2

CHEMICAL ARITHMATICS

We know that atoms of different elements combine in simple whole-number ratios to form molecules. For example, hydrogen and oxygen atoms combine in the mass ratio of 1 : 8 and form water, H_2O. However, it is impossible to deal with individual atoms because they are so tiny that we can neither see nor weigh them. Therefore, we must increase the size of these quantities to the point where we can see them and weigh them. With the help of **mole concept** it is possible to take a desired number of atoms/molecules by weighing (please refer to lesson-1). Now, in order to study chemical compounds and reactions in the laboratory, it is necessary to have adequate knowledge of the quantitative relationship among the amounts of the reacting substances that take part and products formed in the chemical reaction. This relationship is know as stoichiometry. **Stoichiometry** (derived from the Greek *Stoicheion* = element and *metron* = measure) is the term we use to refer to all the quntatitative aspects of chemical compounds and reactions. In the present lesson, you will see how chemical formulae* are determined and how chemical equations prove useful in predicting the proper amounts of the reactants that must be mixed to carry out a complete reaction. In other words we can take reactants for a reaction in such a way that none of the reacting substances is in excess. This aspect is very vital in chemistry and has wide application in industries.

Objectives

After reading this lesson, you will be able to :

- define empirical and molecular formulae;
- differentiate between empirical and molecular formulae;
- calculate precentage by mass of an element in a compound and also work out empirical formula from the percentage composition;
- establish relationship between mole, mass and volume;
- calculate the amount of substances consumed or formed in a chemical reaction using a balanced equation and mole concept, and

* Formulae is plural of formula

- Explain that the amount of limiting reagent present initially limits the amount of the products formed.

2.1 Molcular and Empirical Formulae

In your previous classes, you have studied how to write chemical formula of a sustance. For example, water is represented by H_2O, carbon dioxide is represented by CO_2, methane is represented by CH_4, dinitrogen penta oxide is represented by N_2O_5, and so on. You are aware, formula for a molecule uses a symbol and subscript number to indicate the number of each kind of atoms present in the molcule (subscript 1 is always omitted). Such a formula is called **molecular formula** as it represents a molecule of a substance. A molecule of water consists of two hydrogen atoms and one oxygen atom. So its molecular formula is written as H_2O. Thus a **molecular formula shows the actual number of atoms of different elements in a molecule of a compound.**

There is another kind of formula, the empirical formul of a compound, which gives only relative number of atoms of different elements. These numbers are expressed as the simplest ratio. For example, empirical formula of glucose, which consists of carbon, hydrogen and oxygen in the ratio of 1:2:1 is CH_2O (empirical formulae are also called simplest formulae). Molecular formula of a substance is always an integral multiple of its empirical formula (i.e. molecular formula = X_n where X is empirical formula and n is an integer). For example molecular formula of glucose is $C_6H_{12}O_6$ which is 6 × its empirical formula. Thus, while empirical formula gives only a ratio of atoms, the molecular formula gives the actual number of atoms of each element in an individual molecule. In some cases the ratio of atoms shown in a molecular formula cannot be reduced to smaller integers. In such cases molecular and empirical formulae are the same, for example, sucrose $C_{12}H_{22}O_{11}$ which is popularly known as cane-sugar. In case of certain elements, a molecule consists of several atoms for example P_4, S_8, etc. In such cases, empirical formula will be symbol of the element only.

As you know, common salt, which is chemically called sodium chloride is represented as NaCl. This salt is ionic in nature and does not exist in molecular form. Therefore, NaCl is its empirical formula which shows that sodium and chlorine atoms are present in NaCl in the ratio of 1:1. Similar is the case with all ionic substanes. KCl, $NaNO_3$, MgO are examples of empirical formulae as these are all ionic compounds. Table 2.1 provides a few more examples.

Table 2.1 Molecular and Empirical Formulae

Substance	Molecular formula	Empirical formula
Ammonia	NH_3	NH_3
Carbon dioxide	CO_2	CO_2
Ethane	C_2H_6	CH_3
Fructose	$C_6H_{12}O_6$	CH_2O
Sulphur	S_8	S
Benzene	C_6H_6	CH
Sodium chloride	—	NaCl
Calcium oxide	—	CaO

2.2 Chemical Composition and Formulae

How much carbon is present in one kilogram of methane whose molecular formula is CH_4? How much nitrogen is present in one kilogram of ammonia, NH_3? If we have prepared a substance that is made of 58.8% carbon, 28.4% oxygen, 8.28% nitrogen and 6.56% hydrogen, what is its empirical formula? You have studied **atomic masses, formulae,** and the **mole concept.** Can we solve the problem using these basic concepts? The answer is 'yes'. Atomic masses, formulae and the mole concept are the basic tools needed to solve such problems. What is percentage composition? Let us take up this aspect in a little detail and try to understand.

