SEMESTER–II (Period–V)

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6
TOPIC

Atoms, Molecules, Ions and Stoichiometry

Learning Objectives

Upon completion of this topic, learners will be able to:

- Distinguish the fundamental chemical laws
- Distinguish atoms, molecules and ions
- Discuss the mole concept
- Determine the percent of elements in compounds
- Determine the formula of a compound
- Discuss the kinds and types of chemical reactions
- Analyze the techniques in balancing chemical reactions and
- Determine the limiting reagent/reactant of a chemical reaction.

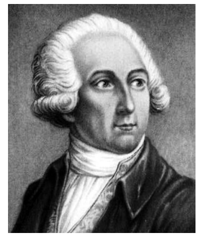
Introduction

The structure of matter has been a subject of speculation from very early times. In the fourth century BC, Greek philosopher, Democritus, suggested that if we go on dividing matter into smaller parts, a stage would be reached when particles obtained cannot be divided further. He called these particles '*atoms*' meaning indivisible. All these concepts were based on philosophical considerations and not much experimental work was done to confirm these ideas. After Democritus' death, little more was done with atomic theory until the end of the eighteenth century, when Antoine Lavoisier introduced modern chemistry to the world. He put forward two important laws of chemical combination which formed the basis of Dalton's atomic theory which was published in 1808.

6.1. FUNDAMENTAL CHEMICAL LAWS

By studying the results of quantitative measurements of many reactions it was observed that whenever substances react, they follow certain laws. These laws are called the **laws of chemical combination**. These laws formed the basis of Dalton's atomic theory of matter.

6.1.1. Law of Conservation of Mass



Antoine Lavoisier
(1743–1794)

This law was stated by the French chemist **Antonie Lavoisier** (1774). This law states that:

In every chemical reaction, the total mass before and after the reaction remains constant.

In other words, *mass can neither be created nor destroyed in a chemical reaction.*

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

$$\begin{array}{ccccc} \text{Mercuric oxide} & \longrightarrow & \text{Mercury} & + & \text{Oxygen} \\ 100\text{ g} & & 92.6\text{ g} & & 7.4\text{ g} \end{array}$$

Law of conservation of mass is also known as *law of indestructibility of matter.*

ACTIVITY 6.1

Demonstration of Law of Conservation of Mass

1. Prepare separately a 5% solution of barium chloride and a 5% solution of sodium sulphate.
2. Take about 20 mL of barium chloride solution in a conical flask.
3. Take sodium sulphate solution in a small test tube. Hang the test tube in the conical flask with the help of a thread. Close the mouth of the flask with a cork (Fig. 6.1).

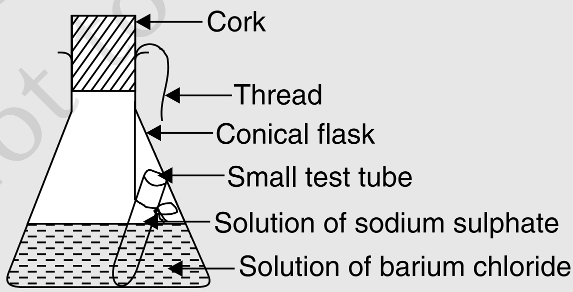


Fig. 6.1. Study of reaction between barium chloride and sodium sulphate.

4. Weigh the flask along with its contents.
5. Now tilt the flask so that the two solutions get mixed.
6. Weigh the flask again along with its contents.

   What do you observe?

It is observed that on mixing the two solutions a chemical reaction takes place which is indicated by the formation of a white precipitate.

Barium chloride + Sodium sulphate $\longrightarrow$ Barium sulphate + Sodium chloride
(White ppt.)

The mass of the flask and its contents remains constant. Thus, during a chemical reaction mass is neither created nor destroyed.

This activity can also be carried out with the following pairs:

(a) Silver nitrate and sodium chloride.

(b) Copper sulphate and sodium carbonate.

Example 6.1: *8.4 g of magnesium carbonate on heating gave 4.4 g of carbon dioxide and 4.0 g of magnesium oxide. Show that these observations are in agreement with law of conservation of mass.*

Solution: Mass of the reactants = 8.4 g

Mass of the products = 4.4 + 4.0 = 8.4 g

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

6.1.2. Law of Definite Proportion

This law deals with the composition of chemical compounds. It was discovered by the French chemists,

A. Lavoisier and Joseph Proust (1799). This law states that:

A pure chemical compound always contains same elements combined together in same proportion by mass.

It implies from this law that in a chemical compound the elements are present in *fixed* and not *arbitrary* ratio by mass. For example, pure water obtained from different sources such as *river*, *well*, *spring*, *sea*, etc., always contains hydrogen and oxygen combined together in the ratio 1 : 8 by mass. Similarly, carbon dioxide can be obtained by different methods such as:

(i) By burning of carbon,

(ii) By heating lime stone, or

(iii) By the action of dilute hydrochloric acid on marble pieces.

It can be shown experimentally that samples of carbon dioxide obtained from different sources contain carbon and oxygen in the ratio of 3 : 8 by mass.

Example 6.2: *A sample of ascorbic acid (vitamin C), synthesized in laboratory contains 15.0 g of carbon in 35.0 g of ascorbic acid. Another sample of ascorbic acid, isolated from lemons contains 42.9% carbon. Show that the data is in accordance with law of constant composition.*

Solution: In the first sample of ascorbic acid,

$$\text{Mass of carbon} = 15.0 \text{ g}$$

$$\text{Mass of ascorbic acid} = 35.0 \text{ g}$$

$$\% \text{ of carbon} = \frac{15.0}{35.0} \times 100 = 42.9\%$$

In the second sample of ascorbic acid,

$$\% \text{ of carbon} = 42.9\%$$

Since both the samples contain same percentage of carbon, the given data is in accordance with law of constant composition.

Example 6.3: *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 3.0 g of carbon is burnt in 50.00 g of oxygen? Which law of chemical combination will govern your answer?*

Solution: Carbon and oxygen combine in the ratio 3: 8 to form carbon dioxide. Even when we use excess of oxygen, carbon and oxygen would combine in the same ratio to form carbon dioxide. Excess oxygen would be left unreacted. Therefore, 3 g of carbon when burnt in 50 g of oxygen would give only 11.0 g of carbon dioxide by reacting with 8 g of oxygen.

The **law of constant composition** governs this answer.

6.1.3. Law of Multiple Proportions

This law was proposed by John Dalton (1803). This law states that:

When two elements combine with each other to form two or more than two compounds, the masses of one element which combine with fixed mass of the other, bear a simple whole number ratio to one another.

For example, carbon and oxygen combine with each other to form carbon monoxide (CO) and carbon dioxide (CO_2).

In carbon monoxide (CO), 12 parts by mass of carbon combine with 16 parts by mass of oxygen.

In carbon dioxide (CO_2), 12 parts by mass of carbon combine with 32 parts by mass of oxygen.

Ratio of the masses of the oxygen which combine with fixed mass of carbon (12 parts) in these compounds is 16 : 32 or 1 : 2, which is a simple whole number ratio.

Example 6.4: *Hydrogen and oxygen are known to form two compounds. The hydrogen content in one of these is 5.93% while in the other it is 11.2%. Show that this data illustrates the law of multiple proportions.*

Solution: *In the first compound*:

$$\text{Hydrogen} = 5.93\%$$
$$\text{Oxygen} = (100 - 5.93)\% = 94.07\%.$$

In the second compound:

$$\text{Hydrogen} = 11.2\%$$
$$\text{Oxygen} = (100 - 11.2)\% = 88.8\%.$$

In the first compound, the number of parts by mass of oxygen that combine with one part by mass of hydrogen $= \dfrac{94.07}{5.93} = 15.86$ parts.

In the second compound, the number of parts by mass of oxygen that combine with one part by mass of hydrogen $= \dfrac{88.8}{11.2} = 7.9$ parts.

The ratio of masses of oxygen that combine with fixed mass (1 part) by mass of hydrogen is 15.86 : 7.9 or 2 : 1.

Since this ratio is a simple whole number ratio, hence the given data illustrates the law of multiple proportions.

