1

STOICHIOMETRY

After completing this lesson, you will be able to:

This is 10 days lesson (period including homework)

- Understand the Given information from which any two of the following may be determined, calculate the theoretical yield, actual yield and percentage yield.
- Interpret a balanced chemical equation in terms of interacting moles, representative particles, masses and volumes of gases at STP.
- Perform stoichiometric calculations with balanced equation using moles, representative particles, masses and volumes of gases (at STP).
- Know the limiting reagent in a reaction, calculate the maximum amount of product (s) produced and the amount of any unreacted excess reagent.
- Construct mole ratios from balanced equations for use as conversion factors in stoichiometric problems.
- Calculate the theoretical yield and the percent yield from the given balanced equation, the amounts of the reactants and the actual yield.

INTRODUCTION

It is observed that a common factor for fireworks manufacturer, an engineer to design car engines and a chemist to manufacture a chemical plant is the amount of materials they need in their specific area of work.

For examples,

1. Fireworks manufacturer is required to mix calculated amounts of chemicals. A wrong calculated amount could result in violent explosions.
2. The engineer must know the relative amounts of fuel and oxygen needed for efficient combustion.
3. A chemist is required to use calculated amounts of reactants to produce exact amount of products.

Such phenomenon is studied through the knowledge of Stoichiometry (Greek words Stoicheion means element and metry means measurement). **The study of relative amounts of substances involved in a chemical reaction is called Stoichiometry. The amounts of reactants and products in a balanced chemical equation are called Stoichiometric Amounts.** Such study is essential when quantitative information about a chemical reaction is required. Moreover it is important to predict yields of chemical products.

1.1 MOLE

The atomic mass, formula mass and molecular mass of a substance expressed in grams is called Mole.

Examples

One mole of O (atom) = 16 g

One mole of O_2 (molecule) = 32 g

One mole of H_2O (molecule) = 18 g

One mole of Na^+ (ion) = 23 g

One mole of NaCl (Formula unit) = 58.5 g

1.2 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 6.023×10^{23}.** One mole of any gas at S.T.P occupies $22.414\ dm^3$ and contains 6.02×10^{23} particles.

$$\text{e.g.}\quad 2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$$

This reaction can also be expressed in terms of Avogadro's number. The equation states that $2\times6.02\times10^{23}$ molecules of hydrogen react with 6.02×10^{23} molecules of oxygen to produce $2\times6.02\times10^{23}$ molecules of water.

The relationship between moles and Avogadro's number in different substances is as follows:

1 mole of O-atoms $= 6.02\times10^{23}$ atoms

1 mole of $O_{2\,(g)}$ $= 6.02\times10^{23}$ molecules

1 mole of $H_2O_{(\ell)}$ $= 6.02\times10^{23}$ molecules

1 mole of $NaCl_{(s)}$ $= 6.02\times10^{23}$ formula units

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

For example, $NaCl \xrightarrow{H_2O} Na^{+1}_{(aq)} + Cl^{-1}_{(aq)}$

It shows that 1 mole of NaCl when dissolves in water gives 1 mole of Na^{1+} ions and 1 mole of Cl^{-1} ions.So According to Avogadro's number we can say that when 6.02×10^{23} formula units of NaCl are dissolved in water, they produce 6.02×10^{23} Na^{1+} and 6.023×10^{23} Cl^{-1} ions.

Mole Ratios

Mole ratio means the ratio of no. of moles of reactants taking part and the no. of moles of products formed. For example, combustion of propane

$$C_3H_{8(g)} + 5O_{2(g)} \longrightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

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.

Example 1.1

Methanol burns according to the following equation.

$$2CH_3OH + 3O_{2(g)} \longrightarrow 2CO_{2(g)} + 4H_2O_{(g)}$$

If 3.50 moles of methanol are burnt in oxygen, calculate

(a) How many moles of oxygen are used?

(b) How many moles of water are produced?

Solution

(a) Moles of methanol = 3.50 moles

Moles of oxygen =?

According to balanced chemical equation

2 moles of CH_3OH = 3 moles of O_2

1 moles of CH_3OH = 3/2 moles of O_2

3.5 moles of CH_3OH = $\frac{3.50 \times 3}{2}$ moles of O_2

= 5.25 moles of O_2

So the number of moles of O_2 consumed = 5.25 moles

(b) No. of Moles of $CH_3OH = 3.5$ moles

No. of Moles of H_2O =?

