Fe
Gold	Aurum	Au
Copper	Cuprum	Cu
Potassium	Kalium	K
Sodium	Natrium	Na
Silver	Argentum	Ag
Mercury	Hydragyrum	Hg
Tin	Stannum	Sn
Lead	Plumbum	Pb
Antimony	Stibium	Sb

It is important to note that the first letter of every chemical symbol is capital letter but, if the symbol consists of two letters, the second letter is not capital letter.

Ex. Symbol for aluminum is Al and not AL

Symbol for lead is Pb and not PB

❖ Significance of the symbol of an element

- Symbol represents name of the element.
- Symbol represents one atom of the element.
- Symbol also represents one mole of atoms of the element. That is, symbol also represents 6.022×10^{23} atoms of the elements
- Symbol represents a definite mass of the element (equal to atomic mass expressed in grams)

➤ Atomic mass

- Atoms are extremely small ; the heaviest atoms have masses of about 10^{-22} g. Even an ultramicrobalance cannot measure the mass of a single atom. However, relative masses of atoms of different elements can be determined. At first, the mass of the lightest atom, hydrogen.
- In 1961, International Union of chemists selected the most stable isotope of carbon (C - 12 isotope) as standard for comparison the atomic masses of various elements. Atomic mass of an element tells us the number of times an atom of the element is heavier than of the mass of an atom of carbon - 12.
- Atomic mass of an element may be defined as the average relative mass of an atom of the element as compared with mass of an atom of carbon (C - 12 isotope) taken as 12 amu.

Atomic Mass

$$= \frac{\text{Mass of 1 atom of the element}}{\frac{1}{12}\text{ of the mass of an atom of carbon} - 12}$$

Ex. The atomic mass of magnesium is 24 u which indicates that one atom of magnesium is 24 times heavier than $\frac{1}{12}$ of a carbon 12 atom.

Atomic Mass of Some Elements

	Element	Symbol	Atomic
1	Hydrogen	H	1 u
2	Carbon	C	12 u
3	Nitrogen	N	14 u
4	Oxygen	O	16 u
5	Sodium	Na	23 u
6	Magnesium	Mg	24 u
7	Aluminium	Al	27 u
8	Phosphorus	P	31 u
9	Sulphur	S	32 u
10	Chlorine	Cl	35.5 u
11	Potassium	K	39 u
12	Calcium	Ca	40 u
13	Iron	Fe	56 u
14	Copper	Cu	63.5 u

❖ Gram atomic mass

Gram atomic mass of an element is defined as that much quantity of the element whose mass expressed in grams is numerically equal to its atomic mass. To find gram atomic mass we keep the numerical value the some as the atomic mass, but simply change the units from u to g. for example, atomic mass of aluminium is 27 u. Its gram atomic mass is 27 g.

- **Gram atomic mass of Isotopes :**

$$\frac{M_1X_1 + M_2X_2}{X_1 + X_2}$$

M_1 & M_2 are relative masses of isotopes and X_1 & X_2 are relative % content

Ex. Chlorine contains two types of atoms having relative masses 35 and 37 and their relative abundance is 3 : 1. In such cases the atomic mass of the element is the average of relative masses of different isotopes of the element.

$$\text{Atomic mass of chlorine} = \frac{35\times3+37\times1}{4} = 35.5$$

Atoms usually exist in two ways.

(i) Molecules (ii) Ions

Molecules

- A combination of atoms is called a molecule. The forces which hold the atoms together in a molecule are called covalent bonds.
- A molecule is the smallest particle of a substance which has the properties of that substance and can exist in the free state.

There are two types of molecules.

Molecules of elements

The molecule of an element contains two (or more) similar atoms chemically combined together.

Ex. A molecule of hydrogen element contains 2 hydrogen atoms combined together, and it is written as H_2.