2.2.1 Percentage Composition

If we know the formula of a compound, we can find out how much of each of the elements is present in a given quantity of the compound. Aluminium is obtained from its oxide. Al_2O_3 (which is found as the ore, bauxite). From the formula we can calculate how much aluminium can be obtained, at least in prinicple, from a given amount of aluminium oxide. Calculation is done by making use of the idea of **percentage composition**

Percentage mass of an element in a compound

$$= \frac{\text{mass of element in one molecular formula or in one empirical formula}}{\text{molecular mass or empirical formula mass of compound}} \times 100$$

$$= \frac{\text{Mass of element in 1mol of compound}}{\text{Molar mass of compound}} \times 100$$

Let us calculate percentage composition of aluminium oxide, Al_2O_3

$$\text{Pecentage of aluminium} = \frac{\text{Mass of aluminium in 1 mol } Al_2O_3}{\text{Molar mass of } Al_2O_3} \times 100$$

Molar mass of Al_2O_3= (2 × 27.0) g + (3 × 16.0) g = 102.0 g

Since 1 mol of Al_2O_3 contains 2 mol of Al atoms, the mass of Al is 2 × 27.0 g = 54.0 g Al

$$\text{Percentage of Aluminium} = \frac{54.0\text{ g}}{102.0\text{ g}} \times 100 = 52.9\ \%$$

We can calculate percentage of oxygen in the same way. One mole of Al_2O_3 contains 3 mole of O atoms, that is, 3 × 16.0 g oxygen therefore

$$\text{Percentage of oxygen} = \frac{3 \times 16.0\text{ g}}{102.0\text{ g}} \times 100 = 47.1\%$$

Example 2.1 : Butanoic acid, has the formula $C_4H_8O_2$. What is the elemental analysis of butanoic acid?

Solution : Molecular formula of the butanoic acid is $C_4H_8O_2$.

In one mole of butanoic acid there are 4 mol of carbon atoms, 8 mol of hydrogen atoms and 2 mol of oxygen atoms. Thus, 1 molar mass of butanoic acid will be equal to the sum of 4 × molar mass of carbon atoms, 8 × molar mass of hydrogen atoms, and 2 × molar mass of oxygen atoms.

Molar mass of butanoic acid = 4 × 12.0 g + 8 × 1.0 g + 2 × 16.0 g = 88.0 g

$$\text{Percentage of C by mass} = \frac{48.0\text{ g}}{88.0\text{ g}} \times 100 = 54.5\%$$

$$\text{Percentage of H by mass} = \frac{8.0\text{ g}}{88.0\text{ g}} \times 100 = 9.1\%$$

$$\text{Percentage of O by mass} = \frac{32.0\text{ g}}{88.0\text{ g}} \times 100 = 36.4\%$$

The percentage of O in butanoic acid can also be calculated as follows :

Percentage of O by mass = 100 – (Percentage of C by mass + Percentage of H by mass)

= 100 – (54.5 + 9.1) = 36.4%

2.3 Determination of Empirical Formulae – Formula Stoichiometry

We have just seen that if we know the formula of a compound we can calculate the percentage composition. Now the question arises, can we determine the formula of the compound if we know the percentage composition of a compound. The answer will be 'yes', but this formula will not be molecular formula; instead it would be **empirical formula** as it would give simplest ratio of different atoms present in a compound. Normally we determine the percentage composition of different elements present in an **unknown compound** and determine its formula. Let us take a simple example of water. Water consists of 11.11% hydrogen and 88.89% oxygen by mass. From the data, we can determine empirical formula of water. Now if we assume that we have a 100.00 g sample of water, then the percentage composition tells us that 100.0 g of water contains 11.11 g of hydrogen atoms and 88.89 g of oxygen atoms.

From the atomic mass table, we find that 1 mol of hydrogn atoms has a mass of 1.0g, and 1 mol of oxygen atoms has a mass of 16.0 g. Now we can write **unit conversion factors** so that the mass of hydrogen can be converted to moles of H atoms and the mass of oxygen can be converted to moles of O atoms. Since 1 mol of H atoms has a mass of 1.0 g we get the conversion factor as

$$\frac{1\text{mol H atoms}}{1.0\text{ g H}}$$

Therefore

$$11.11\text{ g H} = (11.11\text{ g H})\ \frac{1\text{mol H atoms}}{1.0\text{ g H}} = 11.11\text{ mol H atoms}$$

Similarly conversion factor for oxygen will be

$$\frac{1\text{mol O atoms}}{16.0 \text{ g O}}$$

$$\text{Therefore, } 88.89 \text{ g O} = (88.89 \text{ g O}) \frac{1\text{mol O atoms}}{16.0 \text{ g O}} = 5.55 \text{ mol O atoms}$$

Thus in water, the ratio of moles of hydrogen atoms to moles of oxygen atoms is 11.11 : 5.55.

