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3

ATOMS AND MOLECULES

In the previous chapter you learnt about matter. The idea of divisibility of matter was considered long back in India around 500 B.C. Maharishi Kanad, an Indian Philosopher discussed it in his Darshan (Vaisesik Darshan). He said if we go on dividing matter, we shall get smaller and smaller particles. A stage would come beyond which further division will not be possible. He named these particles as 'PARMANU'. This concept was further elaborated by another Indian philosopher, Pakudha Katyayan. Katyayan said that these particles normally exist in a combined form which gives us various forms of matter.

Around the same era, an ancient Greek philosopher Democritus (460 – 370 BC) and Leucippus suggested that if we go on dividing matter, a stage will come when further division of particles will not be possible. Democritus called these individual particles 'atoms' (which means indivisible). These ideas were based on philosophical considerations. Experimental work to validate these ideas could not be done till the eighteenth century. However, today we know what an atom is and how it is responsible for different properties of substances. In this chapter, we shall study about atoms and molecules and related aspects like atomic and molecular masses, mole concept and molar masses. We shall also learn how to write chemical formula of a compound.

OBJECTIVES

After completing this lesson you will be able to :

- *state the law of conservation of mass and law of constant proportions;*
- *list important features of Dalton's atomic theory;*
- *distinguish between atoms and molecules;*
- *define isotopic mass, atomic mass, and molecular mass;*

- *define the mole concept and molar mass;*
- *represent some molecules with the help of a formula;*
- *apply the mole concept to chemical reaction and show a quantitative relationship between masses of reactants and products and*
- *solve simple problems based on various concepts learnt.*

3.1 LAWS OF CHEMICAL COMBINATIONS

There was tremendous progress in Chemical Sciences after 18^{th} century. It arose out of an interest in the nature of heat and the way things burn. Major progress was made through the careful use of *chemical balance* to determine the change in mass that occurs in chemical reactions. The great French Chemist Antoine Lavoisier used the balance to study chemical reactions. He heated mercury in a sealed flask that contained air. After several days, a red substance mercury (II) oxide was produced. The gas remaining in the flask was reduced in mass. The remaining gas was neither able to support combustion nor life. The remaining gas in the flask was identified as nitrogen. The gas which combined with mercury was oxygen. Further he carefully performed the experiment by taking a weighed quantity of mercury (II) oxide. After strong heating, he found that mercury (II) oxide, red in colour, was decomposed into mercury and oxygen. He weighed both mercury and oxygen and found that their combined mass was equal to that of the mercury (II) oxide taken. Lavoisier finally came to the conclusion that ***in every chemical reaction, total masses of all the reactants is equal to the masses of all the products***. This law is known as the **law of conservation of mass.**

There was rapid progress in science after chemists began accurate determination of masses of reactants and products. French chemist Claude Berthollet and Joseph Proust worked on the ratio (by mass) of two elements which combine to form a compound. Through a careful work, Proust demonstrated the fundamental law of definite or constant proportions in 1808. **In a given chemical compound, the proportions by mass of the elements that compose it are fixed, independent of the origin of the compound or its mode of preparation.**

In pure water, for instance, the ratio of mass of hydrogen to the mass of oxygen is always 1:8 irrespective of the source of water. In other words, pure water contains 11.11% of hydrogen and 88.89% of oxygen by mass whether water is obtained from well, river or from a pond. Thus, if 9.0 g of water are decomposed, 1.0 g of hydrogen and 8.0 g of oxygen are always obtained. Furthermore, if 3.0 g of hydrogen are mixed with 8.0 g of oxygen and the mixture is ignited, 9.0 g of water are formed and 2.0 g of hydrogen remains unreacted. Similarly sodium chloride contains 60.66% of chlorine and 39.34% of sodium by mass whether we obtained it from salt mines or by crytallising it from water of ocean or inland salt seas or synthesizing it from its elements sodium and chlorine. Of course, the key word in this sentence is 'pure'. Reproducible experimental results are highlights of scientific thoughts. In fact modern science is based on experimental findings. **Reproducible results indirectly hint for a truth which is hidden**. Scientists always worked for findings this truth and in this manner many theories and laws were discovered. This search for truth plays an important role in the development of science.

3.2 DALTON'S ATOMIC THEORY

The English scientist John Dalton was by no means the first person to propose the existence of atoms, as we have seen in the previous section, such ideas date back to classical times. Dalton's major contribution was to arrange those ideas in proper order and give evidence for the existence of atoms. He showed that the mass relationship expressed by Lavoisier and Proust (in the form of law of conservation of mass and law of constant proportions) could be interpreted most suitably by postulating the existence of atoms of the various elements.

In 1803, Dalton published a new system of chemical philosophy in which the following statements comprise the atomic theory of matter:

1. Matter consists of indivisible atoms.
2. All the atoms of a given chemical element are identical in mass and in all other properties.
3. Different chemical elements have different kinds of atoms and in particular such atoms have different masses.
4. Atoms are indestructible and retain their identity in chemical reactions.
5. The formation of a compound from its elements occurs through the combination of atoms of unlike elements in small whole number ratio.

Dalton's fourth postulate is clearly related to the law of conservation of mass. Every atom of an element has a definite mass. Also in a chemical reaction there is rearrangement of atoms. Therefore after the reaction, mass of the product should remain the same. The fifth postulate is an attempt to explain the law of definite proportions. A compound is a type of matter containing the atoms of two or more elements in small whole number ratio. Because the atoms have definite mass, the compound must have the elements in definite proportions by mass.

