UNIT 10 CHEMICAL REACTIONS

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10.1 INTRODUCTION

The word ‘chemical reactions’ is used to represent changes taking place in our surroundings. In the whole universe, these changes are taking place. Some are slow and some are fast. To represent the reactions, chemical equations are required which are written with the help of symbols and formulae. The reactions always involve energy changes. By studying such reactions mankind can devise various methods to generate more energy. Mole concept is used to measure the amounts and ratios of substances taking part in reactions.

When we study different reactions, we find that many of the reactions do not proceed towards completion. The different factors affecting the state of equilibrium can lead us to device methods to get better yield of compounds.

10.2 OBJECTIVES

At the end of this unit, you will be able to:

- write the formulae of different compounds;
- balance chemical equations for different type of reactions;
- use mole concept in balancing chemical equations;
- use mole concept to calculate the amount of substances used and produced in different reactions; volumes of reactants and products in the case of gases;
Teaching of Chemistry

- differentiate between the different types of reactions;
- balance chemical equations on the basis of oxidation and reduction;
- understand the mechanism of energy changes taking place in different reactions;
- understand the dynamics of various reactions;
- apply the knowledge of equilibrium to get better yield of various compounds in industries.

10.3 REPRESENTING A CHEMICAL REACTION

10.3.1 Concept of Formula

Main Teaching Points

1. Formula: The formula of a compound is the short hand representation of a molecule and consists of atoms of same or different elements. Symbols of various elements are used for writing a formula.

2. Types of Formula
   a) Empirical Formula
   b) Molecular Formula

   a) Empirical Formula represents the simplest whole number ratio of atoms of the various elements present in a molecule of the compound. For example, empirical formula of glucose ($C_6H_{12}O_6$) is $CH_2O$.

   Empirical formula of a compound can be determined by calculating the percentage composition of different elements in a compound.

   b) Molecular Formula

   The molecular formula of a compound represents the actual number of atoms of various elements present in one molecule of the compound. Molecular formula of glucose is ($C_6H_{12}O_6$).

   Atoms do not exist separately but join to form molecules.

Teaching Learning Process

What is Formula?

A molecule of chlorine is made up of two atoms. Therefore, a molecule of chlorine is written as $Cl_2$ and not $2Cl$ (which stands for two atoms of chlorine). Likewise a molecule of phosphorus ($P_4$) contains four atoms.

The molecule of a compound consists of atoms of same or different elements. The different atoms are combined in a certain fixed ratio. Therefore, to write the formula of a compound, the symbols of the elements it is made of are written side by side and the numbers of atoms of each element are indicated by a subscript figure.

Thus, a formula is a brief representation of a chemical compound in terms of symbols of various elements present in it.

Empirical and Molecular Formulae

a) Introduction: Let us take a container full of rubber balls of different colours, in the following proportions: 10% by weight being red balls each weighing 5 grams, 40% by weight being yellow balls each weighing 10 grams and remaining 50% by weight being green balls each weighing 10 grams. The relative number of red, yellow and green balls can be easily found by dividing the percentage of each by its weight.
Red balls = \frac{10}{5} = 2
Yellow balls = \frac{40}{10} = 4
Green balls = \frac{50}{10} = 5

Thus, it is evident that for every 5 green balls, there are 4 yellow and 2 red balls.

In the same manner, by knowing the percentages of different elements in a molecule and their atomic weights, we can find the relative number of various atoms present in it. Percentage composition of an organic compound is determined from the results of quantitative analysis. After knowing this, it is possible to calculate the empirical formula.

b) **Empirical formula** represents the simplest whole number ratio of atoms of the various elements present in a molecule of the compound.

For example, empirical formula of glucose (C\textsubscript{6}H\textsubscript{12}O\textsubscript{6}) is CH\textsubscript{2}O. It represents the atomic ratio C:H:O in glucose as 1:2:1.

In order to determine the empirical formula, the following steps are involved:

i) Divide the % composition of different elements by their respective atomic weights. This gives the relative number of atoms i.e., the atomic ratio.

Thus, atomic ratio = \frac{\text{Percentage of an element}}{\text{Atomic mass of the same element}}

ii) Divide each of these atomic ratios by the smallest number of the atomic ratio. It gives the simplest ratio between the atoms of various elements constituting the compound.

iii) The results in the second step are simple whole numbers. If the figures are not whole numbers, then these are made simple whole number ratios by multiplying them all by the smallest integer.

iv) Symbols of the various elements are written side by side and their respective whole number ratio as sub-scripts on their right.

**Molecular Formula**

The molecular formula of a compound represents the actual number of atoms of various elements present in one molecule of the compound.

For example, molecular formula of glucose is C\textsubscript{6}H\textsubscript{12}O\textsubscript{6} which reveals that one molecule of glucose contains 6 atoms of carbon, 12 atoms of hydrogen and 6 atoms of oxygen.

Molecular formula is either the same or a simple multiple of the empirical formula.

Molecular formula = n \times \text{Empirical formula} where n is a whole number like 1, 2, 3, 4 etc. The value of n is obtained by dividing the molecular weight by empirical formula weight.

n = \frac{\text{Molecular Weight}}{\text{Empirical Formula Weight}}

**Calculation of Empirical and Molecular Formula**

An organic compound contains C = 12.76%
H = 2.13%
Br = 85.11%

Its vapour density (a parallel term for atomic mass in gaseous compounds) is 94. Find its molecular formula.
Teaching of Chemistry

<table>
<thead>
<tr>
<th>Element</th>
<th>%</th>
<th>Atomic weight</th>
<th>Relative number of atoms</th>
<th>Simplest no.</th>
<th>Whole ratio</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>12.76</td>
<td>12</td>
<td>( \frac{12.76}{12} = 1.06 )</td>
<td>( \frac{1.06}{1.06} = 1 )</td>
<td></td>
</tr>
<tr>
<td>H</td>
<td>2.13</td>
<td>1</td>
<td>( \frac{2.13}{1} = 2.13 )</td>
<td>( \frac{2.13}{1.06} = 2 )</td>
<td></td>
</tr>
<tr>
<td>Br</td>
<td>85.11</td>
<td>80</td>
<td>( \frac{85.11}{80} = 1.06 )</td>
<td>( \frac{1.06}{1.06} = 1 )</td>
<td></td>
</tr>
</tbody>
</table>

Therefore empirical formula is \( \text{CH}_2\text{Br} \)

\[
\text{Empirical Formula wt.} = 1 \times 12 + 2 \times 1 + 1 \times 80 \\
= 12 + 2 + 80 \\
= 94
\]

\[
\text{Molecular wt.} = 2 \times \text{v.d} \\
= 2 \times 94 = 188
\]

\[
n = \frac{\text{Molecular Wt.}}{\text{Empirical Formula Wt.}} = \frac{188}{94} = 2
\]

\[
\text{Molecular Formula} = n \times \text{Empirical Formula} \\
= 2 \times [\text{CH}_2\text{Br}] \\
= \text{C}_2\text{H}_4\text{Br}_2
\]

Methodology used: The concept of formula with reference to empirical and molecular formula is made clearer by giving different examples.

Check Your Progress

Notes: a) Write your answers in the space given below.
    b) Compare your answers with those given at the end of the unit.

1. Define and explain the significance of the term formula.
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2. Assign formula to the following compounds:
   a) Ammonium dichromate
   b) Silver Sulphide
   c) Iron Phosphate
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3. Percentage composition of an organic compound is as follows:

\[\text{C} = 10.06\%, \text{H} = 0.84\% \text{ and } \text{Cl} = 89.10\%\]

If its V.D. is 60, calculate its molecular formula.