Since a mole of one element contains the same number of atoms as a mole of another element, the ratio of moles of atoms in a compound is also the ratio of the number of atoms. Therefore, the ratio of hydrogen atoms to oxygen atoms is 11.11:5.55. Now by dividing each by the smaller of the two numbers we can convert both numbers to integers

$$\frac{11.11}{5.55} = 2 \text{ and } \frac{5.55}{5.55} = 1$$

Thus ratio hydrogen and oxygen atoms in water is 2 : 1 and empirical formula of water is H_2O.

Intext Questions 2.1

1. For the compound Fe_3O_4, calculate percentage of Fe and O.

   ....

2. State percent composition for each of the following:

   (a) C in $SrCO_3$ (b) SO_3 in H_2SO_4

   ....

3. What are the empirical formulae of substances having the following molecular formulae?

   H_2O_2, C_6H_{12}, Li_2CO_3, $C_2H_4O_2$, S_8, H_2O, B_2H_6, O_3, S_3O_9, N_2O_3

   ....

4. A compound is composed of atoms of only two elements, carbon and oxygen. If the compound contain 53.1% carbon, what is its empirical formula.

   ....

2.4 Chemical Equation and Reaction Stoichiometry

You have studied that a reaction can be represented in the form of a chemical equation. A balanced chemical equation carries a wealth of information qualitative as well as quantitative. Let us consider the following equation and learn what all information it carries.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) \qquad ...(2.1)$$

(1) **Qualitative Information**

Qualitatively the equation (2.1) tells that iron reacts with oxygen to form iron oxide.

(2) **Quantitative Information**

Quantitatively a balanced chemical equation specifies numerical relationship among the quantities of its reactants and products. These relationships can be expressed in terms of :

(i) **Microscopic quantities,** namely, atoms, molecules and formula units.

(ii) **Macroscopic quantities**, namely, moles, masses and volumes *(in case of gaseous substances)* of reactants and products.

Now let us again take the reaction (2.1) given earlier and get the quantitative information out of it.

2.4.1 Microscopic Quantitative Information

The reaction (2.1)

$$4Fe(s) \quad + \quad 3O_2(g) \quad \rightarrow \quad 2Fe_2O_3(s) \qquad ...(2.1)$$

tells that *4 atoms* of iron react with 3 *molecules* of oxygen to form 2 *formula units* of iron oxide. Often this information is written below each reactant and product for ready reference as shown below:

$$\begin{array}{ccccc} 4Fe(s) & + & 3O_2(g) & \rightarrow & 2Fe_2O_3(s) \\ \textbf{4 atoms of Fe} & & \textbf{3 molecules of } \mathbf{O_2} & & \textbf{2 formula units of } \mathbf{Fe_2O_3} \end{array} \qquad ...(2.1a)$$

2.4.2. Macroscopic Quantitative Information

The microscopic quantitative information discussed in the previous section can be converted into macroscopic information with the help of mole concept which you have learnt in unit 1.

(a) Mole Relationships

We know that Avogadro number of elementary entities like atoms, molecules, ions or formula units of a substance constitute one mole of it. Let us multiply the number of atoms, molecules and formula masses obtained in the previous section (Eq.2.1a) by Avogadro's constant, N_A

$$\begin{array}{lllll} 4\ Fe(s) & + & 3O_2(g) & \rightarrow & 2Fe_2O_3(s) \\ \textbf{4 atoms of Fe} & & \textbf{3 molecules of } \mathbf{O_2} & & \textbf{2 formula units of } \mathbf{Fe_2O_3} \\ 4 \times N_A \text{ atoms of Fe} & & 3 \times N_A \text{ molecules of } O_2 & & 2 \times N_A \text{ formula units of } Fe_2O_3 \\ \textbf{4 mol of Fe} & & \textbf{3 mol of } \mathbf{O_2} & & \textbf{2 mol of } \mathbf{Fe_2O_3} \end{array} \qquad ..(2.1)$$

We may rewrite the above equation as

$$\begin{array}{lllll} 4Fe(s) & + & 3O_2(g) & \rightarrow & 2Fe_2O_3(s) \\ \textbf{4 mol of Fe} & & \textbf{3 mol of } \mathbf{O_2} & & \textbf{2 mol of } \mathbf{Fe_2O_3} \end{array} \qquad ...(2.1b)$$

The above equation (2.1b) gives us the mole relationship between reactants and products. Here *4 mol of Fe* react with *3 mol of* O_2 and produce *2 mol of* Fe_2O_3.

(b) Mass Relationships

The mole relationships which you have learnt in the previous section, can be converted into mass relationship by using the fact that mass of one mole of any substance is equal to its *molar mass* which can be calculated from its formula with the help of relative atomic masses of its constituent elements.

