

Chapter 3

ATOMS AND MOLECULES

Ancient Indian and Greek philosophers have always wondered about the unknown and unseen form of matter. The idea of divisibility of matter was considered long back in India, around 500 BC. An Indian philosopher Maharishi Kanad, postulated that if we go on dividing matter (*padarth*), we shall get smaller and smaller particles. Ultimately, a time will come when we shall come across the smallest particles beyond which further division will not be possible. He named these particles *Parmanu*. Another Indian philosopher, Pakudha Katyayama, elaborated this doctrine and said that these particles normally exist in a combined form which gives us various forms of matter.

Around the same era, ancient Greek philosophers – Democritus and Leucippus suggested that if we go on dividing matter, a stage will come when particles obtained cannot be divided further. Democritus called these indivisible particles atoms (meaning indivisible). All this was based on philosophical considerations and not much experimental work to validate these ideas could be done till the eighteenth century.

By the end of the eighteenth century, scientists recognised the difference between elements and compounds and naturally became interested in finding out how and why elements combine and what happens when they combine.

Antoine L. Lavoisier laid the foundation of chemical sciences by establishing two important laws of chemical combination.

3.1 Laws of Chemical Combination

The following two laws of chemical combination were established after much

experimentations by Lavoisier and Joseph L. Proust.

3.1.1 LAW OF CONSERVATION OF MASS

Is there a change in mass when a chemical change (chemical reaction) takes place?

Activity 3.1

- Take one of the following sets, X and Y of chemicals—

X	Y
(i) copper sulphate	sodium carbonate
(ii) barium chloride	sodium sulphate
(iii) lead nitrate	sodium chloride
- Prepare separately a 5% solution of any one pair of substances listed under X and Y in water.
- Take a little amount of solution of Y in a conical flask and some solution of X in an ignition tube.
- Hang the ignition tube in the flask carefully; see that the solutions do not get mixed. Put a cork on the flask (see Fig. 3.1).

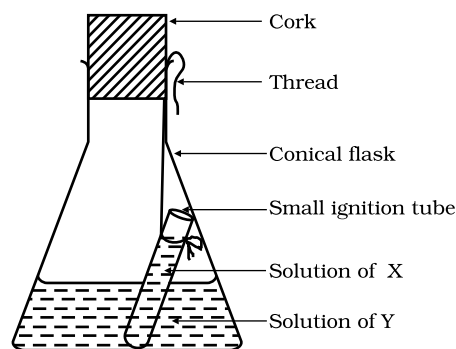


Fig. 3.1: Ignition tube containing solution of X, dipped in a conical flask containing solution of Y.

- Weigh the flask with its contents carefully.
- Now tilt and swirl the flask, so that the solutions X and Y get mixed.
- Weigh again.
- What happens in the reaction flask?
- Do you think that a chemical reaction has taken place?
- Why should we put a cork on the mouth of the flask?
- Does the mass of the flask and its contents change?

Law of conservation of mass states that mass can neither be created nor destroyed in a chemical reaction.

3.1.2 LAW OF CONSTANT PROPORTIONS

Lavoisier, along with other scientists, noted that many compounds were composed of two or more elements and each such compound had the same elements in the same proportions, irrespective of where the compound came from or who prepared it.

In a compound such as water, the ratio of the mass of hydrogen to the mass of oxygen is always 1:8, whatever the source of water. Thus, if 9 g of water is decomposed, 1 g of hydrogen and 8 g of oxygen are always obtained. Similarly in ammonia, nitrogen and hydrogen are always present in the ratio 14:3 by mass, whatever the method or the source from which it is obtained.

This led to the law of constant proportions which is also known as the law of definite proportions. This law was stated by Proust as *“In a chemical substance the elements are always present in definite proportions by mass”*.

The next problem faced by scientists was to give appropriate explanations of these laws. British chemist John Dalton provided the basic theory about the nature of matter. Dalton picked up the idea of divisibility of matter, which was till then just a philosophy. He took the name ‘atoms’ as given by the Greeks and said that the smallest particles of matter are atoms. His theory was based on the laws of chemical combination. Dalton’s

atomic theory provided an explanation for the law of conservation of mass and the law of definite proportions.

John Dalton was born in a poor weaver’s family in 1766 in England. He began his career as a teacher at the age of twelve. Seven years later he became a school principal. In 1793, Dalton left for Manchester to teach mathematics, physics and chemistry in a college. He spent most of his life there teaching and researching. In 1808, he presented his atomic theory which was a turning point in the study of matter.



John Dalton

According to Dalton’s atomic theory, all matter, whether an element, a compound or a mixture is composed of small particles called atoms. The postulates of this theory may be stated as follows:

- All matter is made of very tiny particles called atoms.
- Atoms are indivisible particles, which cannot be created or destroyed in a chemical reaction.
- Atoms of a given element are identical in mass and chemical properties.
- Atoms of different elements have different masses and chemical properties.
- Atoms combine in the ratio of small whole numbers to form compounds.
- The relative number and kinds of atoms are constant in a given compound.

