CHAPTER # 1

STOICHIOMETRY

Q1. Define stoichiometry.
Ans: Stoichiometry:
The study of relative amounts of substances involved in a chemical reaction is called Stoichiometry. Such phenomenon is studied through the knowledge of Stoichiometry (Greek word Stolcheion means element and metry means measurement).

Importance of Stoichiometry:
Stoichiometry is essential when quantitative information about a chemical reaction is required. Moreover it is important to predict yields of chemical products.

Q2. Explain the significance of stoichiometry with the help of example.
OR
How will you explain law of conservation of mass in the case of combustion of hydrogen fuel in rockets?
Ans: Combustion of hydrogen fuel in rockets:
1. Consider a rocket manufacturer uses liquid hydrogen as a fuel. He may have to determine how much fuel is necessary for a particular flight. Hydrogen burns in oxygen (of air) to produce water.

\[ 2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \]

It states that
(i) Two moles of Hydrogen react with one mole of oxygen to form two moles of steam.
(ii) Two molecules of Hydrogen react with one molecule of oxygen to produce two molecules of steam.
(iii) Four grams of hydrogen react with thirty-two grams of oxygen to produce thirty-six grams of water. Here the total mass of reactants is equal to the total mass of products. Thus it confirms the Law of conservation of mass.

2. Another example is the reaction taking place in a gas barbecue. This is the example of combustion to form carbon dioxide and water. The balanced chemical reaction is

\[ \text{C}_3\text{H}_8 + 5 \text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O} \]

Q3. What do you understand by the term Mole?
Ans: Mole:
The atomic mass, formula mass and molecular mass of a substance expressed in grams is called Mole.

\[ \text{Number of moles} = \frac{\text{mass in grams}}{\text{molecular mass}} \]

Example: One mole of O = 16 g
One mole of $O_2$ = 32 g
One mole of $H_2O$ = 18 g

**Explanation of one mole of NaCl:**
The explanation of one mole of NaCl i.e. 58.5 g is quite different as it is ionic in nature and will be called as formula mass which produce ions on dissolving in water. Therefore,

\[ \text{one mole of } NaCl(s) \xrightarrow{H_2O} 1 \text{ mole of } Na^{+}_{(aq)} \text{ ions} + 1 \text{ mole of } Cl^{-}_{(aq)} \text{ ions}. \]

**Q4. How will you explain representative particles by using Avogadro’s number?**

**OR**

**How is mole related to Avogadro’s number?**

**Ans: Representative Particles (Avogadro’s Number):**
“The number of atoms, ions or molecules present in one mole of a substance is called Avogadro’s Number”. Its numerical value is $6.023 \times 10^{23}$. One mole of any gas at S.T.P occupies $22.414 \, dm^3$ and contains $6.02 \times 10^{23}$ particles.

e.g. \[ 2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \]

This reaction can also be expressed in terms of Avogadro’s number. The equation states that $2 \times 6.02 \times 10^{23}$ molecules of hydrogen react with $6.02 \times 10^{23}$ molecules of oxygen to produce $2 \times 6.02 \times 10^{23}$ molecules of water.

**Number of moles of a substance** = \[ \frac{\text{Number of molecules of substance}}{N_A} \]

**Number of moles of a substance** = \[ \frac{\text{Number of molecules of substance}}{6.02 \times 10^{23}} \]

**Relation between moles and Avogadro’s number:**
The relationship between moles and Avogadro’s number in an atom and covalent molecules is as follows:
e.g. 1 mole of $O$ = $6.02 \times 10^{23}$ atoms
1 mole of $O_2(gas)$ = $6.02 \times 10^{23}$ molecules
   = $22.414 \, dm^3$ (volume occupied by 1 mole of gas at S.T.P)
1 mole of $H_2O_{(l)}$ = $6.02 \times 10^{23}$ molecules

**In the case of ionic compounds, the explanation is somewhat different.**

**For example,**

(i) \[ \text{NaCl} \xrightarrow{H_2O} Na^{+}_{(aq)} + Cl^{-}_{(aq)} \]

It shows that 1 mole of NaCl when dissolves in water gives 1 mole of Na$^{+}$ ions and 1 mole of Cl$^{-}$ ions. So According to Avogadro’s number we can say that when $6.02 \times 10^{23}$ formula units of NaCl are dissolved in water there are produced $6.02 \times 10^{23}$ Na$^{+}$ and $6.02 \times 10^{23} \text{Cl}^{-}$ ions.

(ii) \[ C + O_2 \longrightarrow CO_2 \]

This equation shows that 1 mole of C and according to Avogadro’s $6.02 \times 10^{23}$ atoms reacts with 1 mole of $O_2$ i.e $6.02 \times 10^{23}$ molecules of $O_2$ to produce 2 moles or $6.02 \times 10^{23}$ molecules of $CO_2$. 
Q5. Describe construction of mole ratios as conversion factors in stoichiometric calculations.

Ans: Construction of Mole ratios as Conversion Factors in Stoichiometric Calculations:

Mole Ratios:
Mole ratios mean the ratios of number of moles of reactants taking part and the number of moles of products formed.

Example Combustion of propane:
For example, combustion of propane

\[
\text{C}_3\text{H}_8 \quad + \quad 5\text{O}_2 \quad \rightarrow \quad 3\text{CO}_2 \quad + \quad 4\text{H}_2\text{O}
\]

The mole ratios between the reactants and products can be shown as, one mole of \(\text{C}_3\text{H}_8\) reacts with five moles of oxygen to give three moles of \(\text{CO}_2\) and four moles of water. The amount of propane used will not affect these ratios.

Sample Problem No. 1.1

Methanol burns according to the following equation.

\[
2\text{CH}_3\text{OH} \quad + \quad 3\text{O}_2 \quad \rightarrow \quad 2\text{CO}_2 \quad + \quad 4\text{H}_2\text{O}
\]

If 3.50 moles of methanol are burned in oxygen, calculate

(a) How many moles of oxygen are used
(b) How many moles of water are produced

Solution: Mole ratios = conversion factor

The problem can be solved by using correct conversion factors which are obtained from the balanced chemical reaction.

(a) Moles of methanol (Given quantity) = 3.50 moles
Moles of oxygen (Desired) =?

Conversion Factor = Mole ratios \(= \frac{3\text{moles O}_2}{2\text{moles CH}_3\text{OH}}\)

Desired quantity (of \(\text{O}_2\)) = Given quantity \(\times\) conversion factor

\[
= 3.50\text{ moles of CH}_3\text{OH} \times \frac{3\text{ moles O}_2}{2\text{ moles CH}_3\text{OH}}
\]

\[
= \frac{3.50 \times 3}{2} \quad \text{moles of O}_2
\]

Desired quantity = 5.25 moles of \(\text{O}_2\)

So the number of moles of \(\text{O}_2\) consumed (Desired) = 5.25 moles

(b) Given quantity of \(\text{CH}_3\text{OH}\) = 3.50 moles
Desired quantity of \(\text{H}_2\text{O}\) =?