In the reaction that we are discussing, the relative atomic masses of iron and oxygen are 55.8 and 16.0 respectively. Therefore

(i) molar mass of Fe $= 55.8 \text{ g mol}^{-1}$

(ii) molar mass of O_2 $= 2 \times 16.0 = 32 \text{ g mol}^{-1}$

(iii) molar mass of Fe_2O_3 $= (2 \times 55.8 + 3 \times 16.0) \text{ g mol}^{-1}$

$= 159.6 \text{ g mol}^{-1}$

Using these molar masses we can convert the mole relationship given by equation 2.1b into mass relationship as given below :

$4Fe(s)$	+	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$
4 mol Fe		**3 mol O_2**		**2 mol Fe_2O_3**
(4×55.8) g Fe		(3×32) g O^2		(2×159.6) g Fe_2O_3
223.2 g Fe		96 g O_2		319.2 g Fe_2O_3

Thus 223.2 g iron would react with 96 g oxygen and produce 319.2 g iron oxide, We may rewrite the above equation as

$4Fe(s)$	+	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$	(2.1c)
223.2 g Fe		**96 g O_2**		**319.2 g Fe_2O_3**	

(c) Volume Relationships

We know that one mole of *any gas* occupies a volume of 22.7 L* at STP (standard temperature and pressure, 0 °C and 1 bar pressure). We can use this information to arrive at volume relationships between gaseous substances. The reaction that we are considering involves only one gaseous substance, O_2. We may rewrite the equation (2.1b) as

$4Fe(s)$	+	$3O_2(g)$	$\rightarrow$	$2Fe_2O_3(s)$	(2.1b)
4 mol		**3 mol**		**2 mol.**	
		(3×22.7) L at STP			
		68.1 L at STP			

Thus 4 mol of iron would react with 68.1 L of oxygen at STP to produce 2 mol of iron oxide. (The volume relationship becomes more useful for reactions involving 2 or more gaseous substances).

*Earlier, the standard pressure was taken as 1 atmosphere and the volume of one mole of gas at STP was taken as 22.4 L.

We can express microscopic as well macroscopic quantitative relationships involved in the above reaction as shown below:

$4Fe(s)$	+	$3O_2(s)$	$\rightarrow$	$2Fe_2O_3(s)$
4 atoms		**3 molecules**		**2 formula units**
4 mol		3 mol		2 mol
223.2 g		96 g		319.2 g
–		68.1 L at STP		–

We may use even mixed relations. For example, we may say 4 mol of iron would react with 68.1 L (at STP) of oxygen to produce 319.2 g of iron oxide.

Let us understand these relationships with two more examples.

(a) Let us work out the mole, mass and volume relationships for the reaction involved in manufacture of ammonia by Haber's process.

Microscopic relationship	$N_2(g)$	$+ 3H_2(g) \longrightarrow$	$2NH_3(g)$	...(2.2)
Microscopic relationships	1 Molecule	3 Molecules	2 Molecules	
(i) Moles	1 mol	3 mol	2 mol	
(ii) Mass	28 g	$(3 \times 2.0) = 6.0$ g	$(2 \times 17.0) = 34$ g	
(iii) Volume	1×22.7 L = 22.7 L	(3×22.7) = 68.1 L	(2×22.7) = 45.4 L	
or	1 vol	3 vol	2 vol	

(b) Let us take one more reaction, the combustion reaction of butane and work out the different types of relationships. The reaction is :

$2C_4H_{10}(g)$	+	$13O_2(g)$	$\rightarrow$	$8CO_2(g)$	+	$10H_2O(g)$
2 molecules		13 molecules		8 molecules		10 molecules
2 mol		13 mol		8 mol		10 mol
$2 \times (4 \times 12 + 10 \times 1)$ g		(13×32) g		$8 \times (12 + 2 \times 16)$ g		$10 \times (2 \times 1 + 16)$ g
116 g		416 g		352 g		180 g
$2 \times 22.7 = 45.4$ L		$13 \times 22.7 = 295.1$ L		$8 \times 22.7 = 181.6$ L		$10 \times 22.7 = 227$ L
2 vol		13 vol		8 vol		10 vol

Now let us use the mole, mass and volume relationships to make some calculations.

Example 2.2 In the manufacture of ammonia by Haber process, nitrogen reacts with hydrogen at high temprature and high pressure in the presence of a catalyst and gives ammonia.

$$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$$

How much hydrogen would be needed to produce one metric ton of ammonia?

Solution : We should first find out the mass relationships for the reaction.

