SECTION

3 MOLE RATIOS, CHEMICAL FORMULAE AND CHEMICAL EQUATIONS

PHYSICAL CHEMISTRY

Matter and its properties

INTRODUCTION

In this section, you will master essential skills, including IUPAC nomenclature for inorganic compounds, writing compound formulas based on chemical laws, and balancing equations. You will also engage in stoichiometric calculations to deepen your understanding of chemical reactions.

At the end of this section, you should be able to:

- Use IUPAC nomenclature to name inorganic compounds, write the formulae of compounds based on the laws of chemical combination and write balanced chemical equations.
- Perform calculations involving stoichiometric relationship.

Key Ideas:

- ***Oxidation number is t****he charge of an ion in a compound.
- **Empirical formula** is the simplest ratio of atoms of each element in a molecule.
- **Molecular formula** is the actual number of atoms of each element in a molecule.
- **Structural formula** is a formula showing the arrangement of atoms in a molecule.
- **Prefix** indicates the number of atoms of an element in a compound (e.g., mono-, di-, tri-, tetra).
- **Suffix** indicates the type of compound or ion (e.g., -ide, -ate, ite).
- **Reactant** is the substance present at the start of a chemical reaction, which takes part in the reaction itself.
- **Product** is the substance formed as a result of a chemical reaction.

- **Stoichiometry** is the relationship between reactants and products in a chemical reaction.
- ***Coefficient*** is the number placed in front of a chemical formula in an equation to balance the equation.
- **Mole ratio** is the ratio of the number of moles between two or more substances.
- **Yield** refers to the amount (mass) of product obtained from a chemical reaction.
- ***Actual yield*** is the amount of (mass) product obtained from a reaction in practice.
- ***Theoretical yield***: the maximum amount (mass) of product that could be produced from a given amount of reactant, based on stoichiometric calculations.
- ***Percent yield*** is the ratio of the actual yield to the theoretical yield, expressed as a percentage.
- **Limiting reagents** are the reactants in a chemical reaction that are consumed completely, and therefore limit the mass of the product formed.
- **Excess reagents** are the reactants that are left over when the reaction has stopped.

NAMING INORGANIC COMPOUNDS USING IUPAC NOMENCLATURE AND BALANCING THEIR CHEMICAL EQUATIONS

The nomenclature of inorganic compounds is based on the oxidation number system.

The oxidation number of an atom is the number of electrons gained or lost by an atom when forming a compound.

Rules for Assigning Oxidation Number

1. The oxidation number of elements in their elemental state is zero.
2. The oxidation state of oxygen in most compounds is -2 except in peroxide (-1) and superoxide (-½).

3. The oxidation state of hydrogen is (+1). When it is bonded to a non-metal and (-1) when bound to a metal.
4. The oxidation state of an ion is equal to the charge on the ion.
5. For a neutral molecule or polyatomic ion, the sum of the oxidation numbers of all the atoms must be equal to the total charge on it.

Rules for Naming Binary Ionic Compounds

1. Name the cation (positive ion) first followed by the name of the anion (negative ion).
2. The name of the cation is the name of the metal.
3. For metals with atomic numbers above 20, indicate the oxidation state in Roman numerals and brackets.
4. The anions are named by replacing the suffix with 'ide'

Example:

NaCl – Sodium chloride

$FeCl_3$ – Iron (III) chloride

Rules for Naming Simple Acids

1. Use the prefix 'hydro', then the root name of the central atom.
2. Add the suffix 'ic' to the root name.
3. Add the word 'acid'.

Example:

HCl – Hydrochloric acid

HI – Hydroiodic acid

Rules for Naming Oxoacids

1. Use the prefixes 'oxo', 'dioxo', 'trioxo' and 'tetraoxo' to indicate the number of oxygen atoms present.
2. Add the root name of the central atom.
3. Add the suffix 'ate' followed by the oxidation state of the central atom in Roman numerals and brackets.
4. Add the word 'acid'.

Example:

H_2CO_3 – trioxocarbonate (IV) acid

H_2SO_4 – tetraoxosulphate (VI) acid

Rules for Naming Acid Salts

1. Name the metal cation first followed by the name of the oxosalt.
1. If the cation has a relative atomic number above 20, its oxidation state should be indicated in Roman numerals and brackets.
2. Add the word ‘hydrogen’.
3. Name the oxoanion as usual without the word ‘ion’.

Example:

$NaHSO_4$ – Sodium hydrogen tetraoxosulphate (VI)

Rules for Naming Simple Non-Ionic Compounds

1. Name the electropositive element first.
2. Add the root name of the anion.
3. Add the suffix ‘ide’.

Example:

HCl – hydrogen chloride

SiC – Silicon carbide

Rules for Naming Molecular compounds where a pair of elements form different compounds with different number of oxygen atoms

1. Name the electropositive element first.
2. Indicate its oxidation state in Roman numerals and brackets.
3. Name the electronegative element as usual.

Example:

CO_2 – Carbon (IV) oxide

N_2O_3 – Nitrogen (III) oxide

Rules for Naming Hydrated Salt

1. Name the anhydrous part first.
2. Use the prefixes mono, di, tri, tetra, penta, hexa, hepta, octa, etc.to indicate the moles of water of crystallisation.
3. Add the suffix 'hydrate' to it.

Example:

$CuSO_4.5H_2O$ – Copper (II) tetraoxosulphate (VI) pentahydrate

Activity 3.1: IUPAC naming conventions for various groups of inorganic compounds.

Materials needed:

Access to the internet or library resources, textbooks on inorganic chemistry, presentation tools (e.g., PowerPoint, poster boards), markers, paper, and other stationery

Steps:

1. This Activity is best done in a classroom environment.
2. Your teacher will help you form groups of 5.
3. Your group will be assigned a specific type of inorganic compound to research into such as:
   - Simple binary compounds (e.g., $NaCl$, CO_2)
   - Oxides and hydroxides (e.g., H_2O, $NaOH$)
   - Acids (e.g., H_2SO_4, HCl)
   - Bases (e.g., KOH, NH_4OH)
   - Salts (e.g., $NaCl$, K_2SO_4)
4. Research into the IUPAC rules for naming specific compounds.
5. Compile your findings into:
   - The basic principles of IUPAC nomenclature for your compound type.
   - Examples of compounds with their systematic names.

6. Prepare your presentation. This should include:

- An introduction to your inorganic compound assigned to your group.
- A detailed explanation of the IUPAC naming rules of that compound.
- Several examples with explanations.
- A list of exceptions to any IUPAC rules for certain chemicals.
- Visual aids such as diagrams, flowcharts, or models.

A sample outline for your presentation has been given at the end of this activity to guide you.