Conversion factor (or mole ratio) \(= \frac{4\text{ moles H}_2\text{O}}{2\text{ moles CH}_3\text{OH}}\)

Desired quantity (i.e. No of moles of \(\text{H}_2\text{O}\) formed)
\[ = \text{Given quantity of } \text{CH}_3\text{OH} \times \frac{4 \text{ moles H}_2\text{O}}{2 \text{ moles CH}_3\text{OH}} \]
\[ = 3.50 \text{ moles of CH}_3\text{OH} \times \frac{4 \text{ moles H}_2\text{O}}{2 \text{ moles CH}_3\text{OH}} \]

Required \[ = \frac{3.50 \times 4}{2} \text{ moles of H}_2\text{O} \]
Quantity of H\(_2\text{O}\) \[ = 7.00 \text{ moles of H}_2\text{O} \]

Self Check Exercise 1.1

NH\(_3\) is an important raw material in the manufacture of fertilizers. It is obtained by the combination of N\(_2\) and H\(_2\) as shown by the following balanced equation.

\[ \text{N}_2 \quad + \quad 3\text{H}_2 \quad \rightarrow \quad 2\text{NH}_3 \]

How many moles of the following are required to manufacture 5.0 moles of NH\(_3\). (a) Nitrogen (b) Hydrogen

(Ans: (a) N\(_2\) = 2.5 Moles (b) H\(_2\) = 7.5 Moles)

Solution: Stoichiometric Calculation:
(a) From given balanced equation it is clear that:
2 moles of NH\(_3\) = 1 mole of N\(_2\)
1 moles of NH\(_3\) = \(\frac{1}{2}\) moles of N\(_2\)
5 moles of NH\(_3\) = \(5 \times \frac{1}{2}\) moles of N\(_2\) = 2.5 moles of N\(_2\)

(b) From given balanced equation it is clear that:
2 moles of NH\(_3\) = 3 mole of H\(_2\)
1 moles of NH\(_3\) = \(\frac{3}{2}\) moles of H\(_2\)
5 moles of NH\(_3\) = \(5 \times \frac{3}{2}\) moles of H\(_2\) = 7.5 moles of H\(_2\)

Q6. Define molar volume.

Ans: Molar volume:
Molar quantities of gases can be expressed in terms of volumes. It has been experimentally proved that one mole of any gas at STP occupies a volume of 22.4 dm\(^3\). This volume is called molar volume.

Sample Problem No. 1.2

Iron can be produced from iron ore Fe\(_2\)O\(_3\) by reacting the ore with carbon monoxide (CO). Carbon dioxide (CO\(_2\)) is produced in this reaction as a byproduct. What mass of iron can be formed from 425 g of iron ore?

Solution:
The balanced equation can be written as
\[ \text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2 \]

Mass of iron ore = \(\frac{425 \text{ g}}{159.6 \text{ g/moles}^{-1}}\)

\[ = 2.66 \text{ moles of } \text{Fe}_2\text{O}_3 \]

**Number of moles of iron produced:**

\[ \text{i.e Desired quantity} = \text{Given quantity} \times \text{conversion factor} \]

\[ \text{No of moles of } \text{Fe} = \text{No of moles of } \text{Fe}_2\text{O}_3 \times \frac{\text{No of moles of } \text{Fe}}{\text{No of moles of } \text{Fe}_2\text{O}_3} \]

\[ = 2.66 \text{ Moles } \text{Fe}_2\text{O}_3 \times \frac{2 \text{ Moles } \text{Fe}}{1 \text{ Mole } \text{Fe}_2\text{O}_3} \]

\[ = 2.66 \times 2 \text{ Moles of } \text{Fe} = 5.32 \text{ Moles of } \text{Fe} \]

**How to convert number of moles of iron to mass of Fe in grams:**

\[ \text{Desired quantity} = \text{Given quantity} \times \text{conversion factor} \]

\[ \text{(Mass of } \text{Fe produced}) = 5.32 \text{ Mole } \text{Fe} \times \frac{55.9 \text{ g}}{1 \text{ Mole } \text{Fe}} \]

\[ = 5.32 \times 55.9 \text{ g} \]

Mass of iron produced = 297.388 g

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**Self Check Exercise 1.2**

The main engines of the U.S. space shuttle are powered by liquid hydrogen and liquid oxygen. If \(1.02 \times 10^5\) kg of liquid hydrogen is carried on a particular launch, what mass of liquid oxygen is necessary for all the hydrogen to burn. The equation for the reaction is,

\[ 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

*(Ans: \(8.16 \times 10^5\) kg oxygen)*

**Solution:**

\[ \text{Mass of liquid hydrogen} = 1.02 \times 10^5 \text{ kg} = 1.02 \times 10^5 \times 10^3 \text{ g} \]

\[ = 1.02 \times 10^8 \text{ g} \]

**Number of moles of hydrogen** = \(\frac{\text{mass in gram}}{\text{molar mass}}\)

**Number of moles of hydrogen** = \(\frac{1.02 \times 10^8}{2} = 0.51 \times 10^8 \text{ moles} \)

\[ = 5.1 \times 10^7 \text{ moles} \]
From given balanced equation it is clear that:
2 moles of $H_2 = 1$ mole of $O_2$
1 mole of $H_2 = \frac{1}{2}$ mole of $O_2$

$5.1 \times 10^7$ moles of $H_2 = \frac{1}{2} \times 5.1 \times 10^7$ moles of $O_2$

$= 2.55 \times 10^7$ moles of $O_2$

Now, $Mass\ of\ Oxygen = Number\ of\ moles \times Molar\ mass$

$= 2.55 \times 10^7 \times 32\ g = 8.16 \times 10^8\ g$

$Mass\ of\ liquid\ oxygen\ in\ Kg = \frac{8.16 \times 10^8}{10^3} = 8.16 \times 10^5\ Kg$

**Sample Problem No. 1.3**

Calculate the number of molecule of $O_2$ produced by thermal decomposition of 490 grams of $KClO_3$.

**Solution:**
The given mass of $KClO_3 = 490\ g$

Formula mass of $KClO_3 = 122.5\ g\ mole^{-1}$

No. of moles of $KClO_3 = 490 / 122.5 = 4$ moles

According to reaction, $2KClO_3 \rightarrow 2KCl + 3O_2$

Stoichiometrically, 2 moles of $KClO_3 = 3$ moles of $O_2$

4 moles of $KClO_3 = 3/2 \times 4 = 6$ moles of $O_2$

1 mole $= 6.02 \times 10^{23}$ molecules of $O_2$

6 moles $= 6 \times 6.02 \times 10^{23}$ molecules of $O_2$

$= 3.612 \times 10^{24}$ molecules of $O_2$

**Q7. Explain Gay Lussac’s law of combining volume of gases?**

**Ans:** Gay Lussac’s law of combining volume:

According to the Gay Lussacs’ Law of combining volumes, the gases react in simple whole number ratios to produce products.

For example in the reaction: $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$

is telling that one volume of hydrogen gas reacts with one volume of chlorine gas to produce two volumes of hydrogen chloride gas.

**Q8. How will you explain volume of gases at STP?**

**Ans:** Volume of gases at STP:

In stoichiometric calculations the problem can be solved easily if reactants and products are used correctly.

22.414 dm$^3$ of any gas at STP = 1 mole $= 6.02 \times 10^{23}$ molecules.

22.414 dm$^3$ of $H_2$ gas at STP = 2g $= 6.02 \times 10^{23}$ molecules.

22.414 dm$^3$ of $NH_3$ gas at STP = 17g $= 6.02 \times 10^{23}$ molecules.
Molar Volume:
A mole and volume relationship exists between reactants and products provided the gases are at S.T.P. This volume of 22.4 dm³ is called Molar Volume.

Sample Problem No. 1.4

Determine the volume that 2.5 moles of chlorine molecules occupy at STP.

Solution: We know that
22.4 dm³ of Cl₂ (Chlorine) at S.T.P. = 1 mole
Or 1 mole of Cl₂ occupies a volume of 22.4 dm³ at S.T.P.
2.5 mole of Cl₂ occupy a volume of 22.4dm³ × 2.5 = 56 dm³

Self Check Exercise 1.3

(a) How many moles of oxygen molecule are there in 50.0 dm³ of oxygen gas at S.T.P?
(b) What volume does 0.80 mole of N₂ gas occupy at S.T.P?
(Ans: (a) 2.23 moles, (b) 17.93 dm³)

Solution:
(a) We know that
1 mole of O₂ occupies volume of 22.414 dm³ at STP
22.414 dm³ of O₂ at STP = 1 mole of O₂
1 dm³ of O₂ at STP = \( \frac{1}{22.414} \)
50 dm³ of O₂ at STP = \( \frac{1}{22.414} \) × 50 = 2.23 moles
(b) We know that
1 mole of N₂ occupies volume of 22.414 dm³ at STP
1 mole of N₂ = 22.414 dm³ of N₂ at STP
0.8 mole of N₂ = 22.414 × 0.8 = 18 dm³

Ans: Limiting Reactants:
The reactant that is consumed completely in a chemical reaction is called limiting reactant.
Also it can be defined as the reactant which produces the least number of moles of products in a chemical reaction.
The amount left un-used or un-reacted after completion of reaction is called “Reactant in excess”.