6.2. ATOMS, MOLECULES AND IONS

6.2.1. Atoms

Atoms are the building blocks of all matter. An atom is the smallest particle of an element which can ever exist. A chemical reaction involves redistribution of atoms among various species. An atom is neither created nor destroyed during chemical reactions. In other words, in an ordinary chemical reaction, no atom of any element disappears or is changed into an atom of another element. *It is the smallest particle of an element that can take part in chemical reactions.*

Atoms are very small in size. The size of an atom is expressed in terms of atomic radius. Atomic radius is measured in nanometres (nm).

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

Radii of most of the atoms are of the order of 0.1 nm or 10^{-10} m. For example, atomic radius of hydrogen is 0.037 nm while that of gold atom is 0.144 nm.

Atoms cannot be viewed by simple optical microscopes. However, through modern techniques such as scanning tunneling microscope (STM) it is possible to produce magnified images of surfaces of elements showing atoms.

6.2.2. Molecules

As already mentioned an atom is the smallest particle of an element that takes part in chemical reactions. An atom may or may not have independent existence. The smallest particle of an element or a compound which can exist independently is called a **molecule**. Thus,

*A **molecule** is the smallest particle of an element or compound that has independent existence.* A molecule may contain one or more than one atoms.

6.2.2.1. Molecules of Elements

The molecules of elements contain atoms of only one kind.

The number of atoms in a molecule of an element is known as **atomicity of the element**.

Noble gases (He, Ne, Ar, Kr and Xe) have monoatomic molecules.

Hydrogen (H_2), oxygen (O_2), nitrogen (N_2), chlorine (Cl_2), etc., have diatomic molecules. Ozone (O_3) has triatomic molecule.

For phosphorus (P_4) atomicity is four and for sulphur (S_8) atomicity is eight.

Many elements such as carbon, silicon and metals have complex molecular structures consisting of a very large, indefinite number of atoms. These elements are represented by their atomic symbols.

6.2.2.2. Molecules of Compounds

Molecules of compounds contain atoms of two or more different elements. For example, a molecule of ammonia (NH_3), a molecule of methane (CH_4), a molecule of glucose ($C_6H_{12}O_6$).

The molecular formula of a substance tells us about the number of atoms of various elements present in a molecule of the substance.

6.2.3. Ions (or Radicals)

Most of the inorganic compounds which are composed of metals and non-metals, contain oppositely charged parts called ions or radicals. The positively charged part is called **cation** or **electropositive radical** or **basic radical** whereas negatively charged part is called **anion** or **electronegative radical** or **acidic radical**. For example, sodium chloride, is made up of Na^+ ions and Cl^- ions similarly nickel sulphate is made up of Ni^{2+} ions and SO_4^{2-} ions. An ion behaves as single unit in reactions. Thus:

*An **ion** is an atom or group of atoms, carrying positive or negative charge, that behaves as a single unit in reactions.*

The positively charged ion is called **cation**. Whereas the negatively charged ion is called **anion**.

An ion containing only one atom is known as **simple ion** whereas an ion containing two or more than two atoms is known as **polyatomic ion** or **compound ion**. Thus, Na^+ is a simple ion whereas SO_4^{2-} is a compound ion.

The charge on the ion is known as valency of the ion.

6.3. THE MOLE CONCEPT

6.3.1. Mole and Particles

While carrying out investigations we are often interested in knowing the number of atoms and molecules we are dealing with. Sometimes, we have to take the atoms or molecules of different reactants in a definite ratio. For example, let us consider the following reaction:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

In this reaction, one molecule of nitrogen reacts with three molecules of hydrogen. So it would be desirable to take the molecules of nitrogen and hydrogen in the ratio 1 : 3, so that the reactants are completely consumed during the reaction. But we know atoms and molecules are so small in size that it is not possible to count them individually. Moreover, the smallest sample of the substance that we may take contains a very large number of atoms or molecules. In order to deal with such huge numbers a unit of similar magnitude is required. In SI system, a unit

mole was introduced to count entities, at microscopic level, such as atoms, molecules, ions, etc. According to this concept, the number of particles of the substance is related to mass of the substance. For the sake of counting elementary particles the number of atoms in 12 g of carbon-12 isotope has been taken as standard. This number has been experimentally found to be 6.0221367×10^{23} and is called **one mole**. The **mole** may be defined as ***"the amount of the substance that contains as many specified elementary particles as the number of atoms in 12 g of carbon-12 isotope."***

The elementary particles may be atoms, molecules or other discrete particles such as ions, electrons, etc. Mole is simply a unit for counting atoms, molecules, ions, etc. It stands for 6.022×10^{23} particles irrespective of their nature, just as we use one dozen for twelve objects and one score for twenty items irrespective of their nature.

6.3.2. Avogadro Number

The number of entities in one mole is so significant that it has been given a separate name and symbol. It is known as **Avogadro number** or **Avogadro constant** and is denoted as $\mathbf{N_A}$.

$$\boxed{\textbf{One Mole} = 6.022 \times 10^{23} \text{ or } N_A \textbf{ Particles}}$$

- 1 mole oxygen atoms = 6.022×10^{23} O atoms
- 1 mole oxygen molecules = 6.022×10^{23} O_2 molecules
- 1 mole chloride ions = 6.022×10^{23} Cl^- ions
- 1 mole electrons = 6.022×10^{23} electrons

Now, if we multiply the number moles (**n**) of the substance with Avogadro number ($\mathbf{N_A}$) we get total number of particles (**N**) in the given amount of the substance.

$$\boxed{N = n \times N_A}$$

6.3.3. Molar Mass

When we deal with very large number of particles, it is easier to weigh them instead of counting them. Therefore, it is more appropriate to relate the mass of the substance to the number of particles.

It can be easily shown that if we take weights of two elements in the ratio of their atomic masses then they will have equal numbers of atoms. Suppose we take 12.0 g of carbon and 16.0 g of oxygen separately then both the samples will have same number of atoms because the ratio of their weights is same as the ratio of their atomic masses. Now we know that 12.0 g of carbon contains 6.022×10^{23} atoms, therefore, 16.0 g of oxygen will also contain 6.022×10^{23} atoms of oxygen.

Similarly, 32.0 g of oxygen gas contains 6.022×10^{23} oxygen (O_2) molecules and 17.00 g of ammonia will contain 6.022×10^{23} ammonia (NH_3) molecules. From this we can conclude that mass of element in grams which is numerically equal to its atomic mass contains 6.022×10^{23} (N_A) atoms and the mass of substance in grams which is numerically equal to its molecular mass contains 6.022×10^{23} (N_A) molecules. This mass is referred to as the *molar mass*. **Molar mass** of the substance may ***thus, be defined as the mass of one mole, i.e., 6.022 × 10²³ particles of that substance***. Molar mass is represented by **M** and is expressed in the units **g mol^{-1}**. Molar mass in grams is numerically equal to the atomic mass/molecular mass/formula mass expressed in unified mass units (u).

Examples:

- **Atomic oxygen:** Atomic mass = 16 u

  Molar mass of oxygen atoms = **16 g mol^{-1}** and it contains 6.022×10^{23} O atoms
- **Water (H_2O):** Molecular mass = 18 u

  Molar mass of water = **18 g mol^{-1}** and it contains 6.022×10^{23} H_2O molecules.
- **Sodium chloride (NaCl):**

  Formula mass = 58.5 u

  Molar mass of NaCl = **58.5 g mol^{-1}** and it contains 6.022×10^{23} NaCl units.

In short, the relationship between moles, mass and number of particles can be summed up as follows:

Moles	Number of Particles	Mass (g)
1.0	N_A or 6.022×10^{23}	M (molar mass)
n	$n \times N_A$	$n \times$ M

6.3.4. Mole and Volume

From Avogadro's Law it is known that equal volumes of all gases contain equal number of molecules under similar conditions of temperature and pressure. Conversely, we can say that different samples of gases containing equal number of molecules occupy same volume under similar conditions. We know one mole molecules of all gases contain same number (6.022×10^{23}) of molecules, therefore, they occupy same volume under similar conditions of temperature and pressure. The volume occupied by **one mole** molecules of a gaseous substance is called **Molar Volume**. ***One mole molecules of all gases occupy 22.4 litres at 273 K and 1 atm pressure (S.T.P.)***. Hence, molar volume of all gases at S.T.P. is 22.4 litres.

Thus, for the gaseous substances at S.T.P. (273.15 K, 1 atm) conditions,

One mole = 22.4 L (Molar volume)

It may be noted that 1 mole of gas at S.T.P. conditions (273.15 K, 1 bar) occupies a volume of 22.7 litres.