John Dalton (1766-1844)

Fig. 3.1

The Dalton's atomic theory not only explained the laws of conservations of mass and law of constant proportions but also predicted the new ones. He deduced **the law of multiple proportions** on the basis of his theory. The law states that **when two elements form more than one compound, the masses of one element in these compound for a fixed mass of the other element are in the ratio of small whole numbers**. For example, carbon and oxygen form two compounds: Carbon monoxide and carbon dioxide. Carbon monoxide contains 1.3321 g of oxygen for each 1.000g of carbon, whereas carbon dioxide contains 2.6642 g of oxygen for 1.0000 g of carbon. In other words, carbon dioxide contains twice the mass of oxygen as is contained in carbon monoxide (2.6642 g = 2 × 1.3321 g) for a given mass of carbon. Atomic theory explains this by saying that carbon dioxide contains twice as many oxygen atoms for a given number of carbon atoms as does carbon monoxide. The deduction of *law of multiple proportions* from atomic theory was important in convincing chemists of the validity of the theory.

3.2.1 What is an Atom?

As you have just seen in the previous section that an atom is the smallest particle of an element that retains its (elements) chemical properties. An atom of one element is different in size and mass from the atoms of the other elements. These atoms were considered 'indivisible' by Indian and Greek 'Philosophers' in the beginning and the name 'atom' as mentioned earlier, emerged out of this basic philosophy. Today, we know that atoms are not indivisible. They can be broken down into still smaller particles although they lose their chemical identity in this process. But inspite of all these developments atom still remains a **building block** of matter.

3.2.2 What is the size of the atom?

Atoms are very small, they are smaller than anything that we can imagine or compare with. In order to have a feeling of size of an atom you can consider this example: **One teaspoon of water (about 1 mL) contains about three times as many atoms as Atlantic ocean contains teaspoons of water.** Also more than millions of atoms when stacked would make a layer barely as thick as this sheet of paper. Atoms of different elements not only differ in mass as proposed by Dalton but also they differ in size. Now question is why should we bother for the size, mass and other properties of an atom? The reason is simple, every matter we see around us is made of atoms. Is it rectangular, circular or spherical? It is difficult to imagine the real shape of an atom but for all practical purposes it is taken as spherical in size and that is why we talk of its radius. Since size is extremely small and invisible to our eyes, we adopt a scale of nanometer (1nm = 10^{-9} m) to express its size.

You can have a feeling of its size from the following table (Table 3.1).

Table 3.1 : Relative sizes

Radius (in m)	Example
10^{-10}	Atoms of hydrogen
10^{-4}	Grain of sand
10^{-1}	Water melon
0.2×10^{-1}	Cricket ball

You can not see atoms with your naked eyes but by using modern techniques, we can now produce magnified image of surface of elements showing atoms. The technique is known as Scanning Tunneling Microscopy (STM) (Fig. 3.2)

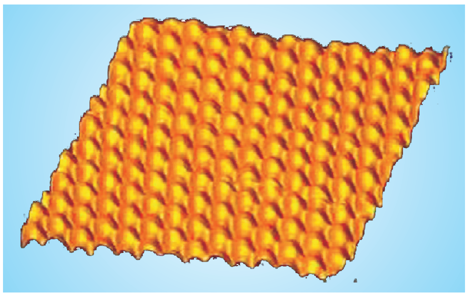


Fig. 3.2: *Image of Copper surface by STM technique. Atom can be seen in magnified image of surfae*

3.2.3 Atomic Mass

Dalton gave the concept of atomic mass. According to him, atoms of the same element have same atomic masses but atoms of different elements have different atomic masses. Since Dalton could not weigh individual atoms, he measured relative masses of the elements required to form a compound. From this, he deduced ***relative atomic masses***. For example, we can determine by experiment that 1.0000 g of hydrogen gas reacts with 7.9367 g of oxygen gas to form water. If we know formula of water, we can easily determine the mass of an oxygen atom relative to that of hydrogen atom.

Dalton did not have a way of determining the proportions of atoms of each element forming water during those days. He assumed the simplest possibility that atoms of oxygen and hydrogen were equal in number. From this assumption, it would follow that oxygen atom would have a mass that was 7.9367 times that of hydrogen atom. This in fact was not correct. We now know that in water number of hydrogen atoms is twice the number of oxygen atoms (formula of water being H_2O). Therefore, relative mass of oxygen atom must be $2\times7.9367 = 15.873$ times that of hydrogen atom. After Dalton, relative atomic masses of several elements were determined by scientists based on hydrogen scale. Later on, ***hydrogen based scale*** was replaced by a scale based on oxygen as it (oxygen) was more reactive and formed a large number of compounds.

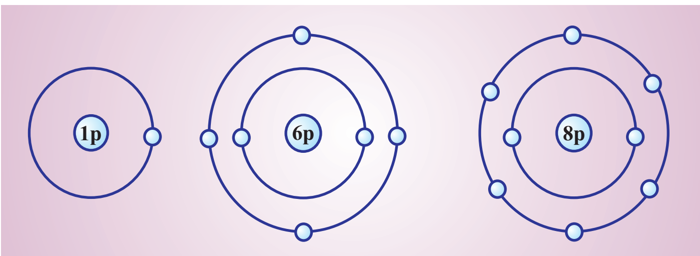
In 1961, C-12 (or $^{12}_{6}C$) atomic mass scale was adopted. This scale depends on measurement of atomic mass by an instrument called ***mass spectrometer***. Mass spectrometer invented early in 20^{th} century, allows us to determine atomic masses precisely. The masses of atoms are obtained by comparison with ***C-12 atomic mass scale***. In fact C-12 isotope is chosen as standard and arbitrarily assigned a mass of exactly ***12 atomic mass units.*** One atomic mass unit (amu), therefore, equals exactly one twelfth of mass of a carbon–12 atom, Atomic mass unit (amu) is now-a-days is written as unified mass unit and is denoted by the letter 'u'.