Q10. How will you identify limiting reactant in a reaction?
Ans: Identification of a Limiting Reactant in a Reaction:
A limiting reactant can be recognized by calculating the number of moles of products formed from data of the given amounts of the reactants using a balanced chemical equation. The reactant, which produces the least amount of products, is the limiting reactant.
Example: For example, 10 moles of \( H_2 \) and 7 moles of \( O_2 \) were reacted to produce \( H_2O \). Which one of the reactant is the limiting reactant? We can calculate as follows:

The reaction is, \[ 2H_2 + O_2 \rightarrow 2H_2O \]

**Stoichiometrically,**

(i) \[ \begin{align*}
2H_2 &= 2H_2O \\
i.e.\ 2\text{ Moles of } H_2 &= 2\text{ Moles of } H_2O \\
10\text{ Moles of } H_2 &= 10\text{ Moles of } H_2O \\
\end{align*} \]

(ii) \[ \begin{align*}
O_2 &= 2H_2O \\
i.e.\ 1\text{ Mole of } O_2 &= 2\text{ Moles of } H_2O \\
so.\ 7\text{ Moles of } O_2 &= 2 \times 7 = 14\text{ Moles of } H_2O \\
\end{align*} \]

Since \( H_2 \) gives the least number of moles of \( H_2O \), i.e. 10 moles, so \( H_2 \) is the limiting reactant.

![Sample Problem No. 1.5](image)

200 g of \( K_2Cr_2O_7 \) was reacted with 200g conc. \( H_2SO_4 \). Calculate

(a) Mass of atomic oxygen produced

(b) Mass of reactant left unreacted

**Solution:**

(a) Mass of \( K_2Cr_2O_7 \) = 200g

Formula Mass of \( K_2Cr_2O_7 \) = 294g mole\(^{-1}\)

No of moles of \( K_2Cr_2O_7 \) = \( \frac{200}{294} \) = 0.68 moles

Mass of \( H_2SO_4 \) = 200g mole\(^{-1}\)

Formula Mass of \( H_2SO_4 \) = 98g

No of moles of \( H_2SO_4 \) = \( \frac{200}{98} \) = 2.04 moles

\( K_2Cr_2O_7 + 4H_2SO_4 \rightarrow K_2SO_4 + Cr_2(SO_4)_3 + 4H_2O + 3(O) \)

<table>
<thead>
<tr>
<th>( K_2Cr_2O_7 )</th>
<th>2 moles</th>
<th>( 0.68 ) moles</th>
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<tbody>
<tr>
<td>1 mole of ( K_2Cr_2O_7 )</td>
<td>=</td>
<td>3 moles of 'O'</td>
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<tr>
<td>0.68 mole of ( K_2Cr_2O_7 )</td>
<td>=</td>
<td>3 \times 0.68 moles of 'O'</td>
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<tr>
<td>4 moles of ( H_2SO_4 )</td>
<td>=</td>
<td>3 moles of 'O'</td>
</tr>
<tr>
<td>2.04 moles of ( H_2SO_4 )</td>
<td>=</td>
<td>( \frac{3}{4} \times 2.04 )</td>
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<tr>
<td></td>
<td>=</td>
<td>1.53 moles</td>
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As \( H_2SO_4 \) is producing small amount so, \( H_2SO_4 \) is the limiting reactant and produced oxygen = 1.53 moles.

Mass in gram = Number of moles \times\) Molecular mass = 16 \times 1.53 = 24.48 g

(b) In this problem \( H_2SO_4 \) is the limiting reactant and \( K_2Cr_2O_7 \) is the reactant in the excess

We have 0.68 moles of \( K_2Cr_2O_7 \) and 2.04 moles of \( H_2SO_4 \)
According to the reaction,
4 moles of H₂SO₄ = 1 mole of K₂Cr₂O₇
2.04 moles of H₂SO₄ = \( \frac{1}{4} \times 2.04 \)
= 0.51 moles of K₂Cr₂O₇

No of moles of K₂Cr₂O₇ left unreacted = 0.68 − 0.51
= 0.17 moles

Mass of K₂Cr₂O₇ = No of moles × Formula Mass of K₂Cr₂O₇
= 0.17 × 294 = 49.98 g

So Mass of K₂Cr₂O₇ left unreacted = 49.98 g

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Sample Problem No. 1.6

20 g of H₂SO₄ on dissolving in water ionizes completely. Calculate
(a) No of H₂SO₄ molecules
(b) No of H⁺ and SO₄²⁻
(c) Mass of individual ion

Solution:

a. Mass of H₂SO₄ = 20 g
Molar Mass of H₂SO₄ = 98.016 g

No of molecules of H₂SO₄ = \( \frac{\text{Mass of H₂SO₄}}{\text{Molar Mass of H₂SO₄}} \times 6.02 \times 10^{23} \)
= \( \frac{20}{98.016} \times 6.02 \times 10^{23} \)
= 1.228 \times 10^{23}

b. H₂SO₄ dissolves in water as follows:
H₂SO₄ → 2H⁺ + SO₄²⁻

According to equation
1 molecule of H₂SO₄ = 2 H⁺ ions
1.228 \times 10^{23} molecules of H₂SO₄ = 2 \times 1.228 \times 10^{23} H⁺ ions

As 1 molecule of H₂SO₄ = 1 SO₄²⁻ ions
So, 1.228 \times 10^{23} molecule of H₂SO₄ = 1.228 \times 10^{23} SO₄²⁻ ions

c. Mass of individual ions

Mass of H⁺ = \( \frac{1.008}{6.02 \times 10^{23}} \times 2.456 \times 10^{23} \)
= 0.411 g

Mass of SO₄²⁻ = \( \frac{96}{6.02 \times 10^{23}} \times 1.228 \times 10^{23} \) = 19 g

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Sample Problem No. 1.7

Magnesium metal reacts with Sulphur to produce MgS. How many grams of magnesium sulphide (MgS) can be made from 1.50g of Mg and 1.50g of sulphur by the reaction

\[
\text{Mg} + \text{S} \rightarrow \text{MgS}
\]
Solution:

Mass of Mg = 1.50g

\[ \text{No. of moles of Mg} = \frac{1.50}{24} = 0.0625 \text{ moles} \]

Mass of S = 1.50g

\[ \text{No. of moles of S} = \frac{1.50}{32} = 0.0467 \text{ moles} \]

Mg + S \rightarrow MgS

i.e., 1 mole of Mg = 1 mole of MgS
so, 0.0625 moles of Mg = 0.0625 moles of MgS
also, 1 mole of Mg = 1 mole of MgS
so, 0.0647 moles of Mg = 0.0647 Moles of MgS

Since S gives the least number of moles of products as compared to Mg, so it is the limiting reactant.

Now we calculate the mass of MgS in grams.

Mass of 1 Mole of MgS = 24 + 32 = 56g

Mass of 0.0467 Mole of MgS = 56 \times 0.0467g = 2.6152g

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**Self Check Exercise 1.4**

(1) Zinc and Sulphur react to form Zinc Sulphide according to the following balanced chemical equation:

\[ \text{Zn} + \text{S} \rightarrow \text{ZnS} \]

If 6.00g of Zinc and 4.00g of Sulphur are available for reaction, then determine

(a) The limiting reactant.
(b) The mass of Zinc Sulphide produced.

