1

MOLE CONCEPT

As you are aware, atoms and molecules are so small that we cannot see them with our naked eyes or even with the help of a microscope. Any sample of matter which can be studied consists of extremely large number of atoms or molecules. In chemical reactions, atoms or molecules combine with one another in a *definite number ratio.* Therefore, it would be pertinent if we could specify the total number of atoms or molecules in a given sample of a substance. We use many *number units* in our daily life. For example, we express the number of bananas or eggs in terms of 'dozen'. In chemistry we use a number unit called **mole** which is very large.

Objectives

After studying this lesson you will be able to:

- state the need of SI units;
- list base SI units;
- explain the relationship between mass and number of particles;
- define Avogadro's constant and state its significance;
- calculate the molar mass of different elements and compounds and
- define molar volume of gases at STP.

1.1 SI Units (Revisited)

Measurement is needed in every walk of life. As you know that for every measurement a 'unit' or a 'reference standard' is required. In different countries, different systems of units gradually developed. This created difficulties whenever people of one country had to deal with those of another country. Since scientists had to often use each other's data, they faced a lot of difficulties. For a practical use, data had to be first converted into local units and then only it could be used.

In 1960, the 'General Conference of Weights and Measures', the international authority on units proposed a new system which was based upon the metric system. This system is called the 'International System of Units' which is abbreviated as SI units from its French name, Le Système Internationale d'Unitès. You have learned about SI units in your earlier classes also and know that they are based upon seven base units corresponding to seven base physical quantities. Units needed for various other physical quantities can be derived from these base SI units. The seven base SI units are listed in Table 1.1

Table 1.1: SI Base Units

Physical Quantity	Name of SI Unit	Symbol for SI unit
Length	Metre	m
Mass	Kilogram	kg
Time	Second	s
Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	Candela	cd

For measuring very large or very small quantities, multiples or sub-multiples of these units are used. Each one of them is denoted by a symbol which is **prefixed** to the symbol of the unit. For example, to measure long distances we use the unit **kilometre** which is a multiple of metre, the base unit of length. Here **kilo** is the prefix used for the multiple $10^{3.}$ Its symbol is k which is prefixed to the symbol of metre, m . Thus the symbol of kilometer is km and

$$1 \text{ km} = 1.0 \times 10^3 \text{ m} = 1000 \text{ m}$$

Similarly, for measuring small lengths we use centimetre (cm) and millimetre (mm) where

$$1 \text{ cm} = 1.0 \times 10^{-2} \text{ m} = 0.01 \text{ m}$$

$$1 \text{ mm} = 1.0 \times 10^{-3} \text{ m} = 0.001 \text{ m}$$

Some prefixes used with SI units are listed in Table 1.2.

Table 1.2: Some prefixes used with SI units

Prefix	Symbol	Meaning	Example
Tera	T	10^{12}	1 terametre (Tm) = 1.0×10^{12} m
Giga	G	10^9	1 gigametre (Gm) = 1.0×10^9 m
Mega	M	10^6	1 megametre (Mm) = 1.0×10^6 m
Kilo	k	10^3	1 kilometre (km) = 1.0×10^3 m
Hecta	h	10^2	1 hectametre (hm) = 1.0×10^2 m
Deca	da	10^1	1 decametre (dam) = 1.0×10^1 m
Deci	d	10^{-1}	1 decimetre (dm) = 1.0×10^{-1} m
Centi	c	10^{-2}	1centimetre (cm) = 1.0×10^{-2} m
Milli	m	10^{-3}	1millimetre (mm) = 1.0×10^{-3} m
Micro	μ	10^{-6}	1micrometre (μm) = 1.0×10^{-6} m
Nano	n	10^{-9}	1nanometre (nm) = 1×10^{-9} m
Pico	p	10^{-12}	1picometre (pm) = 1×10^{-12}m

Before proceeding further try to answer the following questions:

Intext Questions 1.1

1. Name the SI Unit of mass

   ....

2. What symbol will represent 1.0×10^{-6} g ?

   ....

3. Name the prefixes used for (i) 10^2 and (ii) 10^{-9}

   (i) ....