4. Percentage composition of a compound is as follows:

\[\text{C} = 12.76\%, \text{H} = 2.13\%, \text{Br} = 85.11\%\]

Calculate its empirical formula.

10.3.2 Information Derived out of Chemical Equations

**Main Teaching Points**

Chemical reactions can be described by chemical equations. A chemical equation gives us quantitative as well as qualitative information.

We come to know the names and formulae of the reactants and products. It also tells about the number of molecules of the reactants and products, volumes of gaseous reactants and products.
However, a chemical equation can’t tell us:

i) the physical state
ii) concentration
iii) speed of reaction
iv) time required for its completion
v) conditions necessary to carry on the reaction

Teaching Learning Process

It is convenient to represent an element by means of symbols and chemical compounds by their formulae. Similarly chemical reactions may be represented by chemical equations.

A chemical equation is a statement that describes a true chemical change in terms of symbols and formulae. The substances which can bring about chemical change, are known as reactants and the substances which are produced as a result of chemical change, are known as products.

A chemical equation gives us both quantitative and qualitative information.

i) Qualitatively, the chemical equation provides us the names and formulae of the reactants and the products of a chemical reaction.

ii) Quantitatively, it represents
   a) The relative number of molecules of the reactants as well as the products.
   b) The relative masses of the reactants and products.
   c) The relative volumes of gaseous reactants and products.

Limitation of a chemical equation and their removal: The chemical equation can’t tell us:

i) The physical state of the reactants and products.
ii) The concentration of reactants or products.
iii) Speed of reaction.
iv) The time required to complete the reaction.
v) Energy changes taking place in a chemical reaction.
vi) The conditions e.g. temperature, pressure etc. necessary to start and carry on the reaction.

vii) Whether the reaction is reversible or not.
viii) Whether the reaction proceeds with the separation of a precipitate or with evolution of a gas.

However, the equation can be modified or improved to provide information as discussed below:

1) The state of any substance may be indicated by writing abbreviations: (g) for gas, (l) for liquid, (s) for a solid and (aq) for an aqueous solution.

2) The concentration of the reactants is indicated by putting the word ‘conc.’ for concentrated and ‘dil.’ for dilute before the formulae, for e.g.,

   \[ \text{Zn}^{(s)} + \text{dil.H}_2\text{SO}_4 \rightarrow \text{ZnSO}_4 \text{ (aq)} + \text{H}_2 \text{ (g)} \]

3) + Q (in the case of exothermic reactions) and − Q (in the case of endothermic reactions) is written alongside the products in a chemical equation to indicate heat changes.

4) Evolution of gas can be indicated by arrow heads pointing upwards and precipitation can be indicated by arrow heads pointing downwards alongside the formulae of the compound.

5) The conditions of the reaction such as temperature, pressure, catalyst or reversibility etc. are indicated on the arrow head present in between the reactions and the products with suitable symbols. For example, a reversible reaction can be written as

   \[ 2\text{SO}_2 \text{ (g)} + \text{O}_2 \text{ (g)} \rightleftharpoons 2\text{SO}_3 \text{ (g)} + \text{Q} \]
Methodology used: Through various examples, it is made clear how a chemical equation gives the quantitative as well as qualitative information. Special emphasis is given on facts which are not made clear by an equation i.e. physical state of reactants and products, concentration, speed etc. Discussion-cum-demonstration method is used to explain this. Simple reactions between substances, evolution of gases as products, formation of precipitate in some reactions are shown through experiments. For example, reactions between Zn metal and acids like HCl, H₂SO₄, HNO₃ etc. can be demonstrated to show the speed of reaction.

Physical state of different substances can be shown. For example — different metals like Na, Zn, Fe or non-metals like S are shown to explain the solid state.

Check Your Progress

Notes:

a) Write your answers in the space given below.

b) Compare your answers with those given at the end of the unit.

5. Give the significance of the following chemical equations.

i) \[ 2 \text{KClO}_3(s) \rightarrow 2 \text{KCl} (s) + 3\text{O}_2 (g) \]

ii) \[ \text{Zn(s)} + \text{H}_2\text{SO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{H}_2(g) \]

10.3.3 Balancing of Equations

Main Teaching Points

- Balancing of equations means making the number of atoms of different elements on both sides of a chemical equation equal.
- Balancing is essential to satisfy the law of conservation of mass.
- An unbalanced equation is known as skeletal equation.
- Chemical equations can be balanced by either of the following methods.
  a) Hit and Trial Method
  b) By Partial Equation Method
- Hit and Trial method is used to balance the simple equations.
- Partial equation method gives the mechanism of a chemical reaction.
Teaching Learning Process

On the basis of law of conservation of mass, it is essential that the number of atoms of different elements on both the sides of a chemical equation should be equal. If the number of atoms of different elements on both sides of a chemical equation is not equal, then the equation is known as skeletal. Thus, first of all skeletal equation should be written and then atoms of different elements on both sides should be made equal so as to obtain balanced equation.

The chemical equation is balanced by either of the following methods

a) Hit and Trial Method
b) Partial Equation Method

a) Hit and Trial Method: This method is used to balance simple equations. The following steps will be helpful in balancing the equations by this method.

i) Write the skeletal equation i.e., the correct formula for each reactant and product.
ii) Count the atoms of various elements on the two sides of the equation. If different, then they are equalised.
iii) Choose first that atom for balancing which occurs the minimum number of times.
iv) Atoms which occur at maximum number of times on both sides should be chosen last of all for balancing.
v) When any elementary gas is present as reactant or product, it is kept in atomic state. The balanced atomic equation is then made molecular by multiplying it with 2.

Example

1) Action of potassium on water - Potassium reacts with water to produce potassium hydroxide and hydrogen. The skeletal equation is

\[ K + H_2O \rightarrow KOH + H \]

The equation contains hydrogen, an elementary gas, as a product. It is taken in the atomic state in the beginning \( K + H_2O \rightarrow KOH + 2H \) atoms. The number of atoms of each element should be same on both sides of the equation because according to the law of conservation of mass, "Matter is neither created nor destroyed in a chemical change. Total mass of reactants should be same as total mass of products."

There is one ‘K’ atom on both sides. Therefore, there is no need to balance it. Oxygen and hydrogen atoms are also equal on both sides. Thus the equation is already balanced. Therefore, to convert it into molecular equation, it is multiplied throughout by two. It will become:

\[ 2K + 2H_2O \rightarrow 2KOH + H_2 \]

(2) Action of Steam on Hot Iron

\[ Fe + H_2O \rightarrow Fe_3O_4 + H \]

The skeletal equation in the atomic form is

\[ Fe + H_2O \rightarrow Fe_3O_4 + H \]

To balance ‘Fe’ multiply Fe by 3

\[ 3Fe + H_2O \rightarrow Fe_3O_4 + H \]

On the right hand side there are 4 oxygen atoms, therefore, \( H_2O \) should be multiplied by 4 to equalise oxygen atoms on the two sides.

\[ 3Fe + 4H_2O \rightarrow Fe_3O_4 + H \]

To balance ‘H’ atoms ‘H’ is multiplied by 8

\[ 3Fe + 4H_2O \rightarrow Fe_3O_4 + 8H \]
Multiplying the equation by 2 and writing hydrogen in the molecular form, makes the equation balanced.

\[ 6 \text{Fe} + 8\text{H}_2\text{O} \rightarrow 2 \text{Fe}_3\text{O}_4 + 8\text{H} \]
\[ 3 \text{Fe} + 4\text{H}_2\text{O} \rightarrow \text{Fe}_3\text{O}_4 + 4\text{H}_2 \]

When the equation is balanced atoms of each element are equal on both sides of an equation. The law of conservation of mass will be automatically obeyed.

b) Partial Equation Method: Simple equations can be easily balanced by Hit and Trial Method. But with complex reactions, this method does not succeed at all or if at all they can be balanced, they take a very long time. Such equations can be balanced by partial equation method.