You will study in the next chapter that all atoms are made up of still smaller particles.

Questions

1. In a reaction, 5.3 g of sodium carbonate reacted with 6 g of ethanoic acid. The products were 2.2 g of carbon dioxide, 0.9 g water and 8.2 g of sodium ethanoate. Show that these

observations are in agreement with the law of conservation of mass.

sodium carbonate + ethanoic acid
→ sodium ethanoate + carbon dioxide + water

2. Hydrogen and oxygen combine in the ratio of 1:8 by mass to form water. What mass of oxygen gas would be required to react completely with 3 g of hydrogen gas?
3. Which postulate of Dalton's atomic theory is the result of the law of conservation of mass?
4. Which postulate of Dalton's atomic theory can explain the law of definite proportions?

3.2 What is an Atom?

Have you ever observed a mason building walls, from these walls a room and then a collection of rooms to form a building? What is the building block of the huge building? What about the building block of an ant-hill? It is a small grain of sand. Similarly, the building blocks of all matter are atoms.

How big are atoms?

Atoms are very small, they are smaller than anything that we can imagine or compare with. More than millions of atoms when stacked would make a layer barely as thick as this sheet of paper.

Atomic radius is measured in nanometres.
 $1/10^9 \text{ m} = 1 \text{ nm}$
 $1 \text{ m} = 10^9 \text{ nm}$

Relative Sizes	
Radii (in m)	Example
10^{-10}	Atom of hydrogen
10^{-9}	Molecule of water
10^{-8}	Molecule of haemoglobin
10^{-4}	Grain of sand
10^{-2}	Ant
10^{-1}	Watermelon

We might think that if atoms are so insignificant in size, why should we care about them? This is because our entire world is made up of atoms. We may not be able to see them, but they are there, and constantly affecting whatever we do. Through modern techniques, we can now produce magnified images of surfaces of elements showing atoms.

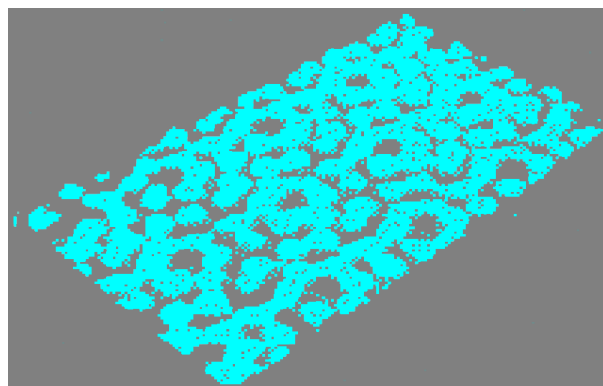


Fig. 3.2: An image of the surface of silicon

3.2.1 WHAT ARE THE MODERN DAY SYMBOLS OF ATOMS OF DIFFERENT ELEMENTS?

Dalton was the first scientist to use the symbols for elements in a very specific sense. When he used a symbol for an element he also meant a definite quantity of that element, that is, one atom of that element. Berzilius suggested that the symbols of elements be made from one or two letters of the name of the element.

	Hydrogen		Carbon		Oxygen
	Phosphorus		Sulphur		Iron
	Copper		Lead		Silver
	Gold		Platina		Mercury

Fig. 3.3: Symbols for some elements as proposed by Dalton

In the beginning, the names of elements were derived from the name of the place where they were found for the first time. For example, the name copper was taken from Cyprus. Some names were taken from specific colours. For example, gold was taken from the English word meaning yellow. Now-a-days, IUPAC (International Union of Pure and Applied Chemistry) approves names of elements. Many of the symbols are the first one or two letters of the element's name in English. The first letter of a symbol is always written as a capital letter (uppercase) and the second letter as a small letter (lowercase).

For example

- (i) hydrogen, H
- (ii) aluminium, Al and not AL
- (iii) cobalt, Co and not CO.

Symbols of some elements are formed from the first letter of the name and a letter, appearing later in the name. Examples are: (i) chlorine, Cl, (ii) zinc, Zn etc.

Other symbols have been taken from the names of elements in Latin, German or Greek. For example, the symbol of iron is Fe from its Latin name ferrum, sodium is Na from natrium, potassium is K from kalium. Therefore, each element has a name and a unique chemical symbol.

the passage of time and repeated usage you will automatically be able to reproduce the symbols).

3.2.2 ATOMIC MASS

The most remarkable concept that Dalton's atomic theory proposed was that of the atomic mass. According to him, each element had a characteristic atomic mass. The theory could explain the law of constant proportions so well that scientists were prompted to measure the atomic mass of an atom. Since determining the mass of an individual atom was a relatively difficult task, relative atomic masses were determined using the laws of chemical combinations and the compounds formed.