   (ii) ....

4. What do the following symbols represent?

   (i) Ms (ii) ms

   (i) ....

   (ii) ....

1.2 Relationship Between Mass and Number of Particles

Suppose you want to purchase 500 screws. How, do you think, the shopkeeper would give you the desired quantity? By counting the screws individually? No, he would give the screws by weight because it will take a lot of time to count them. If each screw weighs 0.8 g, he would weigh 400 g screws because it is the mass of 500 screws ($0.8 \times 500 = 400$ g). You will be surprised to note that the Reserve Bank of India gives the desired number of coins by weight and not by counting.This process of *counting by weighing* becomes more and more labour saving as the number of items to be counted becomes large. We can carry out the reverse process also. Suppose we take 5000 very tiny springs (used in watches) and weigh them. If the mass of these springs is found to be 1.5 g, we can conclude that mass of each spring is $1.5 \div 5000 = 3 \times 10^{-4}$ g.

Thus, we see that mass and number of identical objects or particles are inter-related. Since atoms and molecules are extremely tiny particles it is impossible to weigh or count them individually. Therefore we need a relationship between the mass and number of atoms and molecules (particles). Such a relationship is provided by 'mole concept'.

1.3 Mole – A Number Unit

Mass of an atom or a molecule is an important property. However, while discussing the quantitative aspects of a chemical reaction, the *number* of reacting atoms or molecules is more significant than their masses. Let us understand this with the help of the following activity.

Activity 1.1

Aim: To study whether during a reaction, the reactants react with each other in a simple ratio by mass.

What is required?

China dish, sulphur powder, iron powder, a magnet and a magnifying glass.

What to do?

Mix 1 g each of iron and sulphur powders in a china dish and heat them till the reaction is complete and the mixture becomes a hard mass. Now break it into small pieces. Repeat the procedure with a mixture of 2 g of iron and 1 g of sulphur powder.

What to observe?

- Pieces obtained from the *'reaction mixture containing iron and sulphur in 1:1 ratio by mass* (1 g each) when observed through a magnifying glass show some yellowish particles of sulphur. When a magnet is brought near them, they are not attracted showing that there is no unreacted iron.
- Pieces obtained from the *reaction mixture containing iron and sulphur in 2:1 ratio by mass* (2 g iron and 1 g sulphur) do not show yellow particles of unreacted sulphur but are attracted by the magnet. This shows the presence of some unreacted iron.

Conclusion

You can conclude that iron and sulphur do not react with each other in a simple mass ratio. When taken in 1:1 ratio by mass (Fe:S), some sulphur is left unreacted and when taken in 2:1 ratio by mass (Fe:S) some iron is left unreacted.

Let us now write the chemical equation of this reaction

$$Fe + S \rightarrow FeS$$

From the above chemical equation, it is clear that 1 atom of iron reacts with 1 atom of sulphur to form 1 molecule of iron (II) sulphide (FeS). It means that if we had taken equal *number* of atoms of iron and sulphur, both of them would have reacted completely. Thus we may conclude that *substances react in a simple ratio by number of atoms or molecules*.

From the above discussion it is clear that the *number* of atoms or molecules of a substance is more relevant than their masses. In order to express their number we need a number unit. One commonly used number unit is 'dozen', which, as you know, means a collection of 12. Other number units that we use are 'score' (20) and 'gross'(144 or 12 dozens). These units are useful in dealing with small numbers only. The atoms and molecules are so small that even in the minute sample of any substance, their number is extremely large. For example, a tiny dust particle contains about 10^{16} molecules. In chemistry such large numbers are commonly represented by a unit known as **mole.** Its symbol is 'mol' and it is defined as.

A mole is the amount of a substance that contains as many elementary entities (atoms, molecules or other particles) as there are atoms in exactly 0.012 kg or 12 g of the carbon-12 isotope.

The term mole has been derived from the Latin word *'moles'* which means a 'heap' or a 'pile'. It was first used by the famous chemist Wilhelm Ostwald more than a hundred years ago.