The following steps are used
i) The complex reaction is supposed to proceed in various stages. For these stages simple equations are written. These equations are known as partial equations.
ii) Balance each partial equation by hit and trial method.
iii) Partial equations are then multiplied by suitable whole numbers to cancel the intermediate products which do not occur as a final product.
iv) Add partial equations to get the final balanced equation.

The advantage of this method is that it gives the mechanism of a chemical reaction.

1) Example: Action on chlorine on cold caustic soda solution.

When chlorine is passed through cold caustic soda solution, sodium chloride, sodium hypochlorite and water are produced.

The skeletal equation for the above reaction is:

\[ \text{Cl}_2 + \text{NaOH} \rightarrow \text{NaCl} + \text{NaClO} + \text{H}_2\text{O} \]

The reaction is supposed to take place in a number of steps as shown in the following partial equation.

\[ \text{Cl}_2 + \text{H}_2\text{O} \rightarrow \text{HCl} + \text{HClO} \]
\[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]
\[ \text{HClO} + \text{NaOH} \rightarrow \text{NaClO} + \text{H}_2\text{O} \]

\[ \frac{\text{Cl}_2 + \text{NaOH}}{4} \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{NaClO} \]

In this, each partial equation is already balanced and on adding gives the balanced equation.

2) Example: Reaction of ozone on lead sulphide to form lead sulphate.

i) The reaction may be represented by following two equations:

\[ \text{O}_3 \rightarrow \text{O}_2 + \text{O} \]
\[ \text{PbS} + \text{O} \rightarrow \text{PbSO}_4 \]

ii) The first partial equation is already balanced. To balance the second partial equation, multiply oxygen atom on the L.H.S. by 4.

\[ \text{O}_3 \rightarrow \text{O}_2 + \text{O} \]
\[ \text{PbS} + 4\text{O} \rightarrow \text{PbSO}_4 \]

iii) To remove oxygen which is an intermediate product multiply first partial equation by 4. Thus we get

\[ 4\text{O}_3 \rightarrow 4\text{O}_2 + 4\text{O} \]
\[ \text{PbS} + 4\text{O} \rightarrow \text{PbSO}_4 \]

iv) On adding up the partial equations, we get

\[ \text{PbS} + 4\text{O}_3 \rightarrow \text{PbSO}_4 + 4\text{O}_2 \]
Methodology used: Balancing of equations is made clear through discussion-cum-demonstration method. Special emphasis on hit and trial method and partial equation method is made and home assignments are given to clear the concept of balancing of equation.

Check Your Progress

Notes: a) Write your answers in the space given below.  
    b) Compare your answers with those given at the end of the unit.

6. Balance the following equations:
   a) $I_2 + HNO_3 \rightarrow HIO_3 + NO + H_2O$
   b) $KI + H_2SO_4 \rightarrow KHSO_4 + H_2O + I_2 + SO_2$
   c) $PbS + O_3 \rightarrow PbSO_4 + O_2$
   d) $K_2Cr_2O_7 + H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + H_2 + O_2$
   e) $S + H_2SO_4 \rightarrow SO_2 + H_2O$

7. Write balanced equations for:
   a) Potassium + Water $\rightarrow$ Potassium hydroxide + Hydrogen
   b) Aluminium + Oxygen $\rightarrow$ Aluminium oxide
   c) Iron + Steam $\rightarrow$ Ferric oxide + Hydrogen
   d) Cuprous chloride + Chlorine $\rightarrow$ Cupric chloride

10.4 MOLE CONCEPT

Main Teaching Points
- To understand the mole concept. Visualization of mole as a number i.e., mole is a quantity which depends upon the number of particles.
- Application of mole concept in stoichiometry of chemical reactions.

10.4.1 Mole as a Unit of Amount of Substance
1. Mole is an amount of any substance which contains as many elementary entities as there are atoms present in $0.012 \; \text{kg}$ of $\text{C}_{12}$.
2. One mole of any substance will contain $6.023 \times 10^{23}$ particles.
3. Mole is the unit to express the amount of any substance.

Teaching Learning Process
Weigh 5.6g of iron powder and 5.6g of sulphur powder. Mix them properly and heat them strongly in a test tube. Stop heating when no further reaction takes place and the mixture becomes a hard mass. Cool and test the hard mass so obtained. It will not be attracted by
the magnet and if we observe carefully, we will be able to see some unreacted yellow particles of sulphur.

In the reaction between iron and sulphur, we find that

\[ Fe + S \rightarrow FeS \]

One atom of iron reacts with one atom of sulphur to form one molecule of FeS. As sulphur is left unreacted, so it proves that the number of atoms in 5.6g of iron is less than the number of atoms in 5.6g of sulphur. From the experiment, children will draw the conclusion that in chemical reactions, number of atoms are important for any substance in a given mass and not the mass of substance.

### 10.4.2 Avogadro's Number

Atoms and molecules are extremely small. They can not be even seen with a powerful microscope. Thus they cannot be counted. It is not feasible to work with individual atoms or molecules, because even a small quantity of matter contains very large number of molecules. Therefore, chemicals work with the matter in bulk i.e., the assemblage of a vast number of molecules and not the individual atoms and molecules. So, some bigger unit to count the atoms and molecules is needed.

Commonly, to represent 12 articles, the unit ‘dozen’ is used and for 20 articles the unit ‘score’ is used. In the same way the unit ‘mole’ is used for counting atoms or molecules. Thus **mole represents \(6.023 \times 10^{23}\) particles** and this number is named **Avogadro’s number** which is denoted by \(N\).

The formula mass or molecular weight is given in a.m.u. An amount of a substance whose mass in grams numerically equals the molecular mass is called gram formula or gram-molecular mass. A gram-molecular mass is usually referred to as a gram-mole or simply a mole.

In the case of \(\text{C}_6\text{H}_{12}\text{O}_6\), where the molecular mass is 180 a.m.u., 1 gram molecular weight or 1 gram mole weighs 180 gram. A pile of this glucose weighing 180 grams is 1 gram-mole. Since we are generally concered with gram-mole, it is convenient to omit the prefix gram and to understand that mole means gram mole.

For any substance whose molecular mass is known, 1 mole contains the Avogadro’s number of molecules.

**For example** : 1 mole of atoms means \(6.023 \times 10^{23}\) atoms.

Thus weights of \(6.023 \times 10^{23}\) atoms of a substance is equal to its gram atomic mass and the weights of \(6.023 \times 10^{23}\) molecules of a substance is equal to its gram molecular weight. Thus in general the molecular weight of a substance is represented by the following relationship.

**For example** : \(\text{H}_2\text{O}, \text{CaCO}_3, \text{Na}, \text{Ca}\)

\[
\text{H}_2\text{O} = 18g = 1 \text{ mole} = 6.023 \times 10^{23} \text{ molecules} = 22.4 \text{ litres at N.T.P.}
\]

\[
\text{CaCO}_3 = 100g = 1 \text{ mole} = 6.023 \times 10^{23} \text{ molecules}
\]

\[
\text{Na} = 23g = 1 \text{ g atom} = 6.023 \times 10^{23} \text{ atoms}
\]

\[
\text{Ca} = 40g = 1 \text{ g atom} = 6.023 \times 10^{23} \text{ atoms}
\]

The weight of \(6.023 \times 10^{23}\) atoms of an element is known as the atomic weight of the element.