Let us take the example of a compound, carbon monoxide (CO) formed by carbon and oxygen. It was observed experimentally that 3 g of carbon combines with 4 g of oxygen to form CO. In other words, carbon combines with $\frac{4}{3}$ times its mass of oxygen. Suppose we define the atomic mass unit (earlier abbreviated as 'amu', but according to the latest IUPAC recommendations, it is now written as 'u' – unified mass) as equal to the mass of one carbon atom, then we would assign carbon an atomic mass of 1.0 u and

Table 3.1: Symbols for some elements

Element	Symbol	Element	Symbol	Element	Symbol
Aluminium	Al	Copper	Cu	Nitrogen	N
Argon	Ar	Fluorine	F	Oxygen	O
Barium	Ba	Gold	Au	Potassium	K
Boron	B	Hydrogen	H	Silicon	Si
Bromine	Br	Iodine	I	Silver	Ag
Calcium	Ca	Iron	Fe	Sodium	Na
Carbon	C	Lead	Pb	Sulphur	S
Chlorine	Cl	Magnesium	Mg	Uranium	U
Cobalt	Co	Neon	Ne	Zinc	Zn

(The above table is given for you to refer to whenever you study about elements. Do not bother to memorise all in one go. With

oxygen an atomic mass of 1.33 u. However, it is more convenient to have these numbers as whole numbers or as near to a whole numbers

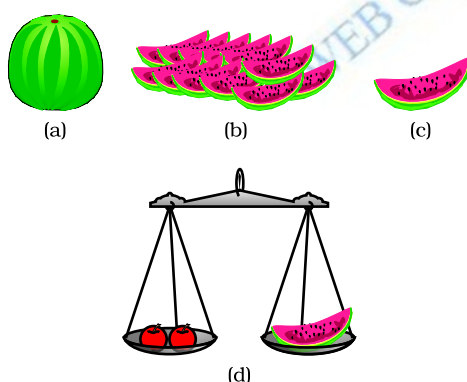
as possible. While searching for various atomic mass units, scientists initially took $1/16$ of the mass of an atom of naturally occurring oxygen as the unit. This was considered relevant due to two reasons:

- oxygen reacted with a large number of elements and formed compounds.
- this atomic mass unit gave masses of most of the elements as whole numbers.

However, in 1961 for a universally accepted atomic mass unit, carbon-12 isotope was chosen as the standard reference for measuring atomic masses. One atomic mass unit is a mass unit equal to exactly one-twelfth ($1/12^{\text{th}}$) the mass of one atom of carbon-12. The relative atomic masses of all elements have been found with respect to an atom of carbon-12.

Imagine a fruit seller selling fruits without any standard weight with him. He takes a watermelon and says, “this has a mass equal to 12 units” (12 watermelon units or 12 fruit mass units). He makes twelve equal pieces of the watermelon and finds the mass of each fruit he is selling, relative to the mass of one piece of the watermelon. Now he sells his fruits by relative fruit mass unit (fmu), as in Fig. 3.4.

Fig. 3.4 : (a) Watermelon, (b) 12 pieces, (c) $1/12$ of



watermelon, (d) how the fruit seller can weigh the fruits using pieces of watermelon

Similarly, the relative atomic mass of the atom of an element is defined as the average

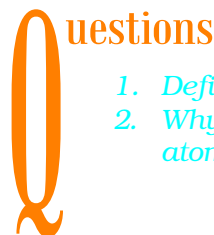
mass of the atom, as compared to $1/12^{\text{th}}$ the mass of one carbon-12 atom.

Table 3.2: Atomic masses of a few elements

Element	Atomic Mass (u)
Hydrogen	1
Carbon	12
Nitrogen	14
Oxygen	16
Sodium	23
Magnesium	24
Sulphur	32
Chlorine	35.5
Calcium	40

3.2.3 HOW DO ATOMS EXIST?

Atoms of most elements are not able to exist independently. Atoms form molecules and ions. These molecules or ions aggregate in large numbers to form the matter that we can see, feel or touch.



Questions

1. Define the atomic mass unit.
2. Why is it not possible to see an atom with naked eyes?

3.3 What is a Molecule?

A molecule is in general a group of two or more atoms that are chemically bonded together, that is, tightly held together by attractive forces. A molecule can be defined as the smallest particle of an element or a compound that is capable of an independent existence and shows all the properties of that substance. Atoms of the same element or of different elements can join together to form molecules.

3.3.1 MOLECULES OF ELEMENTS

The molecules of an element are constituted by the same type of atoms. Molecules of many elements, such as argon (Ar), helium (He) etc. are made up of only one atom of that element. But this is not the case with most of the non-metals. For example, a molecule of oxygen consists of two atoms of oxygen and hence it is known as a diatomic molecule, O_2 . If 3 atoms of oxygen unite into a molecule, instead of the usual 2, we get ozone. The number of atoms constituting a molecule is known as its atomicity.

Metals and some other elements, such as carbon, do not have a simple structure but consist of a very large and indefinite number of atoms bonded together.

