

A Textbook of **Chemistry**

Grade
XI



KHYBER PAKHTUNKHWA TEXTBOOK BOARD,
PESHAWAR

Stoichiometry

After studying this unit, the students will be able to:

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

Teaching

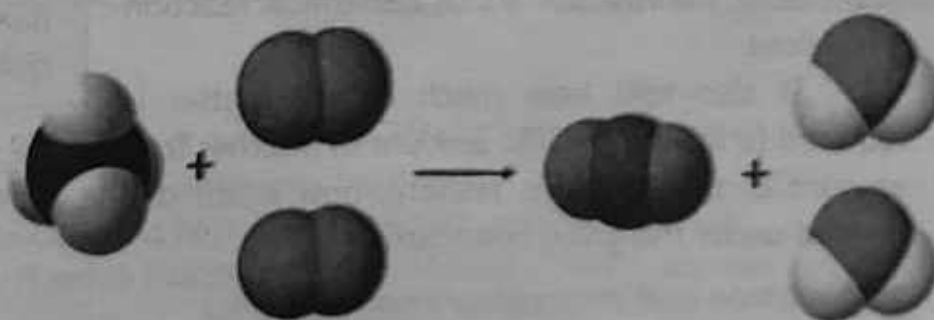
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Assessment

01

Weightage %

08



Introduction

Thousands of different reactions are going on in laboratories, factories, kitchens, car engines, the atmosphere, the earth beneath you and within your body. These chemical reactions have a great effect on your lives. There are many examples of such reactions like conversion of food to energy in the human body, formation of ammonia (fertilizer) by combining nitrogen and hydrogen, preparation of starch by plants from carbon dioxide and water using energy from sunlight etc. The main purpose of chemistry is to understand chemical changes. A chemist calculates the amount of product from the balanced chemical equation. A wrong calculation of amount would result in many problems. Therefore, it is required for a chemist to calculate the exact amount of reactants and products.

As you have learned in previous grades, chemistry involves in the measurements to find out what happens between atoms and molecules during the chemical reaction. In this unit, you will learn how to interpret the balanced chemical equation in terms of interacting with moles, representative particles, masses and volumes of gases and their stoichiometric calculations.

Stoichiometry, (Greek words 'Stoicheion' means elements and 'metron' means measurement). Stoichiometry is quantitative chemistry. It is the study of the relationship between relative amounts of substances involved in a chemical reaction. The formula of a compound provides information about the relative amount of each element present.

Stoichiometry is essential when quantitative information about a chemical reaction is required.

It also tells how much of the matter is required to form a specific amount of another form of the matter or if a particular amount of the matter is present, how much of the product could possibly be formed under the given conditions.

1.1 Mole and Avogadro's number (N_A)

The idea of atomic mass, relative to the mass of the carbon $^{12}_6\text{C}$ isotope, taken as 12 units, has already been discussed in your previous grades. Relative atomic mass is defined as the mass of one atom of an element compared with

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Analytical chemistry is the branch of chemistry, which deals with the study of analysis of obtaining, processing, characterizing the composition of matter and structure of matter, both qualitatively and quantitatively.

the mass of $\frac{1}{12}$ of carbon (C-12).

One gram-atom of any element is the relative atomic mass of the element expressed in grams. For example, the relative atomic mass of chlorine (Cl) is 35.5 atomic mass units (amu) and that of sulphur (S) is 32 amu. Therefore, one gram-atom of chlorine would be 35.5g and one gram-atom of sulphur would be 32 g.

Mole can be defined as the atomic mass, molecular mass, formula mass or ionic mass of a substance expressed in grams.

The concept of the gram-atom is useful because it is impossible to see or weigh individual atoms or even a large number of atoms. We can actually weigh one gram-atom of an element.

One gram-atom of any element contains the same number of atoms. This quantity has been given the name "mole" (mol). It follows that one mole of any element contains the same number of atoms.

Generally speaking, mole is a counting unit just like a dozen (12 similar things) or a gross (144 similar things). Mole, however, is a very large counting unit. It is equal to 6.023×10^{23} similar things, i.e. atoms, molecules, ions, etc. This constant number, present in one mole of a substance, was experimentally determined by an Italian scientist, Amadeo Avogadro and is known as Avogadro's number, represented by N_A . The number of particles (atoms, molecules, ions) in one mole of any substance is called Avogadro's number and is numerically equal to 6.023×10^{23} .

For example,

1 mol of oxygen (O) = 16g of oxygen = 6.023×10^{23} atoms of oxygen

1 mol of oxygen (O_2) = 32g of oxygen = 6.023×10^{23} molecules of oxygen

1 mol of water (H_2O) = 18g of H_2O = 6.023×10^{23} molecules of H_2O

1 mol of sodium chloride (NaCl) = 58.5g of NaCl = 6.023×10^{23} formula units of NaCl

So mole can also be defined as, "the amount (mass) of a substance, which contains Avogadro's number (6.023×10^{23}) of particles (atoms, ions, molecules, formula units)." It establishes a link between mass of substance and number of particles.

1 mole =
 6.023×10^{23}
Avogadro's Number



12 grams of carbon-12 = 1 mole

Figure 1.1 Mole and Avogadro's number (N_A)

Table 1.1 Names of the First 20 Elements

Element	Symbol	Atomic Number	Atomic Mass	Element	Symbol	Atomic Number	Atomic Mass
Hydrogen	H	1	1.008	Helium	He	2	4.002
Lithium	Li	3	6.94	Beryllium	Be	4	9.012
Boron	B	5	10.81	Carbon	C	6	12.01
Nitrogen	N	7	14.00	Oxygen	O	8	15.99
Fluorine	F	9	18.99	Neon	Ne	10	20.18
Sodium	Na	11	22.99	Magnesium	Mg	12	24.30
Aluminium	Al	13	26.98	Silicon	Si	14	28.08
Phosphorus	P	15	30.97	Sulphur	S	16	32.06
Chlorine	Cl	17	35.45	Argon	Ar	18	39.94
Potassium	K	19	39.09	Calcium	Ca	20	40.07

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Several methods are used to find the sample. These may be broadly classified as qualitative analysis and quantitative analysis. Qualitative analysis tells us 'what' is in the sample, while quantitative analysis tell us 'how much' is in the sample. These two types of analysis are often used together and are considered in analytical chemistry.