Remember

The following formulae are quite helpful in solving the problems on mole concept. In these formulae, N_A represents $\mathbf{6.022 \times 10^{23}}$.

1. Mass of **1 atom** of element
$$= \frac{\textbf{Molar mass of atoms}}{\mathbf{N_A}}$$
2. Mass of **1 molecule** of substance $= \dfrac{\textbf{Molar mass}}{\mathbf{N_A}}$
3. **Number of molecules** in **n** moles of substance = $\mathbf{n \times N_A}$
4. **Mass of n moles = n × Molar mass**
5. **Number of moles** in Wg of a substance $= \dfrac{\mathbf{W}}{\textbf{Molar Mass}}$
6. 1 **u** or (1 **amu**) = $\mathbf{1/N_A}$ **g**.
   It is also called **1 avogram** or **1 dalton**.

Summary of various relationships of mole has been illustrated in Fig. 6.2.

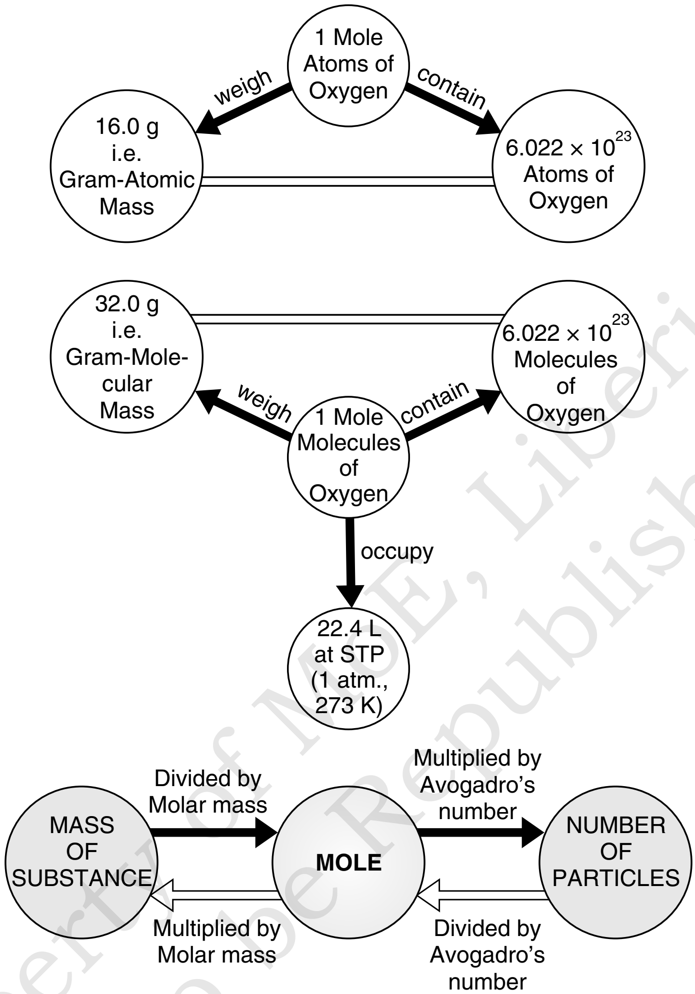


Fig. 6.2. Relationship between mole, number of particles and mass and interconversion of one into the other.

Example 6.5: *(a) Calculate the mass of 2.5 moles of calcium. Atomic mass of calcium is 40. (b) Calculate the mass of 1.5 mole of water (H_2O).*

Solution:

(a) Mass of 1 mole of calcium

$$= \text{molar mass of calcium} = 40 \text{ g mol}^{-1}$$

Mass of 2.5 moles of calcium

$$= 40 \times 2.5 = \mathbf{100\ g}.$$

(b) Molecular mass of water (H_2O)

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

$\therefore$ Molar mass of H_2O = 18 g mol^{-1}.

Mass of 1.5 mole of H_2O

$= 1.5 \times 18 = \mathbf{27\ g.}$

Example 6.6: *Calculate the molar mass of glucose ($C_6H_{12}O_6$).*

Solution: Molecular mass of glucose ($C_6H_{12}O_6$)

$= 6 \times (12.011\ u) + 12 \times (1.008\ u) + 6 \times (16.00\ u)$

$= 72.066\ u + 12.096\ u + 96.00\ u$

$= \mathbf{180.162\ u}$

$\therefore$ Molar mass of glucose = 180.162 g mol^{-1}

Example 6.7: *Calculate mass of the following:*

(i) *one atom of calcium*

(ii) *one molecule of sulphur dioxide (SO_2).*

Solution:

(i) Mass of 6.022×10^{23} atoms of calcium

= gram atomic mass of calcium = 40 g

$\therefore$ Mass of 1 atom of calcium

$$= \frac{40\,\text{g}}{6.022 \times 10^{23}} = \mathbf{6.6 \times 10^{-23}\ g.}$$

(ii) Mass of 6.022×10^{23} molecules of SO_2

= molar mass of SO_2 = 64 g.

$\therefore$ Mass of 1 molecule of SO_2

$$= \frac{64\,\text{g}}{6.022 \times 10^{23}} = \mathbf{1.06 \times 10^{-22}\ g}.$$

Example 6.8: *How many atoms of oxygen are present in 300 g of $CaCO_3$?*

Solution: Molar mass of $CaCO_3$ = 100 g

Now, 1 mole of $CaCO_3$ contains

= 3 mole of oxygen atoms.

or 100 g of $CaCO_3$ contain

$= 3 \times 6.022 \times 10^{23}$ oxygen atoms

∴ 300 g of $CaCO_3$ contain oxygen atoms

$$= \frac{3 \times 6.022 \times 10^{23}}{100} \times 300$$

$$= \mathbf{5.4198 \times 1024 \text{ oxygen atoms}}.$$

6.4. PERCENTAGE COMPOSITION

Percentage composition of a compound refers to the amount of various constituent elements present per 100 parts by mass of the substance. It can be calculated from the formula of the compound.

Knowing the molecular formula of a compound, we can calculate its percentage composition by mass. To do this, we calculate the molecular mass of the compound. From this we can find out mass of one mole of the compound, which is equal to its gram molecular mass. Then we calculate mass of each element in one mole of the compound. The mass percentage of each element is then calculated by the following formula:

Mass per cent of the element X

$$= \frac{\text{Mass of X in one mole}}{\text{Molar mass of the compound}} \times 100$$

Similarly, mass percentages of other elements can be calculated as illustrated in the solved examples given below:

Example 6.9: *Calculate the mass per cent of different elements present in ethanol (C_2H_5OH).*

Solution:

Molar mass of ethanol $= 2 \times 12.01 + 6 \times 1.008 + 16.00$

$= 46.068$ g

Mass per cent of carbon $= \frac{24.02}{46.068} \times 100 = \mathbf{52.14\%}$

Mass per cent of hydrogen $= \frac{6.048}{46.068} \times 100 = \mathbf{13.13\%}$

Mass per cent of oxygen $= \frac{16.00}{46.068} \times 100 = \mathbf{34.73\%}$.

6.5. DETERMINATION OF FORMULA OF A COMPOUND

The **molecular formula** of a compound may be defined as the *formula which gives the actual number of atoms of various elements present in the molecule of the compound*. For example, the molecular formula of the compound glucose can be represented as $C_6H_{12}O_6$. A molecule of glucose contains six atoms of carbon, twelve atoms of hydrogen and six atoms of oxygen.

In order to find out molecular formula of a compound, the first step is to determine its empirical formula from the percentage composition.

The **empirical formula** of a compound may be defined as *the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound*. The empirical formula of the compound glucose ($C_6H_{12}O_6$), is CH_2O which shows that C, H and O are present in the simplest ratio of 1 : 2 : 1.

Empirical formula mass of substance is equal to the sum of atomic masses of all the atoms in the empirical formula of the substance.

6.5.1. Relation between the Two Formulae

Molecular formula is whole number multiple of empirical formula. Thus,

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

where $n = 1, 2, 3...$

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

6.5.2. Steps for Writing the Empirical Formula

The percentage of the elements in the compound is determined by suitable methods and from the data collected, the empirical formula is determined by the following steps:

(i) Divide the percentage of each element by its atomic mass. This will give the relative number of moles of various elements present in the compound.

(ii) Divide the quotients obtained in the above step by the smallest of them so as to get a simple ratio of moles of various elements.