According to balanced chemical equation

2 moles of CH_3OH = 4 moles of H_2O

1 moles of CH_3OH = 4/2 moles of H_2O

3.5 moles of CH_3OH = $\frac{3.50 \times 4}{2}$ moles of H_2O

= 7.00 moles of H_2O

So the number of moles of H_2O produced = 7.00 moles

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.

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

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)

Example 1.2

Iron can be produced from iron ore Fe_2O_3 by reacting the ore with carbon monoxide (CO).Carbon dioxide (CO_2) is produced in this reaction as a by product.What mass of iron can be formed from 425 g of iron ore?

Solution

The balanced equation can be written as

$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$

Mass of iron ore = 425 g (given mass)

No of moles of iron ore = $\dfrac{\text{mass}}{\text{molecular mass}} = \dfrac{425\cancel{g}}{159.6\cancel{g}\text{ moles}^{-1}}$

= 2.66 moles of Fe_2O_3

According to balanced chemical equation

1 mole of Fe_2O_3 = 2 moles of Fe

2.66 moles of Fe_2O_3 = 2.66×2Moles of Fe = 5.32Moles of Fe

Mass of Fe produced = no. of moles of Fe $\times$ Molar mass of iron

= 5.32×55.9 g

Mass of iron produced = 297.388 g

Self Check Exercise 1.2

The main engines of the U.S. space shuttle are powered by liquid hydrogen and liquid oxygen. If 1.02×10^5 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,

$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$

(**Ans:** 8.16×10^5 kg oxygen)

Example 1.3

Calculate the no 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, $2KClO_3 \longrightarrow 2KCl + 3O_2$

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 of O_2 = 6.02×10^{23} molecules of O_2

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

= 3.612×10^{24} molecules of O_2

Example 1.4

20g of H_2SO_4 on dissolving in water ionizes completely. Calculate

a) No of H_2SO_4 molecules
b) No of H^+ and SO_4^{2-}
c) Mass of individual ion

Solution

a. Mass of H_2SO_4 = 20g

Molar Mass of H_2SO_4 = $98.016 g\ mole^{-1}$

No of molecules of H_2SO_4 = $\dfrac{\text{Mass of } H_2SO_4}{\text{Molar Mass of } H_2SO_4} \times 6.02 \times 10^{23}$

$= \dfrac{20}{98.016} \times 6.02 \times 10^{23}$

$= 1.228 \times 10^{23}$

b. H_2SO_4 dissolves in water as follows:

$$H_2SO_4 \longrightarrow 2H^+ + SO_4^{-2}$$

According to equation

1molecule of H_2SO_4 = $2H^+$ ions

1.228×10^{23} molecules of H_2SO_4 $= 2 \times 1.228 \times 10^{23}$ H^+ ions

$= 2.456 \times 10^{23} H^-$ ions

As 1molecule of H_2SO_4 = 1 SO_4^{2-} ions

So, 1.228×10^{23} molecule of H_2SO_4= 1.228×10^{23} SO_4^{2-} ions

c. Mass of individual ions $= \dfrac{\text{No. of ions}}{6.02 \times 10^{23}} \times \text{Molar mass of an ion}$

Mass of H^+ $= \dfrac{2.456 \times 10^{23}}{6.02 \times 10^{23}} \times 1.008$

$= 0.411g$

Mass of SO_4^2 $= \dfrac{1.228 \times 10^{23}}{6.02 \times 10^{23}} \times 96$

$= 19.58g$

Molar Volume

One mole of any gas at STP (standard temperature and pressure) occupies a volume of 22.414 dm^3. This volume is called Molar volume. With the help of this relationship, we can convert the mass of a gas at STP into its volume and vice versa. Hence we can say molar volume of gases is also related with their density at STP (see section 4.4.3).

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

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

22.414 dm^3 of NH_3 gas at STP = 17g = 6.02×10^{23} molecules.

Example 1.5

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

Solution:

We know that

22.414 dm^3 of Cl_2 (Chlorine) at STP = 1 mole

or 1 mole of Cl_2 occupies a volume of 22.414 dm^3 at STP.

2.5 mole of Cl_2 occupy a volume of 22.414dm^3 $\times$ 2.5 = 56.035 dm^3

Self Check Exercise 1.3

(a) How many moles of oxygen molecule are there in 50.0 dm^3 of oxygen gas at S.T.P?

(b) What volume does 0.80 mole of N_2 gas occupy at S.T.P?