Quantitative analysis is very important in chemistry. Chemists usually determine the amount of substance present in the compound. Here are a few sample results of quantitative analysis:

- The NaCl solution concentration is 0.1 M.
- The chemical reaction produces 5.50 moles.
- Ore contains 64.88% silver by mass.

Beside this, quantitative analysis is very important in medical science. Medical sciences use the result of chemical analysis for example, glucose is determined in blood and urine of diabetics. The cholesterol level in blood is determined by quantitative analysis.

1.2 Mole Calculation

Molar Mass

Molar mass is defined as the mass in grams of one mole of a substance.
The term **molar mass** is also used to the mass of a mole of any substance (atoms

molecules, formula units or ions); it is the atomic mass of an atom or the sum of the atomic masses of all the atoms in a molecule, an ion or a formula unit (in grams).

You can find the molar mass (mass of one mole) of any substance using the following steps.

- Write down the formula of the substance; for example, formula of sulphuric acid is H_2SO_4
- Work out its molecular mass; sulphuric acid contains two hydrogen atoms (atomic mass 1 amu), one sulphur atom (atomic mass 32 amu) and four oxygen atoms (atomic mass 16 amu). So, formula mass for sulphuric acid

$$\begin{aligned}(\text{H}_2\text{SO}_4) &= 2\text{H} + \text{S} + 4\text{O} \\ &= (2 \times 1) + (1 \times 32) + (4 \times 16) \\ &= 2 + 32 + 64 = 98 \text{ amu}\end{aligned}$$

- Express this in grams per mole; the molar mass of sulphuric acid is 98 g/mol.

In experimental work, chemists work with varying masses. They cannot always use one mole of a substance. The equation that links the mass of a substance to the number of moles present is,

$$\text{Number of moles (n)} = \frac{\text{mass (in grams)}}{\text{molar mass (in grams per mol)}} \quad (1.1)$$

Using the above equation, you can convert any mass of a particular substance into moles or vice versa.

As one mole of any substance contains 6.023×10^{23} (Avogadro's number) of atoms, molecules or formula units, one can also develop an equation between the number of moles and the number of particles present. This will lead to the following relationship.

$$\text{Number of moles (n)} = \frac{\text{Number of particles (atoms, molecules, formula units)}}{\text{Particles per mole or Avogadro's number (mol}^{-1}\text{)}} \quad (1.2)$$

Or

$$\text{Number of moles (n)} = \frac{\text{Number of particles (atoms, molecules, formula units)}}{6.023 \times 10^{23} \text{ (mol}^{-1}\text{)}} \quad (1.3)$$

The following examples will illustrate the above equations.

Example 1.1

How many moles are there in 60g of sodium hydroxide (NaOH)?

Solution:

The relative formula mass of sodium hydroxide (NaOH) = $\text{Na} + \text{O} + \text{H}$

So,

Molar mass of NaOH = $23 + 16 + 1 = 40 \text{ g/mol}$

$$\text{Number of moles}(n) = \frac{\text{Mass(g)}}{\text{Molar mass of NaOH(g/mol)}}$$

$$\text{number of moles}(n) = \frac{60\text{g}}{40\text{g/mol}} = 1.5\text{mol}$$

Reading Check

1. Define relative atomic mass.
2. What is gram atom?

Practice Problem 1.1

1. Calculate the molecular masses (in amu) of the following compounds:
(a) Sulphur dioxide (SO_2) and (b) Caffeine ($\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$).
(Atomic mass of C = 12, S = 32, N = 14, H = 1, O = 16)
2. Calculate the No. of moles in 5.68g of iron. (Atomic mass of Fe = 55.8)

Example 1.2

What is the mass of 0.5 moles of calcium carbonate (CaCO_3)?

Solution:

The relative formula mass of calcium carbonate (CaCO_3) = Ca + C + 3O

$$\text{Molar mass of CaCO}_3 = 40 + 12 + (16 \times 3) = 100\text{g/mol}$$

$$\text{Number of moles}(n) = \frac{\text{Mass(g)}}{\text{Molar Mass of CaCO}_3(\text{g/mol})}$$

$$0.5 = \frac{\text{mass(g)}}{100\text{g/mol}} \quad \text{or} \quad \text{Mass} = 0.5 \times 100 = 50\text{g}$$

Example 1.3

In a certain experiment, 8.50×10^{25} molecules of water were used. Calculate the number of moles of water.

Solution:

$$\text{Number of H}_2\text{O molecules} = 8.50 \times 10^{25}$$

$$\text{So, Number of moles}(n) = \frac{\text{Number of H}_2\text{O molecules}}{\text{Avogadro's number}}$$

$$\text{Number of moles}(n) = \frac{8.50 \times 10^{25}}{6.023 \times 10^{23}} = 1.41 \times 10^2 \text{mol}$$

Thus, 1.41×10^2 moles of water would have reacted in this particular experiment.

Practice Problem 1.2

How many Cu atoms are present in 0.5 mol of copper?

Example 1.4

How many formula units are present in 125g of hydrated copper sulphate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$)?

Solution:

Mass of hydrated copper sulphate ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$) = 125g

The relative formula mass of hydrated copper sulphate is,

$$\begin{aligned}\text{Formula mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} &= 64 + 32 + (4 \times 16) + [5 \times (2 \times 1 + 16)] \\ &= 64 + 32 + 64 + 5(18) = 250 \text{ amu}\end{aligned}$$

Thus, Molar mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ = 250g/mol

So,

$$\text{Number of moles}(n) = \frac{\text{Mass (g)}}{\text{Molar Mass of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O (g/mol)}}$$

$$\text{Number of moles}(n) = \frac{125\text{g}}{250\text{g/mol}} = 0.5\text{mol}$$

Calculate the number of formula units as,

$$\text{Number of moles}(n) = \frac{\text{number of formula units of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O}}{\text{Avogadro's number}}$$

$$0.5 = \frac{\text{number of formula units of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O}}{6.023 \times 10^{23}}$$

$$\begin{aligned}\text{Number of formula units of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} &= 0.5 \times 6.023 \times 10^{23} \\ &= 3.011 \times 10^{23} \text{ formula units}\end{aligned}$$

Thus, there are 3.011×10^{23} formula units of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ in 125g of the salt.