(Ans. (a) Zinc is the limiting reactant since the whole is consumed.
(b) Mass of Zinc Sulphide produced = 8.94g)

Solution:

Mass of Zn = 6g

Atomic mass of Zn = 65.41

\[ \text{Number of moles of Zn} = \frac{\text{Mass in gram}}{\text{Atomic mass}} = \frac{6}{65.41} = 0.0917 \text{ moles} \]

Mass of S = 4g

Atomic mass of S = 32 g

\[ \text{Number of moles of S} = \frac{\text{Mass in gram}}{\text{Atomic mass}} = \frac{4}{32} = 0.125 \text{ moles} \]

Zn + S \rightarrow ZnS

1 mole of Zn = 1 mole of ZnS

0.0917 moles of Zn = 0.0917 moles of ZnS

Also, 1 mole of S = 1 mole of ZnS

So, 0.125 moles of S = 0.125 moles of ZnS

Since Zn gives the least number of moles of products as compared to S, so it is the limiting reactant. Now we calculate the mass of ZnS in grams.

Mass of 1 mole of ZnS = 65.41 + 32 = 97.41g
Mass of 0.0917 moles of ZnS = 97.41 \times 0.0917 = 8.94 g

Ans. (a) Zinc is the limiting reactant since the whole is consumed.
(b) Mass of Zinc Sulphide produced = 8.94 g

(2) Aluminium reacts with bromine to form Aluminium bromide, as shown by the balanced chemical equation, \(2\text{Al} + 3\text{Br}_2 \rightarrow 2\text{AlBr}_3\)

If 15.8g of Al and 55.6g of Br\(_2\) are available for reaction, then determine
(a) The limiting reactant  (b) The mass of AlBr\(_3\) produced

Ans: (a) Bromine is the limiting reactant.
(b) Mass of AlBr\(_3\) formed = 61.9g

Solution: Mass of Al = 15.8 g
Atomic mass of Al = 27

\[
\text{Number of moles of } \text{Al} = \frac{\text{Mass in gram}}{\text{Atomic mass}} = \frac{15.8}{27.98} = 0.585 \text{ moles}
\]

Mass of Br\(_2\) = 55.6 g
Molar mass of Br\(_2\) = 79.9 \times 2 = 159.8 g/mole

\[
\text{Number of moles of } \text{Br}_2 = \frac{\text{Mass in gram}}{\text{Molar mass}} = \frac{55.6}{159.8} = 0.348 \text{ moles}
\]

Now, \(2\text{Al} + 3\text{Br}_2 \rightarrow 2\text{AlBr}_3\)
2 mole of Al = 2 moles of AlBr\(_3\)
1 mole of Al = \(\frac{2}{2}\) = 1 mole of AlBr\(_3\)
0.585 mole of Al = 0.585 mole of AlBr\(_3\)
Also, 3 moles of Br\(_2\) = 2 moles of AlBr\(_3\)

1 mole of Br\(_2\) = \(\frac{2}{3}\) × 0.348 = 0.232 moles of AlBr\(_3\)

Since Br\(_2\) gives the least number of moles of products as compared to Al, so it is the limiting reactant. Now we calculate the mass of AlBr\(_3\) in grams.

Mass of 1 mole of AlBr\(_3\) = 27 + 79.9 \times 3 = 27 + 239.7 = 266.7 g

Mass of 0.232 moles of AlBr\(_3\) = 0.232 \times 266.7 = 61.9 g

Q11. Differentiate between limiting reactant and reactant in excess in a reaction?

Ans: Difference between limiting reactant and reactant in excess:

<table>
<thead>
<tr>
<th>Limiting reactant</th>
<th>Reactant in excess</th>
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<td>iii. It is usually cheaper.</td>
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<tr>
<td>iv. It is consumed completely in a chemical reaction.</td>
<td>iv. It is not consumed completely in a chemical reaction.</td>
</tr>
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</table>
OR

During a reaction in which 2 reactants are reacted sometimes one component is consumed completely and some amount of other reactant is left unreacted. The reactant which is consumed completely during the reaction is called Limiting Reactant and the reactant whose some amount is left unconsumed is called “Reactant in Excess”.

Sample Problem No. 1.8

Suppose 1.87 moles of ammonium chloride were reacted with 1.35 mole of calcium hydroxide. How many grams of calcium hydroxide are left unreacted in this reaction?

Solution: According to reaction,
\[ \text{Ca(OH)}_2 + 2\text{NH}_4\text{Cl.} \rightarrow \text{CaCl}_2 + 2\text{NH}_4\text{OH} \]
Let us calculate the no. of moles of Ca(OH)_2 in above example that reacts completely with 1.87 moles of NH_4Cl.
2 moles of NH_4Cl. = 1 mole of Ca(OH)_2
1.87 moles of NH_4Cl. = \( \frac{1}{2} \times 1.87 \)
\[ = 0.935 \text{ moles of Ca(OH)}_2 \]
So, no. of moles of Ca(OH)_2 consumed = 0.935 moles
And no. of moles of Ca(OH)_2 initially present = 1.35 moles
So, no. of moles of Ca(OH)_2 un consumed = 1.35 - 0.935 = 0.415 moles

As the molecular mass of Ca(OH)_2 = 74
So the mass of 0.415 moles of Ca(OH)_2 = 74 \times 0.415 = 30.71 g

Result: The excess amount of Ca(OH)_2, which is left unreacted is 30.71 g.
This is also called reactant in excess.

Q12. Differentiate between theoretical yield and actual yield.

Ans:

<table>
<thead>
<tr>
<th>Theoretical yield</th>
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</tr>
</thead>
</table>
| I. “The quantity of product calculated to be obtained from given quantities of initial reactants is called theoretical yield of a reaction”.
| I. The quantity of product that is actually produced in a chemical reaction is called the actual yield. |
| ii. It is calculated from balanced chemical equation | ii. It is calculated from experiments. |
| iii. Theoretical yield is always greater than actual yield. | iii. Actual yield is always lesser than theoretical yield |

Q13. Define percent yield and write its formula.

Ans: Percent Yield:

Percent yield is a measure of the efficiency of a chemical reaction.
Percent yield is calculated to be the experimental yield divided by theoretical yield multiplied by 100%.
\[
\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100
\]

Ans: Quantitative Reaction:
There are many reactions for which the actual yield is almost actually equal to the theoretical yield. Such reactions are quantitative, i.e., they can be used in chemical analysis.

**Sample Problem No. 1.9**

In an industry Copper metal was prepared by the following reaction,

\[ \text{Zn}_{(s)} + \text{CuSO}_4_{(aq)} \rightarrow \text{ZnSO}_4_{(aq)} + \text{Cu}_{(s)} \]

1.274g CuSO₄ when reacted with excess of Zn metal a yield of 0.392g Cu metal was obtained. Calculate the percentage yield.

Solution:

According to reaction, \[ \text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu} \]

Mass of CuSO₄ given \[ = 1.274g \]

Now we convert the no of grams of CuSO₄ into no. of moles.