The relative atomic mass of an element expressed in atomic mass unit is called its ***atomic weight***. Now-a-days we are using *atomic mass* in place of atomic weight.

Further, you have seen that Dalton proposed that masses of all atoms in an element are equal. But later on it was found that all atoms of naturally occurring elements are not of the same mass. We shall study about such atoms in the following section. Atomic masses that we generally use in our reaction or in chemical calculations are ***average atomic masses*** which depend upon relative abundance of isotopes of elements.

3.2.4 Isotopes and Atomic Mass

Dalton considered an atom as an indivisible particle. Later researches proved that an atom consists of several fundamental particles such as : electrons, protons and neutrons. An electron is negatively charged and a proton is positively charged particle. Number of electrons and protons in an atom is equal. Since charge on an electron is equal and opposite to charge of a proton, therefore ***an atom is electrically neutral***. Protons remain in the nucleus in the centre of the atom, and nucleus is surrounded by negatively charged electrons.



(*a*) (*b*) (*c*)

Fig. 3.3: *Arrangement of electrons around nucleus in (a) hydrogen, (b) carbon and (c) oxygen atoms*

The number of protons in the nucleus is called ***atomic number*** denoted by Z. For example in Fig. 3.3, there are 8 protons in the oxygen nucleus, 6 protons in carbon nucleus and only one proton in hydrogen nucleus. Therefore atomic numbers of oxygen, carbon and hydrogen are 8,6 and 1 respectively. There are also neutral particles in the nucleus and they are called 'neutrons'. Mass of a proton and of a neutron is nearly the same.

Total mass of the nucleus = mass of protons + mass of neutrons

Total number of protons and neutrons is called ***mass number (A)***. By convention atomic number is written at the bottom of left corner of the symbol of the atom of a particular element and mass number is written at the top left corner. For example, symbol $^{12}_{6}C$ indicates that there is a total of 12 particles (nucleons) in the nucleus of a carbon atom, 6 of which are protons. Thus, there must be 12-6 = 6 neutrons. Similarly $^{16}_{8}O$ indicates 8 protons and 16 nucleons (8 protons + 8 neutrons). Since atom is electrically neutral, oxygen has 8 protons and 8 electrons in it. Further, atomic number (Z) differentiates the atom of one element from the atoms of the other elements.

An element may be defined as a substance where all the atoms have the same atomic number.

But the nuclei of all the atoms of a given element do not necessarily contain the same number of neutrons. For example, atoms of oxygen, found in nature, have the same number of protons which makes it different from other elements, but their neutrons (in nucleus) are different. This is the reason that the masses of atoms of the same element are different. For example, one type of oxygen atom contains 8 protons and 8 neutrons in one atom, second type 8 protons and 9 neutrons and the third type contains 8 protons and 10 neutrons. We represent these oxygen atoms as $^{16}_{8}O$, $^{17}_{8}O$ and $^{18}_{8}O$ respectively. ***Atoms of an element that have the same atomic number (Z) but different mass number (A) are called isotopes***. In view of difference in atomic masses of the same element, we take average atomic masses of the elements. This is calculated on the basis of the *abundance of the isotopes*. Atomic masses of some elements are provided in Table 3.2.

Example 3.1 : Chlorine is obtained as a mixture of two isotopes $^{35}_{17}Cl$ and $^{37}_{17}Cl$. These isotopes are present in the ratio of 3:1. What will be the average atomic mass of chlorine?

Solution : $^{35}_{17}Cl$ and $^{37}_{17}Cl$ are present in the ratio of 3:1 i.e. out of four atoms, 3 atoms are of mass 35 and one atom of mass 37. Therefore,

$$\text{Average atomic mass} = \frac{35\times3+37\times1}{4} = \frac{142}{4} = 35.5 \text{ u}$$

Thus, average atomic mass of chlorine will be 35.5u.

Table 3.2 : Atomic mass* of some common elements

Elements	Symbol	Mass (u)	Elements	Symbol	Mass (u)
Aluminium	Al	26.93	Magnesium	Mg	24.31
Argon	Ar	39.95	Manganese	Mn	54.94
Arsenic	As	74.92	Mercury	Hg	200.59
Barium	Ba	137.34	Neon	Ne	20.18
Boron	B	10.81	Nickel	Ni	58.71
Bromine	Br	79.91	Nitrogen	N	14.01
Caesium	Cs	132.91	Oxygen	O	16.00
Calcium	Ca	40.08	Phosphorus	P	30.97
Carbon	C	12.01	Platinum	Pt	195.09
Chlorine	Cl	35.45	Potassium	K	39.1
Chromium	Cr	52.00	Radon	Rn	(222)**
Cobalt	Co	58.93	Silicon	Si	23.09
Copper	Cu	63.56	Silver	Ag	107.87
Fluorine	F	19.00	Sodium	Na	23.00
Gold	Au	196.97	Sulphur	S	32.06
Helium	He	4.00	Tin	Sn	118.69
Hydrogen	H	1.008	Titanium	Ti	47.88
Iodine	I	126.90	Tungsten	W	183.85
Iron	Fe	55.85	Uranium	U	238.03
Lead	Pb	207.19	Vanadium	V	50.94
Lithium	Li	6.94	Xenon	Xe	131.30
			Zinc	Zn	65.37

* Atomic masses are average atomic masses covered upto second decimal places. In practice, we use rounded figures.