**Methodology used** : This topic was developed by taking some activities, like demonstrations of basic ideas, collection of various samples, weighing them and thus calculating the number of particles present in a definite mass of the substance. To illustrate the idea of mole as a number we can introduce the idea of dozen, score etc. The demonstrations should be based upon specific objectives.
Check Your Progress

Notes: a) Write your answers in the space given below.
    b) Compare your answers with those given at the end of the unit.

8. Define
   a) Avogadro’s number
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   b) Mole
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9. How many atoms of H and S are contained in 0.40 mole of H₂S?
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10. Arrange the following substances in order of their increasing weights in grams.
    a) one gram atom of silver
    b) one gram atom of nitrogen
    c) one mole atom of calcium
    d) 1 × 10^{23} atoms of carbon
    e) one gram atom of iron
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10.4.3 Stoichiometry

Main Teaching Points
- It deals with the mass relationship in a chemical reaction.
- This aspect of chemistry helps us to determine the number of moles of any known weight of a substance.

Teaching Learning Process
The aspect of chemistry which deals with mass relationship in a chemical reaction is called stoichiometry i.e., application of mole concept.

Example 1: How many moles are there in each of the following:
   i) 3.2g of CH₄
   ii) The amount of CO₂ that contains 0.6 gram-atom of oxygen.
Solution

i) \( \text{CH}_4 = 16g = 1 \text{ mole} = 6.023 \times 10^{23} \text{ atoms} \)

From the above reaction it is clear that 16 gm of \( \text{CH}_4 \) = 1 mole

\[
\therefore \text{3.2g of CH}_4 = 3.2 \times \frac{1}{16} = 0.2 \text{ mole}
\]

ii) \( O = 16 \text{ gram} = 1 \text{ gram atom} = 6.023 \times 10^{23} \text{ atoms} \)

It is clear from the above relationship that

1 gram-atom of oxygen = 16 gram

Example 2 : a) Calculate the volume occupied at S.T.P. by

i) 16 gram of oxygen

ii) 1.5 moles of oxygen

iii) \( 6.023 \times 10^{23} \) molecules of \( \text{CO}_2 \)

b) Calculate the weight of one molecule of \( \text{NH}_3 \)

Solution

a) i) \( \text{O}_2 = 32 \text{ gram} = 1 \text{ mole} = 6.023 \times 10^{23} \text{ molecules} = 22.4 \text{ litres at S.T.P.} \)

From the above relation we have

32 gram of oxygen occupy volume = 22.4 litres at S.T.P.

\[
\therefore 16 \text{ gram of oxygen occupy volume} = \frac{16}{32} \times 22.4 \text{ l. at S.T.P.} = 11.2 \text{ litres at S.T.P.}
\]

ii) Again from the above relationship we have

1 mole of oxygen occupy volume = 22.4 litres at S.T.P.

1.5 mole of oxygen occupy volume = 1.5 \times 22.4 = 33.6 litres at S.T.P.

iii) \( \text{CO}_2 = 44 \text{ gram} = 1 \text{ mole} = 6.023 \times 10^{23} \text{ molecules} = 22.4 \text{ litres at S.T.P.} \)

Thus \( 6.023 \times 10^{23} \) molecules occupy volume = 22.4 litres at S.T.P.

b) \( \text{NH}_3 = 17 g = 1 \text{ mole} = 6.023 \times 10^{23} \text{ molecules} \)

\[
\therefore 6.023 \times 10^{23} \text{ molecules of NH}_3 \text{ weigh} = 17 g
\]

1 molecule of \( \text{NH}_3 \) weigh = \[
\frac{17}{6.02 \times 10^{23}} \text{ g} = 2.822 \times 10^{-23} \text{ gram}
\]

Methodology used : The topic is taught by giving sufficient practice of the problems based on the application of mole concept. A number of written test were given to make the topic clear.

Check Your Progress

Notes : a) Write your answers in the space given below.

b) Compare your answers with those given at the end of the unit.

11. What volume of oxygen will be obtained from 12.26 gram of \( \text{HClO}_3 \) ?
12. Calculate the weight of one molecule of Carbon-dioxide gas?

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10.5 TYPES OF CHEMICAL REACTIONS

Main Teaching Points
Chemical reactions are of the following types
- Combination Reactions
- Decomposition Reactions
- Double Decomposition Reactions
- Reversible Reactions
- Redox Reactions

Teaching Learning Process
Conditions necessary for the reactions to take place
A chemical reaction involves the reaction of the reactants to produce the products. Total mass of the reactants before the reaction is equal to the total mass of products after the reaction. Thus mass is conserved. Therefore, the fundamental basis of chemical reactions is the law of conservation of mass. Atoms are conserved. The only difference may be that their arrangement before and after the reaction might differ. It is evident that in a chemical reaction, bond between atoms in the reacting substances are first broken and then new bonds are formed, after re-arrangement in the molecules of the product.

For example, when two volumes of hydrogen combine with one volume of oxygen two volumes of water are formed. During this change a large amount of heat is evolved. The product water possesses quite different properties from the reactants i.e. hydrogen and oxygen. When this reaction takes place, the bonds between the four atoms of two hydrogen molecules and bonds between the two atoms (one molecule) of oxygen must be broken and these six atoms formed then rearrange to form two water molecules having new bonds, atoms rearrange to form two water molecules, the total number of atoms taking part do not change. Therefore, atoms are neither gained or lost. They simply rearrange. Thus atoms are conserved in chemical reactions.

1. Combination Reactions
When two or more pure substances react to form a single substance, the reaction is said to be combination reaction.

For example: The reaction between ethylene and HBr.
Combination reactions are also called addition reactions.

Addition reaction may also be defined as follows
When a reagent adds to a molecule of another substance to yield a single molecule, the reaction is called addition reaction. Such reactions are most characteristic of molecules having multiple bonds e.g. illustrating the aforementioned example.

\[
\begin{align*}
\text{H} & \quad \text{H} \\
| & | \\
\text{H - C = C + HBr} & | \\
| & | \\
\text{H} & \quad \text{H} \\
\text{Ethylene} & \quad \text{Ethylene bromide}
\end{align*}
\]
2. Decomposition Reactions

Breaking down of bigger molecules into simple molecules, due to the bigger molecule being subjected to high temperature, is known as decomposition reaction.

i) \( \text{C}_2\text{H}_6 \xrightarrow{500^\circ\text{C}} \text{C}_2\text{H}_4 + \text{H}_2 \)  
(Ethane)  (Ethene)

ii) \( \text{CaCO}_3 \xrightarrow{\text{Heat}} \text{CaO} + \text{CO}_2 \)

3. Double Decomposition Reactions

When two chemical compounds react together to exchange their radicals with the formation of two new compounds, the reaction is known as double decomposition reactions.

\[ \text{AgNO}_3 + \text{NaCl} \rightarrow \text{NaNO}_3 + \text{AgCl} \]

The process of double decomposition can be used for the determination of equivalent weight of one of the compounds formed which is insoluble and is thrown out as precipitate.

4. Reversible Reactions

Reactions in which the products of a chemical reaction can interact under a different set of conditions to reform the original substances are called reversible reactions.

These are indicated by putting two half arrow heads ( \( \xrightarrow{\rightleftharpoons} \) ) between the reactants and the products but pointing in opposite directions.