Here you should remember that one mole always contains the **same number of entities**, no matter what the substance is. Thus *mole* is a number unit for dealing with elementary entities such as atoms, molecules, formula units, electrons etc., just as *dozen* is a number unit for dealing with bananas or oranges. In the next section you will learn more about this number.

1.4 Avogadro's Constant

In the previous section we have learned that a mole of a substance is that amount which contains as many elementary entities as there are atoms in exactly 0.012 kilogram or 12 gram of the carbon-12 isotope. This definition gives us a method by which we can find out the amount of a substance (in moles) if we know the number of elementary entities present in it or *vice versa.* Now the question arises how many atoms are there in exactly 12 g of carbon-12. This number is determined experimentally and its currently accepted value is 6.022045×10^{23}. Thus 1 mol = 6.022045×10^{23} entities or particles, or atoms or molecules.

For all practical purposes this number is · rounded off to 6.022×10^{23}.

The basic idea of such a number was first conceived by an Italian scientist Amedeo Avogadro. But, he never determined this number. It was determinned later and is known as *Avogadro's constant* in his honour.

This number was earlier known as *Avogadro's number*. This number alongwith the unit, that is, 6.022×10^{23} mol^{-1} is known as Avogadro constant. It is represented by the symbol N_A. Here you should be clear that mathematically a number does not have a unit. Avogadro's number 6.022×10^{23} will not have any unit but Avogradro's constant will have unit of mol^{-1}. Thus Avogradro's constant, $N_A = 6.022 \times 10^{23}$ mol^{-1}.

Significance of Avogadro's Constant

You know that 0.012 kg or 12 g of carbon –12 contains its *one mole* of carbon atoms. A mole may be defined as the amount of a substance that contains 6.022×10^{23} elementary entities like atoms, molecules or other particles. When we say one mole of carbon –12, we mean 6.022×10^{23} atoms of carbon –12 whose mass is exactly 12 g. This mass is called the *molar mass* of carbon-12. The ***molar mass** is defined as the mass (in grams) of 1 mole of a substance.* Similarly, a *mole of any substance* would contain 6.022×10^{23} particles or elementary entities. The nature of elementary entity, however,depends upon the nature of the substance as given below :

S.No.	Type of Substance	Elementary Entity
1.	Elements like Na, K, Cu which exist in atomic form	Atom
2.	Elements like O, N, H, which exist in molecular form (O_2, N_2, H_2)	Molecule
3.	Molecular compounds like NH_3, H_2O, CH_4	Molecule
4.	Ions like Na^+, Cu^{2+}, Ag^+, Cl^-, O^{2-}	Ion
5.	Ionic compounds like NaCl, $NaNO_3$, K_2SO_4	Formula unit

Formula unit of a compound contains as many atoms or ions of different types as is given by its chemical formula. The concept is applicable to all types of compounds. The following examples would clarify the concept.

Formula	Atoms/ions present in one formula unit
H_2O	Two atoms of H and one atom of O
NH_3	One atom of N and three atoms of H
NaCl	One Na^+ ion and one Cl^- ion
$NaNO_3$	One Na^+ ion and one NO_3^- ion
K_2SO_4	Two K^+ ions and one SO_4^{2-} ion
$Ba_3(PO_4)_2$	Three Ba^{2+} ions and two PO_4^{3-} ions

Now, let us take the examples of different types of substances and correlate their amounts and the number of elementary entities in them.

1 mole C = 6.022×10^{23} C atoms

1 mole O_2 = 6.022×10^{23} O_2 molecules

1 mole H_2O = 6.022×10^{23} H_2O molecules

1 mole NaCl = 6.022×10^{23} formula units of NaCl

1 mole Ba^{2+} ions = 6.022×10^{23} Ba^{2+} ions

We may choose to take amounts other than one mole and correlate them with number of particles present with the help of relation :

Number of elementary entities = number of moles × Avogadro's constant

1 mole O_2 = $1 \times (6.022 \times 10^{23})$ = 6.022×10^{23} molecules of O_2

0.5 mole O_2 = $0.5 \times (6.022 \times 10^{23})$ = 3.011×10^{23} molecules of O_2

0.1 mole O_2 = $0.1 \times (6.022 \times 10^{23})$ = 6.022×10^{22} molecules of O_2

Intext Questions 1.2

1. A sample of nitrogen gas consists of 4.22×10^{23} molecules of nitrogen. How many moles of nitrogen gas are there?

   ....