** Radioactive

INTEXT QUESTIONS 3.1

1. Name the scientists who proposed the law of conservation of mass and law of constant proportions.

2. 12 g of magnesium powder was ignited in a container having 20 g of pure oxygen. After the reaction was over, it was found that 12 g of oxygen was left unreacted. Show that it is according to law of constant proportions.

$$2Mg + O_2 \longrightarrow 2MgO$$

3.3 WHAT IS A MOLECULE?

Dalton proposed in his hypothesis that atoms react to form a molecule which he said as ‘compound atoms’. Today we know what a molecule is. ***A molecule is an aggregate of two or more than two atoms of the same or different elements in a definite arrangement.*** These atoms are held together by chemical forces or *chemical bonds*. (You will study details of molecules in unit of chemical bonding) ***An atom is the smallest particle of a substance but can not exist freely. Contrary to this, a molecules can be considered as the smallest particle of an element or of a compound which can exist alone or freely under ordinary conditions***. A molecule of a substance shows all chemical properties of that substance. To describe the chemical composition of a molecule we take the help of symbols of elements and formulas (described in sec 3.5). Oxygen molecule, with which we are familiar, is made of two atoms of oxygen and therefore it is a *diatomic molecule* (represented by O_2), hydrogen, nitrogen, fluorine, chlorine, bromine and iodine are other examples of diatomic molecules and are represented as H_2, N_2, F_2, Cl_2 , Br_2 and I_2 respectively (Fig. 3.4).

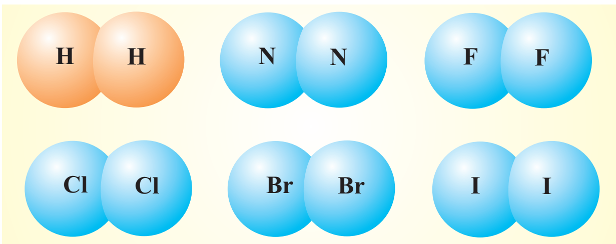
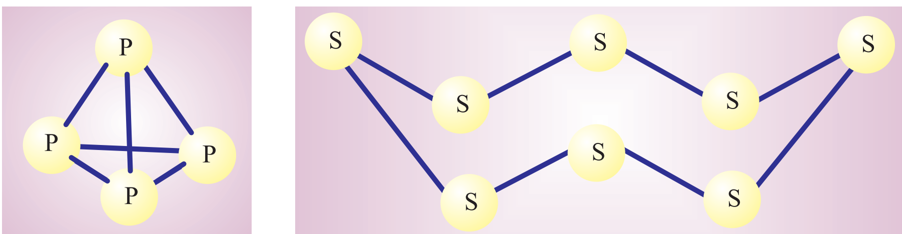


Fig. 3.4 : *Representation of diatomic molecules*

Some other elements exist as more complex molecules. Phosphorus molecule consists of four atoms (denoted by P_4) whereas sulphur exists as eight atom molecule (S_8) at ordinary temperature and pressure Fig. 3.5. A molecule made of four atoms is tetratomic molecule. Normally, molecules consisting of more than three or four atoms are considered under the category of ***polyatomic molecules***. Only a few years back, a form of carbon called buckminsterfullerene having molecular formula C_{60} was discovered which you will study later on in you higher classes.

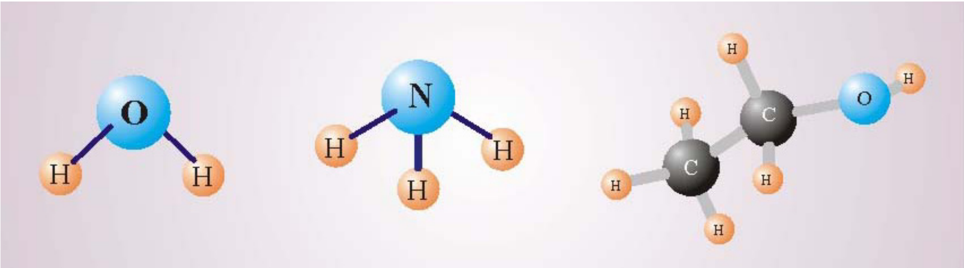


Structure of phosphorus molecule P_4 Structure of a sulphur molecule S_8

Fig. 3.5 : *Molecules of phosphorus and sulphur*

Molecules of compounds are composed of more than one kind of atoms. A familiar example is a water molecule which is composed of more than one kind of atoms.

In one water molecule, there are two atoms of hydrogen and one atom of oxygen. It is represented as H_2O. A molecule of ammonia consists of one nitrogen atom and three hydrogen atoms. A molecule of ethyl alcohol (C_2H_5OH) is composed of nine atoms (2 atoms of carbon, 6 atoms of hydrogen and one atom of oxygen) Fig. 3.6.



Molecule of water　　Molecule of ammonia　　Molecule of ethyl alcohol

Fig. 3.6 : *Molecules of water, ammonia and ethyl alcohol*

3.3.1 Molecular Mass

You have just read that a molecule can be represented in the form of a formula popularly known as ***molecular formula***. Molecular formula may be of an element or of a compound. ***Molecular formula of a compound is normally used for determining the molecular mass of that substance.*** If a substance is composed of molecule (for example : CO_2, H_2O or NH_3), it is easy to calculate the molecular mass. Molecular mass is the sum of atomic masses of all the atoms present in that molecule. Thus the ***molecular mass is the sum of atomic masses of all the atoms present in that molecule.*** The molecular mass of CO_2 is obtained as

$$\begin{array}{ll} C & 1\times 12.0\ u = 12.0\ u \\ 2\ O & 2\times 16.0\ u = 32.0\ u \\ \hline \multicolumn{2}{l}{\text{Mass of } CO_2 \ = \ 44.0\ u} \end{array}$$

Hence, we write molecular mass of CO_2 = 44.0 u

Similarly, we obtain molecular mass of NH_3 as follows:

$$\begin{array}{ll} N & 1\times 14.0\ u = 14.0\ u \\ 3\ H & 3\times 1.08\ u = 3.24\ u \\ \hline \multicolumn{2}{l}{\text{Mass of } NH_3 \ = \ 17.24\ u} \end{array}$$

Molecular mass of ammonia, NH_3 = 17.24 u

For substances which are not molecular in nature, we talk of *formula mass*. For example, sodium chloride (denoted by formula, NaCl) is an ionic substance. For this, we will calculate formula mass, similar to molecular mass. In case of sodium chloride, NaCl;

$$\begin{aligned} \text{Formula mass} &= \text{mass of 1 Na atom} + \text{mass of 1 Cl atom} \\ &= 23\ u + 35.5\ u = 58.5\ u \end{aligned}$$

You will learn more about the molecular and ionic compounds in detail later.