Let us look at the atomicity of some non-metals.

Table 3.3 : Atomicity of some elements

Type of Element	Name	Atomicity
Non-Metal	Argon	Monoatomic
	Helium	Monoatomic
	Oxygen	Diatomic
	Hydrogen	Diatomic
	Nitrogen	Diatomic
	Chlorine	Diatomic
	Phosphorus	Tetra-atomic
	Sulphur	Poly-atomic

3.3.2 MOLECULES OF COMPOUNDS

Atoms of different elements join together in definite proportions to form molecules of compounds. Few examples are given in Table 3.4.

Table 3.4 : Molecules of some compounds

Compound	Combining Elements	Ratio by Mass
Water	Hydrogen, Oxygen	1:8
Ammonia	Nitrogen, Hydrogen	14:3
Carbon dioxide	Carbon, Oxygen	3:8

Activity 3.2

- Refer to Table 3.4 for ratio by mass of atoms present in molecules and Table 3.2 for atomic masses of elements. Find the ratio by number of the atoms of elements in the molecules of compounds given in Table 3.4.
- The ratio by number of atoms for a water molecule can be found as follows:

Element	Ratio by mass	Atomic mass (u)	Mass ratio/atomic mass	Simplest ratio
H	1	1	$\frac{1}{1} = 1$	2
O	8	16	$\frac{8}{16} = \frac{1}{2}$	1

- Thus, the ratio by number of atoms for water is H:O = 2:1.

3.3.3 WHAT IS AN ION?

Compounds composed of metals and non-metals contain charged species. The charged species are known as *ions*. An ion is a charged particle and can be negatively or positively charged. A negatively charged ion is called an 'anion' and the positively charged ion, a 'cation'. Take, for example, sodium chloride (NaCl). Its constituent particles are positively charged sodium ions (Na^+) and negatively charged chloride ions (Cl^-). Ions may consist of a single charged atom or a group of atoms

that have a net charge on them. A group of atoms carrying a charge is known as a polyatomic ion (Table 3.6). We shall learn more about the formation of ions in Chapter 4.

Table 3.5: Some ionic compounds

Ionic Compound	Constituting Elements	Ratio by Mass
Calcium oxide	Calcium and oxygen	5:2
Magnesium sulphide	Magnesium and sulphur	3:4
Sodium chloride	Sodium and chlorine	23:35.5

3.4 Writing Chemical Formulae

The chemical formula of a compound is a symbolic representation of its composition. The chemical formulae of different compounds can be written easily. For this

exercise, we need to learn the symbols and combining capacity of the elements.

The combining power (or capacity) of an element is known as its valency. Valency can be used to find out how the atoms of an element will combine with the atom(s) of another element to form a chemical compound. The valency of the atom of an element can be thought of as hands or arms of that atom.

Human beings have two arms and an octopus has eight. If one octopus has to catch hold of a few people in such a manner that all the eight arms of the octopus and both arms of all the humans are locked, how many humans do you think the octopus can hold? Represent the octopus with O and humans with H. Can you write a formula for this combination? Do you get OH_4 as the formula? The subscript 4 indicates the number of humans held by the octopus.

The valencies of some common ions are given in Table 3.6. We will learn more about valency in the next chapter.

Table 3.6: Names and symbols of some ions

Valency	Name of ion	Symbol	Non-metallic element	Symbol	Polyatomic ions	Symbol
1.	Sodium	Na^+	Hydrogen	H^+	Ammonium	NH_4^+
	Potassium	K^+	Hydride	H^-	Hydroxide	OH^-
	Silver	Ag^+	Chloride	Cl^-	Nitrate	NO_3^-
	Copper (I)*	Cu^+	Bromide	Br^-	Hydrogen carbonate	HCO_3^-
			Iodide	I^-		
2.	Magnesium	Mg^{2+}	Oxide	O^{2-}	Carbonate	CO_3^{2-}
	Calcium	Ca^{2+}	Sulphide	S^{2-}	Sulphite	SO_3^{2-}
	Zinc	Zn^{2+}			Sulphate	SO_4^{2-}
	Iron (II)*	Fe^{2+}				
	Copper (II)*	Cu^{2+}				
3.	Aluminium	Al^{3+}	Nitride	N^{3-}	Phosphate	PO_4^{3-}
	Iron (III)*	Fe^{3+}				

*Some elements show more than one valency. A Roman numeral shows their valency in a bracket.

The rules that you have to follow while writing a chemical formula are as follows:

- the valencies or charges on the ion must balance.
- when a compound consists of a metal and a non-metal, the name or symbol of the metal is written first. For example: calcium oxide (CaO), sodium chloride (NaCl), iron sulphide (FeS), copper oxide (CuO) etc., where oxygen, chlorine, sulphur are non-metals and are written on the right, whereas calcium, sodium, iron and copper are metals, and are written on the left.
- in compounds formed with polyatomic ions, the ion is enclosed in a bracket before writing the number to indicate the ratio. In case the number of polyatomic ion is one, the bracket is not required. For example, NaOH.