Molecules of compounds

The molecule of a compound contains two (or more) different types of atoms chemically combined together

Ex. Hydrogen chloride is a compound. The molecule of hydrogen chloride (HCl) contains two different types of atom. Hydrogen(H) and chlorine atom (Cl)

Molecules of some compounds

Compound	Combining elements	Formula	Ratio by mass
Water	Hydrogen and oxygen	H_2O	1:8
Ammonia	Nitrogen and Hydrogen	NH_3	14:3
Carbon dioxide	Carbon and Oxygen	CO_2	3:8

Atomicity

The number of atoms present in one molecule of an element is called its atomicity.

Ex. Noble gases (helium, neon, argon, krypton, etc.) have one atom each in their molecules such as He, Ne, Ar and Kr. So, the atomicity of noble gases is 1.

Ex. Hydrogen (H_2), nitrogen (N_2), oxygen (O_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2), all have 2 atoms each in their molecules. So, the atomicity of hydrogen, nitrogen, oxygen, chlorine, bromine and iodine is 2 each.

Atomicity of some common elements.

Type of element	Name	Symbol	Atomicity
Non-metal	Helium	He	Monoatomi
	Argon	Ar	Monoatomi
	Neon	Ne	Monoatomi
	Hydrogen	H_2	Diatomic
	Chlorine	Cl_2	Diatomic
	Nitrogen	N_2	Diatomic
	Oxygen	O_2	Diatomic
	Phosphor	P_4	Tetratomic
	Sulphur	S_8	Polyatomic
Metals	Sodium	Na	Monoatomi
	Iron	Fe	Monoatomi
	Aluminiu	Al	Monoatomi
	Copper	Cu	Monoatomi

Ions

An ion is a positively or negatively charged atom (or group of atoms). An ion is formed by the loss or gain of electrons by an atom, so it contains an unequal number of electrons and protons.

Ex. Sodium ion Na^+, magnesium ion Mg^{2+}, chloride ion Cl^-, and oxide ion O^{2-}.

- There are two types of ions : cations and anions.

Cation

A positively charged ion is known as cation. A cation is formed by the loss of one or more electrons by an atom.

Ex. Sodium atom loses 1 electron to form a sodium ion, Na^+, which is cation :

$$\underset{\text{Sodium}}{Na} \xrightarrow{-1\text{ electron}} \underset{\text{(A cation)}}{\underset{\text{Sodium ion}}{Na^+}}$$

The ions of all the metal elements are cations.

Na $\xrightarrow{-1\text{ electron}}$	Na^-
Sodium atom	Sodium ion
Protons = 11 (+ charge)	Protons = 11 (+charge)
Electrons = 11 (– charge)	Electrons = 10 (–charge)
Overall charge = 0	Overall charge = 1+

◈ Anion

A negatively charged ion is known as anion. An anion is formed by the gain of one or more electrons by an atom.

Ex. A chlorine atom gains 1 electron to form a chloride ion, Cl^-, which is an anion.

$$\underset{\text{Chlorine atom}}{Cl} \xrightarrow{+1\text{ electron}} \underset{\text{(An Anion)}}{\underset{\text{Chlorine ion}}{Cl^-}}$$

An anion contains more electrons than a normal atom. A normal atom (or a neutral atom) contains an equal number of protons and electrons. Now, since an anion is formed by the addition of one or more electrons to an atom, therefore, an anion contains more electrons than protons. The ions of all the non metal elements are anions

Cl $\xrightarrow{+1\text{ electron}}$	Cl^-
Chlorine atom	Chlorine ion
Protons = 17 (+ charge)	Protons = 17 (+charge)
Electrons = 17 (– charge)	Electrons = 18 (–charge)
Overall charge = 0	Overall charge = 1–

◈ Simple ions

Those ions which are formed from single atoms are called simple ions.

Ex. Sodium ion, Na^+, is a simple ion because it is formed from a single sodium atom, Na.

◈ Compound ions

Those ions which are formed from groups of joined atoms are called compound ions

Ex. Ammonium ion NH_4^+, is a compound ion which is made up of two types of atoms joined together, nitrogen and hydrogen.