INTEXT QUESTIONS 3.2

1. Nitrogen forms three oxides : NO, NO_2 and N_2O_3. Show that it obeys law of multiple proportions.
2. Atomic number of silicon is 14. If there are three isotopes of silicon having 14, 15 and 16 neutrons in their nuclei, what would be the symbol of the isotope?
3. Calculate molecular mass of the compounds whose formulas are provided below:

   C_2H_4, H_2O and CH_3OH

3.4 MOLE CONCEPT

When we mix two substances, we get one or more new substances. For example, when we mix hydrogen and oxygen and ignite the mixture, we get a new substance-water. This can be represented in the form of a chemical equation,

$$2H_2\ (g) + O_2\ (g) \longrightarrow 2H_2O\ (l)$$

In the above equation, 2 molecules (four atoms) of hydrogen react with 1 molecule (2 atoms) of oxygen and give two molecules of water. We always like to know how many atoms/molecules of a particular substance would react with atoms/molecules of another substance in a chemical reaction. No matter how small they are. The solution to this problem is to have a convenient unit. Would you not like to have a convenient unit? Definitely a unit for counting of atoms/molecules present in a substance will be desirable and convenient as well. This chemical counting unit of atoms and molecules is called ***mole***.

The word mole was, apparently introduced in about 1896 by Wilhelm Ostwald who derived the term from the Latin word 'moles' meaning a 'heap' or 'pile'. The mole whose symbol is 'mol' is the SI (international system) base unit for measuring ***amount of substance***. It is defined as follows:

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

In simple words, mole is the number of atoms in exactly 0.012 kg (12 grams) of C-12. Although mole is defined in terms of carbon atoms but the unit is applicable to any substance just as 1 dozen means 12 or one gross means 144 of anything. Mole is scientist's ***counting unit*** like dozen or gross. By using mole, scientists (particularly chemists) count atoms and molecules in a given substance. Now it is experimentally found that the number of atoms contained in exactly 12 g of C-12 is 602,200 000 000 000 000 000 000 or 6.022×10^{23}. This number is called ***Avogadro's number*** in honour of Amedeo Avogadro, an Italian lawyer and physicist. When this number is divided by 'mole' it becomes a constant and is known as *Avogadro's constant* denoted by symbol, $N_A = 6.02 \times 10^{23}$ mol^{-1}. We have seen that

$$\text{Atomic mass of C} = 12 \text{ u}$$

$$\text{Atomic mass of He} = 4 \text{ u}$$

We can see that one atom of carbon is three times as heavy as one atom of helium. On the same logic 100 atoms of carbon are three times as heavy as 100 atoms of helium. Similarly 6.02×10^{23} atoms of carbon are three times as heavy as 6.02×10^{23} atoms of helium. But 6.02×10^{23} atoms of carbon weigh 12 g, therefore 6.02×10^{23} atoms of helium will weigh $1/3 \times 12$g = 4 g. We can take a few more examples of elements and can calculate the mass of one mole atoms of that element.

3.4.1 Molar Mass

Mass of one mole of a substance is called its molar mass. A substance may be an element or a compound. Mass of one mole atoms of oxygen means mass of 6.02×10^{23} atoms of oxygen. It is found that one mole atoms of oxygen weighs 16.0 g. When we say one mole molecules of oxygen that means 6.02×10^{23} molecules of oxygen (O_2). One mole molecules of oxygen will weigh 32.0 g. Thus,

$$\text{Mass of one mole atoms of oxygen} = 16 \text{ g mol}^{-1}$$

$$\text{Mass of one mole molecules of oxygen} = 32 \text{ g mol}^{-1}$$

When it is not clear whether we are asking for one mole of atoms or one mole of molecules then we take natural form of that substance. For example, one mole of oxygen means one mole of oxygen molecules as oxygen occurs in the form of molecules in nature. In case of compounds, the same logic is applicable. For example, one mole of water means one mole molecules of water which weighs 18 g. Numerically one mole of a substance is equal to atomic or molecular mass of that substance expressed in grams.

Remember, molar mass is always expressed in the unit of g/mol or g mol^{-1}.

For example,

$$\text{Molar mass of nitrogen } (N_2) = 28 \text{ g mol}^{-1}$$

$$\text{Molar mass of chlorine } (Cl_2) = 71 \text{ g mol}^{-1}$$

Table 2.3 Provides molecular and molar mass of a few common substances.

Table 3.3 : Molecular and molar masses

Formula	Molecular Mass (u)	Molar mass (g/mol)
O_2(oxygen)	32.0	32.0
Cl_2(chlorine)	71.0	71.0
P_4 (phosphorus)	123.9	123.9
CH_4 (methane)	16.00	16.0
NH_3 (ammonia)	17.0	17.0
HCl (hydrochloric acid gas)	36.5	36.5
CO_2 (carbon dioxide)	44.0	44.0
SO_2 (sulphur dioxide)	64.0	64.0
C_2H_5OH (ethyl alcohol)	46.0	46.0
C_6H_6 (benzene)	78.0	78.00

Example 3.2 : How many grams are there in 3.5 mol of oxygen?

Solution : For converting mole into mass in grams and vice-versa, we always need a relationship between mass and mole.