2. In a metallic piece of magnesium, 8.46×10^{24} atoms are present. Calculate the amount of magnesium in moles.

   ....

3. Calculate the number of Cl_2 molecules and Cl atoms in 0.25 mol of Cl_2 gas.

   ....

   ....

1.5 Mole, Mass and Number Relationships

You know that 1 mol $= 6.022 \times 10^{23}$ elementary entities

and Molar mass = Mass of 1 mole of substance

= Mass of 6.022×10^{23} elementary entities.

As discussed earlier the elementary entity can be an atom, a molecule, an ion or a formula unit. As far as mole – number relationship is concerned it is clear that one mole of any substance would contain 6.022×10^{23} particles (elementary entities). For obtaining the molar mass, i.e., mole-mass relationship we have to use atomic mass scale.

1.5.1 Atomic Mass Unit

By inernational agreement, a unit of mass to specify the atomic and molecular masses has been defined. This unit is called *atomic mass unit* and its symbol is 'amu'. The mass of one C-12 atom, is taken as exactly 12 amu. Thus, C-12 atom serves as the *standard.* The **Atomic mass unit** is defined as a mass exactly equal to the $1/12^{th}$ of the mass of one carbon-12 atom.

$$1 \text{ amu} = \frac{\text{Mass of one C-12 atom}}{12}$$

Atomic mass unit is also called **unified atomic mass unit** whose symbol is 'u'. Another name of atomic mass unit is **dalton** (symbol Da). The latter is mainly used in biological sciences.

1.5.2 Relative Atomic and Molecular Masses

You are aware that atomic mass scale is a *relative scale* with C-12 atom (also written as ^{12}C) chosen as the standard. Its mass is taken as exactly 12. Relative masses of atoms and molecules are the number of times each atom or molecules is heavier than $\frac{1}{12}$th of the mass of one C-12 atom. Often, we deal with elements and compounds containing isotopes of different elements. Therefore, we prefer to use *average* masses of atoms and molecules. Thus

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

$$\text{and Relative molecular mass} = \frac{\text{Average mass of 1 molecule of the substance}}{\frac{1}{12}\text{th of the mass of one C-12 atom}}$$

Experiments show that one O-16 atom is 1.333 times as heavy as one C-12 atom. Thus

$$\text{Relative atomic mass of O-16} = 1.333 \times 12 = 15.996 \simeq 16.0$$

The relative atomic masses of all elements have been determined in a similar manner. Relative molecular masses can also be determined experimentally in a similar manner . In case we know the molecular formula of a molecule, we can calculate its relative molecular mass by adding the relative atomic masses of all its constituent atoms. Let us calculate the relative molecular mass of water, H_2O.

Relative molecular mass of water, H_2O = (2 × relative atomic mass of H) + (relative atomic mass of O)

$$= (2 \times 1) + (16) = 2 + 16 = 18$$

The relative atomic and molecular masses are just numbers and dimensionless, unit-less quantities.

1.5.3 Atomic, Molecular and Formula Masses

From the definition of atomic mass unit, we can calculate the atomic masses. Let us again take the example of oxygen-16 whose relative atomic mass is 16. By definition:

$$\text{Relative atomic mass of O-16} = 16 = \frac{\text{mass of one O-16 atom}}{\frac{1}{12}\text{th the mass of one C-12 atom}}$$

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

$$\therefore 16 = \frac{\text{mass of one O-16 atom}}{1\text{ amu}}$$

Mass of one O-16 atom = 16 amu

Or Atomic mass of O-16 = 16 amu.

From this example we can see that numerical value of the relative atomic mass and atomic mass is the same. Only, the former has no unit while the latter has the unit *amu*.

Molecular and formula masses can be obtained by adding the atomic or ionic masses of all the constituent atoms or ions of the molecule or formula unit respectively. Let us understand these calculations with the help of following examples.