Molecular mass of CuSO₄ \[ = 63.5 + 32 + 64 = 159.6 \text{ g/mole} \]

159.6g of CuSO₄ \[ = 1 \text{ mole} \]

1.274 g of CuSO₄ \[ = \frac{1}{159.6} \times 1.274 \]

\[ = 7.982 \times 10^{-3} \text{ moles.} \]

Stoichiometrically:

1 mole of CuSO₄ \[ = 1 \text{ mole of Cu} \]

7.98 x 10⁻³ moles of CuSO₄ \[ = 7.982 \times 10^{-3} \text{ moles of Cu.} \]

Hence, Theoretical yield \[ = 0.5072 \text{ g} \]

But Actual yield = 0.392 g

So,

\[ % \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \]

\[ = \frac{0.392}{0.5072} \times 100 = 77.3\% \]

**Sample Problem No. 1.10**

In a reaction, 2.00 moles of CH₄ was reacted with an excess of Cl₂. As a result, 177.0 g of CCl₄ is obtained. What is the (a) theoretical yield (b) actual yield (c) % yield of this reaction?
Solution:

Reaction is, \( \text{CH}_4^{(l)} + 4\text{Cl}_2^{(g)} \rightarrow \text{CCl}_4^{(l)} + 4\text{HCl}^{(g)} \)

Stoichiometrically,

From 2.0 moles of \( \text{CH}_4 \), we would expected to obtained 2.0 moles of \( \text{CCl}_4 \)

(a) Theoretical yield = 2.0 moles of \( \text{CCl}_4 \)

2.0 moles of \( \text{CCl}_4 \) means = 2 \times 154 = 308g

(b) Actual yield = 177.0g of \( \text{CCl}_4 \)

(c) Percent Yield:

\[
\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100
\]

\[
\% \text{ yield} = \frac{177}{308} \times 100 = 57.46 \%
\]

1. The overall balanced equation for the production of ethanol (\( \text{C}_2\text{H}_5\text{OH} \)) from sugar is as follows:

\[
\text{C}_6\text{H}_{12}\text{O}_6^{(aq)} \rightarrow 2\text{C}_2\text{H}_5\text{OH}^{(aq)} + 2\text{CO}_2^{(g)}
\]

(a) What is the theoretical yield of ethanol available from 10.0 g of sugar.

(b) If in a particular experiment, 10.0 g of sugar produces 0.664 g of ethanol, what is the percentage yield?

Ans: (a) Theoretical yield of ethanol = 5.125g

(b) Percentage yield = 12.89 %

Solution:

(a) Mass of \( \text{C}_6\text{H}_{12}\text{O}_6 \) = 10 g

Molar mass of \( \text{C}_6\text{H}_{12}\text{O}_6 \) = 12 \times 6 + 1 \times 12 + 16 \times 6 = 180 g/mole

Number of moles of \( \text{C}_6\text{H}_{12}\text{O}_6 \) = \( \frac{\text{Mass in gram}}{\text{Molar mass}} \) = \( \frac{10}{180} \) = 0.056 mole

1 mole of \( \text{C}_6\text{H}_{12}\text{O}_6 \) = 2 moles of \( \text{C}_2\text{H}_5\text{OH} \)

0.056 moles of \( \text{C}_6\text{H}_{12}\text{O}_6 \) = 2 \times 0.056 moles of \( \text{C}_2\text{H}_5\text{OH} \)

= 0.112 moles of \( \text{C}_2\text{H}_5\text{OH} \)

Molar mass of \( \text{C}_2\text{H}_5\text{OH} \) = 12 \times 2 + 1 \times 6 + 16 \times 1 = 46 g/mole

Mass of \( \text{C}_2\text{H}_5\text{OH} \) = 0.112 \times 46 = 5.125 g

(b) Percentage yield = \( \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \)

Percentage yield = \( \frac{0.664 \text{ g}}{5.125 \text{ g}} \times 100 = 12.89 \% \)

2. Solid carbon dioxide (dry ice) may be used for refrigeration. Some of this carbon dioxide is obtained as a by-product when hydrogen is produced from methane in the following reaction.

\[
\text{CH}_4^{(g)} + 2\text{H}_2\text{O}^{(g)} \rightarrow \text{CO}_2^{(g)} + 4\text{H}_2^{(g)}
\]
(a) What mass of CO₂ should be obtained from the complete reaction of 1250 g of methane.
(b) If the actual yield obtained is 3000 g then what is the percentage yield?

(Ans: a = 3438 g  b = 87.3 %)

Solution: (a)
\[ \text{Given mass of methane} = 1250 \text{ g} \]
\[ \text{Molar mass of CH}_4 = 12 + 4 = 16 \frac{g}{mole} \]
\[ \text{Number of moles of CH}_4 = \frac{\text{Mass in gram}}{\text{Molar mass}} = \frac{1250}{16} = 78.125 \text{ moles} \]

Stoichiometrically,
1 mole of CH₄ = 1 mole of CO₂
78.125 moles of CH₄ = 78.125 moles of CO₂
Molar mass of CO₂ = 12 + 32 = 44 g/mole
Mass of CO₂ obtained = number of moles \times molar mass = 78.125 \times 44 = 3437.5 g

(b) Actual yield = 3000 g
Theoretical yield = 3437.5 g
Percentage yield = \( \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \)
Percentage yield = 87.3 %

---

EXERCISE

MULTIPLE CHOICE QUESTIONS

1. Select the most suitable answer in the following MCQs.
   i. How many molecules are there in one mole of H₂O?
      (a) \( 6.023 \times 10^{19} \)  (b) \( 6.023 \times 10^{23} \)
      (c) \( 1.084 \times 10^{18} \)  (d) none of these
   ii. A flask contains 500 cm³ of SO₂ at STP. The flask contains
       (a) 40 g  (b) 100 g
       (c) 50 g  (d) none of these
   iii. A necklace has 6g of diamond in it. What are the numbers of atoms in it?
       (a) \( 6.02 \times 10^{23} \)  (b) \( 12.04 \times 10^{23} \)
       (c) \( 1.003 \times 10^{23} \)  (d) \( 3.01 \times 10^{23} \)
(a) What mass of CO₂ should be obtained from the complete reaction of 1250 g of methane.

(b) If the actual yield obtained is 3000 g then what is the percentage yield?

(Ans: a = 3438 g  b = 87.3 %)

Solution: (a)

Given mass of methane = 1250 g

Molar mass of CH₄ = 12 + 4 = 16 \frac{g}{mole}

Number of moles of CH₄ = \frac{Mass \text{ in gram}}{Molar \text{ mass}} = \frac{1250}{16} = 78.125 \text{ moles}

Stoichiometrically,

1 mole of CH₄ = 1 mole of CO₂
78.125 moles of CH₄ = 78.125 moles of CO₂
Molar mass of CO₂ = 12 + 32 = 44 g/mole
Mass of CO₂ obtained = number of moles \times molar mass = 78.125 \times 44 = 3437.5 g

(b) Actual yield = 3000 g
Theoretical yield = 3437.5 g
Percentage yield = \frac{Actual \text{ yield}}{Theoretical \text{ yield}} \times 100

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(a) 40 g  
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(c) 50 g  
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iii. A necklace has 6g of diamond in it. What are the numbers of atoms in it?