Practice Problem 1.3

What is the mass of 1.204×10^{22} atoms of lead? (Atomic mass of Pb = 207)

Self-Assessment

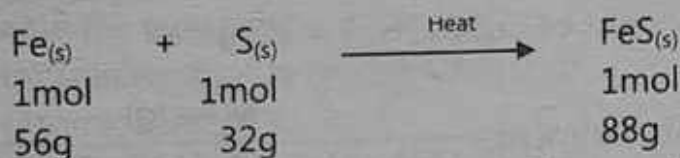
- What is the number of moles in each of the following?
 - 52g of silicon (atomic mass of Si = 28)
 - 1.42g of O_2 (oxygen gas) (atomic mass of O = 16)
 - 3.6×10^{24} atoms of lithium (atomic mass of Li = 7)
 - 4×10^{22} formula units of potassium chloride (KCl). (atomic mass of K = 39, Cl = 35.5)
- Which of the following contains the greatest number of particles?
 - 4g of lithium (atomic mass of Li = 7)
 - 4g of chlorine gas (Cl_2). (atomic mass of Cl = 35.5)
 - 4g of hydrogen gas (H_2). (atomic mass of H = 1)
 - 4g of water (atomic mass of H = 1, atomic mass of O = 16)

1.2.1 The Mole and Chemical Equations

When iron reacts with sulphur, it produces Iron sulphide (FeS), the equation is,



This indicates that equal number of atoms of iron and sulphur are needed to react. We know that 1 mole of iron (56g) and 1 mole of sulphur (32g) contain the same number of atoms. Reaction of these amounts would give 1 mole of iron sulphide (88g).



The mass of the product is equal to the total mass of the reactants. This is in accordance with the law of conservation of mass (matter), which states *that matter (mass) can neither be created nor destroyed during a chemical reaction.*

In chemistry calculations, we frequently need to calculate the relationship between reactants or products in a chemical reaction. For example, if we know the mass of one reactant, we can find how much product will be formed if the reactant is fully converted into products. At the same time, we can know how much of a second reactant is required to fully react with the first reactant. A chemical equation must, therefore, be balanced because it indicates the number of moles of reactants and products involved in the reaction. There are various ways of doing these calculations because quantities can be expressed either in moles or in mass units.

The stoichiometric calculation can be carried out by the following steps.

1. Write correct and balanced chemical equation.
2. Convert the mass of given substance into moles by dividing with molar mass.
3. Calculate the moles of required substance using their coefficient in the balanced chemical equation.
4. Convert the moles of required substance into mass.

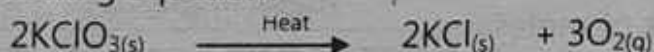
a. **Mole – Mole Conversion**

If you are given the mole of one substance, you can calculate the moles of the other substances.

Example 1.5

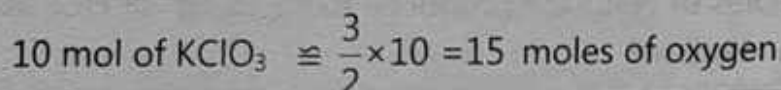
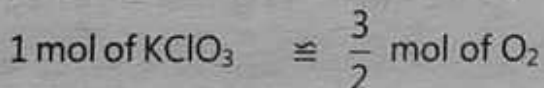
Oxygen can be prepared by the decomposition of potassium chlorate (KClO_3). How many moles of $\text{O}_{2(g)}$ can be formed by taking 10 moles of $\text{KClO}_{3(s)}$

according to the following equation?



Solution:

We have the conversion factor from the balanced chemical equation as:



Practice problem 1.4

How many moles of carbon dioxide will be produced by the complete combustion of 2.0 mol of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) according to the following equation?



Self-Assessment

1. Define molar mass.
2. What is Avogadro's number?
3. What is balanced chemical equation?
4. What are stoichiometric calculations?
5. How many ways can you read the given equation?



b. Mole – Mass Conversion

If you are given the mole of one substance, you can calculate the mass of the other substances.

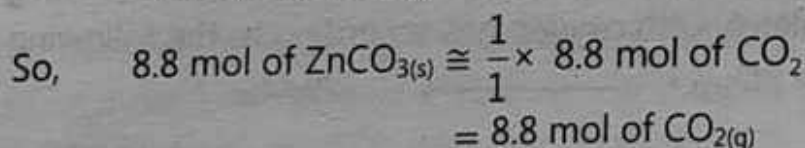
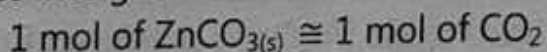
Example 1.6

How many grams of CO_2 can be produced by thermally decomposing 8.8 moles of $\text{ZnCO}_{3(s)}$?

Solution



First, we do a mole – mole ratio calculation and then convert the calculated moles into grams at the end. Thus from balanced equation:



Now convert number of $\text{CO}_{2(g)}$ into grams, using the formula

$$\text{Number of moles} = \frac{\text{mass (in grams)}}{\text{molar mass (in grams per mol)}}$$

$$\text{Molar mass of CO}_{2(g)} = \text{C} + 2\text{O} = 12 + (2 \times 16) = 44 \text{ g/mol}$$

$$\begin{aligned} \text{So, mass of CO}_{2(g)} \text{ produced} &= \text{moles} \times \text{molar mass of CO}_{2(g)} \\ &= 8.8 \text{ mol} \times 44 \text{ g/mol} = 387.2 \text{ g} \end{aligned}$$

Example 1.7

How many moles of NaCl are produced from 15.5 g of HCl, according to the neutralization reaction?

**Solution**

$$\text{The molar mass of HCl} = 1 + 35.5 = 36.5 \text{ g/mol}$$

$$\begin{aligned} \text{Number of moles of HCl (n)} &= \frac{\text{mass (in grams)}}{\text{molar mass (in grams per mol)}} \\ &= \frac{15.5 \text{ g}}{36.5 \text{ g/mol}} = 0.425 \text{ mol} \end{aligned}$$

By using the equivalency (conversion factor) from the balanced chemical equation, we see that,

$$1 \text{ mol of HCl} \cong 1 \text{ mol of NaCl}$$

$$\text{Thus, } 0.425 \text{ mol of HCl} \cong 0.425 \text{ mol of NaCl}$$

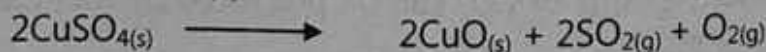
Hence, 15.5 g of HCl will produce 0.425 mol of NaCl.

Practice Problem 1.5

- How many grams of $\text{CO}_{2(g)}$ can be produced by thermally decomposing 6.5 moles of $\text{CaCO}_{3(s)}$?