(**Ans:** (a) 2.23 moles, (b) 17.93 dm^3)

1.3 PERCENTAGE COMPOSITION

The relative amounts of each element in a compound are expressed as the percentage composition. For example the percentage composition by mass in MgO is as follows,

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.

$$\%\text{ of an element} = \frac{\text{mass of element in 1 mole of compound}}{\text{molar mass of compound}} \times 100$$

1.4 LIMITING AND NON 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 reactant left un-used or un-reacted after completion of reaction is called Reactant in excess or Non Limiting reactant.

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 number of moles of products, is the limiting reactant.

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 $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$

Stoichiometrically,

According to balanced chemical equation

(i) — $2\,H_2 = 2\,H_2O$

i.e. 2 moles of H_2 = 2 moles of H_2O

10 moles of H_2 = $\frac{2}{2} \times 10 = 10$ moles of H_2O

(ii) $O_2 = 2H_2O$

i.e. 1 mole of O_2 = 2 moles of H_2O

so, 7 moles of O_2 = $2 \times 7 = 14$ moles of H_2O

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

Example 1.6

200 g of $K_2Cr_2O_7$ were reacted with 200g conc. H_2SO_4 .Calculate

(A) Mass of atomic oxygen produced

(B) Mass of reactant left unreacted

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

Solution:

(a) Mass of $K_2Cr_2O_7$ = 200g

Formula Mass of $K_2Cr_2O_7$ = 294g $mole^{-1}$

No of moles of $K_2Cr_2O_7$ = $\frac{200}{98}$

= 0.68 moles

Mass of H_2SO_4 = 200g $mole^{-1}$

Formula Mass of H_2SO_4 = 98g $mole^{-1}$

No of moles of H_2SO_4 = $\frac{200}{98}$

= 2.04 moles

$$\underset{\text{0.68 moles}}{K_2Cr_2O_7} + \underset{\text{2.04 moles}}{4H_2SO_4} \longrightarrow K_2SO_4 + Cr_2(SO_4)_3 + 4H_2O + 3[O]$$

According to balanced chemical equation

1 mole of $K_2Cr_2O_7$ = 3 moles of [O]

0.68 mole of $K_2Cr_2O_7$ = 3×0.68 moles of [O]

= 2.04 moles of [O]

4 moles of H_2SO_4 = 3 moles of [O]

2.04 moles of H_2SO_4 = $\frac{3}{4}\times2.04$

= 1.53 moles

As H_2SO_4 is producing lesser moles of product so, H_2SO_4 is the limiting reactant.

Mass of atomic oxygen is = No. of moles of oxygen atom x atomic mass

= 1.53 x 16 = 24.4 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 H_2SO_4 = 1 mole of $K_2Cr_2O_7$

2.04 moles of H_2SO_4 = $\frac{1}{4}\times2.04$

= 0.51 moles of $K_2Cr_2O_7$

No of moles of $K_2Cr_2O_7$ left unreacted = 0.68-0.51

= 0.17 moles

Mass of $K_2Cr_2O_7$ = No of moles × Formula Mass of $K_2Cr_2O_7$

= 0.17 × 294 = 49.98

So Mass of $K_2Cr_2O_7$ left unreacted = 49.98g

Example 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?

$$Mg + S \longrightarrow MgS$$

Solution:

Mass of Mg $= 1.50g$

No. of moles of Mg $= \frac{1.50}{24} = 0.0625$ moles

Mass of S $= 1.50g$

No. of moles of S $= \frac{1.50}{32} = 0.0468$ moles

$$Mg + S \longrightarrow MgS$$

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 S = 1 mole of MgS
so, 0.0468 moles of S = 0.0468 Moles of MgS

Since S gives the least No 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 0.0468 Mole of MgS = 56 x 0.0468g = 2.620g

Self Check Exercise 1.4

(1) Zinc and Sulphur react to form Zinc Sulphide according to the following balanced chemical equation $Zn + S \longrightarrow ZnS$

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 = 8.95g)

(2) Aluminium reacts with bromine to form Aluminium bromide, as shown by the balanced chemical equation, $2Al + 3Br_2 \longrightarrow 2AlBr_3$

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

During a reaction in which two 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".

The concept of reactant in excess may be more clear by studying the following sample problem.