Example 1.1 : Calculate the molecular mass of ammonia, NH_3.

Solution : One molecule of NH_3 consists of one N atom and three H atoms.

Molecular mass of NH_3 = (Atomic mass of N) + 3 (Atomic mass of H)

= [14 + (3 × 1)] amu

= 17 amu

Example 1.2: Calculate the formula mass of sodium chloride (NaCl).

Solution : One formula unit of sodium chloride consists of one Na^+ ion and one Cl^- ion.

Formula mass of NaCl = (Ionic mass of Na^+) + (Ionic mass of Cl^-)

= 23 amu + 35.5 amu

= 58.5 amu.

You would have noticed in the above example that ionic mass of Na^+ ion has been taken as 23 amu which is the same as the atomic mass of Na atom. Since loss or gain of few electrons does not change the mass significantly, therefore atomic masses are used as ionic masses. Similarly we have taken ionic mass of Cl^- as 35.5 amu which is the same as the atomic mass of Cl^-.

1.5.4 Molar Masses

We know that molar mass is the mass of 1 mol of the substance. Also, 1 mol of any substance is the collection of its 6.022×10^{23} elementary entities. Thus

Molar mass = Mass of 6.022×10^{23} elementary entities.

(i) Molar mass of an element

You know that the relative atomic mass of carbon–12 is 12. A 12g sample of it would contain 6.022×10^{23} atoms. Hence the molar mass of C-12 is 12 g mol^{-1}. For getting the molar masses of other elements we can use their relative atomic masses.

Since the relative atomic mass of oxygen -16 is 16, a 16 g sample of it would contain 6.022×10^{23} oxygen atoms and would constitute its one mole. Thus, the molar mass of O–16 is 16 g mol^{-1}. Relative atomic masses of some common elements have been listed in Table 1.3

Table 1.3 : Relative atomic masses of some elements (upto 1st place of decimal)

Element	Relative Atomic Mass	Element	Relative Atomic Mass
Hydrogen, H	1.0	Phosphorus, P	31.0
Carbon, C	12.0	Sulphur, S	32.1
Nitrogen, N	14.0	Chlorine, Cl	35.5
Oxygen, O	16.0	Potassium, K	39.1
Sodium, Na	23.0	Iron, Fe	55.9

(ii) Molar mass of a molecular substance

The elementary entity in case of a molecular substance is the molecule. Hence, *molar mass of such a substance would be the mass of its 6.022 ò 10^{23} molecules*, which can be obtained from its relative molecular mass or by multiplying the molar mass of each element by the number of its moles present in one mole of the substance and then adding them.

Let us take the example of water, H_2O. Its relative molecular mass is 18. Therefore, 18 g of it would contain 6.022×10^{23} molecules. Hence, its molar mass is 18 g mol^{-1} . Alternately we can calculate it as :

$$\text{Molar mass of water, } H_2O = (2 \times \text{molar mass of H}) + (\text{molar mass of O})$$
$$= (2 \times 1 \text{ g mol}^{-1}) + (16 \text{ g mol}^{-1})$$
$$= 18 \text{ g mol}^{-1}$$

Table 1.4 lists molecular masses and molar masses of some substances.

Table 1.4 : Molecular masses and molar masses of some substances

Element or Compound	Molecular mass / amu	Molar mass / (g mol^{-1})
O_2	32.0	32.0
P_4	124.0	124.0
S_8	256.8	256.8
H_2O	18.0	18.0
NH_3	17.0	17.0
HCl	36.5	36.5
CH_2Cl_2	85.0	85.0

(iii) Molar masses of ionic compounds

Molar mass *of an ionic compound is the mass of its 6.022 ò 10^{23} formula units.* It can be obtained by adding the molar masses of ions present in the formula unit of the substance. In case of NaCl it is calculated as

Molar mass of NaCl $= \text{molar mass of } Na^+ + \text{molar mass of } Cl^-$

$= (23 \text{ g mol}^{-1}) + (35.5 \text{ g mol}^{-1})$

$= 58.5 \text{ g mol}^{-1}$

Let us take some more examples of ionic compounds and calculate their molar masses.