(iii) Multiply the figures, so obtained by a suitable integer if necessary in order to obtain whole number ratio.

(iv) Finally write down the symbols of the various elements side by side and put the above numbers as the subscripts to the lower right hand corner of each symbol. This will represent the **empirical formula** of the compound.

6.5.3. Steps for Writing the Molecular Formula

The following steps can be followed while writing the molecular formula:

(i) Calculate the empirical formula as described above.

(ii) Find out the empirical formula mass by adding the atomic masses of all the atoms present in the empirical formula of the compound.

(iii) Divide the molecular mass (determined experimentally by some suitable method) by the empirical formula mass and find out the value of *n*.

(iv) Multiply the empirical formula of the compound with n so as to get the **molecular formula** of the compound.

The determination of the empirical formula and molecular formula of a compound involving the above steps is illustrated by the following examples.

Example 6.10: *Write the empirical formula for each of the compounds having molecular formulae*:

(i) C_6H_6 (ii) C_6H_{12} (iii) H_2O_2
(iv) H_2O (v) Na_2CO_3 (vi) B_2H_6
(vii) N_2O_4 (viii) H_3PO_4 (ix) Fe_2O_3
(x) C_2H_2.

Solution: Empirical formula is the simplest whole number ratio of atoms of different elements in the molecule. Therefore, the empirical formula of given compounds are:

(i) CH (ii) CH_2 (iii) HO
(iv) H_2O (v) Na_2CO_3 (vi) BH_3
(vii) NO_2 (viii) H_3PO_4 (ix) Fe_2O_3
(x) CH.

Example 6.11: *What is the simplest formula of the compound which has the following percentage composition: Carbon 80%, Hydrogen 20%? If the molecular mass is 30, calculate its molecular formula.*

Solution: Calculation of empirical formula:

Element	***Percentage***	***At. mass***	***Relative no. of moles***	***Simple ratio***	***Simplest whole no. ratio***
C	80	12	$\frac{18}{12} = 6.66$	$\frac{6.66}{6.66} = 1$	1
H	20	1	$\frac{20}{1} = 20$	$\frac{20}{6.66} = 3$	3

Hence, the empirical formula is CH_3.

6.6. KINDS OF CHEMICAL REACTIONS

Generally, chemical reactions are of two types: (i) Reversible reaction and (ii) Irreversible reaction. Let us discuss these in detail.

6.6.1. Reversible Reaction

The reactions in which products under suitable conditions react to give back reactants are called ***reversible reactions***.

A reversible reaction is represented by putting two arrows pointing in opposite directions between the formulae of the reactants and the products as shown below:

6.6.2. Irreversible Reaction

When a piece of sodium is dropped into water, a violent reaction occurs resulting in the formation of sodium hydroxide and hydrogen gas.

$$2Na(s) + 2H_2O(l) \longrightarrow 2NaOH(aq) + H_2(g)$$

However, it is not possible to carry out the reverse reaction under any known experimental conditions, i.e., reduction of aqueous sodium hydroxide by hydrogen to form sodium and water cannot be achieved.

The reactions in which the products do not react under any condition to give back reactants are called ***irreversible reactions***.

6.7. TYPES OF REACTIONS

There are millions of known chemical reactions. Many of these chemical reactions have common aspects, so they can be grouped into specific classification. The majority of chemical reactions (not all) fall into the following major categories:

- Combination reactions
- Decomposition reactions
- Single replacement reactions
- Double displacement reactions
- Combustion reactions
- Oxidation reduction reactions

6.7.1. Composition/Combination Reactions

ACTIVITY 6.2

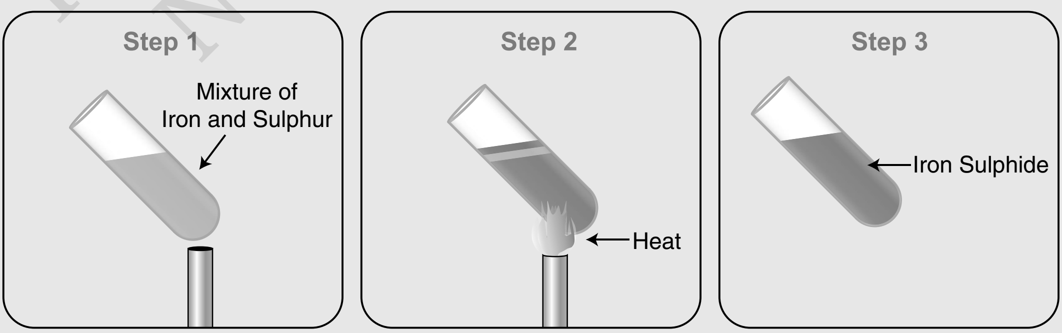
Combination of Iron and Sulphide (Combination Reaction)

Materials Required

Iron powder (filings) (7 grams), sulphur (4 grams), test tube, Bunsen burner, and pair of test tube tongs.

Procedure

1. Prepare a mixture containing iron powder and sulphur powder in the ratio 7:4 by mass.
2. Note the appearance of the pure elements and the mixture.
3. Take about 0.5 g of the mixture into a hard glass test tube.
4. Heat the mixture at the base of the test tube—gently at first and then more strongly (use a blue flame throughout).
5. Allow the test tube to cool down.
6. Observe the product.



Can you write the chemical equation for this reaction?

Safety

- It is advisable to wear protective gloves and heat the mixture using tongs.
- Eye protection required.

In this activity, you will observe that when we start heating a mixture of iron filings and sulphur, the sulphur melts and reacts with the iron filings to form the compound iron (II) sulphide. In this reaction, two different elements combine to form a single product. This is an example of **combination reaction**.

Combination reactions are those reactions in which a single product is produced from two (or more) reactants. The general equation for a combination reaction involving two reactants is

$$X + Y \longrightarrow XY$$

The reactants X and Y can be elements or compounds, or a compound and an element. The product XY is always a compound.

Combination reactions may involve:

- the combination of two elements to form a compound.
- the combination of a compound and an element to form a new compound.
- the combination of two compounds to form a new compound.

Example of some combination reactions are :

(i) $$\underset{\text{Iron}}{Fe\,(s)} + \underset{\text{Sulphur}}{S\,(s)} \xrightarrow{\text{Combination}} \underset{\text{Iron sulphide}}{FeS\,(s)}$$

(ii) $$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \xrightarrow{\text{Combination}} \underset{\text{Sodium chloride}}{2NaCl(s)}$$

(iii) $$\underset{\text{Hydrogen}}{H_2\,(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \xrightarrow{\text{Combination}} \underset{\text{Hydrogen chloride}}{2HCl(s)}$$

6.7.2. Single Replacement Reactions

ACTIVITY 6.3

Reaction of Iron Nails with Copper Sulphate

1. Take three iron nails and clean them by rubbing with sand paper.
2. Take two test tubes marked as (A) and (B). In each test tube, take about 10 ml of copper sulphate solution.
3. Tie two iron nails with a thread and immerse them carefully in the copper sulphate solution in test tube B for about 20 minutes [Fig. 6.3 (a)]. (If the solution is concentrated, leave it for 2 hours) keep one iron nail aside for comparison.

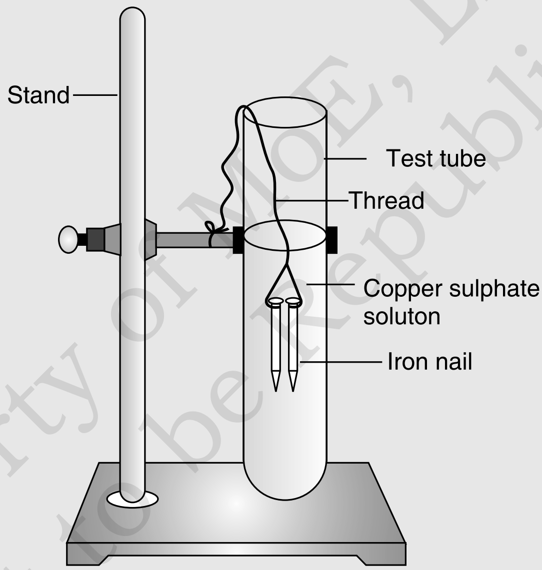


Fig. 6.3. (a) Iron nails dipped in copper sulphate solution

4. After 20 minutes, take out the iron nails from the copper sulphate solution.
5. Compare the intensity of the blue colour of copper sulphate solutions in test tubes (A) and (B), [Fig. 6.3 (b)].
6. Also, compare the colour of the iron nails dipped in the copper sulphate solution with the one kept aside [Fig. 6.3 (b)].