For example

1. When steam is passed over iron, heated at 1000°C, it gets oxidised to magnetic oxide of iron and hydrogen is liberated.

\[ 3\text{Fe}(s) + 4\text{H}_2\text{O}(g) \xrightarrow{\text{Heat}} \text{Fe}_2\text{O}_3 + \text{FeO} + 4\text{H}_2 \]

On the other hand, we pass hydrogen over magnetic oxide of iron, it will be reduced to iron, and steam will be formed.

2. Formation of hydrogen iodide, from hydrogen and iodine vapour in the presence of heated plantinum wire at 800°C.

\[ \text{H}_2 + \text{I}_2 \xrightarrow{800^\circ\text{C}} 2\text{HI} \]

5. Redox Reactions

Redox reactions or oxidation-reduction reactions play a significant role in biological systems and many industrial processes.

Oxidation is the process in which loss of electrons takes place. If one atom, ion or molecule loses electrons there must be another which gains the electrons and is reduced. Thus oxidation of an element occurs at the expense of another which is reduced. The two processes always occur simultaneously or they go hand in hand. Such type of reactions involving oxidation and reduction simultaneously are called redox reactions.

For example

Iron reacts with sulphur to form iron sulphide.

\[ \text{Fe} + \text{S} \rightarrow \text{FeS} \text{ (i.e., Fe}^{2+} \text{ S}^{2-} \text{)} \]

In this reaction, Fe changes into \( \text{Fe}^{2+} \) by losing two electrons due to oxidation and S change into \( \text{S}^{2-} \) by gaining two electrons due to reduction as

\[ \text{Fe} - 2\text{e}^- \rightarrow \text{Fe}^{2+} \text{ (Oxidation)} \]
\[ \text{S} + 2\text{e}^- \rightarrow \text{S}^{2-} \text{ (Reduction)} \]

ii) The reaction between sodium (solid) and chlorine (gas) resulting in the formation of sodium chloride (solid)

\[ \text{i.e., 2Na (s) + Cl}_2(g) \rightarrow 2\text{NaCl (s)} \]

In this reaction Na changes into \( \text{Na}^+ \) by losing one electron due to oxidation and Cl changes into \( \text{Cl}^- \) by gaining one electron due to reduction as:
Teaching of Chemistry

2Na (s) → 2Na⁺ + 2e⁻ (Oxidation)
Cl₂ (g) + 2e⁻ → 2Cl⁻ ( Reduction)

The number of electrons lost by sodium is equal to the number of electrons gained by chlorine.

Methodology used: Types of chemical reactions are explained by lecture-cum-demonstration method. After explaining the various types of chemical reactions, various types of reactions are shown in the chemistry laboratory.

Check Your Progress

Notes: a) Write your answers in the space given below.
      b) Compare your answers with those given at the end of the unit.

13. What are the conditions necessary for reactions to take place?
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14. Define the law of conservation of mass.
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15. Name different types of chemical reactions.
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16. Explain the following types by giving examples:
   a) Redox Reaction
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   b) Double Decomposition Reactions
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   c) Reversible Reactions
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10.6 ENERGY CHANGES DURING THE CHEMICAL REACTIONS

Main Teaching Points

- Various Molecular arrangements have different potential energies.
- Chemical reactions involve molecular re-arrangements.
- During molecular rearrangements, the difference in the two potential energies [pre-reaction and post-reaction], appear as heat and which is why chemical reactions are accompanied by energy changes.
- Exothermic reactions are accompanied by evolution of heat.
- Endothermic reactions are accompanied by absorption of heat.
- The study of energetics of chemical changes is an important aspect of chemistry.

Teaching Learning Process

10.6.1 Why do Energy Changes Takes Place in Chemical Reactions?

Nearly all chemical reactions are accompanied by energy changes. We know that the chemical reactions in combustion of fuels like petrol, coal and natural gas are our major energy sources. Of even great significance is the fact that chemical reactions involved in the digestion of food provide energy for all life. The study of energetics of chemical change is an important aspect of chemistry. It not only finds a large number of applications in the various areas of chemistry but it also plays a useful role in the other branches of science, engineering and technology.

We come across many chemical reactions in which heat is released i.e., exothermic changes and many reactions in which heat energy is absorbed. During the chemical reactions bonds between the atoms in the reactant molecules are rearranged to form product molecules. Energy is needed to break chemical bonds; while energy is released when bonds are formed. In fact, energy change in a chemical reaction is the difference between the energy needed to break the old chemical bonds in the reactant molecules and the energy released when new bonds are formed in the products.

10.6.2 Exothermic Reactions

If a chemical reaction is accompanied by evolution of heat, it is known as an exothermic reaction, \((\text{exo} \rightarrow \text{out})\)

\((\text{thermic} \rightarrow \text{heat})\)

Examples

1. Combustion of Methane

\[
\text{CH}_4 (g) + 2\text{O}_2 (g) \rightarrow \text{CO}_2 (g) + 2\text{H}_2\text{O} (g) + 890.4 \text{ K.J.}
\]

The thermo-chemical equation for the combustion of methane indicates that when one mole or 16 grams of methane reacts with 2 moles or 44 grams of \(\text{O}_2\), two moles or 36 grams of water (in gaseous form) along with one mole of carbon dioxide and 890.4 K.J. of heat is produced.

In this case heat content of reactants is more than products, therefore \(\Delta H\) is negative for this reaction.

Thus heat of the reaction

\[
i.e., \Delta H = -890.4 \text{ KJ}
\]

2. Combustion of Carbon

\[
\text{C(s)} + \text{O}_2(g) \rightarrow \text{CO}_2(g) + 394 \text{ KJ mol}^{-1}
\]

According to this equation, when one gram atom or 12 grams of solid carbon reacts with one mole or 32 grams of oxygen, 1 mole or 44 grams of \(\text{CO}_2\) along with 394 KJ mol\(^{-1}\) of
Teaching of Chemistry

heat is produced. The heat thus produced is also called heat of reaction i.e. \( \Delta H = -394 \text{ KJ mol}^{-1} \)

Thus, the above reaction is represented as,

\[
C(s) + O_2(g) \rightarrow CO_2(g)
\]

\( \Delta H = -394 \text{ KJ mol}^{-1} \)

10.6.3 Endothermic Reactions

If a chemical reaction is accompanied by absorption of heat, it is an endothermic reaction. (endo = in).

Examples

1. Action of steam on carbon

\[
C(s) + H_2O(g) \rightarrow CO(g) + H_2(g) - 131.44 \text{ KJ}
\]

This equation shows that when one mole of steam i.e., 18 grams reacts with one gram atom of C (12 grams) producing 1 mole (28 grams) of carbon monoxide gas and 131.44 KJ. This is an endothermic reaction.

Heat of the reaction i.e.,

\( \Delta H = + 131.44 \text{ KJ mol}^{-1} \)

H in this case is taken positive because heat content of the products are more than the heat content of reactants.

2. Formation of Carbon disulphide

\[
C(s) + 2S(s) \rightarrow CS_2 (l) - 92.09 \text{ KJ}
\]

In the formation of one mole or 76 grams of carbon disulphide from one gram atom (12 grams) of carbon and 2 gram-atoms (64 grams) of sulphur 92.09 KJ of heat is absorbed. So, it is known as Endothermic Reaction and heat of the reaction i.e., \( \Delta H = + 92.09 \text{ KJ mole}^{-1} \)

Some more endothermic reactions

a) \( A(s) \rightarrow 2A(g) + \frac{1}{2} O_2 \)

\( H + 30.56 \text{ KJ Mol}^{-1} \)

b) \( \frac{1}{2} H_2 (g) + \frac{1}{2} I_2 (g) \rightarrow HI (g) \)

\( H = + 25.95 \text{ KJ Mol}^{-1} \)

Example

Take 100 ml. of water in a beaker and note down its temperature. Now dissolve 100g. of glucose in it and again record the temperature. Observe that the temperature decreases and solution becomes cool.