◈ Ionic compounds

The compounds which are made up of ions are known as ionic compounds. In an ionic compound, the positively charged ions (cations) and negatively charged ions (anions) are held together by the strong electrostatic forces of attraction. The forces which hold the ions together in an ionic compound are known as ionic bonds or electrovalent bonds. Since an ionic compound consists of an equal number of positive ions and negative ions, so the overall charge on an ionic compound is zero.

Ex. Sodium chloride (NaCl) is an ionic compound which is made up of equal number of positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-).

Some ionic compound

S.No.	Name	Formula	Ions present
1	Sodium chloride	$NaCl$	Na^+ and Cl^-
2	Potassium chloride	KCl	K^+ and Cl^-
3	Ammonium chloride	NH_4Cl	NH_4^+ and Cl^-
4	Magnesium chloride	$MgCl_2$	Mg^{2+} and Cl^-
5	Calcium chloride	$CaCl_2$	Ca^{2+} and Cl^-
6	Magnesium oxide	MgO	Mg^{2+} and O^{2-}
7	Calcium oxide	CaO	Ca^{2+} and O^{2-}
8	Aluminium oxide	Al_2O_3	Al^{3+} and O^{2-}
9	Sodium hydroxide	$NaOH$	Na^+ and OH^-
10	Copper sulphate	$CuSO_4$	Cu^{2+} and SO_4^{2-}
11	Calcium nitrate	$Ca(NO_3)_2$	Ca^{2+} and NO_3^-

➤ Valency

The combining capacity of an element is known as its valency.

Formulae and valencies of common ions

Electropositive Ions			
Monovalent		**Bivalent**	
Name	**Formula**	**Name**	**Formula**
Potassium	K^{+}	Barium	Ba^{2+}
Sodium	Na^{+}	Calcium	Ca^{2+}
Copper [I]	Cu^{+}	Magnesium	Mg^{2+}
		Manganese [II]	Mn^{2+}
		Zinc	Zn^{2+}
Silver	Ag^{+}	Iron [II]	Fe^{2+}
Ammonium	NH_4^{+}		
Hydrogen	H^{+}	Nickel	Ni^{2+}
		Cobalt	Co^{2+}
		Tin [II]	Sn^{2+}
		Cadmium	Cd^{2+}
		Lead [II]	Pb^{2+}
		Copper [II]	Cu^{2+}
Trivalent		**Tetravalent**	
Name	**Formula**	**Name**	**Formula**
Aluminium	Al^{3+}	Manganese [II]	Mn^{4+}
Chromium	Cr^{3+}	Tin [IV]	Sn^{4+}
Iron [III]	Fe^{3+}	Lead [IV]	Pb^{4+}
Gold	Au^{3+}	Platinum	Pt^{4+}

Electropositive Ions			
Monovalent		**Bivalent**	
Name	**Formula**	**Name**	**Formula**
Hydroxide	OH^{-}	Carbonate	CO_3^{2-}
Hydride	H^{-}	Chromate	CrO_4^{2-}
Fluoride	F^{-}	Dichromate	$Cr_2O_7^{2-}$
Chloride	Cl^{-}	Manganate	MnO_4^{2-}
Bromide	Br^{-}	Tetrathionate	$S_4O_6^{2-}$
Iodide	I^{-}	Sulphide	S^{2-}
Bicarbonate	HCO_3^{-}	Sulphite	SO_3^{2-}
Bisulphite	HSO_3^{-}	Sulphate	SO_4^{2-}
Bisulphate	HSO_4^{-}	Oxide	O^{2-}
Chlorate	ClO_3^{-}	Zincate	ZnO_2^{2-}
Hypochlorite	ClO^{-}	Thiosulphate	$S_2O_3^{2-}$
Nitrite	NO_2^{-}		
Nitrate	NO_3^{-}		
Permanganate	MnO_4^{-}		

Trivalent		Tetravalent	
Name	**Formula**	**Name**	**Formula**
Nitride	N^{3-}	Ferrocyanide	$Fe(CN)_6^{4-}$
Phosphate	PO_4^{3-}		
Ferricyanide	$Fe(CN)_6^{3-}$		
Phosphide	P^{3-}		Pt^{4+}

◈ Variable valency

There are certain elements that exhibit more than one valencies in their ions (or compounds).

