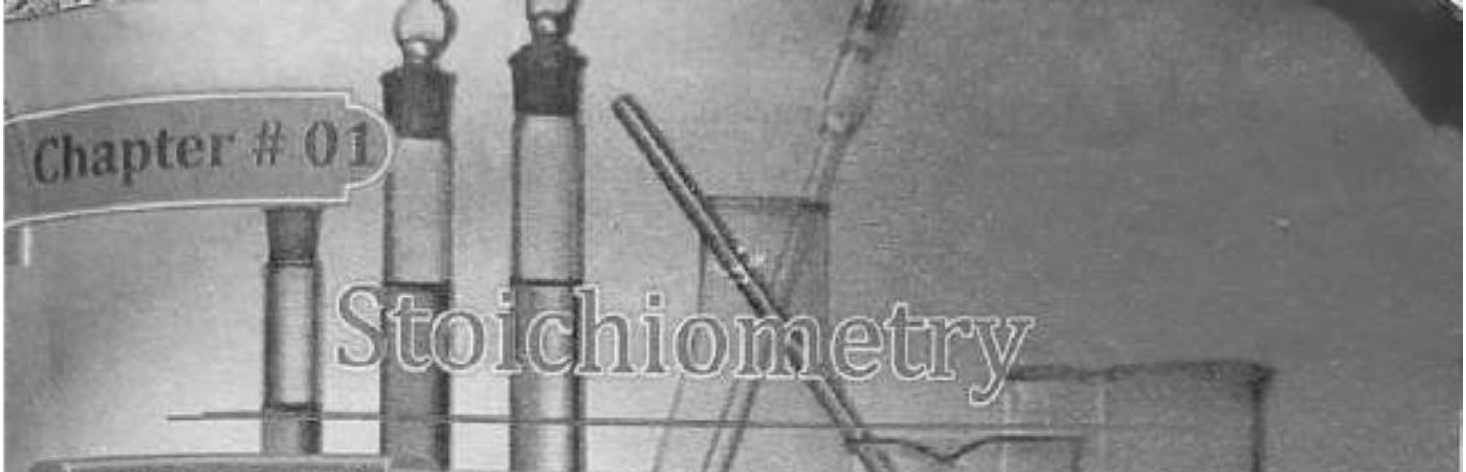


Chapter # 01

Stoichiometry

Major Concepts

1.1. Mole and Avogadro's number
1.2. Mole Calculations
1.3. Percentage Composition
1.4. Excess and Limiting Reagents
1.5. Percentage Yield

Learning Outcomes

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 theoretical yield, actual yield and 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)

Introduction:

Stoichiometry (pronounced as stoy-key-om-eh-tree) is the branch of Chemistry in which we study the relationship between the amounts of reactants and products in a balanced chemical equation.

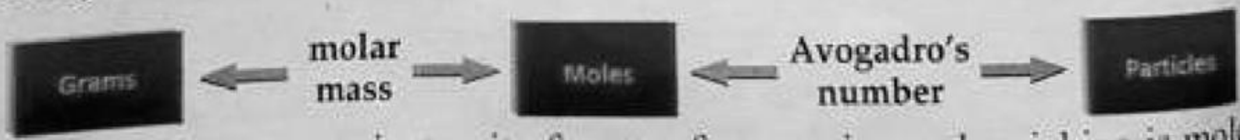
> **Keep in Mind**
>
> A balanced chemical equation has the same number of atoms of each type on both sides of the equation. While writing chemical equations, use the correct formulas for reactants and products and never change the subscripts in the formulas. If you change the subscripts, you change the identity of substances, so we have to balance the equation by changing the number of molecules of each type that appear in the equation.

Stoichiometry (Greek stoicheion, "element" and metron, "measure") means quantitative measure (of reactants and products). Such knowledge plays an important role when calculating the amount of products such as masses, moles, volumes and percentage yield etc. with the help of balanced chemical equation.

Balanced chemical equations have definite ratios of reactants and products just as compounds have definite ratios of elements. Those ratios are used to calculate the mass (or moles) of other substances in a reaction from the mass (or moles) of any one of the substances. It tells you that how to calculate the quantities of substances involved in a reaction. It covers the calculation of the percentage yield of a product from the actual yield and the theoretical yield, based on the amounts of reactants.

1.1 Mole and Avogadro`s Number:

You buy quantity of items in several ways like eggs in dozen (12 numbers), Shoes in pairs (2 numbers), cans in case (24 cans), playing cards in pack (52 numbers), papers in ream (500 sheets), and pencils in gross (144 numbers). Bulk foods, like rice, wheat, sugar, and peanuts are usually purchased by mass, because it is very difficult to count them. All these methods are used by chemists to determine the quantity of matter (counting and weighing).



The most convenient unit of matter for counting and weighing is mole. It connects the macroscale world (so large that can be weighed or count) to the nanoscale world (so small that is inconvenient to weigh or count them).

1.1.1 The Mole:

Mole is a Latin word; it means a 'huge mass'. Its symbol is **mol** and is represented by **n**. One mole is the amount of substance that has as many particles (atom, molecules, ions or formula units) as the number of atoms in exactly 12 g of carbon-12. The mass of one mole of a substance(element or compound) depends on what that substance is, and is equal to the formula mass of that substance in grams. We, therefore, say that the atomic mass, molecular mass, ionic mass or formula mass of a substance when expressed in grams is equal to one mole. Examples are:

One mole of carbon atoms = 12g

One mole of CO_2 molecules = 44g

One mole of NaCl formula units = 58.5g

One mole of SO_4^{2-} ions = 96g

1.1.2 The Avogadro's Number:

The number of particles (atom, molecules, ions or formula units) present in one mole of a substance is called Avogadro's number. It is represent by N_A. The number of particles in one mole of a substance is 6.0221421×10^{23}, which we will usually round to 6.02×10^{23}. Scientists call this value Avogadro's number in the honour of the Italian scientist Amedeo Avogadro (1776-1856). The unit of Avogadro's number is read as either "per mole" or "inverse of mole". Examples are:

Interesting Information

If we were able to count atoms at the rate of 5 million per second, it would take about 4 billion years to count the atoms in one mole.