Molar mass of oxygen (O_2) = 32 g mol^{-1}

Therefore, number of grams of oxygen in 3.5 mol of it

$$= 3.5 \text{ mol of oxygen} \times 32.0 \text{ g mol}^{-1}$$

$$= 112.0 \text{ g of oxygen}$$

Example 3.3 : Find out number of molecules in 27 g of water.

Solution: Mole concept provides a relationship between number of particles and their mass. Thus it is possible to calculate the number of particles in a given mass.

$$\text{Number of mole of } H_2O = \frac{\text{Mass of water } (H_2O)}{\text{Molar mass of } H_2O}$$

$$= \frac{27\text{g}}{18 \text{ g mol}^{-1}} = \frac{3}{2} \text{ mol} = 1.5 \text{ mol}$$

Since 1 mol of water contains 6.02×10^{23} molecules.

Therefore, 1.5 mol of water contains = 6.02×10^{23} molecules mol^{-1} $\times 1.5$ mol

$$= 9.03 \times 10^{23} \text{ molecules of water}$$

INTEXT QUESTIONS 3.3

1. Work out a relationship between number of molecules and mole.
2. What is molecular mass? In what way it is different from the molar mass?
3. Consider the reaction

   $C\ (s) + O_2\ (g) \longrightarrow CO_2\ (g)$

   18 g of carbon was burnt in oxygen. How many moles of CO_2 is produced?
4. What is the molar mass of NaCl?

3.5 WRITING CHEMICAL FORMULA OF COMPOUNDS

As you are aware, a compound is made of two or more than two elements combined in a definite proportion by mass (law of constant proportions). Thus, the number of combining atoms in a compound is fixed. The elements are represented by their symbols (e.g. H for hydrogen, Na for sodium). Similarly a compound is also represented by a shorthand notation known as formula or ***chemical formula***. The formula of a compound indicates **(i) elements constituting the compound and (ii) the number of each constituent element. In other word, the formula of a compound also represents its chemical composition.** The atoms of elements constituting a compound are indicated by their symbols and their number is indicated as a **subscript** on the right hand bottom of the symbol. For example, in the formula of water, H_2O, two atoms of hydrogen are indicated as subscript '2', while oxygen is shown without writing any subscript, which means that the number of oxygen atom is just one.

3.5.1 Valency and Formulation

Every element has a definite capacity to combine with other elements. ***This combining capacity of an element is called its valency.*** You will learn very soon that this combining capacity of elements depends on the electronic configuration of elements. Valencies of a few elements are given in Table 3.4.

Table 3.4 : Valency of elements

Elements	Symbol	Valency	Elements	Symbol	Valency
Hydrogen	H	1	Phosphorus	P	5
Oxygen	O	2	Sodium	Na	1
Carbon	C	4	Magnesium	Mg	2
Nitrogen	N	3	Calcium	Ca	2
Chlorine	Cl	1	Aluminium	Al	3
Bromine	Br	1	Iron	Fe	2
Iodine	I	1	Barium	Ba	2

Most of the simple compounds are made of two elements. Such compounds are called ***binary compounds***. It is easy to write formula of such compounds. When a metal combines with a non-metal, the symbol of the metal-element is written on the left hand side and that of the non-metal element on right hand side. (If both are non-metal, we write more electronegative* element on the right hand side). In naming a compound, the first element is written as such and the name of the second element i.e. more electronegative element, changes its ending to 'ide'. For writing chemical formula, we have to write valencies as shown below and then cross over the valencies of the combining atoms. Formula of the compounds resulting from carbon and chlorine, hydrogen and oxygen, and hydrogen and chlorine can be written as follows:

Element	C	Cl	H	O	H	Cl
Valency	4	1	1	2	1	1
Formula	CCl_4		H_2O		HCl	

Some other examples for writing formula of compounds CaO, NaCl and NH_3 can also be taken for more clarity.

Element	Ca	O	Na	Cl	N	H
Valency	2	2	1	1	3	1
Formula	CaO		NaCl		NH_3	

Thus, we can write formulas of various compounds if we know elements constituting them and their valencies.

Valency, as mentioned, depends on the electronic configuration and nature of the elements. Sometimes an element shows more than one type of valency. We say element has ***variable valency.*** For example nitrogen forms several oxides : N_2O, N_2O_2, N_2O_3, N_2O_4 and N_2O_5. If we take valency of oxygen equal to 2, then valency of nitrogen in the oxides will be 1,2,3,4 and 5 respectively. Valencies are not always fixed. Similar to nitrogen, phosphorus also shows valencies 3 and 5 as reflected in compounds PBr_3 and P_2O_5. In these compounds, there are more than one atom. In such cases, number of atoms is indicated by attaching a numerical prefix (mono, di, tri, etc) as mentioned in Table 2.5.

Table 3.5 : Numerical Prefixes

Number of atoms	Prefix	Example
1	Mono	carbon monoxide, CO
2	Di	carbon dioxide, CO_2
3	Tri	phosphorus trichloride, PCl_3
4	Tetra	carbon tetrachloride, CCl_4
5	Penta	Dinitrogen pentoxide, N_2O_5

Here you would notice that '–o' or '–a' at the end of the prefix is often dropped before another vowel, e.g. monoxide, pentoxide. There is no gap between numerical prefix and the name of the element. The prefix mono is usually dropped for the first element. When hydrogen is the first element in the formula, no prefix is added before hydrogen irrespective of the number. For example, the compound H_2S is named as hydrogen sulphide and not as dihydrogen sulphide.

Thus, we have seen that writing formula of a binary compound is relatively easy. However, when we have to write formula of a compound which involves more than two elements (i.e. of a polyatomic molecule), it is somewhat a cumbersome task. In the following section we shall consider formulation of more difficult compounds.