Ex. Iron can exist as Fe^{2+} or Fe^{3+} in its compounds. In such cases the name of the ion with lower valency ends with a suffix 'ous' while that with higher valency ends with a suffix 'ic'.

Some basic ions exhibiting variable valency

Name of	Formula
Cuprous	Cu^{+}
Cupric	Cu^{2+}
Mercurous	Hg_2^{2+}
Mercuric	Hg^{2+}
Ferrous	Fe^{2+}
Ferric	Fe^{3+}
Plumbous	Pb^{2+}
plumbic	Pb^{4+}
Stannous	Sn^{2+}
Stannic	Sn^{4+}
Aurous	Au^{+}
Auric	Au^{3+}

➤ Chemical formulae

- A compound is represented in the abbreviated form by chemical formula.
- The chemical formula of a compound represents the composition of a molecule of the compound in terms of the symbols of the elements present in it.

Ex. Water is a compound made up of 2 atoms of hydrogen element and 1 atom of oxygen element, so the formula of water is written as H_2O. In the formula H_2O, the subscript 2 indicates 2 atoms of hydrogen. In the formula of water, oxygen O is written without a subscript and it indicates 1 atoms of oxygen.

◈ Formulae of elements

- The chemical formula of an element is a statement of the composition of its molecule in which symbol tells us the element and the subscript tells us how many atoms are present in one molecule. One molecule of hydrogen element contains two atoms of hydrogen, therefore, the formula of hydrogen is H_2.

Ex. The formula H_2 indicates that one molecule of hydrogen element contains 2 atoms of hydrogen. 2 H represents 2 separate atoms of hydrogen; H_2 represents 1 molecule of hydrogen and $2H_2$ represents 2 molecules of hydrogen.

◈ Formulae of compounds

- The chemical formula of a compound is a statement of its composition in which the chemical symbols tell us which elements are present and the subscripts tell us how many atoms of each element are present in one molecule of the compound.

Ex. Water is a compound whose molecule contains 2 atoms of hydrogen and 1 atom of oxygen. So, the formula of water is H_2O.

◈ Rules for writing a chemical formula

- We first write the symbols of the elements which form the compound
- Below the symbol of each element, we write down its valency.
- Finally, we cross-over the valencies of the combining atoms. That is, with first atom we write the valency of second atom (as a subscript); and with the second atom we write the valency of first atom (as subscript).

Ex.

Element C ╳ C

Formula C_1Cl_4 or CCl_4

Valency 4 ╳ 1

Ex.

Element H ╳ C

Valency 1 ╳ 1

Formula HCl

Ex.

Element H O

Valency 1 2

Formula H_2O

Ex.

Element H S

Valency 1 2

Formula H_2S

Ex.

Element N H

Valency 3 1

Formula NH_3

◈ Writing the formula of a compound

Step-1 :

Write the symbols of formulae of the ions of the compound side by side with positive ion on the left hand side and negative ion on right hand side.

Step-2 :

Enclose the polyatomic ion in a bracket.

Step-3 :

Write the valency of each ion below its symbol

Step-4 :

Reduce the valency numerals to a simple ratio by dividing with a common factor, if any.

Step-5 :

Cross the valencies. Do not write the charges positive or negative of the ions.

Ex. Formula of barium nitrate.

Step-1 : Writing the formula of the ions :

Ba^{2+} NO_3^-

Step-2 : Ba^{2+} $(NO_3)^-$

Step-3 : Ba^{2+} $(NO_3)^-$

2 1

Step-4 : Not applicable, because ratio is already simple

Step-5 :

Ba^{2+} $(NO_3)^-$

2 1

Thus, the formula of barium nitrate is $Ba(NO_3)_2$

◈ Significance of the formula of a substance

- Formula represents the name of the substance
- Formula represents one molecule of the substance
- Formula also represents one mole of molecules of the substance. That is, formula also represents 6.022×10^{23} molecules of the substance.
- Formula gives the name of all the elements present in the molecule.
- Formula gives the number of atoms of each element present in one molecule.
- Formula represents a definite mass of the substance