One mole of carbon-12 contains 6.02×10^{23} atoms of carbon-12.

One mole of H_2O contains 6.02×10^{23} molecules of H_2O.

One mole of CaO contains 6.02×10^{23} formula units of CaO.

One mole of CO_3^{2-} contains 6.02×10^{23} ions of CO_3^{2-}.

Table 1.1: The Formula Masses, Molar Masses and Number of Particles of Some Substances

Name of Substance	Formula	Formula Mass (amu)	Mass of one Mole (g/mol)	Number of Particles in one Mole
Oxygen atom	O	16	16	6.02×10^{23} atoms
Oxygen molecule	O_2	32	32	6.02×10^{23} molecules
Water	H_2O	18	18	6.02×10^{23} molecules
Potassium nitrate	KNO_3	101	101	6.02×10^{23} formula units
Carbonate ion	CO_3^{-2}	60	60	6.02×10^{23} ions

Molar Mass and Volume

The volume of one mole of an ideal gas at STP (standard temperature and pressure) is called molar volume.

Its value is equal to $22.414 dm^3$. It is denoted by V_m.

Table 1.2: The Molar Volumes, Molar Masses and Number of Molecules of Some Gases at STP ($0^\circ C$ and 1atm)

Symbol Name	Hydrogen	Methane	Ammonia	Carbon dioxide
Molar mass (g/mol)	2	16	17	44
Number of moles	1 mole	1 mole	1 mole	1 mole
Volume of gas (in dm^3)	22.414	22.414	22.414	22.414
Number of molecules	6.02×10^{23} molecules	6.02×10^{23} molecules	6.02×10^{23} molecules	6.02×10^{23} molecules

The equal volume of all gases at STP has equal number of molecules but they have different masses.

The mass of one mole of a substance in grams is called molar mass.

The unit of molar mass is g/mol.

To determine molar mass, we change the units from atomic mass units to grams of a substance. The substance may be an element or a compound.

For an element, the molar mass in grams per mole is numerically equal to the atomic mass of that element in atomic mass units. For example:

i) Hydrogen atom has an atomic mass of 1.008 amu, so the molar mass of hydrogen atom is 1.008 g/mol and contains 6.02×10^{23} atoms of hydrogen.

ii) Oxygen atom has an atomic mass of 16 amu, so the molar mass of oxygen is 16 g/mol and contains 6.02×10^{23} atoms of oxygen.

If we know the atomic mass of an element, we also know its molar mass.

For a compound, the molar mass in grams per mole is numerically equal to the formula mass of that compound in atomic mass units. For example:

i) Ammonia (NH_3) has a formula mass of 17 amu, so the molar mass of ammonia is 17g/mol and contains 6.02×10^{23} molecules of ammonia.

ii) Potassium nitrate (KNO_3) has a formula mass of 101 amu, so the molar mass of potassium nitrate is 101g/mol and contains 6.02×10^{23} formula units of potassium nitrate.

Example 1.1

Calculate the molar mass of Benzene (C_6H_6).

Solution:

Each mole of C_6H_6 contains six moles of carbon atoms and six moles of hydrogen atoms in the formula.

The molar mass of benzene can be calculated as:

$6 \text{ moles of carbon} \times 12g/mol = 72g$

$6 \text{ moles of hydrogen} \times 1g/mol = 6g$

The molar mass of Benzene(C_6H_6) = 78g

Practice Exercise 1:

Calculate the molar mass of $KMnO_4$.

1.2 Mole Calculations

1.2.1 Calculation of Number of Moles of a Substance

i) If we know the mass of a substance, we can calculate the number of moles by dividing the mass of a substance by molar mass.

$$\text{No. of moles of a substance} = \frac{\text{Given mass of a substance}}{\text{Molar mass of a substance}}$$

Example 1.2

How many moles are present in 20g of NaOH?

Solution:

Mass of NaOH = 20g

Moles of NaOH = ?

$$\text{Moles of NaOH} = \frac{\text{Mass of NaOH}}{\text{Molar mass of NaOH}}$$

$$\text{Moles of NaOH} = \frac{20g}{40gmol^{-1}} = 0.5mol$$

Practice Exercise 2:

The given mass of $KClO_3$ is 12.25g. Calculate the number of moles of potassium chlorate.

ii) If we know the number of moles of a substance, we can calculate the mas by multiplying number of moles with molar mass.

Mass of a substance = Number of moles of a substance × Molar Mass

Example 1.3

Calculate mass of 0.25 moles of H_2SO_4.

Solution:

Moles of H_2SO_4 = 0.25mol

Mass of H_2SO_4 = ?

Mass of H_2SO_4 = Moles of H_2SO_4 × Molar mass of H_2SO_4

$= 0.25 mol \times 98 g mol^{-1}$

$= 24.5 g$

Practice Exercise 3:

What is the mass of 1.50 moles of $Ca(OH)_2$?

iii) If we know about the number of particles of a substance, we can calcula the number of moles by dividing the number of particles by Avogadro number.

$$\text{No. of moles of a substance} = \frac{\text{No. of particles of a substance}}{\text{Avogadro's Number}}$$

Example 1.4

Sucrose, table sugar, contains 3.76×10^{24} molecules of $C_{12}H_{22}O_{11}$. What is the number of moles of sucrose?

Solution:

$$\text{Number of moles of sucrose} = \frac{\text{No. of particles of sucrose}}{\text{Avogadro's Number}}$$

$$\text{Number of moles of sucrose} = \frac{3.76 \times 10^{24}\ \text{molecules}}{6.02 \times 10^{23}\ \text{molecules mol}^{-1}}$$

$= 6.25$ moles

Practice Exercise 4:

An aluminum wire contains 5.5×10^{25} atoms of aluminum. Calculate the number of moles of aluminum.

1.2.2 Calculation of Number of Particles

i) If we know the moles of a substance, we can calculate the number of particles by multiplying number of moles with Avogadro's number.

No. of particles of substance = No. of moles of substance × Avogadro's Number (N_A)

Example 1.5

How many atoms are there in a sodium metal that contains 2.5 moles?

Solution:

Number of moles of sodium = 2.5 mol

Number of atoms of sodium = ?