Example 1.3 : Calculate the molar mass of

(i) K_2SO_4 (ii) $Ba_3(PO_4)_2$

Solution :

(i) Molar mass of K_2SO_4 $= (2 \times \text{molar mass of } K^+) + (\text{molar mass of } SO_4^{2-})$

$= (2 \times \text{molar mass of } K^+) +$
$(\text{molar mass of S} + 4 \times \text{molar mass of O})$

$= [(2 \times 39.1) + (32.1 + 4 \times 16)] \text{ g mol}^{-1}]$

$= (78.2 + 32.1 + 64) \text{ g mol}^{-1} = 174.3 \text{ g mol}^{-1}$

(ii) Molar mass of $Ba_3(PO_4)_2$ $= (3 \times \text{molar mass of } Ba^{2+}) + 2\ (\text{molar mass of } PO_4^{3-})$

$= (3 \times \text{molar mass of } Ba^{2+}) +$
$2\ (\text{molar mass of P} + 4 \times \text{molar mass of O})$

$= [(3 \times 137.3) + 2\ (31.0 + 4 \times 16.0)] \text{ g mol}^{-1}$

$= (411.9 + 190.0) \text{ g mol}^{-1} = 601.9 \text{ g mol}^{-1}$

Now you have learned about the mole, mass and number relationships for all types of substances. The following examples would illustrate the usefulness of these relationships.

Example 1.4 : Find out the mass of carbon -12 that would contain 1.0×10^{19} carbon-12 atoms.

Solution : Mass of 6.022×10^{23} carbon-12 atoms $= 12$ g

Mass of 1.0×10^{19} carbon-12 atoms $= \dfrac{12 \times 1 \times 10^{19}}{6.022 \times 10^{23}}$ g

$= 1.99 \times 10^{-4}$ g

Example 1.5 : How many molecules are present in 100 g sample of NH_3?

Solution : Molar mass of NH_3 $= (14 + 3) \text{ g mol}^{-1} = 17 \text{ g mol}^{-1}$

$\therefore$ 17 g sample of NH_3 contains 6.022×10^{23} molecules

Therefore, 100 g sample of NH_3 would contain $\dfrac{6.022 \times 10^{23} \text{ molecule}}{17\text{g}} \times 100\text{g}$

$= 35.42 \times 10^{23}$ molecules

$= 3.542 \times 10^{24}$ molecules

Example 1.6 : Molar mass of O is 16 g mol^{-1}. What is the mass of one atom and one molecule of oxygen?

Solution : Mass of 1 mol or 6.022×10^{23} atoms of O = 16 g

$$\therefore \quad \text{Mass of 1atom of O} = \frac{16\text{g}}{6.022 \times 10^{23}}$$

$$= 2.66 \times 10^{-23}\ \text{g}$$

Since a molecule of oxygen contains two atoms (O_2),
its mass = $2 \times 2.66 \times 10^{-23}$ g = 5.32×10^{-23} g.

Intext Questions 1.3

1. Calculate the molar mass of hydrogen chloride, HCl.

   ....

2. Calculate the molar mass of argon atoms, given that the mass of single atom is 6.634×10^{-26} kg.

   ....

3. Calculate the mass of 1.0 mol of potassium nitrate, KNO_3 (atomic masses : K = 39 amu; N = 14 amu, O = 16 amu).

   ....

4. The formula of sodium phosphate is Na_3PO_4. What is the mass of 0.146 mol of Na_3PO_4? (atomic masses : Na = 23.0 amu, P = 31.0 amu; O = 16.0 amu).

   ....

1.6 Mass, Molar Mass and Number of Moles

Mass, molar mass and number of moles of a substance are inter-related quantities. We know that :

Molar mass (M) = Mass of one mole of the substance.