7. Deliver your presentation to your peers in class.

8. Be sure to partake in any Questions and Answer sessions where other groups can ask questions or seek clarifications.

Sample Group Presentation Outline

1. *Naming Simple Binary Compounds*

***IUPAC Naming Rules*:**

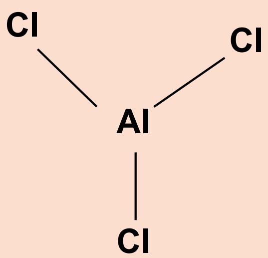
Rule 1: Naming the cation (positive ion) first.

Rule 2: Naming the anion (negative ion) second, modifying its name to end in "-ide".

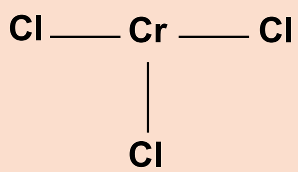
Example: $AlCl_3$ is named aluminium chloride.

$CrCl_3$ is chromium (II) chloride (Cr is in +3 oxidation state)

SO_2 is sulphur (IV) oxide (S is in +4 oxidation state)



Aluminium chloride



Chromium (III) chloride

2. *Naming Oxides and Hydroxides*

***IUPAC Naming Rules*:**

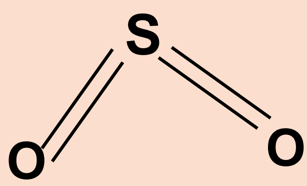
Rule 1: Naming the metal first.

Rule 2: Naming the oxide or hydroxide part.

Example: H_2O is named dihydrogen monoxide,

NaOH is named sodium hydroxide

SO_2 is sulphur (IV) oxide (S is in +4 oxidation state)*



Sulphur (IV) oxide

*It might be worth noting in the presentation that this is commonly called sulphur dioxide and could be considered an exception to the IUPAC rules as it is so common.

3. *Naming Acids*

   ***IUPAC Naming Rules*:**

   Binary Acids: Prefix "hydro-" + root of nonmetal + suffix "-ic" + "acid."

   Oxyacids: Root of polyatomic ion + suffix ("-ate" ions," ions) + "acid."

   Example: HBr is hydrobromic acid, H_2PO_3 is trioxophosphate (VI) acid.

CHEMICAL FORMULAE

A chemical formula is an expression which shows the chemical composition of a compound in terms of the symbols of the atoms involved.

Types of Chemical Formulae

There are three (3) types of chemical formulae:

1. Empirical formula
2. Molecular formula
3. Structural formula

Empirical formula

The empirical formula of a compound is the simplest whole number ratio of atoms present in a compound.

Determination of the empirical formula

1. State the mass or percentage mass of each element.
2. Divide each mass by its relative atomic mass.
3. Determine the simplest whole number ratio of each element by dividing the results by the least ratio number or scale up by X factor to get the simplest whole number ratio.
4. The simplest integer ratio is then used to write the empirical formula as a right-hand subscript.

Molecular formula

Molecular formula is the formula that shows the actual number of atoms of each element in the simplest unit of a substance.

NB: Some compounds have their empirical formula being the same as the molecular formula.

Molecular formula is derived from the empirical formula using the relationship:

$Molecular\ mass = (Empirical\ mass)_n$

The integer **n** obtained is used to multiply each element of the empirical formula to get the molecular formula.

Therefore

$Molecular\ formula = (Empirical\ formula)_n$

Worked Example 3.1

In an experiment to determine the empirical formula of lead sulphide, the following results were obtained:

Mass of lead = 207 g

Mass of sulphur = 32 g

Calculate the empirical formula of lead sulphide [Pb = 207, S = 32].

Solution

		Pb	**S**
Mass of element:		207 g	32 g
Mole of element,	$\left(\frac{m}{Ar}\right)$	$\frac{207}{207}$	$\frac{32}{32}$
Simplest ratio		1	1

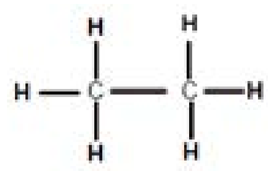
Empirical formula: PbS

Activity 3.2: Trial Question

A compound is found to contain 1.92 g of carbon and 0.48 g of hydrogen. Calculate the empirical formula for the compound.

Structural formula

Structural formula is the formula that shows how the atoms in the molecule or compound are bonded to each other. For example, Ethane, with the molecular formula C_2H_6, has a structure formula:



Percentage Composition of Elements in a Compound

The method of calculating the percentage by mass composition of a compound in terms of its constituent element is as follows:

1. Calculate the molecular mass of the compound.
2. Calculate the mass of the specified elements in the compound, considering the number of atoms of each element in the formula:

   Percentage of element in a compound

   $$= \frac{\text{Relative atomic mass} \times \text{Number of atoms of elements} \times 100}{\text{Relative molecular mass}}$$

Worked Example 3.2

Calculate the percentage by mass of nitrogen and hydrogen in NH_3.

Solution:

By definition,

Percentage of element in a compound

$$= \frac{\text{Relative atomic mass} \times \text{Number of atoms of elements} \times 100}{\text{Relative molecular mass}}$$

$$\% \text{ of nitrogen} = \frac{\text{mass of nitrogen}}{\text{Molar mass of ammonia}} \times 100$$

$$\% \text{ of nitrogen} = \frac{14}{17} \times 100$$

$$= 82.35\ \%$$

$$\% \text{ of hydrogen} = \frac{\text{mass of hydrogen}}{\text{Molar mass of ammonia}} \times 100$$

$$\% \text{ of hydrogen} = \frac{3}{17} \times 100 = 17.65\ \%$$

Activity 3.3: Trial Question

Calculate the percentage by mass of carbon in methane (CH_4).

Activity 3.4: Determining the empirical formula of a named compound

This Activity will guide you to determine the empirical formula of a copper chloride compound using the conservation of mass, the law of definite proportions, and the law of multiple proportions.

Materials needed

Periodic table, calculators, handouts with problem details, whiteboard and markers

Steps:

1. Research and discover the laws of Conservation of Mass, Definite Proportions and Multiple Proportions. You may find the following useful in your search:

   a. ***Conservation of Mass:*** In a chemical reaction, the total mass of reactants equals the total mass of products.

   b. ***Law of Definite Proportions:*** A given chemical compound always contains its component elements in a fixed ratio by mass.

   c. ***Law of Multiple Proportions:*** When two elements form more than one compound, the ratios of the masses of the second element that combine with a fixed mass of the first element are ratios of small whole numbers.

2. Determine the empirical formula of the named compound, e.g. a copper chloride compound with a mass composition of 47.4% copper and 52.6% chlorine.

Given:

Percentage composition: 47.4% Cu, 52.6% Cl

Atomic masses: Cu = 64, Cl = 35.5

3. Calculate the Empirical Formula as follows:

Convert Percentage to Mass:

Assume you have 100 grams of the compound.