Number of atoms of sodium = Number of moles of sodium × N_A

$= 2.5 \text{ mol} \times 6.02 \times 10^{23} \text{ atoms mol}^{-1}$

$= 15.05 \times 10^{23}$ atoms

$= 1.505 \times 10^{24}$ atoms

Practice Exercise 5:

How many molecules are present in 2.50 moles of H_2O_2?

ii) If we know the mass of a substance, we can calculate the number of particles by dividing mass of substance by molar mass and the answer is multiplied by Avogadro's number.

$$\text{No. of particles of substance} = \frac{\text{Given mass of substance (m)}}{\text{Molar mass of substance (M)}} \times \text{Avogadro's Number}(N_A)$$

Example 1.6

How many atoms are present in 50 g of gold ring?

Solution:

Mass of gold = 50g

Molar mass of gold = 197g/mol

Number of atoms of gold = ?

$$\text{Number of atoms of gold} = \frac{\text{Mass of gold}}{\text{Molar mass of gold}} \times \text{Avogadro's Number}$$

$$= \frac{50\text{g}}{197\text{gmol}^{-1}} \times 6.02 \times 10^{23} \text{ atoms mol}^{-1}$$

$$= 0.254 \times 6.02 \times 10^{23} \text{ atoms}$$
$$= 1.529 \times 10^{23} \text{ atoms}$$

Practice Exercise 6:

Calculate number of water molecules in one cup of water having mass equal to 200g.

1.2.3 Calculation of Volume of a Gas

i) If we know the number of moles of a gas, we can calculate the volume of a g at STP by multiplying number of moles of a gas by molar volume.

$$\text{Volume of a gas} = \text{Moles of a gas} \times \text{Molar volume}$$

Example 1.7

What is the volume in dm^3 of 4.75mol of ethane gas?

Solution:

Number of moles of ethane = 4.75 mol

Volume of ethane in dm^3 = ?

Volume of ethane in dm^3 = Moles of ethane × Molar volume

$$= 4.75 \text{ mol} \times 22.414 dm^3 mol^{-1}$$
$$= 106.47 dm^3$$

Practice Exercise 7:

Calculate the volume in decimeter cube of 2.25 moles of laughing gas (N_2O).

ii) If we know the mass of a gas, we can calculate the volume of a gas at ST dividing the mass of a gas by molar mass and the answer is multiplied molar volume?

$$\text{Volume of a gas} = \frac{\text{Mass of a gas}}{\text{Molar mass of a gas}} \times \text{Molar volume}$$

Example 1.8

Argon is placed in a light bulb to minimize the rate of evaporation of the tungst filament. What is the volume in cm^3 of 3.99g of argon?

Solution:

Mass of argon $= 3.99g$

Volume in cm^3 of argon $= ?$

$$\text{Volume of argon} = \frac{\text{Mass of argon}}{\text{Molar mass of argon}} \times \text{Molar volume}$$

$$= \frac{3.99g}{39.9gmol^{-1}} \times 22.414dm^3mol^{-1}$$

$$= 0.1 \times 22.414dm^3$$

$$= 2.214\,dm^3$$

$$\text{Volume of argon in } cm^3 = 2.214dm^3 \times 1000cm^3dm^{-3}$$

$$= 2214cm^3$$

Practice Exercise 8:

Calculate the volume in decimeter cube of 50 g of propane (C_3H_8) gas.

1.2.4 Stoichiometric Calculations

A chemist needs to know how much product is obtained from certain amounts of reactants or a chemist needs to know how much reactants are used to get certain amounts of products. Chemists use stoichiometric calculations to answer these questions.

Principles of Stoichiometric Calculations

Stoichiometric calculations are based on the following principles:

i) Reactants are completely converted into products.

ii) No side reaction occurs.

iii) While doing calculations, the law of conservation of mass and the law of definite proportions are obeyed.

Law of Conservation of Mass

According to law of conservation of mass, matter (mass) can neither be created nor destroyed.

Law of conservation of mass states that: the total mass of reactants is equal to the total mass of products in a balanced chemical equation.

Law of Definite Proportions

According to law of definite proportions, a pure compound always contains the same element combined in the same ratio by mass.

(e.g.) Water has 88.89% oxygen and 11.11% hydrogen by mass, no matter what is its source.

Stoichiometric Relationship

The following types of relationship can be studied with the help of balanced chemical equations.

i) Mass-Mass Relationship

The mass of one substance can be calculated from the given mass of another substance and vice versa.

ii) Mole-Mole Relationship

The moles of one substance can be calculated from the given moles of another substance and vice versa.

iii) Volume-Volume Relationship

The volume of one substance can be calculated from the given volume of another substance and vice versa.

iv) Mole-Mass Relationship

The mass of one substance can be calculated from the given moles of another substance and vice versa.

v) Mass-Volume Relationship

The volume of one substance can be calculated from the given mass of another substance and vice versa.

How to Solve a Stoichiometry Problem

Step 1: Write all the given data with relevant units.

Step 2: Write a balanced chemical equation for the reaction under consideration.

Step 3: Convert the given mass of all the reactants to number of moles.

Step 4: Use the chemical equation to determine the mole ratio of the unknown quantity to the known quantity. The mole ratio can be determined by dividing the coefficient of the unknown by the coefficient of the known.

$$\frac{\text{Moles of unknown}}{\text{Moles of known}}$$

This mole ratio can be used to convert the number of moles of known quantity to a number of moles of unknown. The mole ratio is multiplied with the moles of known calculated in the third step.

$$\frac{\text{Moles of unknown}}{\text{Moles of known}} \times \text{Moles of known} = \text{Moles of unknown}$$

Step 5: Convert moles back to mass by multiplying with molar mass of unknown substance to get answer in grams, if necessary.

Example 1.9 (Mass-Mass Conversion)

Calculate the mass of quicklime (calcium oxide) that is produced by the thermal decomposition of 200g of limestone (calcium carbonate). The equation for this reaction is:

$$CaCO_{3(s)} \xrightarrow{\Delta} CaO_{(s)} + CO_{2(g)}$$

Solution:

Step 1: Write all the given data with relevant units.

Mass of CaO = ?