➤ Molecular mass/Formula mass

Molecular mass expresses as to how many times a molecule of a substance is heavier than $\frac{1}{12}$ th of the mass of an atom of carbon (carbon-12). Thus, Molecular mass =

$$\frac{\text{Mass of a molecule}}{\frac{1}{12}\text{th mass of a carbon atom (carbon} - 12)}$$

Molecular mass = 2 × vapour density

Ex. A molecule of water is 18 times heavier than $\frac{1}{12}$ th of the mass of carbon atom. Therefore, the molecular mass of water is 18 u.

◈ Calculation of molecular mass from atomic masses

The molecules are made up of two or more atoms of different elements. Therefore, the molecular mass may be calculated as the sum of the atomic masses of all the atoms in a molecule of that substance.

Ex. Ammonia has the formula, NH_3. it consists of one atom of N and three atoms of H. The atomic mass of N and H are 14.0 and 1 respectively. Therefore, the molecular mass of NH_3 is

Molecular mass of NH_3 = At. mass of N + 3 × At. mass of H

= 14 + 3 × 1 = 17 u

Ex. Sulphuric acid has the formula H_2SO_4. It consists of two H, one S and four O atoms. The atomic masses of H, S and O are 1,32 and 16 respectively. Therefore, the molecular mass of H_2SO_4 is

Molecular mass of H_2SO_4 =

(2 × at. m. of H) + (1 × at. m. of S) + (4 × at. m. of O)

= (2 × 1) + (1 × 32) + (4 × 16) = 98 u

❖ **Gram molecular mass**

The molecular mass of a substance expressed in gram is called its gram molecular mass.

Ex. Molecular mass of oxygen, O_2 = 32 u

So, gram molecular mass of oxygen, O_2 = 32 grams.

➤ Mole concept

Atoms and molecules are so small in size that they cannot be counted individually. The chemists use the unit mole for counting atoms, molecules or ions. It is represented by n. A mole represents 6.022×10^{23} particles.

Ex. 1 mole of atoms = 6.022×10^{23} atoms.

1 mole of molecules = 6.022×10^{23} molecules

The number of particles present in 1 mole of any substance is fixed i.e. 6.022×10^{23}.

This number is called Avogadro constant or Avogadro number.

it is represented by No.

1 mole of atoms = 6.022×10^{23} atoms = Gram atomic mass or Molar mass of element

$$\text{Number of moles} = \frac{\text{Mass of element}}{\text{Molar mass}}$$

$$n = \frac{m}{M}$$

$$\text{Number of moles} = \frac{\text{Given number of atoms}}{\text{Avogadro number}}$$

$$n = \frac{N}{N_0}$$

No. of moles = n

Given mass = m

Molar mass = M

Given number of particles = N

Avogadro number of particles = N_0

These relations can be interchanged as

Mass of element, m = n × M

or No. of particles of element, N = n × N_0

Similarly,

1 Mole of molecules = 6.022×10^{23} molecules = Gram molecular mass of Molar mass

$$\text{Number of moles} = \frac{\text{Mass of substance}}{\text{Molar mass}}$$

$$n = \frac{m}{M}$$

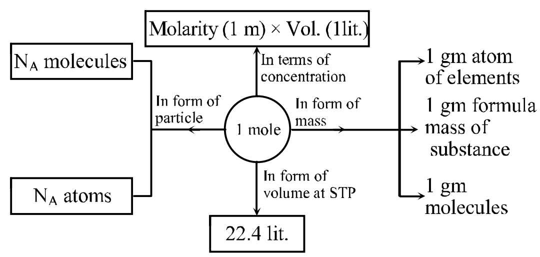
Number of moles

$$= \frac{\text{Given number of molecules}}{\text{Avogadro number}}$$

$$n = \frac{N}{N_0}$$

or m = n × M and N = n × N_0

❖ **Relationship between mole, number of particles and mass and interconversion of one into the other.**



Molarity (M) : - Moles of solute is one litre of solution is known as molarity .

$$M = \frac{\text{Number of moles of solute}}{\text{Volume of solution in litre}}$$

Ex. An ornament of silver contains 20 g of silver. Calculate the moles of silver present (atomic mass of silver = 180 u)

Sol. Moles of silver,

$$n = \frac{m}{M}$$

Mass of silver, m = 20 g,

Molar mass of silver,

M = 108 g

$$\therefore \quad n = \frac{20}{108} = 0.185 \text{ mol.}$$

Ex. How many moles of CO_2 are present in 51.2 g of it ?