Melting of solids, vaporisation of liquids and synthesis of proteins in the living system are endothermic processes.

Methodology used: Energy changes in chemical reactions as well as exothermic and endothermic reactions was explained by lecture-cum-demonstration method. Subject matter was explained by lectures. Lecture method was accompanied by demonstrations. Demonstration of simple experiments involving evolution of heat as well as absorption of heat gives understanding of chemical reactions.
17. What do you understand by the terms exothermic reactions and endothermic reactions? Explain by giving suitable examples.

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10.7 CHEMICAL EQUILIBRIUM

Main Teaching Points

- Equilibrium is a state of reaction at which the rates of two opposing chemical reactions become equal.
- Systems tend to move towards an equilibrium state.
- It can be attained from either direction.
- At the point of equilibrium, the reaction does not stop but appears to have stopped.
- Factors like concentration of the reactants and products, temperature and pressure affect the equilibrium.
- The combined effect of all the above factors is studied by Le-Chatelier's Principle.

Teaching Learning Process

10.7.1 Why is Equilibrium Attained?

Equilibrium state is the state of a reaction at which no further visible changes occur in its macroscopic properties like concentration, colour, pressure etc. The equilibrium state can be achieved in a closed system where transfer of constituent particles does not take place between the reacting system and surroundings.

Take one mole of H₂ in 1 litre container which is colourless and 1 mole of I₂ vapour in another 1 litre container which is purple in colour. Mix them together in a 2 litre container and heat the mixture of H₂ and I₂ to a temperature of 770 K. The children will observe that initially the colour is very dark and during heating the colour becomes light and after sometime there will be no change in colour.

\[
\text{H}_2 + \text{I}_2 \rightarrow 2\text{HI}
\]

(colourless) (violet) (colourless)

This shows that the reaction has stopped.
Now take 21 of HI (colourless) in a 21 container and heat it to a temp. of 770 K. The children will observe that the colourless HI changes to purple colour. The colour becomes dark slowly and ultimately there is no change in colour.

\[ 2 \text{HI} \rightarrow \text{H}_2 + \text{I}_2 \]

(colourless) (colourless) (violet)

The violet colour is due to the decomposition of HI forming I₂ which has violet colour.

Compare the colour of substances in two containers, we find that the colour in two containers is same. This concludes that the reversible reaction takes place in the container and when colour becomes constant, it seems that the reaction has stopped but actually an equilibrium state is attained at which there is no change in colour and also, is there is no change in the concentration of reactants and products.

**Main Characteristics of Chemical Equilibrium**

1. It is dynamic. It is a permanent situation maintained by the equality of the rates of two opposing chemical reactions.
2. The amount of reactants and products remain constant so long as the factors like concentration, temperature, pressure etc. remain unchanged.
3. It can be attained from either direction. The nature and properties of equilibrium state are the same, regardless of the direction from which it is reached.
4. Systems move towards an equilibrium state spontaneously; a system can be removed from equilibrium only by some outside influence and once it is left to itself, the disturbed system returns to an equilibrium state. Systems move towards equilibrium because of an imbalance of reaction rates; at equilibrium these rates are equal and there is no way for the undisturbed system to move away from the equilibrium.
5. A catalyst does not alter the state of equilibrium but it can bring about the equilibrium within less time.
6. It can take place in a closed system only.

**Factors Affecting Equilibrium**

An equilibrium is governed by the following factors:

i) Concentration of the reactants and the products.
ii) Temperature
iii) Pressure

The equilibrium is disturbed when the conditions under which it is occurring are altered.

The combined effect of all the above factors is studied by one principle known as Le-Chatelier's Principle.

According to this principle, if a system at equilibrium is subjected to a disturbance or stress that changes any of the factors that determine the state of equilibrium, the system will react in such a way as to minimise the effect of the disturbance.

The principle may also be stated as

If a system at equilibrium is subjected to a change of concentration, pressure or temperature, the equilibrium shifts in such a direction so as to undo the effect of the change imposed. The principal is very helpful in dealing with chemical equilibria, since it allows us to predict the qualitative response of a system to changes in external conditions. The amounts of reactants reacted and products formed i.e., quantitative mathematical analysis of equilibria is guided by qualitative predictions (temp., pressure etc.)
10.7.2 Applications of the Le-Chatelier's Principle to Some Reactions of Industrial Importance

Main Teaching Points

- Le-Chatelier’s Principle can be used to study important chemical equilibrium reactions of industrial importance.
- Formation of SO₂ and nitric oxide can be influenced by raising temperature or by increasing the concentration of SO₂ or O₂ or both or by increasing the pressure on the system. Formation of SO₃ is favoured by high pressure, low temperature and high concentration of the reactants.
- Formation of nitric oxide is favoured by high temperature, high concentration of reactants. Pressure has no effect on this equilibrium.

Teaching Learning Process

1. Formation of sulphur trioxide (Contact Process for H₂SO₄).

\[ 2\text{SO}_2 (g) + \text{O}_2 (g) \rightleftharpoons 2\text{SO}_3 (g) \quad \Delta H = -45.2 \text{KJ} \]

The formation of SO₃ from SO₂ and O₂ is an exothermic reaction. It is accompanied by a decrease in volume. The reverse reaction must be endothermic.

a) Effect of Temperature

If temperature is raised (by applying heat to the system) then the system will be under stress. Now according to the Le-Chatelier’s Principle, the equilibrium will shift in that direction which tends to minimise the effect of increased heat. That is to say, shifting will be in the direction of the reaction that proceeds with absorption of heat. But a reaction that lowers temperature is endothermic, it is the reverse of the above exothermic reaction i.e.,

\[ 2\text{SO}_3 \rightarrow 2\text{SO}_2 + \text{O}_2 \quad \text{heat energy} \]

In other words the effect of raising the temperature on the above equilibrium is to diminish the yield of SO₃ and to increase the concentration of the reactants (SO₂ + O₂).

b) Effect of Concentration

Increasing the concentration of SO₂ or O₂ or both on the system, imposes a stress. According to the Le-Chatelier’s Principle, to minimise or to counteract this stress, more SO₃ will be formed. Thus the equilibrium is shifted towards the products.

c) Effect of Pressure

The formation of SO₃ is accompanied by a decrease in volume. If the pressure on the system is increased, a stress is imposed. According to the Le-Chatelier’s Principle the equilibrium will shift to that direction which tends to minimise the effect of increased pressure. This stress due to increased pressure can be relieved if volume is decreased. In this equilibrium volume is decreased in the forward direction. Hence equilibrium will shift towards the right side which is occupying the lower volume. Thus the formation of SO₃ is favoured by high pressure.