You will learn later on that there are basically two types of compounds: covalent compounds and ionic or electrovalent compounds. H_2O and NH_3 are covalent compounds. NaCl and MgO are ionic compounds. An ionic compound is made of two charged constituents. One positively charged and other negatively charged. In case of NaCl, there are two ions : Na^+ and Cl^- ion. Charge of these ions in case of electrovalent compound is used for writing formula. It is easy to write formula of an ionic compound only if there is one metal and one non-metal as in the case of NaCl and MgO. If there are more than two elements in an ionic compound, formulation will be a little difficult and in that situation we should know charge of cations and anions.

3.5.2 Formulation of Ionic compounds

Formulation of an ionic compound is easy when we know charge of cation and anion. Remember, in an ionic compound, sum of the charge of cation and anion should be equal to zero. A few examples of cations and anions with their charges are provided in Table 3.6.

Table 3.6 : Charges of some common cations and anions which form ionic compounds

Anions	Charge	Cations	Charge
Chloride ion, Cl^-	–1	Potassium ion, K^+	+1
Nitrate ion, NO_3^-	–1	Sodium ion, Na^+	+1
Hydroxide ion, OH^-	–1	Ammonium ion, NH_4^+	+1
Bicarbonate ion, HCO_3^-	–1	Magnesium ion, Mg^{2+}	+2
Nitrite ion, NO_2^-	–1	Calcium ion, Ca^{2+}	+2
Acetate ion, CH_3COO^-	–1	Lead ion, Pb^{2+}	+2
Bromide ion, Br^-	–1	Iron ion (ous), Fe^{2+}	+2
Iodide ion, I^-	–1	Zinc ion, Zn^{2+}	+2

Sulphite ion, SO_3^{2-}	–2	Copper ion (cupric), Cu^{2+}	+2
Carbonate ion, CO_3^{2-}	–2	Mercury ion (Mercuric), Hg^{2+}	+2
Sulphate ion, SO_4^{2-}	-2	Iron (ic) ion, Fe^{3+}	+3
Sulphide ion, S^{2-}	–2	Aluminium ion, Al^{3+}	+3
Phosphate ion, PO_4^{3-}	–3	Potassium ion, K^+	+1
		Sodium ion, Na^+	+1

Suppose you have to write formula of sodium sulphate which is made of Na^+ and SO_4^{2-} ions. For this the positive and negative charge can be crossed over to give subscripts. The purpose of this crossing over of charges is to find the number of ions required to equate the number of positive and negative charges.

Ion	Na^+	SO_4^{2-}
Charge	1 (positive)	2 (negative)

This gives the formula of sodium sulphate as Na_2SO_4. We can check the charge balance as follows

$$\left.\begin{aligned} 2Na^+ &= 2\times(+1) = +2 \\ 1SO_4^{2-} &= 1\times(-2) = -2 \end{aligned}\right] = 0$$

Thus the compound, Na_2SO_4 is electrically neutral.

Now it is clear that digit showing charge of cation goes to anion and digit showing charge of anion goes to cation. For writing formula of calcium phosphate we take charge of each ion into consideration and write the formula as discussed above.

$$\left(Ca^{2+}\right)_3 \left(PO_4^{3-}\right)_2 = Ca_3\left(PO_4\right)_2$$

Writing formula of a compound comes only by practice, therefore write formulas of as many ionic compounds as possible based on the guidelines given above.

INTEXT QUESTIONS 3.4

1. Write the name of the expected compound formed between
   (i) hydrogen and sulphur
   (ii) nitrogen and hydrogen
   (iii) magnesium and oxygen

2. Propose the formulas and names of the compounds formed between
   (i) potassium and iodide ions
   (ii) sodium and sulphate ions
   (iii) aluminium and chloride ions
3. Write the formula of the compounds formed between
   (i) Hg^{2+} and Cl^-
   (ii) Pb^{2+} and PO_4^{3-}
   (iii) Ba^{2+} and SO_4^{2-}

WHAT YOU HAVE LEARNT

- According to ***law of constant proportions***, a sample of a pure substance always consists of the same elements combined in the same proportion by mass.
- When an element combines with another element and forms more than one compound, then different masses of the one element that combine with the fixed mass of another element are in the ratio of simple whole number or integer. This is the ***law of multiple proportions***.
- John Dalton introduced the idea of an atom as an indivisible particle. An atom is the smallest particle of an element which shows all the properties of that element. An atom can not exist freely and therefore remains in a combined state.
- A molecule is the smallest particle of an element or of a compound which shows all properties of that substance and can exist freely under ordinary conditions.
- A molecule can be represented in the form of a chemical formula using symbols of elements that constitute it.
- Composition of any compound can be represented by its chemical formula.
- Atom of the isotope C-12 is assigned atomic mass unit of 12 and the relative atomic masses of all other atoms of elements are obtained by comparing them with it.
- **The mole is the amount of a substance which contains the same number of particles (atoms, ions or molecule)** as there are atoms in exactly 0.012 kg of carbon-12.
- Avogadro's number is defined as the number of atoms in exactly 0.012 kg (or 12 g) of C-12 and is equal to 6.02×10^{23}. Avogadro's constant is written as 6.02×10^{23} mol^{-1}.

- Mass of one mole atoms or one mole molecules or one mole of formula unit of a substance is its **molar mass**.
- The composition of any compound can be represented by its formula. For writing the formula of a compound, valence or valency of an element is used. This is normally done in case of covalent compounds.
- Valency is the combining capacity of an element and is related to its electronic configuration.
- In ionic compounds, the charge on each ion is used to determine the chemical formula of the compound.