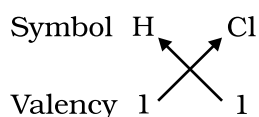
3.4.1 FORMULAE OF SIMPLE COMPOUNDS

The simplest compounds, which are made up of two different elements are called binary compounds. Valencies of some ions are given in Table 3.6. You can use these to write formulae for compounds.

While writing the chemical formulae for compounds, we write the constituent elements and their valencies as shown below. Then we must crossover the valencies of the combining atoms.

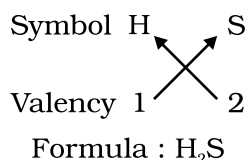
Examples

1. Formula of hydrogen chloride

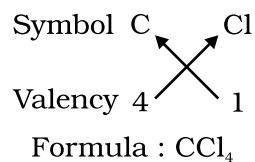


Formula of the compound would be HCl.

2. Formula of hydrogen sulphide

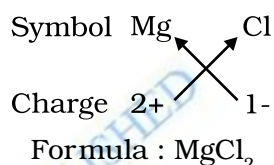


3. Formula of carbon tetrachloride



For magnesium chloride, we write the symbol of cation (Mg²⁺) first followed by the symbol of anion (Cl⁻). Then their charges are criss-crossed to get the formula.

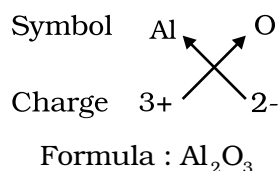
4. Formula of magnesium chloride



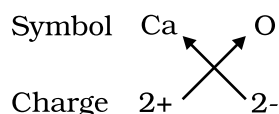
Thus, in magnesium chloride, there are two chloride ions (Cl⁻) for each magnesium ion (Mg²⁺). The positive and negative charges must balance each other and the overall structure must be neutral. Note that in the formula, the charges on the ions are not indicated.

Some more examples

(a) Formula for aluminium oxide:

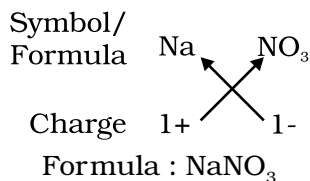


(b) Formula for calcium oxide:

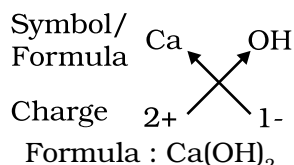


Here, the valencies of the two elements are the same. You may arrive at the formula Ca₂O₂. But we simplify the formula as CaO.

(c) Formula of sodium nitrate:

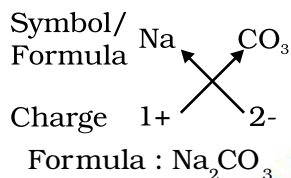


(d) Formula of calcium hydroxide:



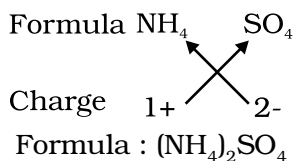
Note that the formula of calcium hydroxide is Ca(OH)₂ and not CaOH₂. We use brackets when we have two or more of the same ions in the formula. Here, the bracket around OH with a subscript 2 indicates that there are two hydroxyl (OH) groups joined to one calcium atom. In other words, there are two atoms each of oxygen and hydrogen in calcium hydroxide.

(e) Formula of sodium carbonate:



In the above example, brackets are not needed if there is only one ion present.

(f) Formula of ammonium sulphate:



Questions

- Write down the formulae of
 - sodium oxide
 - aluminium chloride
 - sodium sulphide
 - magnesium hydroxide
- Write down the names of compounds represented by the

following formulae:

- Al₂(SO₄)₃
 - CaCl₂
 - K₂SO₄
 - KNO₃
 - CaCO₃
- What is meant by the term chemical formula?
 - How many atoms are present in a
 - H₂S molecule and
 - PO₄³⁻ ion?

3.5 Molecular Mass and Mole Concept

3.5.1 MOLECULAR MASS

In section 3.2.2 we discussed the concept of atomic mass. This concept can be extended to calculate molecular masses. The molecular mass of a substance is the sum of the atomic masses of all the atoms in a molecule of the substance. It is therefore the relative mass of a molecule expressed in *atomic mass units (u)*.

Example 3.1 (a) Calculate the relative molecular mass of water (H₂O).
(b) Calculate the molecular mass of HNO₃.