(a) 6.02 \times 10^{23}  
(b) 12.04 \times 10^{23}  
(c) 1.003 \times 10^{23}  
(d) 3.01 \times 10^{23}
iv. What is the mass of aluminium in 204 g of the aluminium oxide, Al₂O₃
(a) 26 g  (b) 27 g  (c) 54 g  (d) 108 g

v. The reactant which is consumed earlier and gives least quantity of product is called.
(a) Reactant  (b) Stoichiometry
(c) Limiting reactant  (d) Stoichiometric amount

vi. Which one of the following compounds contains the highest percentage by mass of nitrogen
(a) NH₃  (b) N₂H₄  (c) NO  (d) NH₄OH

vii. Vitamin-A has a molecular formula C₂₀H₃₀O. The number of vitamin - A molecules in 500 mg of its capsule will be
(a) 6.02×10²³  (b) 1.05×10²³
(c) 3.01×10²²  (d) 3.01×10²¹
(e) none of these

viii. When one mole of each of the following is completely burnt in oxygen, which will give the largest mass of CO₂?
(a) Carbon Monoxide  (b) Diamond
(c) Ethane  (d) Methane

ix. One mole of ethanol and one mole of ethane have an equal
(a) Mass  (b) Number of Atoms
(c) Number of electron  (d) Number of molecules

x. Methane reacts with steam to form H₂ and CO as shown below,
\[ CH₄(g) + H₂O(g) \rightarrow CO(g) + 3H₂(g) \]
What volume of H₂ can be obtained from 100 cm³ of methane at the standard temperature and pressure?
(a) 300 cm³  (b) 200 cm³  (c) 150 cm³  (d) 100 cm³

xii. How many moles of oxygen ate needed for the complete combustion of two moles of butane?
(a) 2  (b) 8  (c) 10  (d) 13

xiii. If four moles of SO₂ are oxidised to SO₃, how many moles of oxygen molecules are required.
(a) 0.5  (b) 1.0  (c) 1.5  (d) 2.0

xiv. The relative atomic mass of Chlorine is 35.5. What is the mass of 2 moles of Chlorine gas?
(a) 142g  (b) 71g  (c) 35.5g  (d) 18.75g

xv. Which of the following statements is incorrect?
(a) 12g of Carbon gas contains one mole of atoms
(b) 28g of Nitrogen gas contains one mole of molecules of N₂
(c) 1 dm$^3$ of a 1.0 Mole dm$^{-3}$ solution of NaCl contains one mole of Chloride ions
(d) None of above

xvi. One mole of propane has the same_________
(a) Number of molecules as one mole of methane (CH$_4$)
(b) Number of C-atoms as in one mole of butane (C$_4$H$_{10}$)
(c) Mass as half a mole of hexane (C$_6$H$_{12}$)
(d) Number of molecules as in one mole of ethane (C$_2$H$_6$)

xvii. What is the mass of one mole of Iodine Molecules?
(a) 254 g (b) 74 g (c) 106 g (d) 127 g

xviii. What volume of SO$_2$ at room temperature and pressure is produced on heating 9.7g of Zinc Sulphide (ZnS) if reaction takes place as follows

$$2\text{ZnS} (s) + 3\text{O}_2 (g) \rightarrow 2\text{ZnO} (s) + 2\text{SO}_2 (g)$$

(a) 1.2 dm$^3$ (b) 2.4 dm$^3$ (c) 3.6 dm$^3$
(d) 4.8 dm$^3$ (e) none of these

Answers

<table>
<thead>
<tr>
<th>i. b</th>
<th>ii. d</th>
<th>iii. d</th>
<th>iv. d</th>
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<td>xviii. e</td>
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2. Answer the following questions briefly:

i. 58.5 amu are termed as formula mass and not molecular mass of NaCl. Why?

Ans: Because NaCl is an ionic substance and molecules of ionic substances are termed as formula units but not as molecular mass. That is why 58.5 amu is the formula mass of NaCl and not molecular mass.

ii. Concept of limiting reactant is not applicable to the reversible reactions. Explain.

Ans: "A reactant which consumes earlier due to its smaller amount and produces least amount of product is called limiting reactant."

During a reversible reaction, reactants are converted into products and products convert back into reactants. So reactants are not completely consumed. As a result a limiting reactant cannot be identified in a reversible reaction.

OR (Second Answer)

Limiting reactant is the reactant which is completely consumed in a chemical reaction whereas in a reversible reaction, there is no reactant which is 100% consumed but is regenerated due to reversible reaction.

Therefore concept a limiting reactant is not applicable to a reversible reaction.

iii. How many covalent bonds are present in 9g of H$_2$O?

Ans: Mass of H$_2$O = 9g

Molecular mass of H$_2$O = 16 + 2 = 18 g/mole
Number of water molecules = \( \frac{\text{Mass of water}}{\text{Molar mass of water}} \times N_A \)

Number of molecules of \( \text{H}_2\text{O} \) = \( \frac{9}{18} \times 6.02 \times 10^{23} = 3.01 \times 10^{23} \) molecules

1 molecule of \( \text{H}_2\text{O} \) = 2 Covalent Bonds

3.01 \( \times \) 10\(^{23} \) molecules of \( \text{H}_2\text{O} \) = 2 \( \times \) 3.01 \( \times \) 10\(^{23} \)

= 6.02 \( \times \) 10\(^{23} \) Covalent Bonds

iv. Differentiate between limiting and non-limiting reactants.

Ans: Difference between limiting reactant and reactant in excess:

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v. How many molecules of water are there in 12 g of ice?

Ans:

\[
\text{Mass of ice } = 12 \text{ g} \\
\text{Molar mass of water } = 16 + 2 = 18 \text{ g/mole} \\
\text{Mass of water } = \frac{12}{18} \times 6.023 \times 10^{23} = 4.01 \times 10^{23} \text{ molecules}
\]

vi. One mole of \( \text{H}_2\text{SO}_4 \) should completely react with two moles of \( \text{NaOH}. \) How does Avogadro’s number help to explain it?

Ans: \( \text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O} \)

In above reaction According to Avogadro’s number 1 mole of or 6.02 \( \times \) 10\(^{23} \) molecules of \( \text{H}_2\text{SO}_4 \) react with 2 moles or \( 2 \times 6.02 \times 10^{23} \) molecules of \( \text{NaOH} \) to produce 1 mole of \( \text{Na}_2\text{SO}_4 \) or 6.02 \( \times \) 10\(^{23} \) molecules of \( \text{Na}_2\text{SO}_4 \) and 2 moles or \( 2 \times 6.02 \times 10^{23} \) molecules of \( \text{H}_2\text{O} \). 

vii. Give reason that 1 mole of different compounds have different masses but have the same number of molecules.

Ans: According to Avogadro’s number 1 mole of different compounds have 6.02 \( \times \) 10\(^{23} \) molecules but have different masses. As different molecules have different atoms having different atomic masses, so their molecular masses may be different but number of molecules in 1 mole must be same i.e. 6.02 \( \times \) 10\(^{23} \) molecules.

Example: 18g of water = 1 mole = 6.023 \( \times \) 10\(^{23} \) molecules.

342g of sucrose = 1 mole = 6.023 \( \times \) 10\(^{23} \) molecules.

viii. 23g of sodium and 238g of uranium have equal number of atoms in them.

Ans: Atomic mass of number is 23 and that of uranium is 238. It means 23g of Na is equal to 1 mole of Na and 238 g of uranium is equal to 1 mole of uranium and
according to Avogadro's Law 1 mole of any element has $6.02 \times 10^{23}$ atoms. It means 1 mole of Na and 1 mole of uranium have equal number of atoms in them that is $6.02 \times 10^{23}$ atoms.

ix. **What is percentage composition? Calculate percentage composition of Glucose.**

**Ans:** Percentage Composition:

A compound may contain certain elements. The percentage of each element in a compound is called percentage composition of that compound. It is calculated as follow:

$$\% \text{ of an element} = \frac{\text{mass of element in compound}}{\text{Molar mass of compound}} \times 100$$

**Percentage Composition of Glucose:**

Molar mass of glucose ($C_6H_{12}O_6$) = $12 \times 6 + 1 \times 12 + 16 \times 6 = 180 \text{ g mole}^{-1}$

$$\% \text{ of an element} = \frac{\text{mass of element in compound}}{\text{Molar mass of compound}} \times 100$$

- $\% \text{ of C} = \frac{72}{180} \times 100 = 40\%$
- $\% \text{ of H} = \frac{12}{180} \times 100 = 6.67\%$
- $\% \text{ of O} = \frac{96}{180} \times 100 = 53.33\%$

x. **Calculate the weight of oxygen gas evolved when 5.0 g of $KClO_3$ are completely decomposed thermally.**

**Ans:** Given Mass of $KClO_3 = 5g$

Formula Mass of $KClO_3 = 39 + 35.5 + 16 \times 3 = 122.5 \text{ g/mole}$

Number of Moles of $KClO_3 = \frac{\text{Mass in gram}}{\text{Molar mass}} = \frac{5}{122} = 0.04 \text{ moles}$

$$2KClO_3 \rightarrow 2KCl + 3O_2$$

According to the above equation

- 2 moles of $KClO_3 = 3$ moles of $O_2$
- 1 moles of $KClO_3 = \frac{3}{2}$ moles of $O_2$
- $0.04 \text{ moles of } KClO_3 = \frac{3}{2} \times 0.04 = 0.06 \text{ moles of } O_2$
- 1 mole of $O_2 = 32 \text{ g of } O_2$
- 0.06 mole of $O_2 = 0.06 \times 32 = 1.92 \text{ g}$

xi. **What is limiting reactant? How will you determine it?**

**Ans:** Excess and Limiting Reactants:

The reactant that is consumed completely in a chemical reaction is called limiting reactant. Also it can be defined as the reactant which produces the least number of moles of products in a chemical reaction.