Mass of Cu = 47.4 grams

Mass of Cl = 52.6 grams

Convert Mass to Moles:

$$\text{Moles of Cu} = \frac{\text{Mass of Cu}}{\text{Atomic mass of Cu}}$$

$$= \frac{47.4 \text{ g}}{64 \text{ g/mol}} = 0.7406 \text{ mol}$$

$$\text{Moles of Cl} = \frac{52.6 \text{ g}}{35.5 \text{ g/mol}} = 1.4817 \text{ mol}$$

4. Determine the simplest ratio:

Divide the moles of each element by the smallest number of moles calculated:

$$\text{Moles of Cu} = \frac{0.7406 \text{ mol}}{0.7406 \text{ mol}} = 1$$

$$\text{Moles of Cl} = \frac{1.4817 \text{ mol}}{0.7406 \text{ mol}} = 2$$

5. Write the Empirical Formula:

The simplest whole-number ratio is 1:2.

Therefore, the empirical formula is $CuCl_2$.

Worked Example 3.3

The molar mass of the compound with the empirical formula CH_2O is 180 g/mol. Determine the molecular formula.

Solution

1. **Calculate the Empirical formula molar Mass:**

   Empirical formula mass of $CH_2O = (1 \times 12) + (2 \times 1) + (1 \times 16)$

   $= 12 + 2 + 16$

   $= 30$

2. Determine the Ratio of Molar Masses to Empirical formula molar mass:

   $$\text{Ratio} = \frac{\text{Molar mass of compound}}{\text{Empirical formula molar mass}}$$

   $$= \frac{180g/mol}{30\text{ g/mol}}$$

   $= 6$

3. Calculate the Molecular Formula:

   Multiply the subscripts in the empirical formula by the ratio calculated:

   Molecular formula $= C_1H_2O_1 \times 6 = C_6H_{12}O_6$

   The molecular formula is $\mathbf{C_6H_{12}O_6}$.

Determining the percentage by mass of an element in a compound.

Worked Example 3.4

Determine the percentage by mass of carbon and hydrogen in C_2H_6. [C = 12 and H = 1.]

Solution

Calculate the Molar Mass of C_2H_6

Find the number of atoms of each element in the compound:

Carbon atoms: 2

Hydrogen atoms: 6

Calculate the total mass of each element in the compound:

The total mass of carbon: $2 \times 12 = 24$ g/mol

Total mass of hydrogen: $6 \times 1 = 6$ g/mol

Calculate the molar mass of C_2H_6

Molar mass of C_2H_6: $24 + 6 = 30$ g/mol

Calculate the percentage by mass of each element

$$\%C = \frac{\text{Total mass of C}}{\text{Molar mass of } C_2H_6} \times 100$$

$$= \frac{24}{30} \times 100$$

$$= 80\%$$

$$\%\text{H} = \frac{\text{Total mass of H}}{\text{Molar mass of } C_2H_6} \times 100$$

$$= \frac{6}{30} \times 100$$

$$= 20\%$$

LAWS OF CHEMICAL COMBINATION

1. The Law of Conservation of Mass

The Law of Conservation of Mass states that mass is not created or destroyed in a chemical reaction, i.e. the total mass of the products made is equal to the total mass of the reactants.

2. The Law of Definite Proportion

It says that the proportion by amounts of each element in a pure compound is always the same, no matter how the compound is prepared.

3. The Law of Multiple Proportion

When two elements combine to form more than one compound, the different masses of one element that combine with a fixed mass of the other element are in a ratio of small whole numbers.

It is defined as an expression which uses chemical symbols in the formula to represent the elements and compounds that occur in a chemical reaction.

Types of Chemical Equations

Note: An equation may fall into multiple categories simultaneously.

1. Combustion

Combustion is a chemical reaction in which a substance reacts with excess or limited oxygen to give oxides of the components of the substance.

Example: $CH_4 + 2O_2 \rightarrow CO_2 + H_2O$

Application in everyday life

Fuel combustion generates heat for homes, electricity for the power grid and to generate movement in engines.

2. Synthesis

Synthesis a reaction in which two or more simple substances combine to form a more complex compound. That is, the reactants combine to form a single product. It can be represented as: $A + B \rightarrow AB$

Example: $Na + Cl \rightarrow NaCl$

Application in everyday life

Synthesis is widely used in chemistry for the formation of salts, organic compounds, biomolecules, medicines, pesticides and polymers.

3. Displacement reaction

A displacement reaction is one in which one atom or ion in a reactant is replaced by another atom or ion of another element.

Example: $Mg + CuSO_4 \rightarrow MgSO_4 + Cu$

Application in everyday life

a. Displacement reactions are essential in various chemical processes.
b. They have practical applications in metallurgy, electrochemistry, and extraction of metals such as gold from their ores.

4. Decomposition

Decomposition is a chemical reaction in which a compound breaks down into two or more simpler substances under certain conditions. That is, a simple reactant undergoes a chemical change to produce multiple products.

It is illustrated as: $AB \rightarrow A + B$

Where AB is the initial reactant and A and B are the products.

The conditions under which the reaction occurs could be heat, light and the use of a catalyst.

Example: $CaCO_{3} \rightarrow CaO + CO_2$

Application in everyday

a. It is important in natural and industrial processes.
b. They are essential in the fields of chemistry, biology and environmental science.

5. Ionic equation

An ionic equation is a chemical equation involving at least one ionic species as a reactant or product, that is, species of dissolved ionic compounds in terms of their free ions; ions that exist in a chemical equation, but are not involved in the overall equation, are **spectator ions**.

Worked Example 3.5

Write the net ionic equation of the reaction,

$$2KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbI_{2(s)} + 2KNO_{3(aq)}$$

Solution:

Write the ionic equation

$$2K^{+}_{(aq)} + 2I^{-}_{(aq)} + Pb^{2+}_{(aq)} + 2NO_{3(aq)} \rightarrow PbI_{2(s)} + 2K^{+}_{(aq)} + 2NO^{-}_{3(aq)}$$

Cancel the spectator ions to yield the net ionic equation

$$\cancel{2K}^{+}_{(aq)} + 2I^{-}_{(aq)} + Pb^{2+}_{(aq)} + 2\cancel{NO}_{3(aq)} \rightarrow PbI_{2(s)} + 2\cancel{K}^{+}_{(aq)} + 2N\cancel{O}_{3(aq)}$$

Write the net ionic equation,

$$2I^{-}_{(aq)} + Pb^{2+}_{(aq)} \rightarrow PbI_{2(s)}$$

Applying and verifying the Laws of Chemical Combination

Verifying the Law of conservation of matter

The Law of conservation of matter states that mass is neither created nor destroyed in a chemical reaction. The total mass of reactants is equal to the total mass of products.