Mass of $CaCO_3$ = 200g

Molar mass of CaO = $56gmol^{-1}$

Molar mass of $CaCO_3$ = $100gmol^{-1}$

Step 2: Write a balanced chemical equation for the reaction.

$$\underset{1mol(100g)}{CaCO_{3(s)}} \xrightarrow{\Delta} \underset{1mol(56g)}{CaO_{(s)}} + CO_{2(g)}$$

Step 3: Use the chemical equation to determine the mass ratio of the unknown quantity to the known quantity.

100g of $CaCO_3$ produces CaO = 56g

1g of $CaCO_3$ produces CaO $= \frac{56g}{100g}$

200g of $CaCO_3$ produces CaO $= \frac{56g}{100g} \times 200g$

$= 112g$

So, the mass of quicklime produced is 112g.

Practice Exercise 9:

How much HCl can be produced when 5g of hydrogen reacts with an excess amount of chlorine? The equation for this reaction is:

$$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)}$$

Example 1.10 (Mole-Mole Conversion)

When 2.5 mol of nitrogen reacts with hydrogen to form ammonia, how many moles of hydrogen are consumed in the process? The equation for this reaction is:

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

Solution:

Step 1: Write all the given data with relevant units.

Number of moles of N_2 = 2.5mol

Number of moles of H_2 = ?

Step 2: Write a balanced chemical equation for the reaction.

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

Step 3: Use the chemical equation to determine the mole ratio of the unknown quantity to the known quantity.

1 mole of N_2 needs H_2 to produce NH_3 $= \frac{3}{1} = 3\text{mol}$

2.5 moles of N_2 needs H_2 to produce $NH_3 = 3 \times 2.5 = 7.5\text{mol}$

So, the moles of hydrogen used in the process are 7.5 moles.

Practice Exercise 10:

How many moles of carbon dioxide are produced when 1.25 moles of glucose are used by a person? The oxygen is in excess. The equation for the reaction is:

$$C_6H_{12}O_{6(s)} + 6O_{2(g)} \longrightarrow 6CO_{2(g)} + 6H_2O_{(l)}$$

Example 1.11 (Mole-Mass Conversion)

What mass of hydrogen can be produced by the decomposition of 4.3 moles of water? The balanced chemical equation for the reaction is:

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

Step 1: Write all the given data with relevant units.

Mass of hydrogen = ?

Moles of water = 4.3mol

Molar Mass of H_2 = 2g mol^{-1}

Step 2: Write a balanced chemical equation for the reaction.

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

Step 3: Use the chemical equation to determine the mole ratio of the unknown quantity to the known quantity.

2 moles of H_2O produces H_2 = 2mol

1 mol of H_2O produces H_2 = $\frac{2}{2}$ = 1mol

4.3 moles of H_2O produces H_2 = $1 \times 4.3 = 4.3$mol

Step 4: Convert moles back to mass of the unknown.

Mass of H_2 = Moles of H_2 × Molar mass of H_2

= 4.3mol × 2g mol^{-1}

= 8.6g

So, the mass of hydrogen produced by the decomposition of water is 8.6g.

Practice Exercise 11:

Calculate the mass of sodium hypochlorite (NaOCl), household bleach, produced by the reaction of 1.75 moles of chlorine with excess sodium hydroxide. The balanced equation is:

$$2\,NaOH_{(aq)} + Cl_{2(g)} \longrightarrow NaOCl_{(aq)} + NaCl_{(aq)} + H_2O_{(l)}$$

Example 1.12 (Mass-Volume Conversion)

Nitrogen peroxide (NO_2) is a reddish brown gas and is used in the manufacture of nitric acid. It can be prepared by the oxidation of nitric oxide (NO):

$$2\,NO_{(g)} + O_{2(g)} \longrightarrow 2\,NO_{2(g)}$$

Determine the volume (in litre) of NO_2 produced by 12 g of NO.

Solution:

Step 1: Write all the given data with relevant units.

Volume of NO_2 = ?

Mass of NO = 12g

Molar mass of NO = 30gmol^{-1}

Step 2: Write a balanced chemical equation for the reaction.

$$2\,NO_{(g)} + O_{2(g)} \longrightarrow 2\,NO_{2(g)}$$

Step 3: Convert the given mass of the reactant to moles.

$$\text{Moles of NO} = \frac{\text{Mass of NO}}{\text{Molar mass of NO}} = \frac{12g}{30g\ mol^{-1}} = 0.4\ mol$$

Step 4: Use the chemical equation to determine the mole ratio of the unknown quantity to the known quantity.

2 mole of NO produces NO_2 $= 2mol$

1 mole of NO produces NO_2 $= \frac{2}{2} = 1mol$

0.4 mole of NO produces NO_2 $= 1 \times 0.4 = 0.4mol$

Step 5: Convert moles to volume of the unknown.

Volume of NO_2 = Mole of NO_2 × Molar volume of NO_2

$= 0.4mol \times 22.414 L\,mol^{-1}$

$= 8.97L$

So, the volume of NO_2 produced by the oxidation of NO is 8.97 L.

Practice Exercise 12:

Methane gas is used as a domestic fuel in the form of natural gas and in the manufacture of urea fertilizer on commercial scale. On combustion, methane gas produces CO_2 and H_2O. Write balanced chemical equation for the reaction. What volume of CO_2 gas is produced when 0.5Kg of methane is burnt in excess oxygen?

1.3 Percentage Composition

The percentage composition is the number of parts by mass of an element in to 100 parts by mass of a compound. Each symbol in the formula of a compound represents the mass of one mole of atoms of the elements. The formula shows mass of one mole of the compound. The chemical formula provides information about the composition of a compound in terms of moles. For example, 1 mole of sodium carbonate, Na_2CO_3 contains 2 moles of sodium, 1 mole of carbon and 3 moles of oxygen atoms. The percentage composition of sodium carbonate tells us the relative masses of Na, C, and O atoms that are present in the compound.

A two-step process is required to calculate the percentage composition of the compound:

i) Calculate the molar mass of a compound.

ii) Divide the total mass of each element in one mole of the compound by the molar mass of the compound and the answer is multiplied by 100.