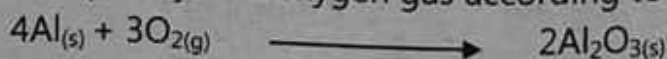
- Calculate the mole of copper oxide obtained from 110 g of CuSO_4 .

**c. Mass – Mass Conversion**

If you are given the mass of one substance, you can calculate the mass of the other substances.

Example 1.8

What mass of aluminium oxide (Al_2O_3) is produced from 14.5g of Al metal when it reacts completely with oxygen gas according to the following equation?

**Solution**

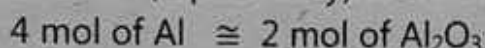
Convert 14.5g of Al into moles,

Molar mass of Al = 27g/mol

Number of moles of Al = $\frac{\text{mass (in grams)}}{\text{Molar mass of Al (in grams per mol)}}$

Number of moles of Al = $\frac{14.5 \text{ g}}{27 \text{ g/mol}} = 0.54 \text{ mol of Al}$

Use conversion factor (equivalency) from the balance chemical equation:



Therefore, $0.54 \text{ mol of Al} \cong \frac{2}{4} \times 0.54 \cong 0.27 \text{ mol of Al}_2\text{O}_3$

Finally, we calculate the mass of Al_2O_3 produced, by converting 0.27 moles of Al_2O_3 into grams. For this, we need molar mass of Al_2O_3 which is,

Molar mass of $\text{Al}_2\text{O}_3 = 2\text{Al} + 3\text{O} = (2 \times 27) + (3 \times 16) = 54 + 48 = 102 \text{ g/mol}$

So, mass of Al_2O_3 produced = mole $\times$ molar mass
 $= 0.27 \text{ mol} \times 102 \text{ g/mol} = 27.54 \text{ g}$

Practice problem 1.6

The oxidation of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) to carbon dioxide (CO_2) and water (H_2O) occurs as:



What is the mass of CO_2 produced, when 856g of $\text{C}_6\text{H}_{12}\text{O}_6$ are decomposed?

1.2.2 Calculations Involving Gases

Not all chemical reactions involve solids and liquids. For those reactions in which gases are involved, it is more convenient to measure volumes than masses. In gaseous reactions, weighing of gases as compared to solids and liquids is relatively difficult. It is much easier to measure the volume of a gas as gases weigh very little. But we need to know how the given volume of a gas is related to the number of atoms or molecules present. For this, we take help from the *Avogadro's law, which states that equal volumes of all gases, at the same temperature and pressure, contain equal numbers of particles (atoms or molecules)*. (You will study this law in more detail in unit 4).

It has been found experimentally that one mole of any gas at standard temperature (0°C or 273K) and pressure (1 atmosphere) or STP (standard temperature and pressure), occupies a volume of 22.4dm^3 . This is

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Units of Volume

1 litre = 1 dm^3

1 litre = 1000ml

$1 \text{ dm}^3 = 1000 \text{ cm}^3$

called the molar volume (volume of one mole) of the gas at STP.

Avogadro's law applies to all gases. So making use of this fact, it is easier to convert the volume of any gas into moles or moles into volume using the following relationship.

$$\text{Number of Moles of gas (n)} = \frac{\text{Volume (dm}^3\text{) at STP}}{\text{molar volume (22.4 dm}^3\text{/mol) at STP}} \quad (1.4)$$

Example 1.9

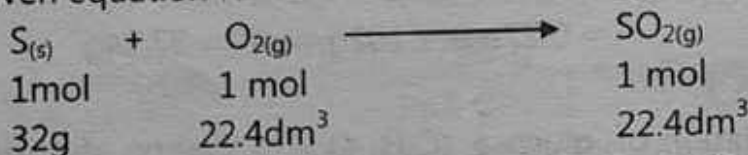
According to the following equation solid sulphur burns in oxygen and produces SO_2 .



If 15g of sulphur is burnt, what volume of $\text{SO}_{2(g)}$ is produced, at STP?

Solution

Given equation for the reaction at STP



$$\text{Number of moles of sulphur burnt} = \frac{\text{mass (in grams)}}{\text{molar mass (in grams per mol)}}$$

Atomic mass of S = 32

$$\text{Number of moles of sulphur burnt} = \frac{15\text{g}}{32\text{ g/mol}}$$

Number of moles of sulphur = 0.469 mol

From the equation 1 mol of S $\cong$ 1 mol of $\text{SO}_{2(g)}$

Therefore, 0.469 of S $\cong$ 0.469 of $\text{SO}_{2(g)}$

To calculate the volume of $\text{SO}_{2(g)}$ produced, we have

$$\text{Number of Moles (n)} = \frac{\text{Volume (dm}^3\text{) at STP}}{\text{molar volume (22.4 dm}^3\text{/mol)}}$$

$$0.469\text{ mol} = \frac{\text{Volume (dm}^3\text{)}}{22.4\text{ dm}^3\text{/mol}}$$

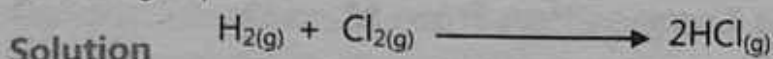
$$\text{Volume (dm}^3\text{)} = 0.469\text{ mol} \times 22.4\text{ dm}^3\text{/mol} = 10.51\text{ dm}^3$$

Thus, burning 15g of sulphur will produce 10.51 dm³ of $\text{SO}_{2(g)}$.

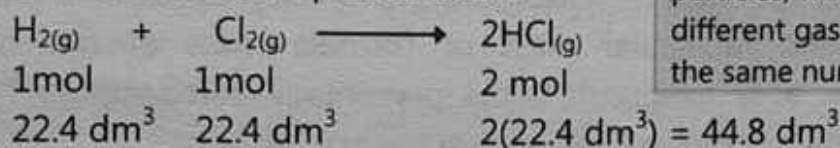
Example 1.10

What volume of chlorine gas would be needed at STP to produce 10 dm³ of hydrogen chloride (HCl) gas in presence of excess of $\text{H}_2(g)$, according to the reaction

following equation.

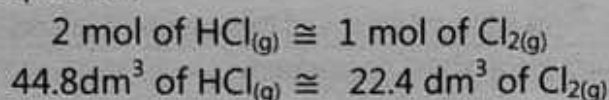


Given the chemical equation at STP



Here, the volumes of the gases involved are in the same ratio, as the numbers of moles, given by the equation.