TERMINAL EXERCISE

1. Describe the following:
   (a) Law of conservation of mass
   (b) Law of constant proportions
   (c) Law of multiple proportions
2. What is the atomic theory proposed by John Dalton? What changes have taken place in the theory during the last two centuries?
3. Write the number of protons, neutrons and electron in each of the following isotopes :

   $^{2}_{1}H$, $^{18}_{8}O$, $^{19}_{9}F$, $^{40}_{20}Ca$
4. Boron has two isotopes with masses 10.13 u and 11.01 u and abundance of 19.77% and 80.23% respectively. What is the average atomic mass of boron?
   (Ans. 10.81 u)
5. Give symbol for each of the following isotopes:
   (a) Atomic number 19, mass number 40
   (b) Atomic number 7, mass number 15
   (c) Atomic number 18, mass number 40
   (d) Atomic number 17, mass number 37
6. How does an element differ from a compound? Explain with suitable examples.
7. Charge of one electron is 1.6022×10^{-19} coulomb. What is the total charge on 1 mol of electrons?

8. How many molecules of O_2 are in 8.0 g of oxygen? If the O_2 molecules were completely split into O (Oxygen atoms), how many mole of atoms of oxygen would be obtained?

9. Assume that human body is 80% water. Calculate the number of molecules of water that are present in the body of a person whose weight is 65 kg.

10. Refer to atomic masses given in the Table (3.2) of this chapter. Calculate the molar masses of each of the following compounds :

    HCl, NH_3, CH_4, CO and NaCl

11. Average atomic mass of carbon is 12.01 u. Find the number of moles of carbon in (a) 2.0 g of carbon. (b) 8.0 g of carbon.

12. Classify the following molecules as di, tri, tetra, penta and hexa atomic molecules:

    H_2, P_4, SF_4, SO_2, PCl_3, CH_3OH, PCl_5, HCl

13. What is the mass of

    (a) 6.02×10^{23} atoms of oxygen

    (b) 6.02×10^{23} molecules of P_4

    (c) 3.01×10^{23} molecules of O_2

14. How many atoms are present in:

    (a) 0.1 mol of sulphur

    (b) 18.g of water (H_2O)

    (c) 0.44 g of carbon dioxide (CO_2)

15. Write various postulates of Dalton's atomic theory.

16. Convert into mole:

    (a) 16 g of oxygen gas (O_2)

    (b) 36 g of water (H_2O)

    (c) 22 g of carbon dioxide (CO_2)

17. What does a chemical formula of a compound represents?

18. Write chemical formulas of the following compounds:

    (a) Copper (II) sulphate

    (b) Calcium fluoride

    (c) Aluminium bromide

    (d) Zinc sulphate

    (e) Ammonium sulphate

ANSWERS TO INTEXT QUESTIONS

3.1

(i) Lavoisier proposed the law of conservation of mass and Proust proposed the law of constant proportions

(ii) In container, 12g of Oxygen was left unreacted. Therefore, amount of unreacted Oxygen = (20 – 12)g = 08g. Thus 12g of magnesium reacted with 8g of oxygen in the ratio 12:8. This is what we expected for MgO i.e 24g of Mg reacted with 16g of Oxygen or 12g of Mg will react with 8g of Oxygen.

3.2

(i) Atomic mass of nitrogen is 14u and that of oxygen is 16u.

In NO, 14g of nitrogen reacted with 16g of oxygen

In NO_2, 14g of nitrogen reacted with 32g of oxygen

In N_2O_3, 28g of nitrogen reacted with 48g of oxygen

or

14g of nitrogen reacted with 24g of Oxygen.

Threfore, amount of oxygen which reacts with 12g of nitrogen in case of NO, NO_2 and N_2O_3 will be in the ratio of 16:32:24 or 2:4:3. This proves the law of multiple proportions.

(ii) Atomic number of Si is 14

Mass number of silicon atoms having 14,15 and 16 neutrons will be 28,29 and 30 respectively and therefore symbols of istopes of silicon will be

$${}^{28}_{14}\text{Si} \qquad {}^{29}_{14}\text{Si} \qquad \text{and} \qquad {}^{30}_{14}\text{Si}$$

(iii) Molecular mass of C_2H_4 = mass of two atom of carbon + mass of 4 atom of hydrogen

$$= 2 \times 12u + 4 \times 1u = 28u$$

Molecular mass of H_2O = mass of two atoms of hydrogen + mass of one atom of oxygen

$$= 2 \times 1u + 1 \times 16u = 18u$$

Molecular mass of CH_3OH = mass of one atom of carbon + mass of 4 atoms of hydrogen + mass of one atom of oxygen

$$= 1 \times 12u + 4 \times 1u + 1 \times 16u = 32u$$

3.3

1. 1 mole of a substance contains 6.023×10^{23} molecules of that substance i.e 1 mole of a substance = 6.023×10^{23} molecules of that substance.

2. Molecular mass is the sum of atomic masses of all the atoms present in that molecule.

Molecular mass is the mass of one molecule whereas molar mass is the mass of 1 mol or 6.023×10^{23} elementary entities (atoms, molecules, ions)

3.

$C(s)$	+	O_2	$\longrightarrow$	CO_2
1 mol of C (12g)		1 mol of O_2(32g)		1 mol of CO_2 (44g)

12 g of carbon gives 1 mole of CO_2

18g of carbon will give 1.5 mole of CO_2

4. Molar mass of NaCl = (23.0 + 35.5)g mol^{-1}

= **58.5** g mol^{-1}

3.4

1. (i) H_2S

(ii) NH_3

(iii) MgO

2. (i) KI, Potassium iodide

(ii) Na_2SO_4, Sodium Sulphate

(iii) $AlCl_3$, Aluminium Chloride

3. (i) $HgCl_2$

(ii) $Pb_3(PO_4)_2$

(iii) $BaSO_4$