Sol. Molecular mass of CO_2 = 12 + 2 + 16 = 44 u

Molar mass of CO_2 (M) = 44 g

Mass of CO_2 (m) = 51.2 g

Moles of CO_2,

$$n = \frac{m}{N} = \frac{51.2}{44} = 1.16 \text{ mol.}$$

Ex. Calculate the mass of

(i) 0.5 moles of N_2 gas

(ii) 0.5 moles of N atoms

Sol. (i) 0.5 moles of N_2 gas

Mass = Molar mass × Number of moles

$m = M \times n$

M = 28 g, n = 0.5

$\therefore$ m = 28 × 0.5 = 14 g

(ii) Mass = Molar mass × Number of moles

$m = M \times n$

n = 0.5 mole, M = 14 g

m = 14 × 0.5 = 7 g

◈ Mass percentage of an element from molecular formula :

The molecular formula of a compound may be defined as the formula which specifies the number of atoms of various elements in the molecule of the compound.

Ex. The molecular formula of glucose is $C_6H_{12}O_6$. This indicates that a molecule of glucose contains six atoms of carbon, twelve atoms of hydrogen and six atoms of oxygen.

The mass percentage of each element is then calculated by the following formula : Mass percentage of element X

$$= \frac{\text{Mass of X in one mole}}{\text{Gram molecular mass}} \times 100.$$

Ex. Calculate the percentage composition (by mass) of formaldehyde (CH_2O).

Sol. Molecular mass of formaldehyde,

$CH_2O = 12 \times 1 + 1 \times 2 + 16 \times 1 = 30$

Mass of one mole of formaldehyde = 30 g

1 Mole of CH_2O contains 1 mole (12 g) of carbon. 2 moles of hydrogen (2 g) and 1 mole of oxygen (16 g)

$$\text{Percentage of carbon} = \frac{12g}{30g} \times 100 = 40.0\%$$

$$\text{Percentage of hydrogen} = \frac{2g}{30g} \times 100 = 6.7\%$$

$$\text{Percentage of oxygen} = \frac{16g}{30g} \times 100 = 53.3\%$$

◈ Empirical formula

The empirical formula of a compound may be defined as the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound.

Ex. The empirical formula of the compound glucose ($C_6H_{12}O_6$), is CH_2O which shows that C, H, and O are present in the simplest ratio of 1 : 2 : 1.

◈ Rules for writing the empirical formula

The empirical formula is determined by the following steps :

- Divide the percentage of each elements by its atomic mass. This gives the relative number of moles of various elements present in the compound.
- Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.
- Multiply the figures, so obtained by a suitable integer, if necessary, in order to obtain whole number ratio.
- Finally write down the symbols of the various elements side by side and put the above numbers as the subscripts to the lower right hand corner of each symbol. This will represent the empirical formula of the compound.

Ex. A substance, on analysis, gave the following composition : Na = 43.4%, C = 11.3%, O = 45.3%. Calculate its empirical formula

[Atomic masses = Na = 23, C = 12, O = 16]

Sol.

Element	Symbol	%	Atomic mass	Relative number of moles	Simple ratio of moles	Simplest whole no. ratio
Sodium	Na	43.4	23	$\frac{43.4}{23} = 1.88$	$\frac{1.88}{0.94} = 2$	2
Carbon	C	11.3	12	$\frac{11.3}{12} = 0.94$	$\frac{0.94}{0.94} = 1$	1
Oxygen	O	45.3	16	$\frac{45.3}{16} = 2.83$	$\frac{2.83}{0.94} = 3$	3

Therefore, the empirical formula is Na_2CO_3

❖ Determination molecular formula :

Molecular formula = Empirical formula × n

$$n = \frac{\text{Molecular formula}}{\text{Empirical formula}}$$

Ex. What is the simplest formula of the compound which has the following percentage composition : Carbon 80%, Hydrogen 20%, If the molecular mass is 30, calculate its molecular formula.