$N_2(g)$	+	$3H_2(g) \longrightarrow$	$2NH_3(g)$
1 mol		**3 mol**	**2 mol**
$1 \times 28g = 28$ g		$3 \times 2g = 6.0$ g	$2 \times 17g = 34$ g

Notes

We know that :

$$1 \text{ metric ton} = 1000 \text{ kg} = 10^3 \text{ kg} = 10^6 \text{ g}$$

From the mass relationship 34 g NH_3 requires 6.0 g H_2 for its manufacture.

$$\therefore\ 10^6\text{g } NH_3 \text{ would require } \frac{6.0 \times 10^6}{34} \text{ g} = 1.76 \times 10^5\text{g of } H_2.$$

Thus 1 metric ton of ammonia will be obtained by using 1.176×10^5g of Hydrogen.

Example 2.3 In a rocket motor fuelled by butane, C_4H_{10}, how many kg of O_2 should be provided with each kg of butane to provide for complete combustion?

Solution : The combustion reaction of butane is

$$2C_4H_{10}(g) \ + \ 13O_2(g) \longrightarrow 8CO_2(g) \ + \ 10H_2O(g)$$

2 mol **13 mol**

2 ò 58 = 116 g **13 ò 32 = 416 g**

Thus, to completely burn 116 g butane, oxygen required is 416g.

Therefore, to completely burn 1 kg (1000 g) butane, oxygen required will be

$$= \frac{416 \times 1000}{116} \text{ g } O_2$$

$$= 3586 \text{ g } O_2$$

$$= 3.586 \text{ kg } O_2 \approx 3.59 \text{ kg } O_2$$

Example 2.4 When lead sulphide; PbS and lead oxide, PbO, are heated together the products are lead metal and sulphur dioxide, SO_2 ,

$$PbS(s) + 2PbO(s) \xrightarrow{\text{heat}} 3Pb\ (1) + SO_2(g)$$

If 14.0 g of lead oxide reacts according to the above equation, how many (a) moles of lead (b) grams of lead, (c) atoms of lead and (d) grams of sulphur dioxide are formed?

(Atomic mass : Pb = 207.0, S = 32.1 ; O =16.0)

Solution : For each part of the question we will use the balanced equation

$$PbS(s) \ + \ 2PbO(s) \xrightarrow{\text{heat}} 3Pb\ (1) + SO_2(g)$$

1mol **2mol** **3 mol** **1mol**

Now formula mass of PbO = (207.0 + 16.0) = 223.0 amu

Thus, one mole of lead oxide formula units have a mass of 223.0 g. Therefore, 14.0 g of PbO is $\dfrac{14.0 \text{ g PbO}}{223.0 \text{ g mol}^{-1} \text{ PbO}} = 6.28 \times 10^{-2}$ mol PbO

(a) The balanced equation shows that 2 mol of PbO form 3 mol of Pb. Therefore, 6.28×10^{-2} mol of PbO form

$$6.28 \times 10^{-2}\,\text{mol PbO} \times \frac{3\,\text{mol Pb}}{2\,\text{mol PbO}} = 9.42 \times 10^{-2}\,\text{mol Pb}$$

(b) The atomic mass of Pb is 207.0 ; this tells us that one mol of lead has a mass 207.0 g. Thus, 9.42×10^{-2} mol of Pb has a mass of

$$9.42 \times 10^{-2}\,\text{mol Pb} \times \frac{207.0\,\text{g Pb}}{1\text{mol Pb}} = 19.5\,\text{g Pb}$$

(c) 9.42×10^{-2} mol of Pb is

9.42×10^{-2} mol of Pb $\times$ 6.022×10^{23} atoms mol^{-1} = 5.67×10^{22} Pb atoms

(d) The balanced equation shows that 2 mol of PbO form 1 mol of SO_2.

Therefore, 6.28×10^{-2} mol of PbO formula unit forms

$$6.28 \times 10^{-2}\,\text{mol PbO} \times \frac{1\text{mol}\,SO_2}{2\,\text{mol PbO}}$$

$$= 3.14 \times 10^{-2}\,\text{mol}\,SO_2$$

Now the relative molecular mass of SO_2 = 32.1 + 2(16.0) = 64.1

Molar mass of SO_2 = 64.1 g mol^{-1}

Therefore, 3.14×10^{-2} mol of SO_2 molecules have a mass of 3.14×10^{-2} mol $\times$ 64.1 g mol^{-1}
= 2.01 g

Intext Questions 2.2

(1) How many grams of NH_3 can be made according to the reaction

$N_2(g) + 3H_2(g) \longrightarrow 2NH_3\,(g)$

from (a) 0.207 mol of N_2 (b) 22.6 g of H_2

....

(2) In reaction

$C_2H_4(g) + 3O_2(g) \longrightarrow 2CO_2(g) + 2H_2O(\ell)$

How many (a) moles of O_2 are consumed and (b) moles of H_2O are formed when $4.16 \text{ X } 10^{-2}$ mol of C_2H_4 react?

....