Example 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 the reaction,

$$Ca(OH)_2 + 2NH_4Cl \longrightarrow CaCl_2 + 2NH_4OH$$

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 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 $Ca(OH)_2$ unreacted = 1.35-0.935

= 0.415 moles

Molar mass of $Ca(OH)_2$ = 74

Mass of $Ca(OH)_2$ = No. of moles × Molar mass

Mass of $Ca(OH)_2$ = 0.415 × 74

= 30.71g

Result: The excess amount of $Ca(OH)_2$, which is left unreacted is 30.71g.

1.5 THEORETICAL YIELD, ACTUAL YIELD AND PERCENT YIELD

"The quantity of a product calculated from a balanced chemical equation is called theoretical yield of the reaction". The quantity of a product that is actually produced in a chemical reaction is called the actual yield.

The percent yield can be calculated as,

$$\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

There are many reactions for which the actual yield is almost equal to the theoretical yield. Such reactions are quantitative, i.e., they can be used in chemical analysis. On the other hand, for some reactions, particularly those involving organic compounds, the actual yield of a reaction is less than the theoretical yield and percent yield, less than 100 %.

The actual yield may be less than the theoretical yield due to following reasons,

i) Side reactions may produce by-products
ii) Some reactions are reversible
iii) Mechanical loss takes place due to filtration, distillation, and separation by separating funnel, washing and crystallization etc.

Example 1.9

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

$$Zn_{(s)} + CuSO_{4(aq)} \longrightarrow ZnSO_{4(aq)} + Cu_{(s)}$$

1.274g $CuSO_4$ when reacted with excess of Zn metal a yield of 0.392g of Cu metal is obtained. Calculate the percentage yield.

Solution:

Mass of $CuSO_4$ given = 1.274g

Now we convert the no of grams of $CuSO_4$ into no. of moles.

Molecular mass of $CuSO_4$ = 63.5 + 32 + 64 = 159.5 g/mole

159.5g of $CuSO_4$ = 1 mole

1.274 g of $CuSO_4$ $= \frac{1}{159.5} \times 1.274$

$= 7.98210^{-3}$ moles.

Stoichiometrically

1 mole of $CuSO_4$ = 1 mole of Cu

7.98 x 10^{-3} moles of $CuSO_4$ = 7.982 $\times$ 10^{-3} moles of Cu.

as, 1 mole of Cu = 63.5 g

so, 7.982 x 10^{-3} moles of Cu $= 63.5 \times 7.982 \times 10^{-3}$

= 0.5072 g of Cu.

Hence, Theoretical yield = 0.5072 g

Actual yield = 0.392 g

So, % yield $= \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

$= \frac{0.392}{0.5072} \times 100 = 77.3\%$

Example 1.10

In a reaction, 2.00 moles of CH_4 were reacted with an excess of Cl_2. As a result, 177.0 g of CCl_4 is obtained. What is the

(a) theoretical yield **(b)** actual yield **(c)** % yield of this reaction?

Solution:

Reaction is, $CH_{4(g)} + 4Cl_{2(g)} \longrightarrow CCl_{4(l)} + 4HCl_{(g)}$

Stiochiometrically,

From 2.0 moles of CH_4, we would expected to obtained 2.0 moles of CCl_4

(a) Theoretical yield = 2.0 moles of CCl_4

Mass of CCl_4 = No. of moles x molar mass

= 2 x 154 = 308g

(b) Actual yield = 177.0g of CCl_4

(c) Percent Yield:

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

% yield $= \frac{177}{308} \times 100$

= 57.46 %

Self Check Exercise 1.5

1. The overall balanced equation for the production of ethanol (C_2H_5OH) from glucose is as follows:

$$C_6H_{12}O_{6(aq)} \longrightarrow 2C_2H_5OH_{(aq)} + 2CO_{2(g)}$$

(a) What is the theoretical yield of ethanol available from 10.0 g of glucose?

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

(**Ans:** (a) Theoretical yield of ethanol = 5.152g (b) Percentage yield = 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.

$$CH_{4(g)} + 2H_2O_{(g)} \longrightarrow CO_{2(g)} + 4H_{2(g)}$$

(a) What mass of CO_2 should be obtained from 1250 g of methane?

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

(**Ans:** a = 3437.5 g b = 87.3 %)

References for additional information

- James Brady and John R. Holum, Chemistry, The studies of matter and its changes.
- Theodore L. Brown, H. Eugene LeMay Jr and Bruce E. Bursten, Chemistry, The central Science.
- Rose Marie Gallagher and Paul Ingram, Complete chemistry.
- Graham Hill and John Holman, Chemistry in Context
- E. N. Ramesden, Calculations for A-Level chemistry.