Sol. Calculation of empirical formula :

Element	%	Atomic mass	Relative number of moles	Simple ratio of moles	Simplest whole no. ratio
C	80	12	$\frac{80}{12} = 6.66$	$\frac{6.66}{6.66} = 1$	1
H	20	1	$\frac{20}{1} = 20$	$\frac{20}{6.66} = 3$	3

$\therefore$ Empirical formula is CH_3.

Calculation of molecular formula :

Empirical formula mass = 12 × 1 + 1 × 3 = 15

$$n = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{30}{15} = 2$$

Molecular formula

= Empirical formula × 2 = $CH_3 \times 2 = C_2H_6$.

Ex. On heating a sample of $CaCO_3$, volume of CO_2 evolved at NTP is 112 cc. Calculate

(i) Weight of CO_2 produced

(ii) Weight of $CaCO_3$ taken

(iii) Weight of CaO remaining

Sol. (i) Mole of CO_2 produced $\frac{112}{22400} = \frac{1}{200}$ mole

mass of $CO_2 = \frac{1}{200} \times 44 = 0.22$ gm

(ii) $CaCO_3 \longrightarrow CaO + CO_2$

(1/200 mole)

mole of $CaCO_3 = \frac{1}{200}$ mole

$\therefore$ mass of $CaCO_3 = \frac{1}{200} \times 100 = 0.5$ gm

(iii) mole of CaO produced = $\frac{1}{200}$ mole

mass of CaO = $\frac{1}{200} \times 56 = 0.28$ gm

* Interesting by we can apply

Conversation of mass or wt. of CaO

= wt. of $CaCO_3$ taken – wt. of CO_2 produced

= 0.5 – 0.22 = 0.28 gm

Ex. If all iron present in 1.6 gm Fe_2O_3 is converted in form of $FeSO_4.(NH_4)_2SO_4.6H_2O$ after series of reaction. Calculate mass of product obtained.

Sol. If all iron will be converted then no. of mole atoms of Fe in reactant & product will be same.

$\therefore$ Mole of $Fe_2O_3 = \frac{1.6}{160} = \frac{1}{100}$

mole atoms of Fe = $2 \times \frac{1}{100} = \frac{1}{50}$

mole of $FeSO_4.(NH_4)_2SO_4.6H_2O$ will be same as mole atoms of Fe because one atom of Fe is present in one molecule.

$\therefore$ Mole of $FeSO_4.(NH_4)_2.SO_4.6H_2O = \frac{1}{50}$

$\therefore$ Mass = $\frac{1}{50}$ × Molecule wt.

$= \frac{1}{50} \times 342 = 7.84$ gm.

Ex. Calculate mass of hydrazine N_2H_4 obtained when 1.12 litre of N_2 taken at NTP reacts with H_2 according to $N_2 + 2H_2 \longrightarrow N_2H_4$.

Sol. Moles of N_2 taken = $\frac{1.12}{22.4} = \frac{1}{20}$

$N_2 + 2H_2 \longrightarrow N_2H_4$

(1/20 mole) (1/20 mole)

mass of $N_2H_4 = \frac{1}{20} \times 32 = 1.6$ gm

Ex. Calculate mass of Na_2SO_4 obtained when 100 ml of 0.2 M H_2SO_4 is completely neutralised by NaOH.

Sol. Mole of H_2SO_4 taken =

Molarity × Vol. (lit) = $0.2 \times \frac{100}{1000} = 0.02$

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

0.02 mole 0.02 mole

Mole of Na_2SO_4 obtained = 0.02

mass of $Na_2SO_4 = 0.02 \times 142 = 2.84$ gm