2.5 Limiting Reagents

We generally find that substances which react with each other are not present in exactly the same proportionin a reaction mixture as stated by a balanced chemical equation. For example, if 2 mol each of hydrogen and oxygen are mixed and a spark is passed through the mixture, water is formed, according to the equation

$$\underset{\textbf{2 mol}}{2H_2} \quad + \quad \underset{\textbf{1mol}}{O_2} \quad \longrightarrow \quad \underset{\textbf{2 mol}}{2H_2O}$$

Here, 2 mol of hydrogen react with only 1 mol of oxygen, and 1 mol of oxygen therefore remains unreacted. In this example hydrogen is said to be the **limiting reagent or reactant** because its amount becomes zero and the reaction therefore stops before the other reactant; that is, the oxygen is used up. The amount of hydrogen present initially limits the amount of product that is formed.

Example 2.5 3 mol of sulphur dioxide SO_2 is mixed with 2 mol of oxygen O_2, and after reaction is over sulphur trioxide, SO_3 is obtained.

(i) Which is the limiting reagent?

(ii) What is the maximum amount of SO_3 that can be formed?

Solution : (i) We must first write the balanced equation

$$2SO_2 + O_2 \rightarrow 2SO_3$$

According to the above equation

(a) 2 mol of SO_3 can be formal from 2 mol of SO_2.

$\therefore$ Amount of SO_3 that can be formed from 3 mol of SO_2.

$$= (3 \text{ mol } SO_2) \times \frac{2\text{mol } SO_3}{2\text{mol } SO_2} = 3 \text{ mol } SO_3$$

(b) 2 mol of SO_3 can be formed from 1 mol of O_2. Therefore, the amount of SO_3 that can be formed from 2 mol of O_2.

$$= (2 \text{ mol } O_2) \times \frac{2\text{mol } SO_3}{1\text{mol } O_2} = 4 \text{ mol } SO_3$$

According to the definition, the limiting reactant is that reactant which gives the smallest amount. In this case SO_2 is the limiting reactant.

(ii) The maximum amount of product that can be obtained is the amount formed by the limiting reagent. Thus a the maximum amount of SO_3 that can be obtained is 3 mol.

Example 2.6 2.3 g of sodium metal is introduced into a 2L flask filled with chlorine gas at STP (273 K, 1bar). After the reaction is over, find :

(i) What is the limiting reagent in this reaction?

(ii) How many moles of sodium chloride are formed?

(iii) Which substance is left unconsumed at the end of the reaction? Find out its mass in grams.

(iv) What percentage of the substance present in excess is converted into sodium chloride?

(Given : Na = 23, Cl = 35.5)

Solution :

	2 Na (s)	+ Cl_2(g)	$\longrightarrow$	2NaCl (s)
	2mol	**1 mol**		**2 mol**
or		22.7 L at STP		

(i) Moles of sodium introduced = $\dfrac{2.3\,g}{23\,g\,mol^{-1}}$ = 0.1 mol

From the above equation, it is clear that 2 mol NaCl is formed from 2 mol Na

Therefore 0.1 mol Na can produce = $\dfrac{2 \times 0.1}{2}$ = 0.1 mol NaCl

Molar volume at STP = 22.7 L

Therefore moles of chlorine in 2 L volume at STP = $\dfrac{2L}{22.7\,L\,mol^{-1}}$ = 0.088 mol

From equation : 1 mol Cl_2 can produce 2 mol NaCl

Therefore 0.088 mol Cl_2 can produce 2 × 0.088 = 0.176 mol NaCl.

Since sodium produces less amount of NaCl, it is the limiting reagent.

(ii) Sodium being the limiting reagent, as calculated in (i), the moles of NaCl produced = 0.1 mol

(iii) From above equation, 2 mol NaCl is produced from 1 mol Cl_2

Therefore 0.1 mol NaCl is produced from $\dfrac{1 \times 0.1}{2}$ = 0.05 mol Cl_2

Initial moles of Cl_2 = 0.088 mol

Moles of Cl_2 left unconsumed = (0.088 – 0.05) mol = 0.038 mol

Therefore, mass of Cl_2 left unconsumed = 0.038 g × 71.0 g mol^{-1} = 2.698 g

(because molar mass of Cl_2 = 2 × 35.5 = 71.0 g mol^{-1})

(iv) Moles of Cl_2 consumed = 0.05 mol out of 0.088 mol

∴ Percent of Cl_2 consumed and converted into NaCl = $\dfrac{0.05}{0.088}$ × 100 = 56.8 %

Example 2.7 : 2.0 g mixture of $MgCO_3$ and $CaCO_3$ are heated till no further loss of weight takes place. The residue weighs 1.04g. Find the percentage composition of the mixture. (Mg = 24, Ca = 40, C = 12, O = 16)