1. **Select the most suitable answer in the following MCQs.**

i. How many molecules are there in one mole of H_2O?

(a) 6.023×10^{19} (b) 6.023×10^{23}

(c) 1.084×10^{18} (d) none of these

ii. A flask contains 500 cm^3 of SO_2 at STP. The flask contains;

(a) 40 g (b) 100 g

(c) 50 g (d) 1.427g

iii. A necklace has 6g of diamond in it. What are the numbers of atoms in it?
(a) 6.02×10^{23} (b) 12.04×10^{23}
(c) 1.003×10^{23} (d) 3.01×10^{23}

iv. What is the mass of aluminium in 204 g of the aluminium oxide, Al_2O_3 ?
(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 in excess (b) Stoichiometry
(c) Limiting reactant (d) Stoichiometric amount

vi. Which one of the following compounds contains the highest percentage by mass of nitrogen?
(a) NH_3 (b) N_2H_4
(c) NO (d) NH_4OH

vii. Vitamin-A has a molecular formula $C_{20}H_{30}O$. The number of vitamin-A molecules in 500 mg of its capsule will be;
(a) 6.02×10^{23} (b) 1.05×10^{21}
(c) 3.01×10^{22} (d) 3.01×10^{23}

viii. When one mole of each of the following is completely burnt in oxygen, which will give the largest mass of CO_2?
(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 H_2 and CO as shown below,

$$CH_{4(g)} + H_2O_{(g)} \longrightarrow CO_{(g)} + 3H_{2(g)}$$

What volume of H_2 can be obtained from 100 cm^3 of methane at the standard temperature and pressure?
(a) 300 cm^3 (b) 200 cm^3
(c) 150 cm^3 (d) 100 cm^3

xi. The Avogadro's constant is the number of
(a) Atoms in 1g of Helium (b) Molecules in 35.5g of Chloride
(c) Electrons needed to deposit 24g of Mg
(d) Atoms in 24g of Mg

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 SO_2 are oxidised to SO_3, 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 N_2
(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 H-atoms as one mole of methane (CH_4)
(b) Number of C-atoms as in one mole of butane (C_4H_{10})
(c) Mass as half a mole of hexane (C_6H_{14})
(d) Number of molecules as in one mole of ethane (C_2H_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:

$$2ZnS_{(s)} + 3O_{2(g)} \longrightarrow 2ZnO_{(s)} + 2SO_{2(g)}$$

(a) 1.2 dm^3 (b) 2.4 dm^3
(c) 3.6 dm^3 (d) 4.8 dm^3

2. Answer the following questions briefly:

i. 58.5 amu are termed as formula mass and not molecular mass of NaCl. Why?
ii. Concept of limiting reactant is not applicable to the reversible reactions. Explain.
iii. How many covalent bonds are present in 9g of H_2O?
iv. Differentiate between limiting and non-limiting reactants.
v. How many molecules of water are there in 12 g of ice?
vi. Which contains more atoms, 1 mole of Fe or 1 mole of H_2? Justify your stand.
vii. Give reason that 1 mole of different compounds have different masses but have the same number of molecules.
viii. 23g of sodium and 238g of uranium have equal number of atoms in them.
ix. Calculate the weight of oxygen gas evolved when 5.0 g of $KClO_3$ are completely decomposed thermally.
x. What is the relationship between mass and volume of a gas at S.T.P?
xi. The actual yield is lesser than the theoretical yield. Give reasons?
xii. What are the representative particles in one mole of a gas at S.T.P?
xiii. What is Stoichiometry and Stoichiometric amounts?

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

(b) The liquid $CHBr_3$ has a density of 2.89 g cm^{-3}. What volume of this liquid should be measured to contain a total of 4.8×10^{24} molecules of $CHBr_3$. (**Ans:** 696.8 cm^3)

4. (a) Differentiate between actual yield and theoretical yield.

(b) The following reaction never goes to completion. Therefore less amount of NH_3 is obtained than expected theoretically,

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)}$$

42.0 g of H_2 produces 120.2 g of NH_3. Calculate the percentage yield of NH_3.

(**Ans:** 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 $(C_6H_{12}O_6)$ 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 %)

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)

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?

$$AgNO_3 + NaCl \longrightarrow AgCl + NaNO_3$$

(**Ans:** 101.24 g)