Worked Example 3.6

Consider the reaction between nitrogen and hydrogen to form ammonia:

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

Solution:

Calculate the mass of reactants:

Mass of $N_2 = (2 \times 14) = 28$ g

Mass of 3 $H_2 = 3(2 \times 1) = 6$ g

Total mass of reactants $= 28 + 6 = 34$ g

Calculate the mass of products:

Total mass of 2 $NH_3 = 2\left[(1 \times 14) + (3 \times 1)\right] = 34$ g

Verify the Law:

Total mass of reactants = Total mass of product

$$34 \text{ g} = 34 \text{ g}$$

Verifying the Law of Definite Proportions

The Law of definite proportions states that a chemical compound always contains the same proportion of elements by mass.

Example: methane (CH_4) always contains C and H in a mass ratio of 3:1. Here is how to verify this:

Calculate the mass ratio of elements in CH_4:

Molar mass of $CH_4 = (1 \times 12) + (4 \times 1)$

$$= 16$$

Mass of C in $CH_4 = 1 \times 12 = 12$ g
Mass of H in $CH_4 = 4 \times 1 = 4$ g

Determine the mass ratio:
Mass ratio of C to H $= \frac{12}{4} = \frac{3}{1}$

Verifying the Law of Multiple Proportions

The Law of multiple proportions states that when two elements combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.

Example: copper forms two oxides: copper (I) oxide (Cu_2O) and copper (II) oxide (CuO.

Compare the two Compounds:
Molar mass of $Cu_2O = (2 \times 64) + 16 = 144$ g mol^{-1}
Molar mass of $CuO = (1 \times 64) + 16 = 80$ g mol^{-1}

Fixed Mass of oxygen:
Mass of Oxygen (O) in both compounds = 16 g

Variable mass of Cu
Mass of Cu in $Cu_2O = 128$ g
Mass of Cu in $CuO = 80$ g

Find out the ratio:
Ratio of the masses of copper that combine with a fixed mass of oxygen (16 g):

$$\frac{128}{64} = \frac{16}{8} = \frac{2}{1}$$

The masses of copper in a simple whole number ratio of 2:1.

Activity 3.5: Trial Questions

1. Discuss the rules to be followed in balancing chemical equations.
2. Write and balance the following types of reactions:

a. Combustion Reaction

Example: to balance the chemical equation in which Ethyne (C_2H_2) combusts in oxygen (O_2) to form carbon dioxide (CO_2) and water (H_2O), First, write the unbalanced equation:

$$C_2H_2 + O_2 \rightarrow CO_2 + H_2O$$

Count the atoms of each element:

Reactants: C = 2, O = 2, H = 2

Products: C = 1, O = 3, H = 2

Balance the equation:

$$2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$$

Check your work:

Reactants: C = 4, O = 10, H = 4

Products: C = 4, O = 10, H = 4

b. Synthesis Reaction

Example: Aluminium (Al) reacts with chlorine (Cl_2) to form aluminium chloride ($AlCl_3$)

Write the unbalanced equation:

$$Al + Cl_2 \rightarrow AlCl_3$$

Count the atoms of each element:

Reactants: Al = 1, Cl = 2

Products: Al = 1, Cl = 3

Balance the equation:

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

Check your work:

Reactants: Al = 2, Cl = 6

Products: Al = 2, Cl = 6

c. Displacement (Replacement) Reaction

Example: Magnesium chloride ($MgCl_2$) reacts with potassium phosphate (K_3PO_4) to form magnesium phosphate ($Mg_3(PO_4)_2$) and potassium chloride (KCl)

Write the unbalanced equation:

$$MgCl_2 + K_3PO_4 \rightarrow Mg_3(PO_4)_2 + KCl$$

Count the atoms of each element:

Reactants: Mg = 1, Cl = 2, K = 3, P = 1, O = 4

Products: Mg = 3, Cl = 1, K = 1, P = 2, O = 8

Balance the equation:

$$3MgCl_2 + 2K_3PO_4 \rightarrow Mg_3(PO_4)_2 + 6KCl$$

Check your work:

Reactants: Mg = 3, Cl = 6, K = 6, P = 2, O = 8

Products: Mg = 3, Cl = 6, K = 6, P = 2, O = 8

d. Decomposition Reaction

Example: Decomposition of calcium carbonate ($CaCO_3$) to form calcium oxide (CaO) and oxygen (CO_2).

Write the unbalanced equation:

$$CaCO_3 \rightarrow CaO + CO_2$$

Count the atoms of each element:

Reactants: Ca = 1, C = 1, O = 3

Products: Ca = 1, C = 1, O = 3

Balance the equation:

$$CaCO_3 \rightarrow CaO + CO_2$$

Check your work:

Reactants: Ca = 1, C = 1, O = 3

Products: Ca = 1, C = 1, O = 3

e. Ionic Equation

Example: Sodium sulphate (Na_2SO_4) reacts with barium chloride ($BaCl_2$) to form barium sulphate ($BaSO_4$) and sodium chloride (NaCl).

Write the balanced molecular equation:

$$Pb\left(NO_3\right)_2(aq) + BaCl_2(aq) \rightarrow Ba\left(NO_3\right)_2\left(aq\right) + PbCl_2\left(s\right)$$

Write the complete ionic equation:

$$Pb^{2+}(aq) + 2NO_3^-(aq) + Ba^{2+}(aq) + 2Cl^-\left(aq\right) \rightarrow Ba^{2+}(aq) + 2NO_3^-(aq) + PbCl_2\left(s\right)$$

Write the net ionic equation:

$$Pb^{2+}(aq) + 2Cl^-\left(aq\right) \rightarrow PbCl_2\left(s\right)$$

Activity 3.6: Trial Question

Balance the following reactions:

1. $H_2 + O_2 \rightarrow H_2O$
2. $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$

Activity 3.7: Demonstrating that mass is conserved in a chemical reaction

Scenario Experiment: Reaction between $Na_2CO_3$and $CaCl_2$.

Prepare 250 cm^3 of 1 mol dm^{-3} Na_2CO_3 and 250 cm^3 of 1 mol dm^{-3} $CaCl_2$ solutions.

Materials needed:

Sodium carbonate solution (Na_2CO_3), calcium chloride solution ($CaCl_2$), 2 beakers, a stirring rod, pipettes, an electronic balance

Steps:

Preparation

1. Measure 25 cm^3 of 1 mol dm^{-3} Na_2CO_3 solution into a 250 cm^3beaker labelled A.
2. Measure 25 cm^3 of 1 mol dm^{-3} $CaCl_2$ solution into another beaker labelled B.
3. Record the initial mass of each solution using a balance.

Reaction

4. Slowly add the $CaCl_2$ solution to the Na_2CO_3 solution while stirring.

5. Observe the reaction and take note of any changes.

6. Record the final mass of the mixture using a balance.

$Na_2CO_3 + CaCl_2 \rightarrow CaCO_3\ (\text{precipitate}) + 2NaCl$

7. Record your observations and measurements in the table below:

Measurement Description	Mass (g)
Mass of empty Beaker A	
Mass of empty Beaker B	
Mass of Beaker A with sodium carbonate	
Mass of Beaker B with calcium chloride	
Total mass of reactants (A + B)	
Total mass of products (precipitate + solution).	