Mathematically,

$$\text{Percentage composition of an element} = \frac{x\,(\text{Molar mass of an element})}{\text{Molar mass of a compound}} \times 100$$

Where, x is the number of moles of the element present in one mole of the compound.

Example 1.13

What is the percentage composition of each element in sodium carbonate?

Solution:

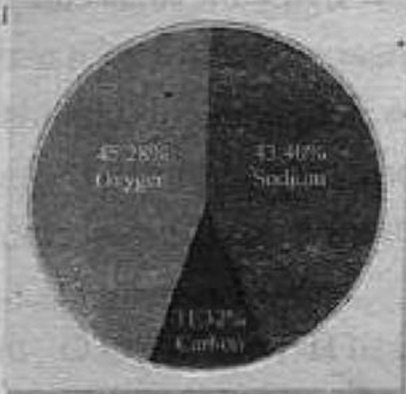
The chemical formula of sodium carbonate is Na_2CO_3. It contains 2 moles of sodium, 1 mole of carbon, and 3 moles of oxygen.

Molar mass of $Na_2CO_3 = 23\times2+1\times12+16\times3 = 106gmol^{-1}$

$$\%\text{age of Na} = \frac{2(23gmol^{-1})Na}{106gmol^{-1}} \times 100 = 43.40\%$$

$$\%\text{age of C} = \frac{1(12gmol^{-1})C}{106gmol^{-1}} \times 100 = 11.32\%$$

$$\%\text{age of O} = \frac{3(16gmol^{-1})O}{106gmol^{-1}} \times 100 = 45.28\%$$



Remember:

The sum of percentages of all the elements present in the formula should be equal to $100\% \pm 0.02\%$.

Practice Exercise 13:

Calculate the percentage of nitrogen in ammonia (NH_3) and nitric acid (HNO_3).

1.4 Excess and Limiting Reagents

In many chemical processes, the quantities of the reactants used are usually not present in the proportions indicated by the balanced chemical equation. Because the main objective of the reaction is:

i) To produce maximum amount of product, frequently a large amount of inexpensive reactant is supplied to ensure that whole of the mass of expensive reactant is completely converted to the desired product.

ii) To increase the rate of reaction.

At the end of the reaction only one of the reactants may be consumed completely while the other reactants will remain unreacted. The reactant that is completely consumed at the completion of reaction is called limiting reagent (or limiting reactant). The maximum amount of product formed depends upon the amount of limiting reactant in the reaction mixture. When this reactant is consumed completely, the reaction stops and no further products are formed. The reactants which are in larger amounts and remain unreacted at the end of the reaction are called "excess reagents" (or excess reactants). The concept of the limiting reagent is analogous to the relationship between the numbers of frames and wheels to make bicycles. Suppose you want to make some bicycles. Each is made from one bicycle frame and two bicycle wheels. You have 70 bicycle frames and 100 bicycle wheels. How many bicycles can you make? The answer is 50. When you run out of bicycle wheels you must stop making bicycles. Cycle wheels are the "limiting reagent" in the language of chemistry, because they limit the number of bicycles. The component (wheels) which produces the fewer number of bicycles is the limiting component while the component (frames) left behind is the excess component.

Now consider a chemical reaction between hydrogen and oxygen. If we react 6g of H_2 with 32g of O_2, then we will get 36g of water as:

$$\underset{6g}{2H_2} + \underset{32g}{O_2} \longrightarrow \underset{36g}{2H_2O}$$

When the reaction between hydrogen and oxygen proceeds to completion, 32g of oxygen consumes completely and the reaction stops and no more products are formed. In this reaction, formation of water is limited by oxygen (O_2) and hydrogen (H_2) is in excess. Out of 6g of hydrogen, 2g remains unreacted. Here H_2 is an excess reactant and O_2 is limiting reactant.

Identification of Limiting Reagent

You can determine the limiting reagent with the help of the following steps:

i) Write the balanced chemical equation.

ii) Convert the given mass of all the reactants to moles.

iii) Calculate the amount of product (in moles or grams, as required) from each reactant with the help of balanced chemical equation.

iv) The reactant that gives the least amount of product is the limiting reagent.

v) Calculate the amount of product formed by limiting reactant.

vi) The reactant that is left over after the completion of reaction is excess reagent.

vii) If you want to find the amount of excess reagent, then subtract the amount used from the starting quantity of the reactant.

Example 1.14

Natural gas consists primarily of methane (CH_4). The complete combustion of methane (CH_4) gives carbon dioxide (CO_2) and water.

$$CH_{4(g)} + 2O_{2(g)} \xrightarrow{\text{Ignition}} CO_{2(g)} + 2H_2O_{(g)}$$

a) How many grams of CO_2 can be produced when 30g of CH_4 and 50g of O_2 are allowed to combine?

b) How many grams of excess reagent are left unreacted after the completion of reaction?

Solution (a):

Mass of methane (CH_4) = 30g
Mass of oxygen (O_2) = 50g
Mass of CO_2 in grams = ?
Mass of excess reactant left behind in grams = ?

Step 1: Write balanced chemical equation:

$$CH_{4(g)} + 2O_{2(g)} \xrightarrow{\text{Ignition}} CO_{2(g)} + 2H_2O_{(g)}$$

Step 2: Convert the given mass of both the reactants in to their moles:

$$\text{Moles of } CH_4 = \frac{\text{given mass of } CH_4}{\text{molar mass of } CH_4} = \frac{30g}{16gmol^{-1}} = 1.875 \text{ mol}$$

$$\text{Moles of } O_2 = \frac{\text{given mass of } O_2}{\text{molar mass of } O_2} = \frac{50g}{32gmol^{-1}} = 1.563 \text{ mol}$$

Step 3: Calculate the number of moles of product from each reactant:

➤ Compare the number of moles of CH_4 with those of CO_2
From the balanced chemical equation we know:
1 mole of methane produces CO_2 = 1mol
1.875 mole of methane produces CO_2 = 1×1.875 mol
= 1.875 mol of CO_2

➤ Compare the number of moles of O_2 with those of CO_2
From the balanced chemical equation we know:
2 moles of oxygen produces CO_2 = 1 mol

1 moles of oxygen produces $CO_2 = \frac{1mol}{2mol}$

1.563 moles of oxygen produces $CO_2 = 0.5 \times 1.563$mol
= 0.7815 moles of CO_2

- From the above calculation, it is clear that the limiting reactant is O_2 because it produces fewer amounts (moles) of product (CO_2) than CH_4.
- Notice that the limiting reactant is not necessarily the reactant which is present in small amount.