**Determination of Limiting Reactant:**

The following three steps should be followed to find out the limiting reactant

(i) Calculate the number of moles from the given amount of reactant.
(ii) Find out the number of moles of product with the help of a balanced chemical equation.
(iii) Identify the reactant which produces the least amount of product as limiting reactant.

xii. **Define Theoretical yield and actual yield.**

**Ans:** **Theoretical Yield:**
"The amount of the products calculated from a balanced chemical equation is called theoretical yield."

i. It is also known as calculated yield or expected yield.

ii. Theoretical yield of a reaction is always greater than the actual yield of the same reaction.

**Actual yield:**
"The amount of the products obtained with a given amount of the reactant in an actual experiment is called actual yield."

i. It is also known as experimental yield.

ii. The actual yield of a chemical reaction is always less than the theoretical yield.

xiii. **What is conversion factor?**

**Ans:** **Conversion factor or Mole ratio:**
Conversion factor means the ratios of number of moles of reactants taking part and the number of moles of products formed.

**Example Combustion of propane:**
For example, combustion of propane

\[ C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O \]

The mole ratios between the reactants and products can be shown as, one mole of \( C_3H_8 \) reacts with five moles of oxygen to give three moles of \( CO_2 \) and four moles of water. The amount of propane used will not affect these ratios.

**OR**
The simplest stoichiometry problems are mole-to-mole conversions, and for those you need only a balanced chemical equation. The coefficients in the balanced equation are used to construct a conversion factor between moles of one substance and moles of a different substance in the same reaction.

Mole ratios of reactants and products as given by a balanced chemical equation is called conversion factor e.g. in reaction,

\[ C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O \]

1 mole of \( C_3H_8 \) (propane) reacts with 5 moles of \( O_2 \) to produce 3 moles of \( CO_2 \) and 4 moles of \( H_2O \). Conversion factor is not effected by amount of reactants or products.

xiv. **What is the relationship between mass and volume of a gas at S.T.P.?**

**Ans:** **Molar Volume:**
A mole and volume relationship exists between reactants and products provided the gases are at S.T.P. This volume of 22.414 dm\(^3\) is called Molar Volume.
In stoichiometric calculations the problem can be solved easily if reactants and products are used correctly.

\[
\begin{align*}
22.414 \text{ dm}^3 \text{ of } H_2 \text{ gas at STP} & = 2g \quad = 6.02 \times 10^{23} \text{ molecules.} \\
22.414 \text{ dm}^3 \text{ of } NH_3 \text{ gas at STP} & = 17g \quad = 6.02 \times 10^{23} \text{ molecules.}
\end{align*}
\]
xv. **The actual yield is lesser than the theoretical yield. Give reasons.**

**Ans:** Actual yield of a reaction is less than the theoretical yield because of following reasons.

i. The reaction may not go to completion and may reduce the yield of product.

ii. Material may be lost in handling

iii. Side reactions may produce by-products

iv. Reactions are reversible

v. Mechanical loss takes place due to filtration, distillation, and separation by separating funnel, washing and crystallization etc.

vi. Inexperience worker or method may be faulty

xvi. **What are the representative particles in one mole of a gas at S.T.P.?**

**Ans:** One mole of any gas at STP occupies 22.414 dm³ and contains $6.02 \times 10^{23}$ particles.

**Examples:**

(i) $2\text{g of } \text{H}_2 = 1 \text{ mole of } \text{H}_2 = 22.414 \text{ dm}^3$

(ii) $16\text{g of CH}_4 = 1 \text{ mole of CH}_4 = 22.414 \text{ dm}^3$

According to Avogadro's Law

There are $6.02 \times 10^{23}$ Number of particles present in 22.414dm³ (1 mole) of a gas.

xvii. **What is Stoichiometry and Stoichiometric amounts.**

**Ans:** Stoichiometry:

The study of relative amounts of substances involved in a chemical reaction is called Stoichiometry. Such phenomenon is studied through the knowledge of Stoichiometry (Greek word *Stoicheion* means element and *metry* means measurement).

**Stoichiometric Amounts:**

The amounts of the reactants or the products as given by the balanced chemical equation are called stoichiometric amounts.

$$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$$

2 moles 1 mole 2 moles

4 g 32 g 36 g

3. **(a) What is Avogadro’s number? Give examples. How will be explain moles with the help of Avogadro’s Number.**

**(b) The liquid CHBr₃ has a density of 2.89 g dm⁻³. What volume of this liquid should be measured to contain a total of 2.4 x 10²⁴ molecules of CHBr₃.**

**(Ans: 348.4 cm³)**

**Ans:** (a) What is Avogadro’s number? Give examples. How will be explain moles with the help of Avogadro’s Number.

The number of atoms, ions, molecules or formula units present in 1 mole of a substance in called Avogadro's number and its numerical value is $6.02 \times 10^{23}$

e.g. 1 mole of Na(23g) = $6.02 \times 10^{23}$ atoms of Na

1 mole of NaCl = $6.02 \times 10^{23}$ formula units of NaCl

1 mole of NaCl = $6.02 \times 10^{23}$ Na⁺ ions

1 mole of NaCl = $6.02 \times 10^{23}$ Cl⁻ ions
Definition of Mole in terms of Avogadro's number:
6.02 × 10²³ atoms of an element, 6.02 × 10²³ molecules of a compound or
6.02 × 10²³ formula units of an ionic substance is called a mole.

e.g. 6.02 × 10²³ atoms of Na = 1 mole of Na
6.02 × 10²³ molecules of H₂O = 1 mole of H₂O
6.02 × 10²³ formula units of NaCl = 1 mole of NaCl

(b) The liquid CHBr₃ has a density of 2.89 g dm⁻³. What volume of this liquid should be measured to contain a total of 4.8 × 10²⁴ molecules of CHBr₃.

Solution: (Ans: 696.8 dm³)

\[ \text{Density of CHBr}_3 = 2.98 \text{ g dm}^{-3} \]

\[ \text{Volume} = ? \]

\[ \text{Molecules of CHBr}_3 = 4.8 \times 10^{24} \]

\[ \text{Molar mass of CHBr}_3 = 12 + 1.008 + 239.7 = 252.7 \]

As we know that

\[ \text{Number of molecules} = \text{moles} \times N_A \]

\[ \text{Number of molecules} = \frac{\text{Mass in gram}}{\text{Molar mass}} \times N_A \]

\[ \text{Mass in gram} = \frac{252.7 \times 4.8 \times 10^{24}}{6.022 \times 10^{23}} \]

\[ \text{Mass in gram} = 2014.2 \text{ gram} \]

\[ \text{Volume} = \frac{\text{Mass in gram}}{\text{Density}} \]

\[ \text{Volume} = \frac{2014.2}{2.89} = 696.8 \text{ dm}^3 \]

4. (a) Differentiate between actual yield and theoretical yield. How will you explain the percentage yield of the substance with the help of;

(b) The following reaction never goes to completion. Therefore less amount of NH₃ is obtained than expected theoretically,

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \]

42.0 g of H₂ produces 120.2 g of NH₃. Calculate the percentage yield of NH₃. (Ans: 50.5 %)

Ans: (a) Differentiate between actual yield and theoretical yield.