If total mass of reactants = total mass of products

Then mass is conserved in the chemical reaction.

Activity 3.8: Trial Question

Verify the Law of Conservation of Mass with the reaction:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

STOICHIOMETRY

Welcome to the world of stoichiometry, where you will explore the relationships between reactants and products in chemical reactions. You will learn to calculate reactant amounts (and masses), predict product yields, balance equations, and understand the mole concept. Get ready to enhance your problem-solving skills, critical thinking, and understanding of chemistry. Let us dive in and uncover the secrets of chemical reactions!

Stoichiometry is the relationship between quantities of reactants and products in a chemical reaction.

Mole Ratio

The relative quantity of any two substances that take part in a chemical reaction is termed as **mole ratio**. The stoichiometric coefficients in the balanced equation are considered as the number of moles of each reactant or product that would either be consumed or produced under ideal conditions.

Worked Example 3.7

In the reaction,

2C	+	O_2	→	2CO
2 mol		1 mol		2 mol

The mole ratio between carbon and oxygen is written as:

$$\frac{n(C)}{n(O_2)} = \frac{2}{1}$$

$$n = \frac{N}{L}$$

Using Stoichiometric Quantities to Calculate Numbers of Entities Chemical Reactions

1. Write the mole ratio using the stoichiometric coefficient of the known substance and that of the substance being calculated.
2. Calculate the number of moles of the substance being calculated.
3. Use the calculated moles and the relation $N = n \times L$ to calculate the number of entities.

Worked Example 3.8

Consider the equation: $N_2 + 3H_2 \rightarrow 2NH_3$

Calculate the number of molecules of ammonia gas produced if 3.01×10^{23} molecules of Hydrogen react with Nitrogen gas. [$L = 6.02 \times 10^{23}$]

Solution

$N_2 + 3H_2 \rightarrow 2NH_3$

$N(H_2) = 3.01 \times 10^{23}$, n = ?

Determine the number of moles of Hydrogen using the formula

$$n = \frac{N}{L}$$

$$n(H_2) = \frac{3.01 \times 10^{23}}{6.02 \times 10^{23}} = 0.5 \text{ mol}$$

Write mole ratio between ammonia and Hydrogen

$$\frac{n(NH_3)}{n(H_2)} = \frac{2}{3}$$

Calculate the number of moles of ammonia produced:

$$n(NH_3) = \frac{2}{3} \times n(H_2)$$

$$n(NH_3) = \frac{2}{3} \times 0.5 = 0.33 \text{ mol}$$

Solve for the number of molecules of product using the formula

$$N = n \times L$$

$$N(NH_3) = 0.33 \times 6.02 \times 10^{23}$$

$$= 1.99 \times 10^{23} \text{ molecules}$$

Calculating the mass of a substance

1. Write the correct balanced equation.
2. Convert the quantity of the known substance into the number of moles using the correct mole formula
   $n = \frac{m}{M}$.
3. Write the mole ratio using the stoichiometric coefficient of the known substance and substance being sought.
4. Calculate the number of moles of the substance being sought.

5. Use the calculated moles and the relation $m = n \times M$ to calculate the mass of the substance.

Worked Example 3.9

Consider the reaction,

$N_2 + 3H_2 \rightarrow 2NH_3$

Calculate the mass of ammonia produced if 7 g of Nitrogen reacts with excess Hydrogen gas.

[A_r: N = 14, H = 1]

Solution

Given the equation,

$N_2 + 3H_2 \rightarrow 2NH_3$

$m(N_2) = 7g$, $m(NH_3) = ?$

Determine moles of Nitrogen using its mass and molar mass and the formula

$$n = \frac{m}{M}$$

$$n(N_2) = \frac{7}{2 \times 14} = 0.25 \text{ mol}$$

Determine the mole ratio between Nitrogen and Ammonia

$$\frac{n(NH_3)}{n(N_2)} = \frac{2}{1}$$

$$n(NH_3) = 2 \times n(N_2)$$

$$n(NH_3) = 2 \times 0.25$$

$$= 0.5 \text{ mol}$$

$$A_r\ (NH_3) = 14 + 3(1)$$

$$= 17 \text{ g mol}^{-1}$$

$$m = 0.5 \times 17$$

$$= 8.5 \text{ g of Ammonia}$$

Calculate the concentration of substance (analyte)

1. Write the correct balanced equation.
2. Observe the units closely and convert any volumes or concentrations as required; for example, converting cm^3 to dm^3.
3. Convert the quantity of the known substance into mole using the correct mole formula

   $c = \frac{n}{V}$.
4. Write the mole ratio using the stoichiometric coefficient of the known substance and substance being sought.
5. Calculate the number of moles of the substance being sought.

   $c = \frac{n}{V}$
6. Use the calculated moles and the relation to calculate the concentration of the unknown substance.

Worked Example 3.10

1. Consider the reaction,

   $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

   Given that 20 cm^3 of H_2SO_4 reacts completely with 25 cm^3 of 0.5 mol dm^{-3} NaOH, calculate the concentration of H_2SO_4.

Solution

$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

$c(NaOH) = 0.5 \text{ mol dm}^{-3}$

$V(NaOH) = 25 \text{ cm}^3 = 0.025 \text{ dm}^3$

$c(H_2SO_4) = ?$

$V(H_2SO_4) = 20 \text{ cm}^3 = 0.020 \text{ dm}^3$

Determine the number of moles of Sodium hydroxide solution using the formula, $n = c \times V$

$n = 0.025 \times 0.5$

$= 0.0125 \text{ mol}$

Write the mole ratio:

$$\frac{n(H_2SO_4)}{n(NaOH)} = \frac{1}{2}$$

$$n(H_2SO_4) = \frac{1}{2} \times n(NaOH)$$

$$n(H_2SO_4) = \frac{1}{2} \times 0.0125$$

$$= 0.00625 \text{ mol}$$

Determine the concentration of sulphuric acid using the formula, $C = \frac{n}{V}$:

$$C = \frac{0.00625}{0.020}$$

$$= 0.313 \text{ mol dm}^{-3}$$

Calculating for the volume of substance

Procedure

1. Write the correct balanced equation.
2. Convert the quantity of the known substance into a number of moles using the correct mole formula

   $n = \frac{m}{M}$

   $n = \frac{N}{L}$

   $n = \frac{V}{V_m}$
3. Write the mole ratio using the stoichiometric coefficients of the known substance and substance being calculated.
4. Calculate the number of moles of the substance being sought.
5. Use the calculated number of moles and the relationship $V = n \times V_m$ to calculate the volume of the gas.

Worked Example 3.11

10.5 g of methane reacts with excess Oxygen to produce carbon dioxide and water. Calculate the volume of carbon dioxide gas produced at s.t.p. [A_r:C = 12, H = 1, O = 16, V_m = 22.4 dm^3/mol].