Step 4: Now multiply the moles of CO_2 with its molar mass to get amount of carbon dioxide produced at the end of the reaction.

$$\text{Mass of } CO_2 \text{ in grams} = \text{Moles of } CO_2 \times \text{Molar mass of } CO_2$$
$$= 0.7815 \text{ moles} \times 44 \text{g mol}^{-1}$$
$$= 34.39\text{g}$$

(b)

Step 5: The quantity of limiting reactant can also be used to calculate the quantity of excess reactant used:

2 moles of O_2 reacts with moles of CH_4 = 1 mol

1 mole of O_2 reacts with moles of CH_4 $= \frac{1}{2}$ mol

1.563 moles of O_2 reacts with moles of $CH_4 = \frac{1}{2} \times 1.563\text{mol}$

$= 0.7815$ mol

Step 6: The mass of methane (excess reagent) is equal to the starting quantity minus the amount used during the reaction.

$$\text{Number of moles of } CH_4 = \text{Starting quantity} - \text{Quantity use}$$
$$= 1.875\text{mol} - 0.7815\text{mol}$$
$$= 1.0935 \text{ mol}$$

$$\text{Mass of } CH_4 \text{ (excess reagent)} = \text{Moles of } CH_4 \times \text{Molar mass}$$
$$= 1.0935 \text{ mol} \times 16\text{g mol}^{-1}$$
$$= 17.5\text{g}$$

Practice Exercise 14:

Which of the following reaction mixtures could produce the greatest amount of product when they combine according to the reaction given below?

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

a) 1 mole of N_2 and 3 moles of H_2
b) 2 mole of N_2 and 3 moles of H_2

c) 1 mole of N_2 and 5 moles of H_2
d) 3 mole of N_2 and 3 moles of H_2
e) Each produce the same amount of product

Society, Technology and Science

Chemistry is a Quantitative Science

- Stoichiometry is very important in medical sciences and is used to:
  - Determine the glucose level in the blood of diabetics.
  - Determine the steroid and other stimulants in the urine of athletes. Athletes use steroids and other stimulants to enhance performance and increase strength.
  - Determine the cholesterol level in the blood of patients. Cholesterol is a form of fat that's not all bad. But cholesterol can have harmful effects.
- It is helpful in determining the amount of drugs to give a patient. The medicine has no effect when given in small amounts and can cause toxic state or death when given in large amounts. For example, paracetamol is used as a pain killer and to decrease fever. An overdose may result a blood thinning, organ damage and severe liver damage.
- It is the stoichiometry that enables the pilots to determine the distance that a plane will travel before needing to be refueled.

1.5 Percentage Yield

The amount of product either calculated from balanced chemical equation or actually obtained from a reaction is called yield or chemical yield.

The amount of product calculated from balanced chemical equation is called theoretical yield or expected yield while the amount of product actually obtained from a chemical reaction is called actual yield or practical yield. If the actual yield is very low, the final cost can be very high.

In most chemical reactions, the amount of product obtained (actual yield) is always less than theoretical yield due to following reasons:

i) In some reactions the products formed may react further among themselves or with the reactants (Side reactions take place) that give products (by-products) other than the main product.

ii) Many reactions are reversible. They do not go to completion. In these reactions the products formed react to produce the original reactants.

iii) Impurities in the reactants. Suppose you want to prepare chlorine gas from 10g of NaCl that contains some impurities (the substances other than sodium chloride). So you do not know exactly how much pure sodium chloride you have. If you calculate the amount of product from equation, it should be greater than the amount of product actually obtained from a reaction. It means that there is not 10g of pure NaCl in the sample.

iv) Moreover, many reactions simply are not complete; some amount of reactant remains unreacted at the end of the reaction.

v) Loss of product during separation, filtration, washing, drying, distillation, crystallization etc.

For these and other reasons, it is useful to point out a difference between the theoretical and the actual yields of a chemical reaction and to calculate the percentage yield. Actual yield divided by theoretical yield and the answer is multiplied by 100 is called percent yield.

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

Significance of % age yield:

i) Percentage yield shows efficiency of reaction.

ii) Greater the % age yield, higher will be the efficiency of reaction and vice versa.

Example 1.15

Lithium on heating with nitrogen produces lithium nitride:

$$6Li + N_2 \longrightarrow 2Li_3N$$

When 30g of lithium reacts with an excess of nitrogen then how much lithium nitride is produced? If the actual yield of lithium nitride is 38g, what is the percent yield of the reaction?

Solution:

Given mass of Lithium = 30g

Actual yield of Li_3N = 38 g

Mass of lithium nitride produced = ?

Percent yield of the reaction = ?

Before going to calculate percent yield we must know about theoretical yield and actual yield.

i) Theoretical yield:

Equation:

	$6Li$	$+ N_2$	$\longrightarrow$	$2Li_3N$
	6 moles	Excess		2 moles
or	42g			70g

According to the above balanced chemical equation:

42g of lithium produces lithium nitride = 70g

1g of lithium produces lithium nitride $= \frac{70g}{42g}$

30g of lithium produces lithium nitride $= \frac{70g}{42g} \times 30g$

$= 50g$

So,

Theoretical yield of lithium nitride is 50g.

ii) Actual yield of lithium nitride

Actual yield of lithium nitride is 38g.

iii) Percent yield of lithium nitride

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

$$\%\text{ yield of lithium nitride} = \frac{38g}{50g} \times 100$$

$= 76\%$

Practice Exercise 15:

Iron sulphide is produced by heating iron with sulphur:

$$Fe_{(s)} + S_{(s)} \longrightarrow FeS_{(s)}$$

When 28kg of iron is combined with excess of sulphur, 40 kg of iron sulphide (FeS) is formed. Calculate the percentage yield of the FeS.