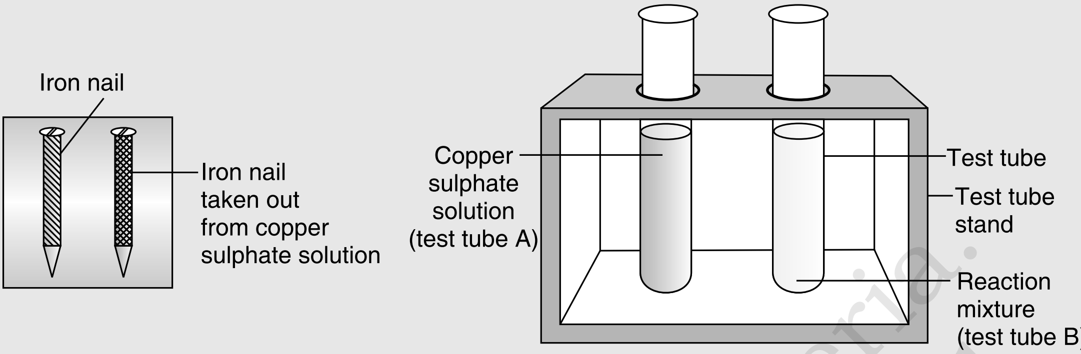


Fig. 6.3. (b) Iron nails and copper sulphate solutions compared before and after the experiment

Single replacement reactions are those reactions in which one element displaces another element from a compound. The general equation for this type of reaction is

$$X + YZ \longrightarrow XZ + Y$$

In reactants, X is an element and YZ is a compound. In products, XZ is a compound and Y is an element.

In single replacement reactions, more active metals displace less active metals (or hydrogen) from their compounds. Some common elements are arranged in decreasing order of their ability to replace element (metal ion) in aqueous solution. This series is known as activity series.

K > Na > Ca > Mg > Al > Zn > Cr > Fe > Ni > Sn > Pb > H > Cu > Ag > **Au**

Note: Potassium (K) is the *most reactive* metal and gold (Au) is the *least reactive* metal.

Example of some single replacement reactions are:

(i) $\underset{\text{(Blue solution)}}{\underset{\text{Copper sulphate}}{CuSO_4\,(aq)}} + \underset{\text{(Grey)}}{\underset{\text{Iron}}{Fe(g)}} \longrightarrow \underset{\text{(Greenish solution)}}{\underset{\text{Iron sulphate}}{FeSO_4(aq)}} + \underset{\text{(Red-brown)}}{\underset{\text{Copper}}{Cu\,(s)}}$

(ii) $\underset{\text{(Colourless solution)}}{\underset{\text{Silver nitrate}}{2AgNO_3\,(aq)}} + \underset{\text{(Red-brown)}}{\underset{\text{Copper}}{Cu(s)}} \longrightarrow \underset{\text{(Blue solution)}}{\underset{\text{Copper nitrate}}{Cu(NO_3)_2(aq)}} + \underset{\text{(Greyish white)}}{\underset{\text{Silver}}{2Ag\,(s)}}$

(iii) $\underset{\text{(Blue solution)}}{\underset{\text{Copper sulphate}}{CuSO_3\,(aq)}} + \underset{\text{(Silvery-white)}}{\underset{\text{Zinc}}{Zn(s)}} \longrightarrow \underset{\text{(Colourless solution)}}{\underset{\text{Zinc sulphate}}{ZnSO_4\,(aq)}} + \underset{\text{(Red-brown)}}{\underset{\text{Copper}}{Cu\,(s)}}$

6.7.3. Double Replacement Reactions

ACTIVITY 6.4

Formation of Precipitate

1. Take about 3 ml of sodium sulphate solution in a test tube.
2. In another test tube, take about 3 ml of barium chloride solution.
3. Mix the two solutions (Fig. 6.4).

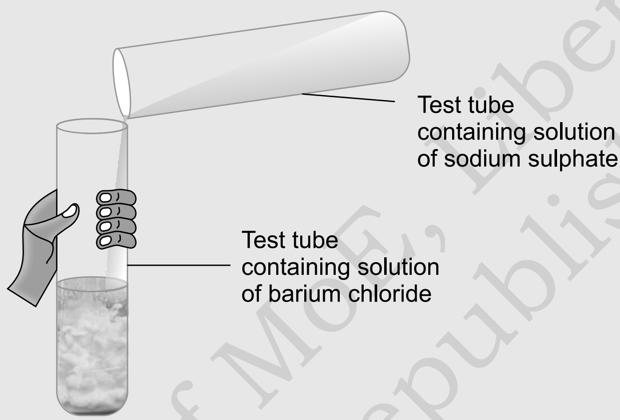


Fig. 6.4. Formation of Precipitate

What do you observe?

In Activity 6.4, you will observe that a white substance, which is insoluble in water, is formed.

This insoluble substance formed is known as a *precipitate*. Any reaction that produces a precipitate can be called a ***precipitation reaction***.

Reaction:

$$\underset{\text{Sodium Sulphate}}{Na_2SO_4\,(aq)} + \underset{\text{Barium chloride}}{BaCl_2(aq)} \longrightarrow \underset{\text{Barium sulphate}}{BaSO_4(s)} + \underset{\text{Sodium chloride}}{2NaCl\,(aq)}$$

Double displacement reactions are those reactions in which two compounds react by exchange of ions to form two new compounds. These reactions are also known as **metathesis reactions** or **exchange reactions**.

The general equation for a double displacement reaction is.

$$AX + BY \longrightarrow AY + BX$$

where A and B are positive ions (cations).

X and Y are negative ions (anions).

Another example of double displacement reaction is when we mix a solution of silver nitrate ($AgNO_3$) and sodium chloride, silver chloride and sodium nitrate is formed.

$$AgNO_3\ (aq) + NaCl\ (aq) \longrightarrow AgCl\ (s) + NaNO_3\ (aq)$$

Double displacement reactions result in the removal of ions from the solutions.

Examples of some double replacement reactions are:

(i)	$BaCl_2(aq)$	+	$CuSO_4(aq)$	$\longrightarrow$	$BaSO_4(s)$	+	$CuCl_2(aq)$
	Barium chloride		Copper sulphate		Barium sulphate (White ppt.)		Copper chloride
(ii)	$CuSO_4(aq)$	+	$H_2S\ (g)$	$\longrightarrow$	$CuS(s)$	+	$H_2SO_2(aq)$
	Copper sulphate		Hydrogen sulphide		Copper sulphide (Black ppt.)		Sulphuric acid
(iii)	$Ba(OH)_2(aq)$	+	$2HCl(aq)$	$\longrightarrow$	$BaCl_2(aq)$	+	$2H_2O(l)$
	Base		Acid		Salt		Water
(iv)	$2NaOH(aq)$	+	$H_2SO_4(aq)$	$\longrightarrow$	$Na_2SO_4(aq)$	+	$H_2O(l)$
	Base		Acid		Salt		Water

6.7.4. Decomposition Reactions

ACTIVITY 6.5

Decomposition of Silver Chloride

1. Take about 2 g of silver chloride in a China dish.
2. Note the colour of silver chloride.

3. Place the China dish in sunlight for 10–30 minutes.
4. Observe the colour of silver chloride after 30 minutes.

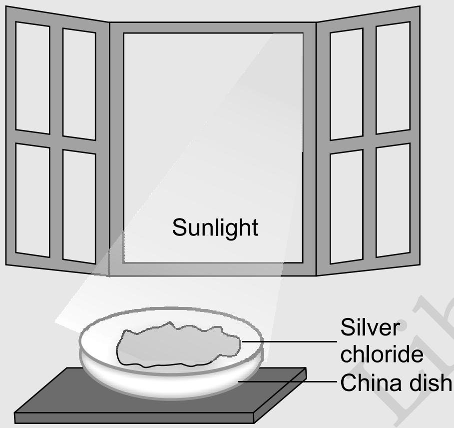


Fig. 6.5. Decomposition of $AgCl_2$

In Activity 6.5, you will observe that white silver chloride turns grey in sunlight. This is due to the decomposition of silver chloride into silver and chlorine by light.