From equation,



Therefore, $10\text{ dm}^3 \text{ of HCl}_{(g)} = \frac{22.4\text{ dm}^3}{44.8\text{ dm}^3} \times 10\text{ dm}^3 \text{ Cl}_2 = 5\text{ dm}^3 \text{ of Cl}_{2(g)}$

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Equal volumes of different gases contain the same number of particles, whereas, equal mass of different gases do not contain the same number of particles.

Practice Problem 1.7

1. Calculate the No. of molecules present in 11.5 dm^3 of N_2 at STP.
2. What is the volume, in dm^3 , of 3 moles of oxygen gas (O_2) at STP?
3. In an experiment, hydrochloric acid was reacted with calcium carbonate at STP and 80 cm^3 of carbon dioxide was produced. Calculate the number of molecules of carbon dioxide given off.

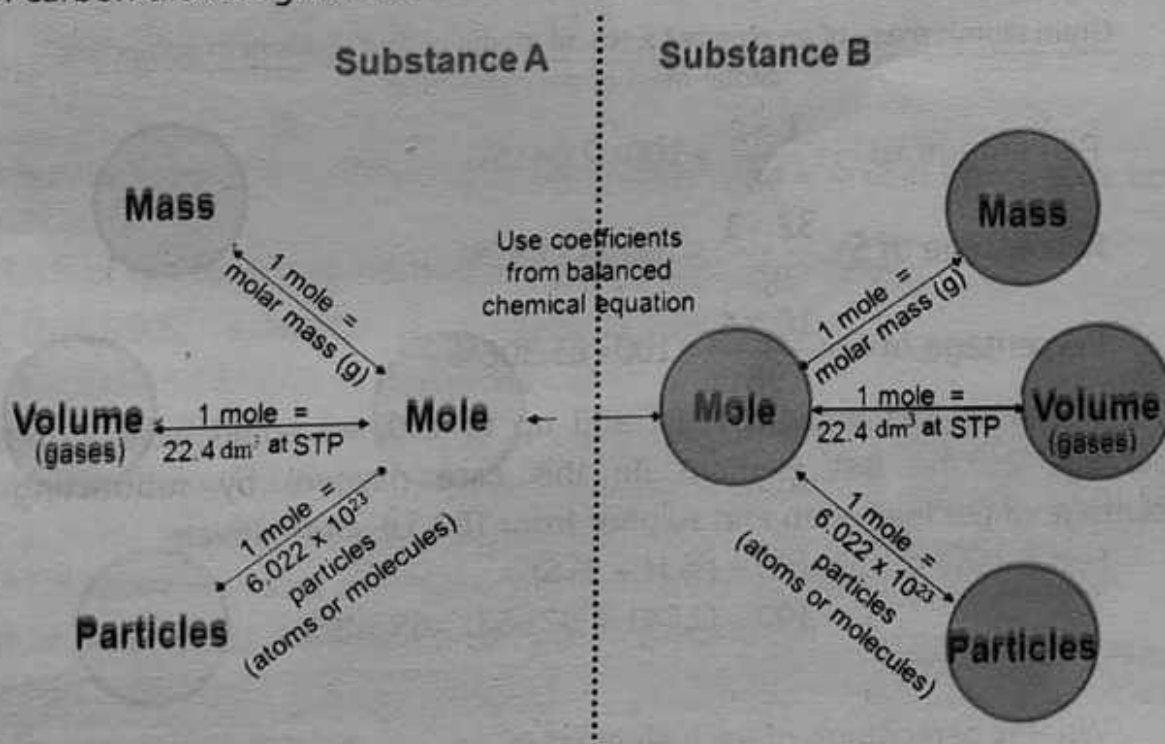


Figure 1.2 Representation of different Conversions

1.3 Percentage Composition

a. Determining Percent Composition from Mass Data

The percent by mass of each element in a compound is called percentage composition of a compound. Once the elements present in a compound are identified and the molecular mass or formula mass of the compound is determined, it is easy to calculate the percentage composition. The following mathematical relationship will help you in this connection.

$$\text{Percentage of an Element} = \frac{\text{Total mass of the element in compound}}{\text{Total mass of the compound}} \times 100 \quad (15)$$

Or

$$= \frac{\text{Gram atomic mass of an element} \times \text{No. of atoms of that element in compound}}{\text{Molar mass of the compound}} \times 100 \quad (16)$$

Sum of the individual percentages of all the elements must be equal to 100.

Example 1.11

Calculate the percentage composition of sulphuric acid (H_2SO_4).

Solution

From the formula, we get:

$$\begin{aligned} \text{Molar mass of } \text{H}_2\text{SO}_4 &= (2 \times 1) + (1 \times 32) + (4 \times 16) \\ &= 2 + 32 + 64 = 98 \text{ g/mol} \end{aligned}$$

Thus, percentage of each element is calculated as follows,

$$\frac{\text{Gram atomic mass of an element} \times \text{No. of atoms of that element in compound}}{\text{Molar mass of the compound}} \times 100$$

$$\text{Percentage of H} = \frac{1 \times 2}{98} \times 100 = 2.041\%$$

$$\text{Percentage of S} = \frac{32 \times 1}{98} \times 100 = 32.653\%$$

$$\text{Percentage of O} = \frac{16 \times 4}{98} \times 100 = 65.306\%$$

Since, the percentages must add up to 100, we could have found the percentage of the last element (in this case oxygen) by subtracting the percentage of the hydrogen and sulphur from 100. i.e. alternatively,

$$\begin{aligned} \text{Percentage of O} &= 100 - (\% \text{ H} + \% \text{ S}) \\ &= 100 - (2.041 + 32.653) = 65.306\% \end{aligned}$$

Practice Problem 1.8

What is percentage of each element in,
(a) Benzene (C_6H_6) (b) Glucose ($\text{C}_6\text{H}_{12}\text{O}_6$)

b. Determining Masses from Percent Composition Data

You can also perform the reverse calculation, determining the mass of an element in a given sample, if we know the total mass of the sample and its percent composition.

Example 1.12

You have a 10.0g sample of a metal alloy that contains only aluminium and zinc. If the sample is 36% aluminium by mass, what masses of Al and Zn are present?