Summary of Facts and Concepts

- The study of quantities of materials consumed and produced in chemical reactions is called Stoichiometry.
- Calculations using balanced equations are called stoichiometric calculations.
- When you know the quantity of one substance in a reaction, you can calculate the quantity of any other substance consumed or created in a reaction.
- The amount of substance that has as many particles as the number of atoms in exactly 12 g of carbon-12 is called mole. Therefore, 12 g of carbon contains 6.02×10^{23} carbon atoms.
- One mole of atoms of any element contains 6.02×10^{23} atoms, regardless of the type of element.
- The number of particles in one mole of a substance is called Avogadro's number. Its value is equal to 6.02×10^{23}.

- The volume occupied by one mole of an ideal gas at STP is called molar volume and this amount is equal to 22.414 dm^3.
- The limiting reactant is the reactant that is consumed first in a chemical reaction and produces the smallest yield. When this reactant is consumed completely, the reaction stops and no further products are formed. The reactant that is left after the reaction has stopped is known as excess reagent.
- The amount of product which is calculated using the balanced equation is called theoretical yield. The amount of product obtained when the reaction takes place is called actual yield. The ratio of actual yield to theoretical yield is called percent yield. Percent yields may range from a fraction of 1 percent to 100 percent.

Questions and Problems

Q.1: Four answers are given for each question. Select the correct answer.

i) The branch of Chemistry which describes the relationship between the amounts of reactants and products in a balanced chemical equation is called:

(a) Physical chemistry (b) Biochemistry
(c) Stoichiometry (d) Organic chemistry

ii) What are the number of covalent bonds in 68g of H_2S gas?

(a) 3.01×10^{23} (b) 6.02×10^{23}
(c) 2.41×10^{24} (d) 24.1×0^{24}

iii) The mass of O_2 required to burn 0.1 mole of C_2H_5OH is:

(a) 32g (b) 3.2g
(c) 5.6g (d) 9.6g

Equation: $C_2H_5OH + 3O_2 \longrightarrow 2CO_2 + 3H_2O$

iv) The volume occupied by 1.4g of N_2 at STP is:

(a) $2.24dm^3$ (b) $22.4dm^3$
(c) $1.12dm^3$ (d) $112dm^3$

v) A beaker contains 9 g of water. The number of hydrogen atoms is:

(a) 6.02×10^{23} (b) 3.01×10^{23}
(c) 6.02×10^{24} (d) 3.01×10^{24}

vi) One mole of diamond chain and one mole of gold ring have same number of:

(a) neutrons (b) protons
(c) electrons (d) atoms

vii) The largest number of molecules are present in:

(a) 4.8g of C_2H_5OH (b) 3.6g of H_2O

(c) 2.8g of CO (d) 5.4g of N_2O_5

viii) Limiting reactant is that:

(a) Which remains unreacted

(b) Which gives maximum amount of product

(c) Which gives minimum amount of product

(d) Which has low-price

ix) The amount of product obtained practically is called:

(a) Expected yield (b) Theoretical yield

(c) Actual yield (d) fractional yield

x) The reactant which is in larger amount and remains unreacted is called:

(a) Limiting reactant (b) Excess reactant

(c) Expensive reactant (d) Restricting reactant

Q.2: Fill in the blanks with suitable words given in the brackets:

i) A balanced chemical equation has the __________ number of atoms of each element on both sides of the equation. (same / different)

ii) There are __________ molecules in two moles of CH_3OH. (6.02×10^{23} / 12.04×10^{23})

iii) The number of atoms in one mole of neon is __________. (6.02×10^{23} / 12.04×10^{23})

iv) The mass of __________ moles of N_2 is 56g. (one / two)

v) The space occupied by 0.50 moles of Cl_2 at STP is __________. ($11.207dm^3$ / $22.414dm^3$)

vi) The equal volume of all gases at STP has __________ number of molecules but they have __________ masses. (equal / different)

vii) The percentage of nitrogen in N_2O_4 is __________. (30.43% / 69.57%)

viii) 1 mole of Cu_2O has __________ atoms of copper and __________ atoms of oxygen. (6.02×10^{23} / 12.04×10^{23})

ix) Limiting reactant gives __________ amount of product. (Minimum / Maximum)

x) Actual yield is always __________ than theoretical yield. (less / more)

Q.3: Label the following statements as True or False.

i) Stoichiometry tells you that how to calculate the quantities of substances involved in a reaction.

ii) The stoichiometric calculations can be performed only when Avogadro's law is obeyed.

iii) One atom of Mg is twice in mass as compared to one carbon atom.

iv) The reactants are on the right side of arrow in a chemical equation.

v) Avogadro's number is represented by N_A.

vi) The number of hydrogen atoms in 1.5 moles of H_2S is equal to the number of hydrogen atoms in 1.5 moles of HI.

vii) The molar mass of PO_4^{-3} ion is 95g mol^{-1}.

viii) Ionic compounds consist of molecules.

ix) The amount of product calculated from balanced chemical equation is called actual yield.

x) Greater is the percentage yield; higher will be the efficiency of reaction.

Q.4: What is stoichiometry? Why is stoichiometry important? Give some examples.

Q.5: Give the principles and relationships of stoichiometric calculations.

Q.6: How can you solve a Stoichiometry Problem?

Q.7: a) Define and explain mole and Avogadro's number with examples.

b) Define and explain molar mass and molar volume with examples. How can we calculate the molar mass of a substance?

c) What does the mole have in common with the pair, the dozen and the gross?

Q.8: Explain the following:

a) What is the mass, in grams, of one mole of C^{12}? 12g

b) How many carbon atoms are present in a mole of C^{12}? 6.6

c) What is Avogadro's number, and how is it related to the mole?

d) Avogadro's number of atoms of different elements has different masses.

e) Why chemists use mole as a unit? There partical are very small that why they use a unit

f) Which would have a higher mass: a mole of Na atoms or a mole of C atoms?

g) Which would contain more atoms: a mole of Na atoms or a mole of C atoms? they are same

h) How many atoms are present in 1 molar mass of sulphuric acid?

Q.9. What is percentage composition? How can you calculate the percentage composition of a compound?