2. Formation of nitric oxide (Birkeland - Eyde Process)

The formation of nitric oxide from nitrogen and oxygen is an endothermic reaction.

\[ \text{N}_2 + \text{O}_2 \rightarrow 2\text{NO}; \quad \Delta H = +180.83 \text{KJ} \]

a) Effect of Temperature

Forward reaction is accompanied by an absorption of heat. If we increase the temperature of the equilibrium mixture, the system is under stress. According to the Le-Chatelier’s Principle, the stress due to the added heat can be relieved by shifting the equilibrium towards the products. The yield of nitric oxide will therefore, be higher with increasing temperature and decrease of temperature favours dissociation of nitric oxide.
b) Effect of Concentration

If we add nitrogen or oxygen or both to the equilibrium mixture, the system is disturbed and is under stress. According to Le-Chatelier's principle, the stress imposed can be relieved if \( \text{N}_2 \) and \( \text{O}_2 \) combine to yield more nitric oxide. Thus increasing the concentration of oxygen, nitrogen or both will favour the formation of nitric oxide.

c) Effect of Pressure

In the formation of nitric oxide, no change in volume takes place because one volume of \( \text{N}_2 \) reacts with one volume of \( \text{O}_2 \) to yield 2 volumes of nitric oxide. If the pressure is increased, the stress due to increased pressure can be relieved if the equilibrium were shifted to the side occupying lower volume. But the volume is not changed (lowered in either side). Hence pressure will have no effect on this equilibrium.

Methodology used: The idea of equilibrium was developed through demonstrations followed by discussions. The effect of different factors like temperature, pressure, concentration, was explained through demonstrations.

Check Your Progress

Notes: a) Write your answers in the space given below.
      b) Compare your answers with those given at the end of the unit.

18. The formation of ammonia according to the equation:
   \[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 + 93.77 \text{ KJ} \]
   will be favoured by
   a) Low temperature
   b) Low Pressure
   c) Low temperature and high pressure
   d) Low temperature and low pressure

19. For the reaction \( \text{N}_2 + \text{O}_2 \rightarrow 2\text{NO} \), which of the following conditions favour forward reaction?
   a) High temperature
   b) High pressure
   c) Low temperature
   d) Low concentration of \( \text{N}_2 \)

10.8 LET US SUM UP

Chemical reactions are represented by chemical equations, which consist of chemical formulae of reactants and products, their physical states, concentration and other conditions. Balancing of equations means making the number of atoms of different elements on both sides of the reaction equal. It is important to get complete information of relative masses or volumes of substances involved in the reaction. An unbalanced equation is known as skeletal equation.

Mole is a unit of a certain amount of substance containing \( 6.023 \times 10^{23} \) number of particles. To clarify it further, we can say that

1 mole of atoms of a substance = atomic mass of the substance.

Chemical reactions are of various types such as combination reactions, decomposition reactions, double decomposition reactions, redox reactions and reversible reactions. Almost all reactions lead to energy changes. Hence, reactions are also classified as exothermic reactions which lead to evolution of heat and endothermic reactions which lead to absorption of heat. Reactions reach equilibrium when the rate of forward reaction (i.e.
formation of products) becomes equal to backward reaction (i.e. formation of reactants). All reaction systems tend to move towards an equilibrium state. But, only, those reactions achieve an equilibrium which have a closed system i.e. transfer of constituent particles does not take place between the reaching system and surroundings.

Hence, in this unit, chemical equation, types of chemical reactions and chemical equilibrium have been discussed in detail. You would also appreciate that these topics have been elaborated along with relevant pedagogical strategies.

### 10.9 UNIT-END EXERCISES

1. Describe the concept of formula. How does a chemical reaction represent changes taking place in surroundings?
2. What is mole concept? How is it used to calculate the amount of substances used in different reactions?
3. What is the role of energy in a chemical reaction?
4. Explain oxidation and reduction.
5. How is an equation balanced? Explain with suitable examples?
7. Describe the law of conservation of mass. What is a reversible reaction?
8. Explain the following by giving examples:
   1. Combination reactions
   2. Decomposition reactions
9. Why is the study of energetics of chemical changes an important aspect of Chemistry?
10. What are the factors which affect the equilibrium of reactions? Elaborate.

### 10.10 ANSWERS TO CHECK YOUR PROGRESS

1. **FORMULA** : A formula is a brief representation of a chemical compound in terms of symbols of various elements present in it.

   **Significance of a Formula** : The formula of a substance has qualitative and quantitative significance. For example, formula \( \text{H}_2\text{SO}_4 \) signifies:
   - One molecule of sulphuric acid.
   - A molecule of sulphuric acid consists of two atoms of hydrogen one atom of sulphur and four atoms of oxygen.
   - It signifies that two parts by weight of hydrogen, 32 parts by weight of sulphur and 64 parts by weight of oxygen combine to give 98 parts by weight of sulphuric acid.
   - It signifies that molecular mass of sulphuric acid is 98.

2. a) \((\text{NH}_4)_2\text{Cr}_2\text{O}_7\)
b) \(\text{Ag}_2\text{S}\)  
c) \(\text{Fe}_3(\text{PO})_4\) (Ferrous phosphate) \(\text{FePO}_4\) (Ferric phosphate)

3. | Element | % age of element | Atomic mass | \(\frac{\% \text{age of element}}{\text{atomic mass}} = \frac{10.6}{12} = 0.83\) | Relative no. of atoms | Simple atomic ratio |
<table>
<thead>
<tr>
<th></th>
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<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>10.06</td>
<td>12</td>
<td>0.83</td>
<td></td>
<td>(\frac{0.83}{1} = 1)</td>
</tr>
</tbody>
</table>
Teaching of Chemistry

Hence its empirical formula = CHCl₃

Empirical formula mass = 12 + 1 + (35.53) 
= 119.5

V.D.(given) = 60

Molecular mass = 2 v.d.
= 2 x 60
= 120

\[ N = \frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{120}{119.5} = 1 \]

Molecular formula of the organic compound = \( n \times \text{e.f.} \)
= 1 \times \text{CHCl₃}
= \text{CHCl₃} Ans.

4. Determination of the empirical formula:

<table>
<thead>
<tr>
<th>Element</th>
<th>%age of element</th>
<th>Atomic mass</th>
<th>( \frac{% \text{ of atom of element}}{\text{mass of atom of element}} )</th>
<th>%age of no. of atoms</th>
<th>Relative atomic ratio</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>12.76</td>
<td>12</td>
<td>( \frac{12.76}{12} ) = 1.06</td>
<td>1.06</td>
<td>1.06 = 1</td>
</tr>
<tr>
<td>H</td>
<td>2.13</td>
<td>1</td>
<td>( \frac{2.13}{1} ) = 2.13</td>
<td>2.13</td>
<td>2.13 = 2</td>
</tr>
<tr>
<td>Br</td>
<td>85.11</td>
<td>80</td>
<td>( \frac{85.11}{80} ) = 1.06</td>
<td>1.06</td>
<td>1.06 = 1</td>
</tr>
</tbody>
</table>

The empirical formula is CH₃Br.

5. i) \( 2\text{KClO}_3 (s) \xrightarrow{\text{MnO}_2} 2\text{KCl}(s) + 3\text{O}_2(g) \). The equation gives us the following information:

a) It shows the decomposition of solid potassium chlorate by the action of heat into solid potassium chloride and oxygen gas.

b) It shows that 2 molecules of potassium chlorate decompose into 2 molecules of potassium chloride and 3 molecules of O₂ gas.

c) The decomposition takes place in the presence of manganese dioxide which acts as a cata. st.

d) \( 245\text{g KClO}_3 \) decompose to gives 149g of potassium chloride and 96g of oxygen gas.

ii) Significance of Zn (s) + H₂SO₄ (aq) \( \rightarrow \) ZnSO₄ (aq) + H₂ (g)

a) It shows that zinc metal reacts with sulphuric acid to produce zinc sulphate and hydrogen.

b) It tells the physical states of reactants as well as products.