Molar mass of water is 18 g mol^{-1}. If we have 18 g of water, we have 1mol of it. Suppose we have 36 g water (18×2), we have 2 mol of it. In general in a sample of water of mass ($n \times 18$) g, the number of moles of water would be n. We may generalize the relation as

$$\text{Number of moles (amount) of a substance} = \frac{\text{mass of the substance}}{\text{molar mass of the substance}}$$

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

or

$$m = n \times M$$

These relations are useful in calculations involving moles of substances.

Example 1.7 : In a reaction, 0.5 mol of aluminium is required. Calculate the amount of aluminium required in grams? (atomic mass of Al = 27 amu)

Solution :

Molar mass of Al $= 27 \text{ g mol}^{-1}$

Required mass $=$ no. of moles $\times$ molar mass

$= (0.5 \text{ mol}) \times (27 \text{ g mol}^{-1})$

$= 13.5 \text{ g}$

1.7 Molar Volume, V_m

Molar volume *is the volume of one mole of a substance.* It depends upon temperature and pressure. It is related to the density, by the relation.

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

In case of gases, we use their volumes at **standard temperature and pressure (STP).** For this purpose **0 ^{0}C** or **273 K** temperature is taken as the **standard temperature** and **1bar** pressure is taken as the **standard pressure**. At STP, the molar volume of an ideal gas is 22.7 litre*. You will study that gases do not behave ideally and therefore their molar volume is not exactly 22.7 L. However, it is very close to 22.7 L and for all practical purposes we take the molar volume of all gases at STP as 22.7 L mol^{-1}.

Intext Questions 1.4

1. How many moles of Cu atoms are present in 3.05 g of copper (Relatine atomic mass of Cu = 63.5).

   ....

2. A piece of gold has a mass of 12.6 g. How many moles of gold are present in it?

   (Relative atomic mass of Au = 197)

   ....

3. In a combustion reaction of an organic compound, 2.5 mol of CO_2 were produced. What volume would it occupy at STP (273K, 1bar) ?

   ....

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* Earlier 1 atmosphere pressure was taken as the standard pressure and at STP (273K, 1atm) the molar volume of an ideal gas was taken as 22.4 L mol^{-1}. The difference in the value is due to the change in the standard pressure (1bar) which is slightly less than 1atm.

What You Have Learnt

- Mole is the amount of a substance which contains as may elementary entities as there are atoms present in 0.012 kg or 12 g of C-12. Thus mole denotes a number.
- The number of elementary entities present in one mole of a substance is 6.022×10^{23}.
- The mass of one mole of a substance is called its molar mass It is numerically equal to relative atomic mass or relative molecular mass expressed in grams per mole (g mol^{-1}) or kilogram per mole (kg mol^{-1}).
- Molar volume is the volume occupied by one mole of a substance. One mole of an ideal gas at standard pressure and temperature, STP (273 K and 1 bar) occupies 22.7 litres.
- In ionic substances, molar mass is numerically equal to the formula mass of the compound expressed in grams.
- If the molar mass of a substance is known, then the amount of a substance present in a sample having a definite mass can be calculated. If *M* is the molar mass, then, the amount of substance *n*, present in a sample of mass *m* is expressed as $n = \dfrac{m}{M}$.

Terminal Exercise

1. How many atoms are present in a piece of iron that has a mass of 65.0 g/ (atomic mass; Fe = 55.9 amu).
2. A piece of phosphorus has a mass of 99.2 g. How many moles of phosphorus, P_4 are present in it? (atomic mass, P = 31.0 amu)
3. Mass of 8.46×10^{24} atoms of fluorine is 266.95 g. Calculate the atomic mass of fluorine.
4. A sample of magnesium consists of 1.92×10^{22} Mg atoms. What is the mass of the sample in grams? (atomic mass = 24.3 amu)
5. Calculate the molar mass in g mol^{-1} for each of the following:

   (i) Sodium hydroxide, NaOH

   (ii) Copper Sulphate $CuSO_4.5H_2O$.

   (iii) Sodium Carbonate, $Na_2CO_3.10H_2O$
6. For 150 gram sample of phosphorus trichloride (PCl_3), calculate each of the following:

   (i) Mass of one PCl_3 molecule.

   (ii) The number of moles of PCl_3 and Cl in the sample.

(iii) The number of grams of Cl atoms in the sample.