Q.10: a) What are limiting and excess reagents? How will you determine them?

b) Can there be a limiting reagent if only one reactant is present?

c) We use expensive reactants in small amounts and inexpensive in large amounts, why?

Q.11: Give an everyday example that illustrates the limiting reagent concept.

Q.12: Define theoretical, actual yield and percentage yield. How do we calculate the percentage yield of a chemical reaction?

Q.13: Actual yield of a reaction is always smaller than that of theoretical yield, why?

Q.14: How much iron is required to produce 162.3g of $FeCl_3$ when chlorine is in excess?

$$2Fe + 3Cl_2 \longrightarrow 2FeCl_3$$

Q.15: Calcium metal reacts with oxygen to form calcium oxide, CaO.

a) Write a balanced equation for the reaction.

b) How many grams of oxygen are required to react with 35g of Ca?

c) How many grams of CaO will be produced?

Q.16: Urea, a fertilizer, can be prepared in the laboratory by the combination of ammonia and carbon dioxide according to the following balanced equation:

$$2NH_{3(g)} + CO_{2(g)} \longrightarrow (NH_2)_2CO_{(aq)} + H_2O_{(l)}$$

(a) Calculate the number of moles of urea formed by the combination of 2.75 moles of ammonia.

(b) Calculate the number of moles of carbon dioxide needed to combine with 2.75 moles of ammonia.

Q.17: Oxygen gas can be prepared by the thermal decomposition of potassium chlorate ($KClO_3$) in the laboratory:

$$2KClO_{3(s)} \longrightarrow 2KCl_{(s)} + 3O_{2(g)}$$

How many grams of oxygen can be prepared from 5 moles of $KClO_3$?

Q.18: Acetylene gas (C_2H_2) is used for welding and for the artificial ripening of fruits. Acetylene gas reacts with hydrogen gas to form ethane gas (C_2H_6). The balanced chemical equation for the reaction is:

$$C_2H_{2(g)} + 2H_{2(g)} \longrightarrow C_2H_{6(g)}$$

If 10 grams of hydrogen gas reacts with excess acetylene, how many [
of ethane gas can be produced?

Q.19: Octane (C_8H_{18}) is a component of gasoline that burns according t
equation:

$$2C_8H_8 + 20O_2 \longrightarrow 16CO_2 + 8H_2O$$

How many moles of O_2 are needed to burn 2.5mol of C_8H_{18}?

Q.20: Calculate how many moles of oxygen are required to make 27(
aluminum oxide. The balanced equation is:

$$4Al_{(s)} + 3O_{2(g)} \longrightarrow 2Al_2O_{3(s)}$$

Q.21: What are the Molar Masses of the following Compounds?

i) NH_3 ii) H_2O
iii) H_2SO_4 iv) C_2H_5OH
v) NH_4Br

Q.22: How many moles are in each of the following samples?

i) 10g of Ca ii) 100g of Hg
iii) 30g of NaOH iv) 94.5g of HNO_3
v) 402g of $HClO_4$

Q.23: How many grams are in a mole of each of the following substances?

i) N_2 ii) I_2
iii) CaO iv) Kbr
v) $CoCl_2.6H_2O$

Q.24: How many moles are in 25g of each of the following substances?

i) F_2 ii) Ni
iii) CS_2 iv) $C_2F_2Cl_2$
v) $FeSO_4$

Q.25: How many particles are present in 2.5 moles of carbonate ion, CC
2×10^3 moles of aspirin ($C_9H_8O_4$)?

Q.26: What is the volume of 1.25 moles of SO_2 and 0.50 moles of argon gas

Q.27: Butane gas is a fuel and used in pocket lighters. Its molecular forn
C_4H_{10}. Calculate its percentage composition.

Q.28: Calculate the percentage of oxygen in each of the following:

i) SO_2 ii) SiO_2
iii) $(NH_2)_2CO$ iv) SO_4^{2-}
v) $C_6H_{12}O_6$

Q.29: Hydrogen and bromine react to produce hydrobromic acid:

$H_2 + Br_2 \longrightarrow 2HBr$

How many grams of HBr are formed from reaction of 4.50g of H_2 with 15.75g of Br_2? Which reactant is limiting?

Q.30: How much water can be made from 50g of hydrogen and 50g of oxygen by the reaction?

$2H_2 + O_2 \longrightarrow 2H_2O$

Q.31: When 30g of silica reacts with 12g of carbon, silicon carbide is produced; this is an important ceramic material.

$SiO_2 + 3C \longrightarrow SiC + 2CO$

(Silicon carbide)

Calculate which reactant is limiting and how much silicon carbide is produced.

Q.32: When 227g of NH_4NO_3 decomposes, how much N_2O and H_2O are formed?

$NH_4NO_3 \longrightarrow N_2O + 2H_2O$

Q.33: 4.64g of iron reacts completely with an excess of $CuSO_4$ and gives 5g of copper.

$Fe + CuSO_4 \longrightarrow FeSO_4 + Cu$

a) What is the theoretical yield of the reaction?

b) What is the percentage yield of the reaction?

Q.34: Hydrazine is used as fuel in rockets and can be prepared by the following reaction:

$ClNH_2 + 2NH_3 \longrightarrow N_2H_4 + NH_4Cl$

Hydrazine

145g of hydrazine is produced by the combination of 280g of $ClNH_2$ with an excess of NH_3. Calculate the percentage yield of this reaction.

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Ans:- Data: given mass of H_2 = 4.5g

given mass of Br_2 = 15.75 g

Molar mass of H_2 = 2 g mol⁻¹

Molar mass of Br_2 = 160 g mol⁻¹

Mass of HBr = ?

limiting reactant = ?

No of moles of H = 4.5/2 = 2.25 mole

No of moles of Br = given mass / Molar mass = 15.75/160 = 0.084 moles

$H_2 + Br_2 \rightarrow 2HBr$

moles of unknown / moles of known × given moles

2/1 × 2.25

2/1 × 2.25 = 4.5 moles HBr.

moles of unknown / moles of known × given moles

2/1 × 0.0984

2 × 0.0984

0.1968 HBr

∴ mass of HBr = No of moles of HBr × Molar mass of HBr

= 0.1968 × 81 = 15.9 g of HBr

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