Reactants – Zinc – solid
H₂SO₄ – aqueous

Products ZnSO₄ – aqueous
H₂ – gas

c) 69g of Zinc react with 98g of H₂SO₄ to produce 165g of ZnSO₄ and 2g of hydrogen.

6 a) \( \text{I}_2\text{IOHNO}_3 \rightarrow 1\text{ONOO}^- + 4\text{H}_2\text{O} + 2\text{HIO}_3 \)

b) \( 2\text{KI} + 3\text{H}_2\text{SO}_4 \rightarrow 2\text{KHSO}_4 + 2\text{H}_2\text{O} + \text{I}_2 + \text{SO}_2 \)

c) \( \text{PbS} + \text{O}_3 \rightarrow \text{PbSO}_4 + 4\text{O}_2 \)

d) \( 2\text{K}_2\text{Cr}_2\text{O}_7 + 8\text{H}_2\text{SO}_4 \rightarrow 2\text{K}_2\text{SO}_4 + 2\text{Cr}_2(\text{SO}_4)_3 + 8\text{H}_2\text{O} + 3\text{O}_2 \)
7. a) Potassium + Water $\rightarrow$ Potassium hydroxide + Hydrogen
   Equation: $2K + 2H_2O \rightarrow 2KOH + H_2$

b) Aluminium + Oxygen $\rightarrow$ Aluminium oxide
   Equation: $4Al + 3O_2 \rightarrow 2Al_2O_3$

c) Iron + Steam $\rightarrow$ Ferric oxide + Hydrogen
   Equation: $2Fe + 3H_2O \rightarrow Fe_2O_3 + 3H_2$

d) Cuprous chloride + Chloride $\rightarrow$ Cupric chloride
   Equation: $2CuCl + Cl_2 \rightarrow 2CuCl_2$

8. a) Avogadro's Number: The number of molecules of any gas present in one gram mole of the gas is called Avogadro's number and is symbolized as $N$.
   One gram mole of any gas at S.T.P. occupies volume equal to 22.4 litres.
   22.4 litres of all the gases at S.T.P. = $6.023 \times 10^{23}$ molecules.

b) Mole: Mole is the amount of any substance which contains as many elementary entities as there are atoms present in 0.012 kg of Cl$_2$.

9. $H_2S = 34g = 1$ mole $6.023 \times 10^{23}$ molecules
   1 mole of $H_2S$ contains $= 2$ gram atoms of hydrogen
   0.4 mol of $H_2S$ contains $= 2 \times 0.4 = 0.8$ gram atoms of hydrogen
   $H = 1$ gm = 1 gram atom $= 6.023 \times 10^{23}$ atoms
   1 gram atom of Hydrogen $= 6.023 \times 10^{23}$ atoms of hydrogen
   0.8 gram atom of hydrogen $= 6.023 \times 10^{23} \times 0.8 = 4.8184 \times 10^{23}$ atoms of hydrogen.

Similarly
   1 mole of $H_2S$ contains $= 1$ gm atom of $S = 6.023 \times 10^{23}$ atoms.
   then 1 gram atom of $S = 6.023 \times 10^{23}$ atoms of $S$
   0.4 gram atom of $S = 0.4 \times 6.023 \times 10^{23}$ atoms of $S$
   $= 2.4092 \times 10^{23}$ atoms of $S$

10. a) one gram atom of Ag = 108 gram.
    b) one gram atom of Nitrogen = 14g
    c) one mole atom of Ca = 40g
    d) $1 \times 10^{23}$ atoms of C = 1.992g.
    e) one gram atom of Iron = 56g.
    Arranging them in the increasing order of weight
   1) weight of $1 \times 10^{23}$ atoms of C = 1.992g
   2) weight of one gram atom of N = 14g
   3) weight of one mole atom of Ca = 40g
   4) weight of one gram atom of Iron = 56g
   5) weight of one gram atom of Ag =108g

11. $2KClO_3 \rightarrow 2KCl + 3O_2$
    $(245g) \quad (149g) \quad (96g)$
    $245g$ of $KClO_3$ yield $= 96g$ of Oxygen
    $1g$ of $KClO_3$ yield $= \frac{96}{245} \ g$ of Oxygen
    $12.26 \ g$ of $KClO_3$ yield $= \frac{96 \times 12.26}{245} \ = 4.8 \ g$
32g of Oxygen occupy volume = 22.4 litres at S.T.P.

4.8g of Oxygen occupy volume = \( \frac{22.4 \times 4.8}{32} \) = 3.36 litres

Hence 12.26 g of KClO₃ Proves 3.36 litres of Oxygen

12. \( \text{CO}_2 = 44g = 1 \text{ mole } \times 6.023 \times 10^{23} \) molecules

\( \text{6.023} \times 10^{23} \) molecules of \( \text{CO}_2 \) Weigh = 44g

1 molecule of \( \text{CO}_2 \) Weigh = \( \frac{44}{6.023 \times 10^{23}} \) = \( 7.3 \times 10^{-23} \) g

13. The total mass of the reactants before the reaction is equal to the total mass of products after the reaction. Atoms are conserved. The only difference may be that their arrangement before and after the reaction changes. Bond between atoms in the reacting substances are first broken and then new fundamental basis of chemical reactions is the law of Conservation of Mass.

14. Law of Conservation of Mass states that total mass of the reactants before the reaction is equal to the total mass of products after the reaction.

15. Types
   1) Combination Reaction
   2) Decomposition Reaction
   3) Double Decomposition Reaction
   4) Reversible Reaction
   5) Redox Reaction

16. a) **Redox Reaction**: It is type of reaction in which when one part of compound undergoes oxidation the other part undergoes reduction.

\[
\text{Oxidation - 2e}^- \quad \downarrow
\]

For example: \( \text{Cu}^{2+} \text{SO}_4^{2-} + \text{Fe} \rightarrow \text{Cu} + \text{FeSO}_4 \)

\[\text{Reduction + 2 e}^- \quad \uparrow\]

Here copper is reduced from \( \text{Cu}^{++} \) to \( \text{Cu} \) atoms and \( \text{Fe} \) is Oxidised.

b) **Double decomposition Reaction**: When two chemical compounds react together to exchange their radicals with the formation of two new compounds, the reaction is known as double decomposition reaction.

c) **Reversible reaction**: Reactions in which the products of a chemical reaction, can interact under a different set of conditions to reform the original substances are called reversible reactions.

\[\text{for e.g. 3 Fe(s) + 4H}_2\text{O(g) } \rightarrow \text{Fe}_3\text{O}_4 + 4\text{H}_2(g)\]

17. **Exothermic Reactions**: The reaction which is accompanied by the evolution of heat, is known as an exothermic reaction.

For example

1) Combustion of methane is an exothermic reaction

\[
\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + 890.4 \text{ KJ}
\]

2) Combustion of Carbon

\[
\text{C(s) + O}_2 \rightarrow \text{CO}_2 + 394 \text{ KJ}
\]
Endothermic Reactions: If a chemical reaction is accompanied by absorption of heat it is known as an endothermic reaction.

For example
1) Dissolution of glucose in water
2) Dissolution of NH₄Cl in water
3) Reaction between Ba(OH)₂ and NH₄Cl

\[
\text{Ba(OH)}_2 + \text{NH}_4\text{Cl} \rightarrow \text{BaCl}_2 + 2\text{NH}_4\text{OH}
\]

18. c) Low temperature and high pressure.
19. a) High Temperature favours forward reaction.