(iv) The number of molecules of PCl_3 in the sample.

7. Find out the mass of carbon-12, that would contain 1×10^{19} atoms.
8. How many atoms are present in 100 g sample of C-12 atom?
9. How many moles of $CaCO_3$ would weigh 5 g?
10. If you require 1.0×10^{23} molecules of nitrogen for the reaction $N_2 + 3H_2 \rightarrow 2NH_3$.

    (i) What is the mass (in grams) of N_2 required?

    (ii) How many moles of NH_3 would be formed in the above reaction from 1.0×10^{23} molecules of N_2?

    (iii) What volume would NH_3 gas formed in (ii) occupy at STP?

Answers to Intext Questions

1.1

1. Kilogram
2. μg
3. (i) h (ii) n
4. (i) Megasecond, 10^6 s

   (ii) millisecond, 10^{-3} s.

1.2

1. Moles of N_2 gas = $\dfrac{4.22 \times 10^{23} \text{ molecules}}{6.022 \times 10^{23} \text{ molecules mol}^{-1}}$ = 0.70 mol
2. Amount of magnesium (moles) = $\dfrac{8.46 \times 10^{24} \text{ atoms}}{6.022 \times 10^{23} \text{ atoms mol}^{-1}}$ = 14.05 mol
3. No. of Cl_2 molecules in 0.25 mol Cl_2 = $0.25 \times 6.022 \times 10^{23}$ molecules

   $= 1.5055 \times 10^{23}$ molecules

   Since each Cl_2 molecule has 2 Cl atoms, the number of Cl atoms = $2 \times 1.5055 \times 10^{23}$ = 3.011×10^{23} atoms.

1.3

1. Molar mass of hydrogen chloride = molar mass of HCl

   = 1 mol of H + 1 mol of Cl

   $= 1.0 \text{ g mol}^{-1} + 35.5 \text{ g mol}^{-1}$

   $= 36.5 \text{ g mol}^{-1}$

2. Molar mass of argon atoms = mass of 1 mol of argon

$= \text{mass of } 6.022 \times 10^{23} \text{ atoms of argon.}$

$= 6.634 \times 10^{-26} \text{ kg} \times 6.022 \times 10^{23} \text{ mol}^{-1}$

$= 39.95 \times 10^{-3} \text{ kg mol}^{-1}$

$= 39.95 \text{ g mol}^{-1}$

3. Molar mass of KNO_3 = mass of 1 mol of K + mass of 1 mol of N + mass of 3 mol of O.

Since molar mass of an element is numerically equal to its atomic mass but has the units of g mol^{-1} in place of amu $= 39.1 \text{ g} + 14.0 \text{ g} + 3 \times 16.0 \text{ g}$

$\therefore$ Molar mass of KNO_3 $= 39.1 \text{ g} + 14.0 \text{ g} + 48.0 \text{ g} = 101.1 \text{ g mol}^{-1}$

4. Mass of 1 mol of Na_3PO_4 = 3 × (mass of 1mol of Na) + mass of 1 mol of P + 4 × (mass of 1mol of oxygen)

$= 3\,(23.0 \text{ g}) + 31.0 \text{ g} + 4(16.0) \text{ g}$

$= 69.0 \text{ g} + 31.0 \text{ g} + 64.0 \text{ g} = 164.0 \text{ g}$

$\therefore$ Mass of 0.146 mol of Na_3PO_4 $= 0.146 \times 164.0 \text{ g} = 23.94 \text{ g}$

1.4

1. Moles of Cu atoms in 3.05 g copper $= \dfrac{3.05 \text{ g}}{63.5 \text{ g mol}^{-1}} = 0.048 \text{ mol}$

2. Moles of gold, Au $= \dfrac{12.6 \text{ g}}{197 \text{ g mol}^{-1}} = 0.064 \text{ mol}$

3. Molar volume of any gas at STP (298 K, 1 bar) = 22.7 L

$\therefore$ Volume occupied by 2.5 mol CO_2 at STP $= 2.5 \times 22.7 \text{ L} = 56.75 \text{ L}$