Reaction:

$$2AgCl(s) \xrightarrow{\text{Sunlight}} 2Ag(s) + Cl_2(g)$$

Decomposition reactions are those reactions in which a single reactant breaks down into two (or more) simpler substances (elements or compounds). The general equation for a decomposition reaction in which there are two products, is

$$XY \longrightarrow X + Y$$

The reactant XY is always a compound. The products X and Y may be elements or compounds.

In other words, decomposition reactions are opposite of combination reactions. These reactions often involve an energy source such as heat, light, or electricity which breaks apart the bonds of compounds.

The products of decomposition reactions may be

- two elements
- one (or more) elements and one (or more) compounds
- two (or more) compounds.

Examples of some decomposition reactions are :

(i) $$\underset{\text{Water}}{2H_2O(l)} \xrightarrow[\text{Decomposition}]{\text{Electricity}} \underset{\text{Hydrogen}}{2H_2(g)} + \underset{\text{Oxygen}}{O_2(g)}$$

(ii) $$2H_2O_2(l) \xrightarrow{\text{light}} 2H_2O(l) + O_2(g)$$

(iii) $$\underset{\text{Hydrogen peroxide}}{2H_2O_2(l)} \xrightarrow{MnO_2} \underset{\text{Steam}}{2H_2O(g)} + \underset{\text{Oxygen}}{O_2(g)}$$

6.7.5. Combustion Reactions

ACTIVITY 6.6

Burning of Magnesium Ribbon

1. Take a piece of magnesium ribbon and hold it with a pair of tongs. Light magnesium ribbon.
2. The magnesium ribbon starts burning with a dazzling flame.

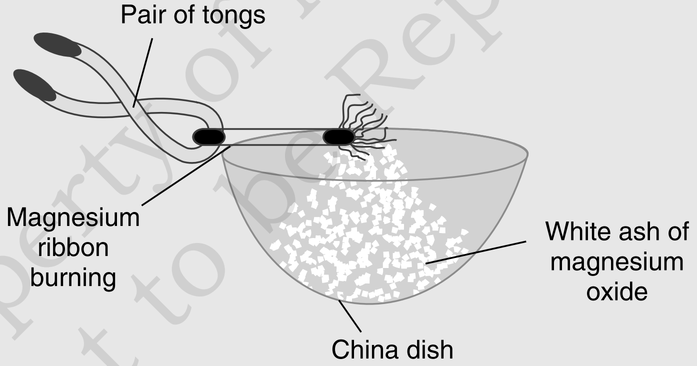


Fig. 6.6. Burning of magnesium ribbon in air

Explanation

You would observe that magnesium ribbon soon changes into white powder. This white powdery substance is magnesium oxide which is formed as a result of combination reaction.

Reaction

$$2Mg\ (s) + O_2\ (g) \longrightarrow 2MgO\ (s)$$

This type of reaction refers to the reaction of an element or compound with oxygen. Combustion usually releases a lot of heat energy. It is also referred to as *burning*.

Note: All combustion reactions are exothermic reactions.

6.7.5.1. Elements on Combustion

When elements undergo combustion, generally only one product is formed.

For example,

- Burning of coal

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

- Formation of water

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

- Burning of magnesium

$$2Mg\,(s) + O_2\,(g) \longrightarrow 2MgO\,(s)$$

6.7.5.2. Compounds on Combustion

When compounds undergo combustion, two or more products are formed. When carbon-hydrogen (hydrocarbon) or carbon-hydrogen-oxygen compounds undergo combustion in an excess of oxygen, the products are carbon dioxide and water. For example,

- Burning of natural gas

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

- Glucose reacts with oxygen

$$\underset{\text{(glucose)}}{C_6H_{12}O_6\,(aq)} + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(l)$$

ACTIVITY 6.7

Oxidation and Reduction Reactions

1. Take 1g of copper powder in the China dish.
2. Put the tripod stand over the burner and the wire gauze on top of it.
3. Now put the China dish containing copper powder over the wire gauze.

4. Turn on the burner and heat the China dish.
5. Take the black copper(II) oxide powder and heat it again with hydrogen gas.

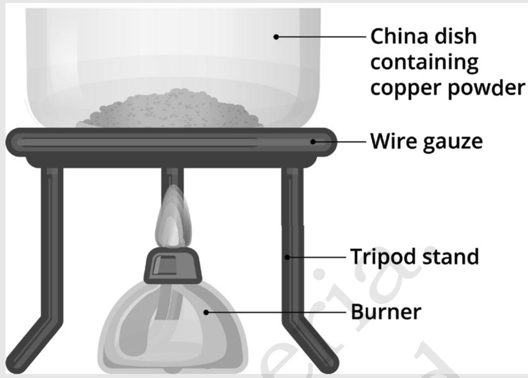


Fig. 6.7. Heating copper powder

Observation:

- After heating, we observed that the copper powder becomes coated with black colour.
- This black coloured powder is copper(II) oxide.
- The black powder of copper(II) oxide turns brown.

Result:

After heating, copper becomes copper oxide because of oxidation; when a chemical reaction gains oxygen, it is said to be oxidised.

The black powder of copper(II) oxide turns brown again, and we get copper powder back. This reaction happens because of the reduction. When there is a loss of oxygen in a chemical reaction, it is said to be reduced.

Reactions:

Oxidation reaction:

$$2Cu + O_2 \xrightarrow{\triangle} 2CuO$$

Redox reaction:

$$CuO + H_2 \xrightarrow{\triangle} Cu + H_2O$$

Reduction

$$CuO + H_2 \xrightarrow{\text{Heat}} Cu + H_2O$$

Oxidation

In this activity, the copper(II) oxide is loses oxygen during the reduction reaction. The copper(II) oxide is reduced to copper. Thus, ***reduction*** *is a process of losing oxygen.*

The hydrogen gains oxygen, and is being oxidised. Thus, ***oxidation*** *is the process of receiving oxygen.*

An oxidation-reduction reaction occurs when one reactant receives oxygen while the other reactant loses oxygen or when one reactant is oxidised while the other reactant is reduced at the same time. The oxidation-reduction reaction is also known as the redox reaction.

Here are some more examples of reduction reactions:

(i) *Reaction of PbO and carbon*:

Oxidation

$$PbO + C \longrightarrow Pb + CO$$

Reduction

Here, oxygen is being removed from lead oxide (PbO) and is being added to carbon (C). Thus, PbO is reduced while C is oxidised.

(ii) *Reaction of H_2S and Cl_2*:

Oxidation

$$Cl_2 + H_2O \longrightarrow 2HCl + S$$

Reduction

6.8. WRITING AND BALANCING CHEMICAL EQUATIONS

A chemical change can be represented by using symbols and formulae of various species involved in the change. Such a representation is known as chemical equation.

Representation of a chemical change in terms of symbols and formulae of the reactants and products is known as ***chemical equation*** *of the reaction.*

For example, when zinc carbonate is heated, zinc oxide and carbon dioxide are formed. This chemical change may be represented by the chemical equation given below:

$$ZnCO_3 \longrightarrow ZnO + CO_2$$

6.8.1. Writing Chemical Equation

Following are various steps of writing a chemical equation:

Step I: Writing Skeletal Equation

In order to write a chemical equation for a reaction, the first step is to write symbols and formulae of various reactants and products. For example, when zinc carbonate reacts with dilute hydrochloric acid it gives zinc chloride, carbon dioxide and water. This is represented by the following **'word equation'**.

In a word equation, we express the chemical change or chemical reaction by writing names of the reactants on left hand side and names of the products on right hand side. An arrow is put between reactants and products.

$$\text{Zinc carbonate} + \text{Hydrochloric acid} \longrightarrow \text{Zinc chloride} + \text{Carbon dioxide} + \text{Water}$$

Using symbols and formulae for various reactants and products we get,

$$ZnCO_3 + 2HCl \longrightarrow ZnCl_2 + CO_2 + H_2O$$

- Reactants are written on left hand side while products are written on right hand side.
- The symbols and formulae of various reacting substances (reactants) are separated by a plus sign (+) between them.
- The symbols and formulae of various substances formed (**products**) are also separated by a plus sign (+) between them.
- An arrow ($\longrightarrow$) is put between reactants and products.