Solution : Mixture of $MgCO_3$ and $CaCO_3$ taken = 2.0 g

Let the mass of $MgCO_3$ be = x g

Therefore the mass of $CaCO_3$ = $(2.0 - x)$ g

The decomposition reactions are

$$\underset{\substack{\textbf{(24+12+48) g} \\ \textbf{84 g}}}{MgCO_3\,(s)} \rightarrow \underset{\substack{\textbf{(24+16) g} \\ \textbf{40 g (Residue)}}}{MgO(s)} + CO_2(g) \qquad \text{(i)}$$

$$\underset{\substack{\textbf{(40 + 12 + 48) g} \\ \textbf{100 g}}}{CaCO_3\,(s)} \rightarrow \underset{\substack{\textbf{(40 + 16) g} \\ \textbf{56 g (Residue)}}}{CaO(s)} + CO_2(g) \qquad \text{(ii)}$$

From the equation (i)

84 g $MgCO_3$ leaves a residue = 40 g

x g $MgCO_3$ will leave residue = $\dfrac{40x}{84}$ g

From the equation (ii)

100 g $CaCO_3$ leaves a residue = 56 g

$(2.0 - x)$ g $CaCO_3$ will leave residue = $\dfrac{56 \times (2.0 - x)}{100}$ g

Total mass of the residue = $\dfrac{40x}{84} + \dfrac{56 \times (2.0 - x)}{100}$ = 1.04 g (given)

$$40 \times 100x + 84 \times 56 \times 2 - 84 \times 56x = 84 \times 100 \times 1.04$$

$$4000x + 9408 - 4704x = 8736$$

$$9408 - 8736 = (4704 - 4000)x$$

$$672 = 704x$$

Therefore, mass of $MgCO_3$ in the mixture = $x = \dfrac{672}{704}$ = 0.96 g

Therefore, percentage of $MgCO_3$ = $\dfrac{0.96}{2.0} \times 100$ = 48 %

and percentage of $CaCO_3$ = 100 – 48 = 52 %

What You Have Learnt

- A chemical formula is used not only to represent the name of a compound but also to indicate its composition in terms of (i) relative number of atoms and (ii) relative number of moles of atoms.
- A molecular formula of a substance shows(i) the number of atoms of different elements in one molecule.(ii) the number of moles of atoms of different elements in one mole of molecule.
- An empirical formula shows only a ratio of (i) number of atoms, and (ii) moles of atoms in a compound.
- Molecular formula is always an integral multiple of the empirical formula.
- The empirical formula of a compound can be determined from its chemical analysis.
- In order to determine a compound's molecular formula, molecular mass also must be known.
- Stoichiometry is the quantitative study of the composition of chemical compounds (compound or formula stoichiometry) and of the substances consumed and formed in chemical reactions (reaction or equation stoichiometry).
- Chemical equations specify not only the identities of substances consumed and formed in a reaction, but also the relative quantities of these substances in terms of (a) atoms, molecules, and formula units and (b) moles of these entities.
- A balanced chemical equation demonstrates that all the atoms present in the reactants are accounted for in the product; atoms are neither created nor destroyed in a reaction.
- The stoichiometric ratios among the moles of reactants shown in a balanced equation are useful for determining which substance is entirely consumed and which substance(s) is (are) left over.

Terminal Exercise

1. Write empirical formulae of the following compounds:

   CO, Na_2SO_3, C_4H_{10}, H_2O_2, KC1

   ....

2. The empirical formula of glucose is CH_2O which has a formula mass of 30 amu. If the molecular mass of glucose is 180 amu. Determine the molecular formula of glucose

   ....

3. What is ratio of masses of oxygen that are combined with 1.0 gram of nitrogen in the compound NO and N_2O_3?

   ....

4. A compound containing sulphur and oxygen on analysis reveals that it contains 50.1% sulphur and 49.9% oxygen by mass. What is the simplest formula of the compound?

....

5. Hydrocarbons are organic compound composed of hydrogen and carbon. A, 0.1647 g sample of a pure hydrocarbon on burning in a combustion tube produced 0.5694 g of CO_2 and 0.0845 g of H_2O. Determine the percentage of these elements in the hydrocarbon.

....

6. On combustion 2.4 g of a compound of carbon, hydrogen and oxygen gave 3.52 g of CO_2 and 1.44 g of H_2O. The molecular mass of the compound was found to be 60.0 amu.

(a) What are the masses of carbon, hydrogen and oxygen in 2.4 g of the compound?

....

(b) What are the empirical and molecular formulae of the compound?

....

7. (i) What mass of oxygen is required to react completely with 24 g of CH_4 in the following reaction?

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(1)$$

(ii) How much mass of CH_4 would react with 96 g of oxygen.

8. In the reaction $H_2 + C1_2 \rightarrow 2HC1$

How many grams of chlorine, $C1_2$ are needed to react completely with 0.245 g of hydrogen, H_2, to give hydrogen chloride, HC1? How much HC1 is formed?

....

9. 3.65 g of H_2 and 26.7 g of O_2 are mixed and reacted. How many grams of H_2O are formed?