Solution

Use the problem-solving strategy:

1. Write the correct balanced equation.

   $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

2. Convert the quantity of the known substance into a number of moles using the correct mole formula

   $m(CH_4) = 10.5g,$

   $m(CH_4) = 12 + (1 \times 4) = 16 \text{ g mol}^{-1}$

   $$n = \frac{m}{M}$$
   $$= \frac{10.5}{16}$$
   $$= 0.66 \text{ mol}$$

3. Write the mole ratio using the stoichiometric coefficients of the known substance and substance being calculated.

   $$\frac{n(CO_2)}{n(CH_4)} = \frac{1}{1}$$

   $n(CO_2) = n(CH_4)$

4. Calculate the number of moles of the substance being sought for.

   $n(CO_2) = n(CH_4) = 0.66 \text{ mol}$

   $\therefore n(CO_2) = 0.66 \text{ mol}$

5. Use the calculated number of moles and the relationship $V = n \times V_m$ to calculate the volume of the gas.

   $$n(CO_2) = \frac{V}{V_m}$$

   $$V = n(CO_2) \times V_m$$
   $$= 0.66 \times 22.4$$
   $$= 14.78 \text{ dm}^3$$

Activity 3.9: Perform simple chemical reactions, write and balance equations, and perform stoichiometric calculations.

Materials needed:

- Chemicals**:** Sodium hydroxide (NaOH), Sodium bicarbonate ($NaHCO_3$), hydrochloric acid (HCl), Magnesium ribbon.
- Equipment**:** test tubes, test tube rack, beakers, measuring cylinders, safety goggles, gloves, lab coat.

Safety instructions:

- Follow the teacher's instructions and ask questions if unsure.
- Wear safety goggles, gloves, and a lab coat always.
- Handle all chemicals with care.

Steps:

Reaction between Sodium bicarbonate ($NaHCO_3$) and hydrochloric acid (0.100 mol dm^{-3} HCl):

1. Measure 50 cm^3 of HCl using a measuring cylinder and pour it into a beaker.

2. Add one teaspoon of Sodium bicarbonate to the beaker with HCl.

3. Observe the reaction and record your observations.

4. Write the balanced chemical equation for the reaction.

5. Calculate the number of moles of $NaHCO_3$ used if 4.0 g of $NaHCO_3$ were reacted.

6. Using the balanced equation, determine the mole ratio of reactants to products.

7. Calculate the:

  a. number of moles of CO_2 produced.

  b. mass CO_2 produced.

  c. volume of CO_2 produced.

  d. concentration of CO_2 in g dm^{-3}

  e. concentration of CO_2 in mol dm^{-3}.

  f. concentration of CO_2 in ppm (remember that ppm = gdm^{-3} x 1000).

Steps:

Reaction between hydrochloric acid and Sodium hydroxide:

1. Measure 25 cm^3 of 0.100 mol dm^{-3} HCl using a measuring cylinder and pour it into a beaker **A**.
2. Measure 25 cm^3 of NaOH using a measuring cylinder into another beaker **B**.
3. Add the solution in beaker **B** to the solution in beaker **A**.
4. Observe the reaction and record your observations.
5. Write the balanced chemical equation for the reaction.
6. Calculate the:
   - **a.** number of moles of HCl used if 25 cm^3 of 0.10 mol dm^{-3} HCl solution was reacted.
   - **b.** number of moles of NaCl produced.
   - **c.** mass of NaCl produced.
   - **d.** concentration of NaCl in g dm^{-3}
   - **e.** concentration of NaCl in mol dm^{-3}.
   - **f.** concentration of NaCl in ppm.

Steps:

Reaction of Magnesium with hydrochloric acid:

1. Place 5 g of Magnesium ribbon into a test tube.
2. Add 20 cm^3 of hydrochloric acid to the test tube.
3. Observe the reaction and note any changes.
4. Write the balanced chemical equation for the reaction.
5. Calculate the:
   - **a.** number of moles of Magnesium used.
   - **b.** number of moles of H_2 produced.
   - **c.** mass H_2 produced.
   - **d.** volume of H_2 produced.
   - **e.** concentration of H_2 in g dm^{-3}.

LIMITING AND EXCESS REAGENTS

In a reacting system involving two reactants with initial quantities given or having information to determine their initial quantities, the reactant that is completely used up is called the **limiting reagent**. The reagent that is not completely used up is called the **excess reagent**.

The maximum quantity of the products formed is determined by the limiting reagent.

Procedure for Determining the Limiting and Excess Reagents

1. Calculate the initial quantity of each reactant in moles.
2. If the stoichiometric ratio of the reactants is 1:1, the reagent with a lower value is the limiting reagent, and the other is the excess reagent.
3. If the stoichiometric ratio of the reactant is not 1:1, then write a mole ratio between the two reactants and solve for one of them.
4. Compare the calculated number of moles with the initial quantity of moles. If the calculated number of moles is greater than the initial amount, then it is the limiting reagent, and if it is less, it is the excess reactant.

Worked Example 3.12

Consider the reaction,

$N_2 + 3H_2 \rightarrow 2NH_3$

If 12.0 g of Nitrogen and 8.0 g of Hydrogen react in the formation of ammonia,

a. Determine the limiting reagent.
b. Calculate the mass of ammonia (NH_3) produced. [A_r: N = 14, H = 1]

Solution:

$N_2 + 3H_2 \rightarrow 2NH_3$

$m(N_2) = 12g, m(H_2) = 8g$

a. Determine the initial moles of both reactants

$$n\left(N_2\right) = \frac{m}{M}$$
$$= \frac{12}{28} = 0.43 \text{ mol}$$

$$n(H_2) = \frac{m}{M} = \frac{8}{2} = 4\ mol$$

b. Determine the limiting reagent

$$\frac{n(N_2)}{n(H_2)} = \frac{1}{3}$$

$$n(N_2) = \frac{1}{3} \times n(H_2)$$

$$n(N_2) = \frac{1}{3} \times 4$$

$$= 1.33 \text{ mol}$$

1.33 moles of N_2 are required to react with 4 moles of H_2, but we only have 0.43 moles, so N_2 is the limiting reactant.

c. Write the mole ratio between the limiting reagent and the product (NH_3)

$$\frac{n(NH_3)}{n(N_2)} = \frac{2}{1}$$

$$n(NH_3) = \frac{2}{1} \times n(N_2)$$

$$n(NH_3) = 2 \times 0.43$$

$$= 0.86 \text{ mol}$$

Mass of NH_3, $m = n \times M$

$= 0.86 \times 17$

$= 14.62$ g

Activity 3.10: Calculations for limiting reagents and excess reactants

Materials needed:

- Worksheet, Periodic tables, calculator

Given the reaction: $2H_2 + O_2 \rightarrow 2H_2O$

If 3 mol of H_2 react with 2 mol of O_2, determine the:

a. Limiting reagent.

b. Excess reactant.

c. Mass of water formed.