Solution

You are told that the sample is 36 % Aluminium by mass. Because the only other component is zinc, it must make up the remaining 64% of the mass. You can multiply each of these percentages by 10.0 grams to find the masses of each element.

$$\% \text{ of Aluminium} = 36$$

$$\% \text{ of Zn} = 100 - 36 = 64\%$$

$$\text{Mass of sample} = 10\text{g}$$

To find masses of Al and Zn,

$$\text{Mass of Al} = \frac{36}{100} \times 10 = 3.6\text{g}$$

$$\text{Mass of Zn} = \frac{64}{100} \times 10 = 6.4\text{g}$$

Practice problem 1.9

- Calculate the percent composition by mass of H, P and O in Phosphoric acid (H_3PO_4).
- Which of the following has highest percentage of nitrogen
 - $(\text{NH}_4)_2\text{SO}_4$
 - NH_4Cl
 - NH_4NO_3
 - $(\text{NH}_4)_3\text{PO}_4$

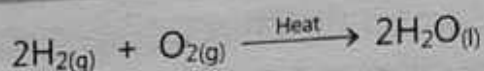
1.4 Excess and Limiting Reagents

Limiting Reagent: A reactant that is consumed earlier during a chemical reaction is called limiting reagent. It can also be defined as the reagent which produces the least number of moles of products in a chemical reaction. It limits or determines the amount of products formed in the reaction.

Excess Reagent: The reactant whose some amount remains unreacted or unused after the completion of the reaction is called excess reagent. Consider the reaction of hydrogen and oxygen to form water.

Reading Check

What is percentage composition?



A brief inspection of this balanced chemical equation indicates that one mole of oxygen, (if it is completely consumed) requires two moles of hydrogen to form 2 moles of water. If one carefully measures 1 mole of oxygen and 2 moles of hydrogen, ignites them, one would get 2 moles of water. When the stoichiometric amounts (as given in the balanced chemical equation) are reacted together, there will be no reactants left at the end of the reaction.

In actual practice, however, due to one reason or the other, one may not take the amount of the reactants in stoichiometric ratios, as demanded by the balanced chemical equation. For example, one may take 5 moles of oxygen for 4 moles of hydrogen to form water. Since the above balanced chemical equation indicates that,

$$\begin{aligned} 2 \text{ moles of H}_2 &\cong 1 \text{ mole of O}_2 \\ \text{Therefore, 4 moles of H}_2 &\cong \frac{1}{2} \times 4 = 2 \text{ moles of O}_2 \end{aligned}$$

Thus, out of 5 moles of oxygen available, only 2 moles will react with 4 moles of hydrogen and the remaining 3 moles of oxygen will be left unreacted. The unreacted oxygen is termed as "reactant (reagent) in excess" and hydrogen which is consumed earlier, is termed as "limiting reagent". In a chemical reaction the maximum amount of the product formed is controlled by the amount of "limiting reagent" because as soon as the limiting reagent is consumed, the reaction stops and no more products are formed.

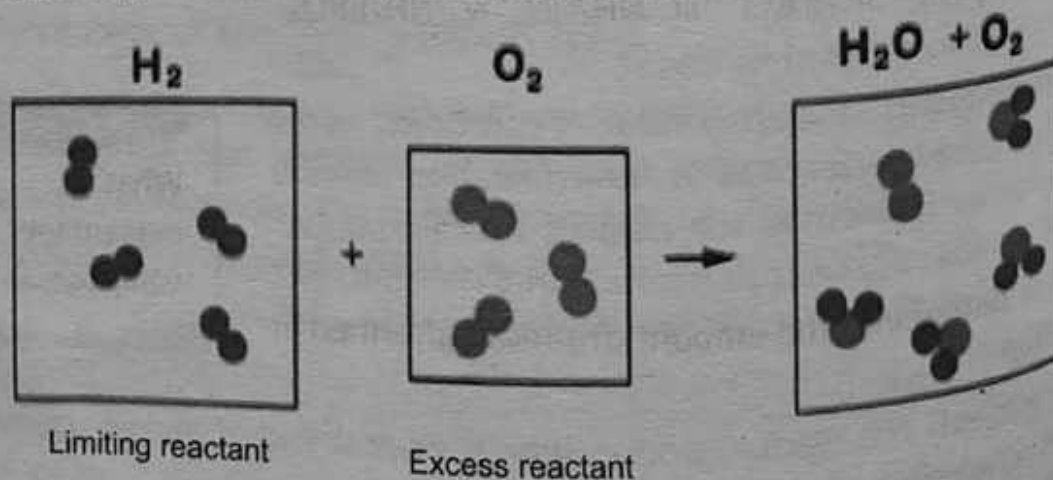


Figure 1.3 Limiting and excess reagent

Tidbit

In the chemical industry, you need to know the limiting reactant in order to get the maximum yield of product at minimum cost.

In many situations, it is inconvenient, or time consuming to weigh the exact amounts of two or more reactants needed to produce a given amount of the product. Thus, almost, all reactions that go to completion have a limiting reagent. Sometimes, it is obvious while in other situation it needs calculation to identify it. Following are the steps that help in identifying the limiting reagent.

1. The amounts of the reactants (if given in mass units) are first converted into moles.
2. Using balanced chemical equations, the moles of the required product are calculated from the available moles of each reactant.
3. The reactant, which gives the least number of moles of the required product, will be the limiting reagent.

Alternatively, comparison of the moles of reactants with help of balanced chemical equation also helps in identifying the limiting reagent.

Example 1.13

Carbonic acid (H_2CO_3) can be prepared according to the following reaction.

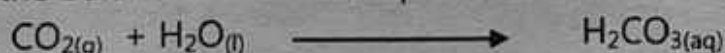


If 120g of CO_2 is dissolved in 80g of water then,

- (a) Identify the limiting reagent
- (b) Calculate the maximum amount (in grams) of the product formed.

Solution

- (a) Given the balanced chemical equation



First we convert the amounts of both the reactants into moles,

Mass of $\text{CO}_2 = 120\text{g}$

Molar mass of $\text{CO}_2 = (1 \times 12) + (2 \times 16) = 12 + 32 = 44\text{g/mol}$

Moles of $\text{CO}_2 = \frac{120\text{g}}{44\text{g/mol}} = 2.72 \text{ moles}$

And mass of $\text{H}_2\text{O} = 80\text{g}$

Molar mass of $\text{H}_2\text{O} = (2 \times 1) + (1 \times 16) = 18\text{g/mol}$

Moles of $\text{H}_2\text{O} = \frac{80\text{g}}{18\text{g/mol}} = 4.44 \text{ moles}$

Now calculate the number of moles of the product (H_2CO_3) from the number of moles of each of the reactants.