....

10. Caustic soda NaOH can be commercially prepared by the reaction of Na_2CO_3 with slaked line, $Ca(OH)_2$. How many grams of NaOH can be obtained by treating 2.0 kg of Na_2CO_3 with $Ca(OH)_2$?

....

11. A portable hydrogen generator utilizes the reaction

$$CaH_2 + H_2O \rightarrow Ca(OH)_2 + 2H_2$$

How many grams of H_2 can be produced by a 100 g cartridge of CaH_2?

....

12. The reaction $2Al + 3MnO \rightarrow Al_2O_3 + 3Mn$ proceeds till the limiting substance is consumed. A mixture of 220 g Al and 400 g MnO was heated to initiate the reaction. Which initial substance remained in excess and by how much ? (Al = 27, Mn = 55).

....

13. $KClO_4$ may be prepared by means of following series of reactions

$Cl_2 + 2KOH \rightarrow KCl + KClO + H_2O$

$3KClO \rightarrow 2KCl + KClO_3$

$4KClO_3 \rightarrow 3KClO_4 + KCl$

How much Cl_2 is needed to prepare 400 g $KClO_4$ by the above sequence?

(K = 39, Cl = 35.5, O = 16, H = 1)

....

14. 2.0 g of a mixture of Na_2CO_3 and $NaHCO_3$ was heated when its weight reduced to 1.876 g. Determine the percentage composition of the mixture.

15. Calculate the weight of 60 % sulphuric acid required to decompose 150 g of chalk (calcium carbonate). Given Ca = 40, C = 12, O = 16, S = 32)

Answers to Intext Questions

2.1

(1) Molar mass of Fe_3O_4 $= 3 \times 56.0 + 4 \times 16.0$

$= (168.0 + 64.0) = 232.0 \text{ g mol}^{-1}$

Percentage of Fe $= \dfrac{168.0}{232.0} \times 100 = 72.41\%$

Percentage of O $= \dfrac{64.0}{232.0} \times 100 = 27.59\%$

(2) (a) Molar mass of $SrCO_3 = 87.6 + 12.0 + 48.0 = 147.6 \text{ g mol}^{-1}$

Percentage of carbon C in $SrCO_3 = \dfrac{12.0}{147.6} \times 100 = 8.13\%$

(b) Molar mass of $H_2SO_4 = 2.0 + 32.1 + 64.0 = 98.1 \text{ g mol}^{-1}$

Molar mass of $SO_3 = 32.1 + 48.0 = 80.1 \text{ g mol}^{-1}$

Percentage of SO_3 in $H_2SO_4 = \dfrac{80.1 \times 100}{98.1} = 81.65\%$

3.

Substance	Empirical formula
H_2O_2	HO
C_6H_{12}	CH_2
Li_2CO_3	Li_2CO_3
$C_2H_4O_2$	CH_2O
S_8	S
H_2O	H_2O
B_2H_6	BH_3
O_3	O_3
S_3O_9	SO_3
N_2O_3	N_2O_3

4. Percentage of carbon = 53.1%

Percentage of Oxygen = 46.9%

Suppose we take 100 g of the substance then moles of carbon = $\frac{53.1}{12.0}$ g = 4.43 mol

$$\text{mole of oxygen} = \frac{46.0}{16.0} = 2.93 \text{ mol}$$

$$\text{molar ratio of C and O} = \frac{4.43}{2.93} : \frac{2.93}{2.93}$$

$$= 1.50 : 1 \text{ or } 3 : 2$$

Empirical formula of the compound is C_3O_2

2.2

1. In equation

$$\underset{\textbf{1 mol}}{N_2(g)} + \underset{\textbf{3 mol}}{3H_2(g)} \rightarrow \underset{\textbf{2 mol}}{2NH_3(g)}$$

0.207 mol of N_2 gives 0.414 mol of NH_3

0.414 mol of NH_3 = 0.414 mol × 17.0 g mol^{-1} = 7.038 g of NH_3

22.6 g of hydrogen = $\frac{22.6}{2.0}$ = 11.3 mol of hydrogen

11.3 mol of hydrogen will give $\frac{2}{3}$ × 11.3 mol of NH_3 = 7.53 mol

Therefore, mass of NH_3 = 7.53 mol × 17.0 g mol^{-1} = 128.01 g

2. $C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(g)$

1 mol **3 mol** **2 mol** **2 mol**

(a) 4.16×10^{-2} mol of C_2H_4 will consume $3 \times 4.16 \times 10^{-2}$ mol of oxygen

$= 12.48 \times 10^{-2} = 1.248 \times 10^{-1}$ mol of O_2

(b) moles of H_2O formed = $2 \times 4.16 \times 10^{-2}$ mol

$= 8.32 \times 10^{-2}$ mol of H_2O