Solution:

- (a) Atomic mass of hydrogen = 1u,
oxygen = 16 u

So the molecular mass of water, which contains two atoms of hydrogen and one atom of oxygen is

$$= 2 \times 1 + 1 \times 16$$

$$= 18 \text{ u}$$

- (b) The molecular mass of HNO₃ = the atomic mass of H + the atomic mass of N + 3 × the atomic mass of O
- $$= 1 + 14 + 48 = 63 \text{ u}$$

3.5.2 FORMULA UNIT MASS

The formula unit mass of a substance is a sum of the atomic masses of all atoms in a formula unit of a compound. Formula unit mass is calculated in the same manner as we calculate the molecular mass. The only

difference is that we use the word formula unit for those substances whose constituent particles are ions. For example, sodium chloride as discussed above, has a formula unit NaCl. Its formula unit mass can be calculated as–

$$1 \times 23 + 1 \times 35.5 = 58.5 \text{ u}$$

Example 3.2 Calculate the formula unit mass of CaCl_2 .

Solution:

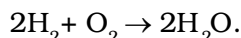
$$\begin{aligned} &\text{Atomic mass of Ca} \\ &+ (2 \times \text{atomic mass of Cl}) \\ &= 40 + 2 \times 35.5 = 40 + 71 = 111 \text{ u} \end{aligned}$$

Questions

1. Calculate the molecular masses of H_2 , O_2 , Cl_2 , CO_2 , CH_4 , C_2H_6 , C_2H_4 , NH_3 , CH_3OH .
2. Calculate the formula unit masses of ZnO , Na_2O , K_2CO_3 , given atomic masses of $\text{Zn} = 65 \text{ u}$, $\text{Na} = 23 \text{ u}$, $\text{K} = 39 \text{ u}$, $\text{C} = 12 \text{ u}$, and $\text{O} = 16 \text{ u}$.

3.5.3 MOLE CONCEPT

Take an example of the reaction of hydrogen and oxygen to form water:



The above reaction indicates that

- (i) two molecules of hydrogen combine with one molecule of oxygen to form two molecules of water, or
- (ii) 4 u of hydrogen molecules combine with 32 u of oxygen molecules to form 36 u of water molecules.

We can infer from the above equation that the quantity of a substance can be characterised by its mass or the number of molecules. But, a chemical reaction equation indicates directly the number of atoms or molecules taking part in the reaction. Therefore, it is more convenient to refer to the quantity of a substance in terms of the number of its molecules or atoms, rather than their masses. So, a new unit “mole” was introduced. One mole of any species (atoms,

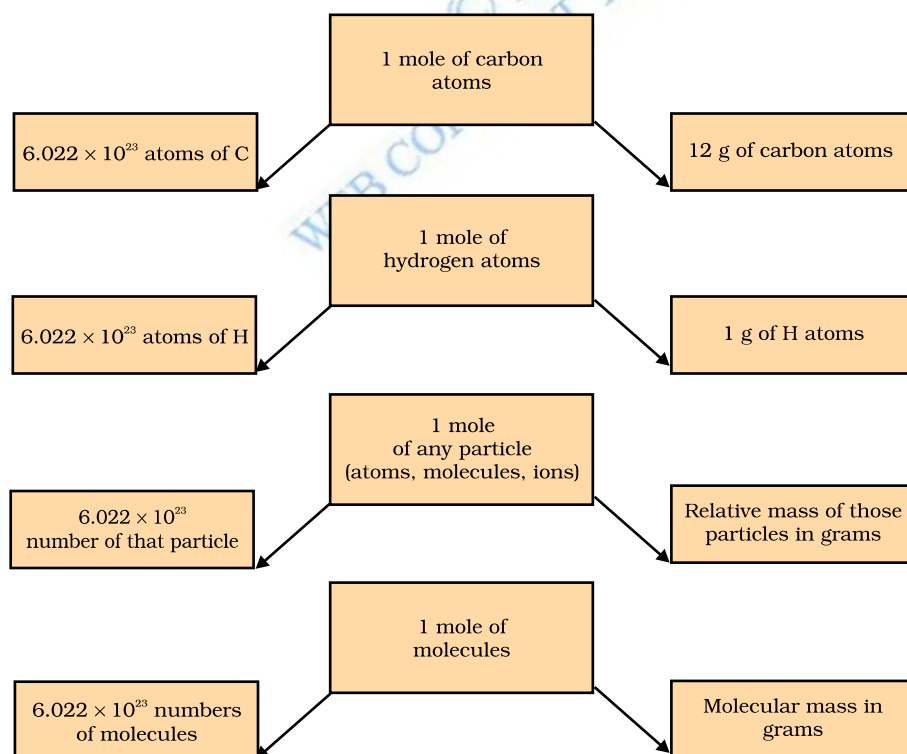


Fig. 3.5: Relationship between mole, Avogadro number and mass

molecules, ions or particles) is that quantity in number having a mass equal to its atomic or molecular mass in grams.

The number of particles (atoms, molecules or ions) present in 1 mole of any substance is fixed, with a value of 6.022×10^{23} . This is an experimentally obtained value. This number is called the Avogadro Constant or Avogadro Number (*represented by* N_0), named in honour of the Italian scientist, Amedeo Avogadro.

1 mole (of anything) = 6.022×10^{23} in number,
as, 1 dozen = 12 nos.

1 gross = 144 nos.