Steps to solve:

1. Identify the limiting and excess reactants:

Use the balanced equation to find the mole ratio.

$$\frac{n(O_2)}{n(H_2)} = \frac{1}{2}$$

$$n(O_2) = \frac{1}{2} \times n(H_2)$$

$$= \frac{1}{2} \times 3 = 1.5 \text{ mol}$$

Compare the mole ratio with the actual moles available

Since 2 mol of O_2 are available, but only 1.5 mol is needed to react with the available H_2, O_2 is in excess.

∴ The limiting reagent is H_2

The excess reactant is O_2 := 2 – 1.5 = 0.50 mol

Mass of H_2O produced:

Use the moles of the limiting reagent to find the moles of H_2O formed:

$$\frac{n(H_2O)}{n(H_2)} = \frac{2}{2} = 1$$

$$n(H_2O) = n(H_2) = 3 \text{ mol}$$

Convert moles of H_2O to gram

$$M(H_2O) = (2 \times 1) + 16$$

$$= 18 \text{ g mol}^{-1}$$

$$m(H_2O) = n(H_2O) \times M(H_2O)$$

$$= 3 \text{ mol} \times 18 \text{ g mol}^{-1}$$

$$= 54.0 \text{ g}$$

Consider the reaction: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$.

If 10 moles of Iron and 8 moles of Oxygen are available, determine the

(a) limiting reagent.

(b) excess reactant.

(c) mass of water formed.

Steps to solve:

1. Identify limiting and excess reactants:

Use the balanced equation to find the mole ratio.

$$\frac{n(O_2)}{n(Fe)} = \frac{3}{4}$$

$$n(O_2) = \frac{3}{4} \times n(Fe)$$

$$= \frac{3}{4} \times 10 = 7.5 \text{ mol}$$

Compare the mole ratio with the actual moles available.

Since 8 mol of O_2 are available, but only 7.5 mol are needed to react with the available Fe,

. O_2 is in excess.

∴ Limiting reagent is Fe

Excess reactant is O_2 := 8 – 7.5 = 0.50 mol

Mass of Fe_2O_3 produced:

Use the moles of the limiting reagent to find moles of Fe_2O_3 formed.

$$\frac{n(Fe_2O_3)}{n(Fe)} = \frac{2}{4} = \frac{1}{2}$$

$$n(Fe_2O_3) = \frac{1}{2} n(Fe)$$

$$= \frac{1}{2} \times 10 = 5 \text{ mol}$$

Convert moles of Fe_2O_3 to gram

$$M(Fe_2O_3) = (2 \times 56) + (3 \times 16)$$

$$= 160 \text{ g mol}^{-1}$$

$$m(Fe_2O_3) = n(Fe_2O_3) \times M(Fe_2O_3)$$

$$= 5 \text{ mol} \times 160 \text{ g mol}^{-1}$$

$$= 800.0 \text{ g}$$

Calculate the volume of Chlorine required to react completely with 50 **cm^3 of 1.0 mol dm^3 Sodium Bromide (NaBr) solution.**

Steps to solve:

Writing and balancing equations

Unbalanced: $Cl_2 + NaBr \rightarrow NaCl + Br_2$

Balanced: $Cl_2 + 2\ NaBr \rightarrow 2NaCl + Br_2$

Find the number of mol of NaBr

V(solution) = 50 cm^3 = 0.050 dm^3, c(NaBr) = 0.10 mol dm^{-3}

$$\begin{aligned} n(NaBr) &= c(NaBr) \times V(solution) \\ &= 0.10 \text{ mol dm}^{-3} \times 0.050 \text{ dm}^3 \\ &= 0.050 \text{ mol} \end{aligned}$$

Use the mole ratio to determine moles of Chlorine (Cl_2) needed:

$$\frac{n(Cl_2)}{n(NaBr)} = \frac{1}{2}$$

$$\begin{aligned} n(O_2) &= \frac{1}{2} \times n(NaBr) \\ &= \frac{1}{2} \times 0.050 \\ &= 0.025 \text{ mol} \end{aligned}$$

Calculate the volume of Chlorine Gas (Cl_2):

Assuming the reaction occurs at standard temperature and pressure, where 1 mole of any gas occupies 22.4 dm^3.

$$\begin{aligned} V(Cl_2) &= n(Cl_2) \times V_m \\ &= 0.025 \text{ mol} \times 22.4 \text{ dm}^3 \text{ mol}^{-1} \\ &= 0.56 \text{ dm}^3 \end{aligned}$$

PERCENTAGE YIELD OF THE PRODUCT OF A REACTION

In a chemical reaction, the calculated amount of a product is usually not obtained due to the following:

1. The reaction may be reversible.
2. Some reactants may undergo side reactions.
3. Some products cannot be separated or recovered from the mixture.

The Actual Yield is the amount of a product obtained from a chemical reaction in practice.

The percentage yield of a reaction is the percentage of the product obtained compared to their theoretical maximum yield calculated from the balanced equation.

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

Worked Example 3.13

Magnesium metal reacts with hydrochloric acid to produce Magnesium Chloride and Hydrogen gas.

a. If 12 g of Magnesium reacts with excess HCl, calculate the maximum theoretical mass of Magnesium Chloride formed.

b. If 42.0 g of purified anhydrous Magnesium Chloride was obtained, calculate the percentage yield. [A_r: Mg = 24, Cl = 35.5]

Solution:

Use the problem-solving approach:

a. Analyse the question:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$

$m(Mg) = 12\ g$

$m(MgCl_2) = ?$

b. Determine the number of moles of Mg:

$$n(Mg) = \frac{m}{M}$$

$$= \frac{12}{24} = 0.50\ mol$$

c. Write the mole ratio between $MgCl_2$ and Mg:

$$\frac{n(MgCl_2)}{n(Mg)} = \frac{1}{1}$$

$$n(MgCl_2) = n(Mg)$$

$$= 0.5 \text{ mol}$$

d. Determine the theoretical mass of $MgCl_2$ formed:

$$m = n \times M.$$

But, $M\ (MgCl_2) = 24 + (2 \times 35.5) = 95 \text{ g mol}^{-1}$

$$m = 0.5 \times 95$$

$$= 47.5 \text{ g}$$

e. Determine the percentage yield:

$$\text{Percentage yield} = \frac{\text{Actual amount obtained}}{\text{maximum theoritical yield}} \times 100$$

$$= \frac{42}{47.5} \times 100 = 88.42\ \%$$

Activity 3.11: Determining Actual Yield, Theoretical Yield, and Percentage Yield

Step 1: Write the Balanced Chemical Equation

- Identify the reactants and products.
- Ensure the equation is balanced to reflect the conservation of mass.