Equations such as given above in which no attempt has been made to equalize the number of atoms of various elements on both the sides of the equation are known as **skeletal equations**. Thus:

*A **skeletal equation** is an equation in which various reactants and products are represented by their respective formulae but no attempt is made to equalize the number of atoms of various elements on both the sides of the equation.*

Step II : Balancing of Chemical Equation

We know from Dalton's atomic theory that atoms can neither be created nor destroyed. Moreover, from law of conservation of mass we know that mass is neither created nor destroyed during a chemical reaction.

Therefore, number of atoms of various elements on both the sides of the chemical equation must be equal.

After writing skeletal equation, the next step is to equalize the number of atoms of various elements on both the sides of the equation by multiplying various formulae by appropriate coefficients. This process is known as **balancing of chemical equations** which is discussed later in this section in details.

In order to equalize the number of atoms of various elements, the various species are multiplied by appropriate numbers. For example, formation of ammonia is represented by the balanced chemical equation given below:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

An equation in which number of atoms of each element is equal on both the sides of the equation is known as ***balanced chemical equation****.*

Sometimes the skeletal equation and the balanced equation may be identical. Some examples are given below:

$$CaCO_3 \xrightarrow{\text{Heat}} CaO + CO_2$$

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2$$

6.8.2. Balancing of Simple Chemical Equations

The balancing of chemical equations is based upon Dalton's atomic theory and law of conservation of mass. According to Dalton's atomic theory atom is the smallest unit of an element that takes part in chemical reactions and that during chemical reactions atoms are neither created nor destroyed. Therefore, **the number of atoms of each element should remain same before and after the reaction**.

To make the number of atoms of all the elements equal on both the sides in a skeletal equation is known as balancing.

During balancing, the symbols and formulae of various species given in the equation are multiplied with appropriate coefficients. A coefficient is a small whole number, like the coefficients used in algebraic equations. The simple equations can be balanced by **Hit and trial** method. The following steps may be followed while balancing the chemical equations by this method:

Step 1: Write the correct skeleton equation. The skeleton equation contains the formulae of only one molecule of each reactant and product. Once you are satisfied that the skeleton equation is correct, do not change the subscripts in any of the formulae.

Step 2: Start with the compound that has the maximum atoms or the maximum kinds of atoms and the atoms present in it are balanced first.

Step 3: Balance elements that appear only once on each side of the arrow first. Then balance elements that appear more than once on a side.

Step 4: Elementary substances are balanced last of all.

Step 5: If required the whole equation is multiplied by some suitable number in order to make all the coefficients whole numbers.

6.8.2.1. Balancing of Ionic Equations

A molecular equation is balanced on the basis of **mass or material balancing** by equating the number of atoms of various elements on both the sides of the chemical equation. However, in case of ionic equations we have to consider not only mass or material balancing but also **charge balancing**.

For example, consider the following equation:

$$Zn + Ag^{+} \longrightarrow Zn^{2+} + Ag$$

This equation is balanced as far as mass balancing is considered. However, there is + 1 charge on reactant side and + 2 charge on product side. In order to balance charges Ag^{+} on reactant side and Ag on product side are multiplied by 2. Thus, the correct balanced equation is

$$Zn + 2Ag^{+} \longrightarrow Zn^{2+} + 2Ag$$

Thus, **a balanced chemical equation must satisfy mass balance as well as charge balance**. The application of these rules is illustrated in the following solved examples.

Example 6.12: *Aluminium reacts with oxygen on heating to form aluminium oxide. Write balanced equation for the reaction.*

Solution: Aluminium + Oxygen $\longrightarrow$ Aluminium oxide

The skeleton equation for the reaction is

$$Al + O_2 \longrightarrow Al_2O_3$$

Start with Al_2O_3 because it contains maximum number of atoms and most kinds of atoms. To balance aluminium atoms multiply Al by 2.

$$2Al + O_2 \longrightarrow Al_2O_3$$

In order to balance oxygen atoms, multiply O_2 by $\frac{3}{2}$ so that number of O atoms becomes three on both the sides.

$$2Al + \frac{3}{2}\,O_2 \longrightarrow Al_2O_3$$

In order to make all the coefficients whole numbers, the whole equation is multiplied by 2.

$$2 \times \left[2Al + \frac{3}{2}O_2 \longrightarrow Al_2O_3 \right]$$

$$\mathbf{4Al + 3O_2 \longrightarrow 2Al_2O_3}$$

The equation is balanced. There are 4 aluminium atoms and 6 oxygen atoms on each side.

6.9. LIMITING REAGENTS/REACTANTS

In many situations during the chemical reactions, one of the reactants is taken in excess. This is to ensure the completion of reaction. However, on completion of reaction, some of the reactant, taken in excess, is left over. For example, consider the combustion of hydrogen

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

Suppose that 2 moles of H_2 and 2 moles of O_2 are available for reaction. It follows from the equation that only 1 mole of O_2 is required for complete combustion of 2 moles of H_2; 1 mole of O_2 will, therefore, be left over on completion of the reaction. The amount of the product obtained is determined by the amount of the reactant that is completely consumed in the reaction. This reactant is called the *limiting reagent*. Thus, **limiting reagent** may be defined as *the reactant which is completely consumed during the reaction*. The reactant which is not completely consumed is referred to as **excess reactant**.

In the above example, H_2 is the limiting reagent and O_2 is excess reactant. The amount of H_2O formed will, therefore, be determined by the amount of H_2. Since 2 moles of H_2 are taken, it will form 2 moles of H_2O on combustion.

Example 6.13: *How much magnesium sulphide can be obtained from 2.00 g of magnesium and 2.00 g of sulphur by the reaction Mg + S → MgS? Which is the limiting reagent? Calculate the amount of one of the reactants which remains unreacted. [Atomic masses: Mg = 24.3, S = 32.1]*

Solution: First of all each of the masses are expressed in moles :

$$2.00 \text{ g of Mg} = \frac{2.00}{24.3} = 0.0824 \text{ moles of Mg}$$

$$2.00 \text{ g of S} = \frac{2.00}{32.1} = 0.0624 \text{ moles of S}$$

From the equation, $Mg + S \rightarrow MgS$, it follows that one mole of Mg reacts with one mole of S. We are given more moles of Mg than of S. Therefore, Mg is in excess and some of it will remain unreacted when the reaction is over. **S is the limiting reagent** and will control the amount of product. From the equation we note that one mole of S gives one mole of MgS, so 0.0624 mole of S will react with 0.0624 mole of Mg to form 0.0624 mole of MgS.

Molar mass of MgS = 56.4 g

$\therefore$ Mass of MgS formed

= 0.0624 × 56.4 g = **3.52 g of MgS**

Mole of Mg left unreacted

= 0.0824 – 0.0624 moles of Mg

= 0.0200 moles of Mg

Mass of Mg left unreacted

= moles of Mg × molar mass of Mg

= 0.0200 × 24.3 g of Mg = **0.486 g of Mg.**

EXPERIMENT 6.1

Aim: ***To explore a reaction with a limiting reactant.***

You will need: Zinc metal (powder), hydrochloric acid, 3 flasks and 3 balloons.

Safety: Goggles must be worn and hold the balloons on the test tubes tightly while the reaction takes place.

Procedure:

1. Take three flasks of equal sizes and label them as A, B and C.
2. Weigh the following three amounts of Zn powder: 7.000 grams, 3.270 grams, and 1.310 gram.
3. Take three balloons and label them as 1, 2 and 3. Put the three different masses of zinc powder into the balloons using a small plastic funnel.

   Note: Make sure the zinc powders goes to the bottom of the balloon.
4. Using the graduated cylinder and pipette, accurately measure and transfer 0.100 mole of hydrochloric acid into each of the three flasks.
5. Attach the filled balloons to the mouth of the flasks.

   Note: Make sure that the contents of the balloon and flasks are not mixed.

6. After the balloons are securely attached to the flasks, lift the balloons on the flasks so that the contents of the balloon mix with the contents of the flask.

   Note: Make sure the balloons are held on tightly to the flasks.

   What do you observe?