Besides being related to a number, a mole has one more advantage over a dozen or a gross. This advantage is that mass of 1 mole of a particular substance is also fixed.

The mass of 1 mole of a substance is equal to its relative atomic or molecular mass in grams. The atomic mass of an element gives us the mass of one atom of that element in atomic mass units (u). To get the mass of 1 mole of atom of that element, that is, molar mass, we have to take the same numerical value but change the units from 'u' to 'g'. Molar mass of atoms is also known as gram atomic mass. For example, atomic mass of hydrogen = 1 u. So, gram atomic mass of hydrogen = 1 g.

1 u hydrogen has only 1 atom of hydrogen

1 g hydrogen has 1 mole atoms, that is,

6.022×10^{23} atoms of hydrogen.

Similarly,

16 u oxygen has only 1 atom of oxygen,

16 g oxygen has 1 mole atoms, that is,

6.022×10^{23} atoms of oxygen.

To find the gram molecular mass or molar mass of a molecule, we keep the numerical value the same as the molecular mass, but simply change units as above from u to g. For example, as we have already calculated, molecular mass of water (H_2O) is 18 u. From here we understand that

18 u water has only 1 molecule of water,

18 g water has 1 mole molecules of water, that is, 6.022×10^{23} molecules of water.

Chemists need the number of atoms and molecules while carrying out reactions, and for this they need to relate the mass in grams

to the number. It is done as follows:

1 mole = 6.022×10^{23} number

= Relative mass in grams.

Thus, a mole is the chemist's counting unit.

The word "mole" was introduced around 1896 by Wilhelm Ostwald who derived the term from the Latin word *moles* meaning a 'heap' or 'pile'. A substance may be considered as a heap of atoms or molecules. The unit mole was accepted in 1967 to provide a simple way of reporting a large number– the massive heap of atoms and molecules in a sample.

Example 3.3

1. Calculate the number of moles for the following:

(i) 52 g of He (finding mole from mass)

(ii) 12.044×10^{23} number of He atoms (finding mole from number of particles).

Solutions:

No. of moles = n

Given mass = m

Molar mass = M

Given number of particles = N

Avogadro number of particles = N_0

(i) Atomic mass of He = 4 u

Molar mass of He = 4g

Thus, the number of moles

$$= \frac{\text{given mass}}{\text{molar mass}}$$

$$\Rightarrow n = \frac{m}{M} = \frac{52}{4} = 13$$

(ii) we know,

1 mole = 6.022×10^{23}

The number of moles

$$= \frac{\text{given number of particles}}{\text{Avogadro number}}$$

$$\Rightarrow n = \frac{N}{N_0} = \frac{12.044 \times 10^{23}}{6.022 \times 10^{23}} = 2$$

Example 3.4 Calculate the mass of the following:

- 0.5 mole of N_2 gas (mass from mole of molecule)
- 0.5 mole of N atoms (mass from mole of atom)
- 3.011×10^{23} number of N atoms (mass from number)
- 6.022×10^{23} number of N_2 molecules (mass from number)

Solutions:

- mass = molar mass $\times$ number of moles
 $\Rightarrow m = M \times n = 28 \times 0.5 = 14 \text{ g}$
- mass = molar mass $\times$ number of moles
 $\Rightarrow m = M \times n = 14 \times 0.5 = 7 \text{ g}$
- The number of moles, n

$$= \frac{\text{given number of particles}}{\text{Avogadro number}} = \frac{N}{N_0}$$

$$= \frac{3.011 \times 10^{23}}{6.022 \times 10^{23}}$$

$$\Rightarrow m = M \times n = 14 \times \frac{3.011 \times 10^{23}}{6.022 \times 10^{23}}$$

$$= 14 \times 0.5 = 7 \text{ g}$$
- $n = \frac{N}{N_0}$

$$\Rightarrow m = M \times \frac{N}{N_0} = 28 \times \frac{6.022 \times 10^{23}}{6.022 \times 10^{23}}$$

$$= 28 \times 1 = 28 \text{ g}$$

Example 3.5 Calculate the number of particles in each of the following:

- 46 g of Na atoms (number from mass)
- 8 g O_2 molecules (number of molecules from mass)
- 0.1 mole of carbon atoms (number from given moles)

Solutions:

- The number of atoms

$$= \frac{\text{given mass}}{\text{molar mass}} \times \text{Avogadro number}$$

$$\Rightarrow N = \frac{m}{M} \times N_0$$

$$\Rightarrow N = \frac{46}{23} \times 6.022 \times 10^{23}$$

$$\Rightarrow N = 12.044 \times 10^{23}$$
- The number of molecules

$$= \frac{\text{given mass}}{\text{molar mass}} \times \text{Avogadro number}$$

$$\Rightarrow N = \frac{m}{M} \times N_0$$
 atomic mass of oxygen = 16 u
 $\therefore$ molar mass of O_2 molecules
 $= 16 \times 2 = 32 \text{ g}$

$$\Rightarrow N = \frac{8}{32} \times 6.022 \times 10^{23}$$

$$\Rightarrow N = 1.5055 \times 10^{23}$$

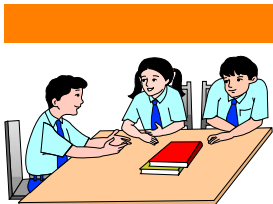
$$\approx 1.51 \times 10^{23}$$
- The number of particles (atom) = number of moles of particles $\times$ Avogadro number

$$N = n \times N_0 = 0.1 \times 6.022 \times 10^{23}$$

$$= 6.022 \times 10^{22}$$

Questions

- If one mole of carbon atoms weighs 12 grams, what is the mass (in grams) of 1 atom of carbon?
- Which has more number of atoms, 100 grams of sodium or 100 grams of iron (given, atomic mass of Na = 23 u, Fe = 56 u)?



What you have learnt

- During a chemical reaction, the sum of the masses of the reactants and products remains unchanged. This is known as the Law of Conservation of Mass.
- In a pure chemical compound, elements are always present in a definite proportion by mass. This is known as the Law of Definite Proportions.
- An atom is the smallest particle of the element that cannot usually exist independently and retain all its chemical properties.
- A molecule is the smallest particle of an element or a compound capable of independent existence under ordinary conditions. It shows all the properties of the substance.
- A chemical formula of a compound shows its constituent elements and the number of atoms of each combining element.
- Clusters of atoms that act as an ion are called polyatomic ions. They carry a fixed charge on them.
- The chemical formula of a molecular compound is determined by the valency of each element.
- In ionic compounds, the charge on each ion is used to determine the chemical formula of the compound.
- Scientists use the relative atomic mass scale to compare the masses of different atoms of elements. Atoms of carbon-12 isotopes are assigned a relative atomic mass of 12 and the relative masses of all other atoms are obtained by comparison with the mass of a carbon-12 atom.
- The Avogadro constant 6.022×10^{23} is defined as the number of atoms in exactly 12 g of carbon-12.
- The mole is the amount of substance that contains the same number of particles (atoms/ ions/ molecules/ formula units etc.) as there are atoms in exactly 12 g of carbon-12.
- Mass of 1 mole of a substance is called its molar mass.



Exercises

1. A 0.24 g sample of compound of oxygen and boron was found by analysis to contain 0.096 g of boron and 0.144 g of oxygen. Calculate the percentage composition of the compound by weight.
2. When 3.0 g of carbon is burnt in 8.00 g oxygen, 11.00 g of carbon dioxide is produced. What mass of carbon dioxide will

be formed when 3.00 g of carbon is burnt in 50.00 g of oxygen?
Which law of chemical combination will govern your answer?

3. What are polyatomic ions? Give examples.
4. Write the chemical formulae of the following.
 - (a) Magnesium chloride
 - (b) Calcium oxide
 - (c) Copper nitrate
 - (d) Aluminium chloride
 - (e) Calcium carbonate.
5. Give the names of the elements present in the following compounds.
 - (a) Quick lime
 - (b) Hydrogen bromide
 - (c) Baking powder
 - (d) Potassium sulphate.
6. Calculate the molar mass of the following substances.
 - (a) Ethyne, C_2H_2
 - (b) Sulphur molecule, S_8
 - (c) Phosphorus molecule, P_4 (Atomic mass of phosphorus = 31)
 - (d) Hydrochloric acid, HCl
 - (e) Nitric acid, HNO_3
7. What is the mass of—
 - (a) 1 mole of nitrogen atoms?
 - (b) 4 moles of aluminium atoms (Atomic mass of aluminium = 27)?
 - (c) 10 moles of sodium sulphite (Na_2SO_3)?
8. Convert into mole.
 - (a) 12 g of oxygen gas
 - (b) 20 g of water
 - (c) 22 g of carbon dioxide.
9. What is the mass of:
 - (a) 0.2 mole of oxygen atoms?
 - (b) 0.5 mole of water molecules?
10. Calculate the number of molecules of sulphur (S_8) present in 16 g of solid sulphur.
11. Calculate the number of aluminium ions present in 0.051 g of aluminium oxide.

(Hint: The mass of an ion is the same as that of an atom of the same element. Atomic mass of Al = 27 u)



Group Activity

Play a game for writing formulae.

Example 1 : Make placards with symbols and valencies of the elements separately. Each student should hold two placards, one with the symbol in the right hand and the other with the valency in the left hand. Keeping the symbols in place, students should criss-cross their valencies to form the formula of a compound.

Example 2 : A low cost model for writing formulae: Take empty blister packs of medicines. Cut them in groups, according to the valency of the element, as shown in the figure. Now, you can make formulae by fixing one type of ion into other.

For example:



Formula for sodium sulphate:

2 sodium ions can be fixed on one sulphate ion.

Hence, the formula will be: Na_2SO_4

Do it yourself :

Now, write the formula of sodium phosphate.