Actual Yield:

The quantity of product that is actually produced in a chemical reaction is called the actual yield.

Theoretical Yield:

“The quantity of product calculated to be obtained from given quantities of initial reactants is called theoretical yield of a reaction”.

The theoretical yield is not an estimate, but the calculated amount of the yield based on the best of conditions for the reaction being carried to completion.

The actual yield is the measured amount from the actual experiment. This is often less than the ideal theoretical yield because of other factors that affect the reaction, mainly that the reagent used in the reaction is consumed so that you have less material to progress the reaction to its full theoretical yield.
(b) The following reaction never goes to completion. Therefore less amount of NH₃ is obtained than expected theoretically.

\[ \text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g) \]

42.0 g of H₂ produces 120.2 g of NH₃.

(Ass: 50.5 %)

**Calculate the percentage yield of NH₃.**

**Ans.** Expected theoretical yield can be calculated as:

Mass of H₂ = 42g

Number of moles of H₂ = \( \frac{\text{Mass in gram}}{\text{Molar mass}} = \frac{42}{2} = 21 \) moles

According to reaction equation:

\[ \text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3 \]

3 moles of H₂ = 2 moles of NH₃

1 mole of H₂ = \( \frac{2}{3} \) moles of NH₃

21 moles of H₂ = \( \frac{2}{3} \times 21 = 14 \) moles of NH₃

Mass of NH₃ in grams = 14 \times 17 = 238 g

But actual yield = 120.2 g

\[ \text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \]

% yield = \( \frac{120.2}{238} \times 100 = 50.5 \% \)

5. (a), What do you know about percentage composition? How will you determine the percentage of each element in the substance?

(b) Glucose (C₆H₁₂O₆) is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass percentage of each element in glucose.

(Ans: C = 40 %, H = 6.66 %, O = 53.33 %)

**Ans:**

(a) What do you know about percentage composition? How will you determine the percentage of each element in the substance?

**Method to determine the percentage composition of a known compound:**

i) calculate the molar mass of compound

ii) calculate the percentage of each element in one mole of the compound. This is done by dividing the mass of each element in one mole of the compound by the molar mass multiplied by 100.

\[ \% \text{ of an element} = \frac{\text{mass of element in compound}}{\text{Molar mass of compound}} \times 100 \]

(b) Glucose (C₆H₁₂O₆) is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass percentage of each element in glucose. (Ans: C = 40 %, H = 6.66 %, O = 53.33 %)

**Solution:** Molar Mass of Glucose (C₆H₁₂O₆) = 72 + 12 + 96 = 180 g/mole

\[ \% \text{ of an element} = \frac{\text{mass of element in compound}}{\text{Molar mass of compound}} \times 100 \]
\[ \% \text{ of } C = \frac{22}{180} \times 100 \quad = 40\% \]
\[ \% \text{ of } H = \frac{12}{180} \times 100 \quad = 6.66\% \]
\[ \% \text{ of } O = \frac{96}{180} \times 100 \quad = 53.33\% \]

6. (a) How will you calculate the theoretical yield and percentage yield in a balanced chemical equation?

(b) A small piece of pure Al Metal having a volume of 2.50 cm\(^3\) is reacted with excess of HCl. What is the weight of H\(_2\) liberated? The density of Al is 2.70 g cm\(^{-3}\). (Ans: 0.752g)

Ans: (a) How will you calculate the theoretical yield and percentage yield in a balanced chemical equation?

**Theoretical Yield:**

"The quantity of product calculated to be obtained from given quantities of initial reactants is called theoretical yield of a reaction."

**Method of finding theoretical yield:**

i. find the number of moles of reactant which is not an excess.

ii. Find the number of moles of required product from the balanced chemical equation.

iii. Find the mass of product by using formula: Number of moles = \( \frac{\text{Mass in gram}}{\text{Molar mass}} \)

iv. The mass of the product is the theoretical yield.

\[
\text{Theoretical yield} = \frac{\text{Actual yield}}{\% \text{ yield}} \times 100
\]

**Percent Yield:** Percent yield is a measure of the efficiency of a chemical reaction. Percent yield can be calculated from the balanced chemical equation it is done by divided the theoretical yield with actual yield multiplying it by 100.

\[ \% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 \]

(b) A small piece of pure Al Metal having a volume of 2.50 cm\(^3\) is reacted with excess of HCl. What is the weight of H\(_2\) liberated? The density of Al is 2.70 g cm\(^{-3}\). (Ans: 0.752g)

\[
\text{Volume of Al} = 2.50 \text{ cm}^3 \quad , \quad \text{Density of Al} = 2.70 \text{ g cm}^{-3}
\]

\[
\text{Density} = \frac{\text{mass}}{\text{volume}}
\]

\[
\text{Mass of Al} = \text{Density} \times \text{Volume}
\]

\[
= 2.70 \times 2.50 = 6.75 \text{ g}
\]

\[
\text{Moles of Al} = \frac{\text{Mass in gram}}{\text{Molar mass}} = \frac{6.75}{27} = 0.25 \text{ moles}
\]

According to Reaction:

\[
2\text{Al} + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2
\]

2 moles of Al = 3 moles of H\(_2\)  
1 mole of Al = \( \frac{3}{2} \)
0.125 moles of Al = \frac{3}{2} \times 0.25 = 0.375 \text{ moles}

molecular mass of H$_2$ = 2

Mass of H$_2$ in grams = 0.375 \times 2 = 0.752 \text{ g}

7. How much Silver Chloride will be formed by mixing 120.0 g of Silver Nitrate with a solution of 52.0 g of NaCl.

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

(Ans: 101.24 g)

**Solution:**

Given mass of AgNO$_3$ = 120 g

Molar mass of AgNO$_3$ = 107.87 \times 1 + 14 \times 1 + 16 \times 3 = 169.87 \text{ g/mole}

Number of moles AgNO$_3$ = \frac{\text{Mass in gram}}{\text{Molar mass}} = \frac{120}{169.868} = 0.71 \text{ moles}

Mass of NaCl = 52 g

Molar mass of NaCl = 23 + 35.5 = 58.5 \text{ g/mole}

Number of moles NaCl = \frac{\text{Mass in gram}}{\text{Molar mass}} = \frac{52}{58.5} = 0.89 \text{ moles}

According to the balanced chemical equation.

1 mole of AgNO$_3$ = 1 mole of AgCl

0.71 mole of AgNO$_3$ = 0.71 mole of AgCl

Also

1 mole of NaCl = 1 mole of AgCl

0.89 mole of AgNO$_3$ = 0.89 mole of AgCl

Since, AgNO$_3$ produces least amount of product therefore, it is the limiting reactant.

Number of moles AgCl produced = 0.71 moles of AgCl

Molar mass of AgCl = 107.87 + 35.5 = 143.37 \text{ g/mole}

Mass of AgCl produced = 0.71 \times 143.37 = 101.7 \text{ g}

8. Which contains more atoms, 1 mole of Fe or 1 mole of H$_2$? Justify your stand.

(Ans: H$_2$)

**Solution:**

1 mole of Fe atoms = 6.02 \times 10^{24} \text{ Fe atoms}

1 mole of H$_2$ atoms = 2 \times 6.02 \times 10^{24} \text{ H atoms} = 12.04 \times 10^{24} \text{ H atoms}

Therefore 1 mole of H$_2$ contains more atoms than 1 mole of Fe.