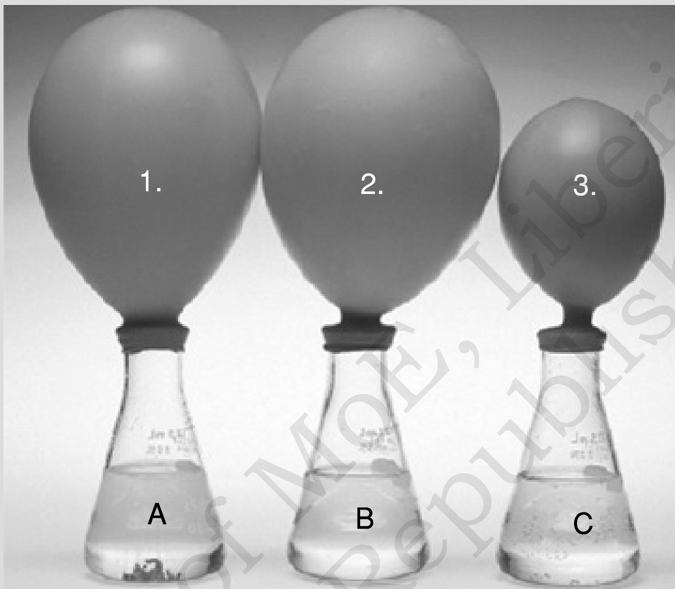


Fig. 6.8. Reaction with limiting reactant

Observation:

When the reactants are combined, the following reaction occurs:

$$Zn\ (s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

The H_2 gas inflates the balloon attached to the flask. The results are as follows:

- **Flask A:** Balloon inflates completely, but some Zn remains when inflation ceases.
- **Flask B:** Balloon inflates completely. No Zn remains.
- **Flask C:** Balloon does not inflate completely. No Zn remains.

Explanation:

In flask A, there is four times the stoichiometric quantity of Zn present, so the balloon inflates to a certain extent because all of the HCl reacts to form H_2 gas; and excess Zn that did not react is visible in the bottom of the flask.

The flask B, contains stoichiometric equivalent quantities of both the reactants, so the balloon inflates to the same extent as the flask A because all of the HCl reacts to form H_2 gas; all of the Zn is used up. It takes longer for the balloon to inflate to the same extent as the first balloon because the reaction slows down considerable as the concentration of Zn and HCl approaches zero towards the end of this reaction.

In flask C, there is one quarter of the stoichiometric quantity of Zn, so the balloon is noticeably smaller than the other two since the Zn is used up before all the HCl is converted to H_2 gas.

GLOSSARY

- **Anion:** A negatively charged ion.
- **Atom:** The smallest particle of an element that takes part in chemical reactions.
- **Cation:** A positively charged ion.
- **Combustion reactions:** A combustion reaction is a reaction of a substance with oxygen and is generally accompanied by rapid release of heat to produce a flame.
- **Decomposition reactions:** The reaction in which a single compound breaks up into two or more simpler substances.
- **Displacement reactions:** The chemical reactions in which one element displaces another element from a compound and takes its place.
- **Endothermic:** A chemical reaction in which heat energy is absorbed.
- **Exothermic:** A chemical reaction in which heat energy is released.
- **Ion:** An atom or group of atoms carrying positive or negative charge.
- **Molar mass:** The mass of one mole, i.e., 6.023×10^{23} particles of that substance.
- **Molar volume:** The volume occupied by one mole molecules of a gaseous substance.
- **Molecule:** The smallest particle of a substance that has independent existence.
- **Oxidation:** It is the process of gain of oxygen or loss of hydrogen.
- **Reduction:** It is the process of loss of oxygen or gain of hydrogen.

SUMMARY

- Law of Conservation of Mass states that "In every chemical reaction, the total mass before and after the reaction remains constant."
- Law of Constant Composition states that "A pure chemical compound always contains same elements combined together in the same proportion by mass."
- Law of multiple proportion states that "When two elements combine with each other to form two or more than two compounds, the masses of one element which combine with fixed mass of the other, bear a simple whole number ratio to one another."
- The molecular formula of a compound may be defined as the formula which gives the actual number of atoms of various elements present in the molecule of the compound.
- The empirical formula of a compound may be defined as the formula which gives the simplest whole number ratio of atoms of the various elements present in the molecule of the compound.
- The reactions in which the products do not react under any condition to give back reactants are called irreversible reactions.
- The reactions in which products under suitable conditions react to give back reactants are called reversible reactions.
- An equation in which number of atoms of each element is equal on both the sides of the equation is known as balanced chemical equation.
- To make the number of atoms of all the elements equal on both the sides in a skeletal equation is known as balancing.
- Limiting reagent may be defined as the reactant which is completely consumed during the reaction.

EVALUATION

I. Multiple Choice Questions

1. Who proposed law of definite proportions?

(a) John Dalton (b) Joseph Proust

(c) Robert Boyle (d) None of these

2. Molar volume of all gases at S.T.P. is:

(a) 24.2 litres
(b) 22.4 litres
(c) 22.5 litres
(d) None of these

3. Which of the following contains one mole molecules of the substance?

(a) 16 g Oxygen
(b) 7 g Nitrogen
(c) 2 g Hydrogen
(d) 36 g Water

4. When a single product is produced from two or more reactants, the reaction is

(a) Metathesis reaction
(b) Decomposition reaction
(c) Combination reaction
(d) Displacement reaction

5. Combination reactions may involve

(a) Combination of two elements
(b) Combination of two compounds
(c) Combination of one element and one compound
(d) All of these

6. Name the reaction in which energy in the form of heat, light and electricity is required to complete the reaction.

(a) Combination
(b) Decomposition
(c) Single replacement
(d) Double replacement

7. These reactions are also known as Metathesis reaction.

(a) Combustion reaction
(b) Decomposition reaction
(c) Combination reaction
(d) Double displacement reaction

8. This is the most reactive metal.

(a) Potassium
(b) Gold
(c) Aluminium
(d) Iron

9. It is the process of gain of oxygen or loss of hydrogen.

(a) Oxidation
(b) Reduction
(c) Oxidation-Reduction
(d) None of these

10. Which of the following is a characteristic of a reversible reaction?

(a) It never proceeds to completion in a closed container.
(b) It proceeds only in forward direction.
(c) Number of moles of reactants and products are equal.
(d) It can be influenced by a catalyst.

II. State True or False

1. Law of conservation of mass is also known as law of indestructibility of matter.
2. Carbon dioxide contains carbon and oxygen in the ratio of 3 : 5.
3. Atomic radius is measured in nanometres.
4. The balancing of chemical equations is based upon Dalton's atomic theory and law of conservation of mass.
5. In a chemical reaction, the reactant which is completely consumed is called excess reactant.
6. Gold is the least reactive metal.
7. Reduction is the process of loss of oxygen or gain of hydrogen.
8. All combustion reactions are endothermic.
9. In a chemical equation, reactants are written on the right hand side.
10. Cations are also called basic radicals.

III. Answer the Following Questions

1. State law of conservation of mass.
2. On heating, potassium chlorate decomposes to potassium chloride and oxygen. In one experiment 30.0 g of potassium chlorate generates 14.9 g of potassium chloride and 9.6 g of oxygen. What mass of potassium chlorate remain sun decomposed?
3. State law of definite proportion.
4. 1.375 g of cupric oxide was reduced by heating in a current of hydrogen and the weight of copper that remained was 1.098 g. In another experiment, 1.179 g of copper was dissolved in the nitric acid and the resulting copper nitrate converted into cupric oxide by ignition. The weight of cupric oxide formed was 1.476 g. Show that these results illustrate the law of definite proportion.
5. State law of multiple proportion.
6. Carbon and oxygen are known to form two compounds. The carbon content in one of these is 42.9% while in the other it is 27.3%. Show that this data is in agreement with law of multiple proportions.
7. Define an atom.
8. Why is it not possible to see an atom with naked eyes?
9. Define molecule.
10. What is the difference between an atom and a molecule?
11. What is the difference between a molecule of an element and a molecule of a compound?
12. What is an ion?

13. Define mole.

14. Calculate the number of gram-atoms and gram-moles in 25.4 mg of iodine (I_2). Atomic mass of I = 127 u.

15. What is molecular formula? Give an example.

16. What are the steps for writing molecular formula?

17. What is empirical formula? Give an example.

18. What are the steps for writing empirical formula?

19. Find empirical formula for each of the compound having molecular formulae:

(i) H_2O (ii) Na_2CO_3

(iii) N_2O_4 (iv) Fe_2O_3

20. Explain reversible reaction citing an example.

21. Explain irreversible reaction citing an example.

22. What is a chemical equation? Give an example.

23. What are the two steps to write a chemical equation?

24. What do you mean by skeletal equation?

25. What is a balanced chemical equation? Why should the chemical equations be balanced?

26. What do you mean by the term limiting reagents?