From balanced chemical equation, we have,



Therefore, $2.72 \text{ mol of CO}_2 \cong 2.72 \text{ mol of H}_2\text{CO}_3$

Similarly, $1 \text{ mol of H}_2\text{O} \cong 1 \text{ mol of H}_2\text{CO}_3$

Therefore, $4.44 \text{ mol of H}_2\text{O} \cong 4.44 \text{ mol of H}_2\text{CO}_3$

Since 120g of CO_2 produces less number of moles (2.72 moles) of carbonic acid, so CO_2 is the limiting reagent and H_2O is the reagent in excess.

(b) The amount of product (H_2CO_3) is controlled by the amount of limiting reagent. Therefore, the maximum amount (in moles) of the product formed will be 2.72 moles.

We have the molar mass of $\text{H}_2\text{CO}_3 = (2 \times 1) + (1 \times 12) + (3 \times 16) = 62 \text{ g/mol}$

$$\text{Therefore, moles of H}_2\text{CO}_3 = \frac{\text{Mass (g)}}{\text{Molar mass}}$$

$$2.727 \text{ mol} = \frac{\text{Mass (g)}}{62 \text{ g/mol}}$$

$$\begin{aligned} \text{Mass of H}_2\text{CO}_3 &= 2.727 \text{ mol} \times 62 \text{ g/mol} \\ &= 169.07 \text{ g} \end{aligned}$$

Thus, the maximum amount of H_2CO_3 produced will be 169.07g.

Practice Problem 1.10

If 0.600 mol of chlorine gas is reacted with 0.500 mol of aluminium metal to produce aluminium chloride, which reactant is in excess? How many moles of aluminium chloride can be produced during the reaction?

1.5 Theoretical Yield and Actual Yield as Percentage

Yield: The yield of a chemical reaction is the amount of product that is produced from given amount of reactants.

When we perform stoichiometric calculations, we are trying to find out the theoretical yield based on the amount of reactants available.

Theoretical yield: The amount of the product calculated on the basis of a balanced chemical equation is called theoretical yield of the reaction. Actual yields can only be determined by performing the experiment and measuring the final mass of the product.

It is the calculated quantity of the product that we should get, using the starting amounts of the reactants. However, in actual practice, you will often get a less amount of the yield (product) than the theoretical yield. This is called actual yield of the reaction.

Actual yield: The actual yield is the amount of product that is actually produced during a chemical reaction by performing an experiment. Other names, used for actual yield, are experimental yield or practical yield.

The actual yield of a reaction is always less than the theoretical yield. There are number of reasons. Few are given below.

- The reaction may have not gone to completion, i.e. all the amount of the reactants may not have converted into products due to reversibility of the reaction.
- Some of the amounts of the reactants may have converted into some other products by a side reaction or chain reaction.
- It may be difficult to recover all of the products from the reaction medium and some of the products might be lost during recovery. This is called mechanical loss of the product (Human error).
- Reaction conditions (like temperature, pressure, pH etc) might have been disturbed.

Example 1.14

When 1.92 g of magnesium was heated in excess of oxygen. Calculate the theoretical yield of magnesium oxide (MgO).

Solution:



Molar mass of Mg = 24g/mol

$$\text{Number of moles of Mg} = \frac{\text{Mass of Mg}}{\text{Molar mass of Mg}} = \frac{1.92}{24} = 0.08 \text{ mol}$$

According to above equation,

$$2 \text{ mol of Mg} \cong 2 \text{ mol of MgO}$$

$$\text{So, } 0.08 \text{ moles of Mg} \cong 0.08 \text{ mol of MgO}$$

$$\text{Molar mass of MgO} = 24 + 16 = 40 \text{ g/mol}$$

$$\text{Theoretical yield of MgO} = \text{Number of moles of MgO} \times \text{molar mass of MgO}$$

$$\text{Theoretical yield of MgO} = 0.08 \times 40 = 3.2\text{g}$$

Percent Yield

The efficiency of a chemical reaction is determined with the help of percent yield, which is actually a comparison of actual yield and theoretical yield.

Percent yield can be calculated using the following expression:

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

Example 1.15

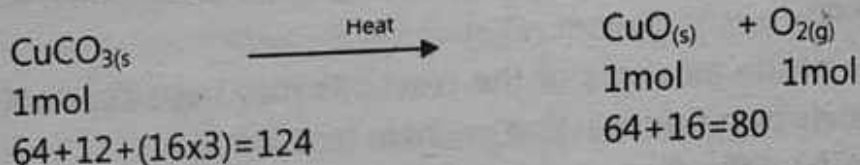
Heating 24.8g of copper carbonate (CuCO_3) in a crucible produced only 13.9g of copper oxide (CuO). What is the percentage yield of copper oxide?

Solution

The actual yield of CuO = 13.9g

Theoretical yield is calculated from the balanced chemical equation

which is:



$$\text{Moles of CuCO}_3 = \frac{24.8}{124} = 0.2$$

According to balanced chemical equation,

1 mol of $\text{CuCO}_3 \cong 1$ mol of CuO

0.2 mol of $\text{CuCO}_3 \cong 0.2$ mol of CuO

So,

$$\text{Mass of CuO} = \text{Theoretical yield} = 0.2 \times 80 = 16.0\text{g}$$

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

$$\text{Percentage Yield of Copper oxide} = \frac{13.9}{16.0} \times 100$$

$$\text{Percentage Yield of copper oxide} = 86.87\%$$

Practice Problem 1.11

If 1.274 g of copper sulphate (CuSO_4) reacts with excess of zinc metal, 0.392 g copper metal was obtained during the chemical reaction. Calculate the percentage yield of Cu metal.

**Self-Assessment**

1. Show the stepwise sequence that how mass (reactants) into mass (products) can be converted.
2. Define the molar volume of the gas. Convert 5dm^3 into cubic centimetre (cm^3).
3. Define limiting reagent.
4. What steps are used for identifying the limiting reagent?
5. What is percentage yield?
6. Why actual yield of a reaction is always less than the theoretical yield?

KEY POINTS

- Stoichiometry is quantitative chemistry, which deals with the calculations based on balanced chemical equations.
- Mole is the atomic mass, molecular mass or formula mass of an element, molecule or ionic compound, respectively, expressed in grams.
- Avogadro's number is the number of particles in one mole of substance. It is a constant number, equal to 6.023×10^{23} .
- The idea of mole can be applied to chemical equations, to calculate the amounts of reactants consumed or that of the products formed.
- Stoichiometric problems can be solved, using mole- mole, mole – mass and mass – mass conversions.
- When gases are involved, the concept of molar volume of the gas at STP (22.4dm^3) is helpful to calculate the number of moles and volumes.
- Percentage composition of a compound is the number of grams of each element per 100g of the compound.
- Limiting reagent in a chemical reaction is the one which is consumed earlier or which gives least amount of product from a balanced chemical equation.
- Excess reagent is the one, which is left behind unreacted, after the reaction is over.
- Theoretical yield is the amount of the product calculated on the basis of balanced chemical equation.
- Actual yield is the amount of product that is actually produced experimentally during a chemical reaction.
- The percent yield of a reaction tells us what percentage of the possible amount of product was actually obtained.
- Chemists use percentage yield to express the efficiency of a chemical reaction.

EXERCISE

Choose the Correct Option.

- The mass of an atom(element) compared with the mass of one atom of C - 12 is called,
 - One mole
 - Gram atomic mass
 - Atomic number
 - Relative atomic mass
- Which of the following is not true for a mole?
 - It is a counting unit
 - It is the gram atomic or gram formula mass of a substance
 - It contains 6.023×10^{23} particles
 - It contains different number of particles for different substances
- What is the mass (in grams) of 5 moles of water (H_2O)?
 - 90g
 - 36g
 - 18g
 - 100g
- The number of molecules in 22g of CO_2 is
 - 6.023×10^{23}
 - 3.011×10^{23}
 - 6.023×10^{21}
 - 6.023×10^{22}
- What will be the values of temperature and pressure at standard condition (STP),
 - $100^\circ C$, 1 atm
 - 298K, 1atm
 - 273K, 760mm Hg
 - $0^\circ C$, 760cm Hg
- The molar volume of SO_2 gas at STP is,
 - $64 dm^3$
 - $24 dm^3$
 - $22.4 dm^3$
 - $22.4 cm^3$
- The percentage of Ca in $CaCO_3$ is,
 - 12%
 - 10%
 - 48%
 - 40%
- Given the equation: $CO_{2(g)} + C_{(s)} \longrightarrow 2CO_{(g)}$

Which of the following equivalences is not correct for the reaction,

 - $1 \text{ mol } CO_2 \cong 2 \text{ mol } CO$
 - $1 \text{ mol } C \cong 56 \text{ g } CO$
 - $44 \text{ g } CO_2 \cong 28 \text{ g } CO$
 - $44 \text{ g } CO_2 \cong 12 \text{ g } C$
- A limiting reactant is one
 - Which is present in maximum amount
 - Which produces minimum No. of moles of product
 - Which produces maximum No. of moles of product
 - Does not affect the amount of product

- ### Short Questions

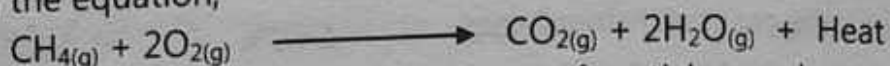
- ### Numerical Questions

- $$2\text{Mg}_{(s)} + \text{CO}_{2(g)} \longrightarrow 2\text{MgO}_{(s)} + \text{C}_{(s)} \quad \text{Ans: 2.79g}$$

Descriptive Questions

1. (a) Define and explain mole and Avogadro's number with examples.

(b) Given the equation,



How can this equation be read in terms of particles, moles and masses?

- (c) Define molar volume of a gas. What will be the volume of 60g of NH_3 at STP? (Atomic mass of N = 14, H = 1) Ans: 79.05 dm³

2. (a) What do you mean by percentage composition of a compound? How the percentage of an element is calculated in a compound.

(b) Calculate the percentage composition of each element in the following compounds. (Atomic mass of Al = 27, K = 39, Mg = 24, Mn = 55, Na = 23, O = 16, S = 32)

(i) MgSO_4

(ii) KMnO_4

(iii) $\text{NaAl}(\text{SO}_4)_2$

Ans:

(i) % Mg = 20, % S = 26.66, % O = 53.33

(ii) % K = 24.68, % Mn = 34.81, % O = 40.50

(iii) % Na = 9.50, % Al = 11.15, % S = 26.44, % O = 52.89

3. When steam is passed through red-hot carbon (a mixture of H_2 and CO gas) called water gas, is formed.



- (a) Which is the limiting reagent if 24.5g of carbon is mixed with 1.89 moles of water vapours?
- (b) Calculate the amount (in grams) of the excess reagent left unreacted. (Atomic mass of C = 12, O = 16, H = 1)

Ans: (a) Water is the limiting reagent. (b) 1.812g

4. (a) Calculate the percentage yield if 6.53g of hydrogen gas is produced when 5 mole of zinc is consumed in the reaction. (Atomic mass of Zn = 65.38, Cl = 35.5, H = 1)



Ans: 65.3%

- (b) The percentage yield of the following reaction is 85%.



How many grams of $\text{AlCl}_3(\text{s})$ will be actually obtained from 100g of aluminium metal? (Atomic mass of Al = 27, Cl = 35.5) Ans: 420.19g

(c) Given the equation,



at STP. How many moles of NH_3 would be formed if 6.3dm^3 of N_2 gas react with an excess of H_2 ? (Atomic mass of $\text{N}=14$, $\text{H}=1$) Ans: 0.56 mol

5. Given the equation,



(a) How many moles of water will be obtained by burning 5.6 moles of O_2 in an excess of H_2 ?

(b) How many moles of O_2 would be needed to react with 58.5g of H_2 to form water?

(c) How many grams of H_2 would be needed to form 120g of H_2O ?

(Atomic mass of $\text{O}=16$, $\text{H}=1$)

Ans: (a) 11.2 mol, (b) 14.625 mol, (c) 13.33g

Project

- Take a glass and add 250cm^3 of water to it. Find out the number of water molecules present in it.
- Take another glass and add 250cm^3 of water to it. Then add 200g of sugar into it. Now find out the number of sugar molecules in the given sample. Also, find out the total number of molecules present in sugar solution (water + sugar) in the glass.