  Example: $2H_2 + O_2 \rightarrow 2H_2O$

↓

Step 2: Convert Reactant Quantities to Moles

- Use molar masses to convert grams of reactants to moles.

  Given: 10g H_2

  Molar Mass H_2: 2 g/mol

  Moles of $H_2 = \frac{10\text{g}}{2 \text{ g/mol}} = 5 \text{ mol}$

↓

Step 3: Use Stoichiometry to Determine Theoretical Yield

- Use the balanced chemical equation to find the mole ratio between reactants and products.
- Calculate the moles of product expected from the given moles of reactants.
- Convert moles of product to grams using its molar mass.

Mole Ratio: $H_2 : H_2O = 2 : 2 = 1:1$

$n(H_2O) = n(H_2) = 5 \text{ mol } H_2O$

Molar Mass $H_2O = 18$ g/mol

Mass of $H_2O = 5 \text{ moles} \times 18 \text{ g/mol} = 90\text{g}$

Step 4: Measure the Actual Yield

- Perform the experiment and measure the actual amount of product obtained.

Actual Yield from Experiment = 85g H_2O

Step 5: Calculate Percentage Yield

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

$$\text{Percentage yield} = \frac{85 \text{ g}}{90 \text{ g}} \times 100 = 94.4\%$$

Worked Example 3.14

In a laboratory experiment, students aimed to produce 25 grams of copper sulphate ($CuSO_4$) according to the reaction: $Cu + H_2SO_4 \rightarrow CuSO_4 + H_2$

If they actually obtained 20 grams of copper sulphate, calculate the percentage yield of the reaction.

Solution:

- Identify the theoretical yield and actual yield:

  Actual yield: The amount of product actually obtained, which is 20 grams of $CuSO_4$.

  Theoretical yield: The amount of product that was aimed for, which is 25 grams of $CuSO_4$.

- Use the formula for percentage yield:

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

$$= \frac{20\text{ g}}{25\text{ g}} \times 100 = 80\%$$

Worked Example 3.15

In a chemical synthesis, 30 grams of Calcium carbonate ($CaCO_3$) reacts with excess hydrochloric acid (HCl) to produce Calcium Chloride ($CaCl_2$), Carbon dioxide (CO_2), and water (H_2O). If 22 grams of Calcium Chloride are obtained in the reaction, calculate the percentage yield.

Solution:

Write the balanced chemical equation: $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

Convert the mass of $CaCO_3$ to moles:

$$M(CaCO_3) = 40 + 12 + (3 \times 16)$$

$$= 100\text{ g mol}^{-1}$$

$$n(CaCO_3) = \frac{m(CaCO_3)}{M(CaCO_3)}$$

$$= \frac{30\ g}{100\text{ g mol}^{-1}} = 0.30\text{ mol}$$

Use stoichiometry to find the moles of $CaCl_2$ produced:

The balanced equation shows a 1:1 mole ratio between $CaCO_3$ and $CaCl_2$.

$$n(CaCl_2) = n(CaCO_3) = 0.030\text{ mol}$$

Convert the moles of $CaCl_2$ to grams:

$$M(CaCl_2) = 40 + (2 \times 35.5)$$

$$= 111\text{ g mol}^{-1}$$

$$\text{Theoretical yield} = 0.30 \text{ mol} \times 111 \text{ g mol}^{-1}$$
$$= 33.3 \text{ g}$$

Actual yield of $CaCl_2 = 22$ grams

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$
$$= \frac{22 \text{ g}}{33.3 \text{ g}} \times 100 = 66.07\%$$

Worked Example 3.16

In the preparation of aspirin ($C_9H_8O_4$), a student reacted 20 grams of salicylic acid ($C_7H_6O_3$) with excess acetic anhydride ($C_4H_6O_3$). The theoretical yield of aspirin is 24 grams. If the student obtained 18 grams of aspirin, calculate the percentage yield.

Solution:

Balanced chemical equation:

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2$$

Identify the theoretical yield and actual yield:

Actual Yield: 18 grams of aspirin ($C_9H_8O_4$)

Theoretical Yield: 24 grams of aspirin ($C_9H_8O_4$).

Use the formula for percentage yield:

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$
$$= \frac{18 \text{ g}}{24 \text{ g}} \times 100 = 75\%$$

REVIEW QUESTIONS

Review Questions 3.1

Calculate the empirical formula for the following compounds:

1. A compound contains 3.24 g of sulphur and 3.24 g of oxygen.
2. A sample of a compound contains 4.2 g of nitrogen and 12 g of oxygen.
3. A compound consists of 6.7 g of phosphorus and 8.5 g of oxygen.
4. Write the chemical formulae of the following compounds
   a. Potassium chloride
   b. Iron (II) bromide
   c. Copper (II) tetraoxosulphate (VI)

Review Questions 3.2

1. Calculate the percentage by mass of oxygen in water (H_2O).
2. Determine the percentage by mass of sodium in sodium chloride (NaCl).
3. Calculate the percentage by mass of calcium in calcium carbonate ($CaCO_3$).
4. Determine the percentage by mass of carbon in glucose ($C_6H_{12}O_6$).

Review Questions 3.3

1. A compound contains 54.5% carbon, 9.1% hydrogen, and 36.4% oxygen. The molar mass of the compound is 88 g/mol. Determine the empirical and molecular formulas.
   a. Convert Percentage to Mass (Assume 100 g of compound):

      Mass of C = ________ g

      Mass of H = ________ g

      Mass of O = ________ g
   b. Convert mass to moles:

      Moles of C = ________

Moles of H = ________

Moles of O = ________

c. Determine the simplest Ratio:

Moles of C = ________

Moles of H = ________

Moles of O = ________

d. Write the Empirical Formula:

Empirical formula =

e. Calculate the Empirical Formula Molar Mass:

Empirical formula molar mass = ________ g/mol

f. Determine the Ratio of Molar Masses:

Ratio = ________

g. Calculate the Molecular Formula:

Molecular formula =

2. An organic compound of relative molecular mass 46 on analysis was found to contain 52.0% carbon, 13.3% hydrogen and the remaining being oxygen. Determine its

a. Empirical formula

b. Molecular formula

[H=1.0, C=12.0, O=16.0]

3. Determine the percentage by mass of nitrogen and hydrogen in ammonia (NH_3).

[N = 14, H = 1]

Review Questions 3.4

Balance the following reactions:

1. $Fe + O_2 \rightarrow Fe_2O_3$

2. $N_2 + H_2 \rightarrow NH_3$

3. $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

4. $KClO_3 \rightarrow KCl + O_2$

5. $Mg + HCl \rightarrow MgCl_2 + H_2$
6. $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
7. $C_6H_{12}O_6 + O_2 \rightarrow CO_2 + H_2O$
8. $Al + HCl \rightarrow AlCl_3 + H_2$
9. Balance the following chemical equation and use it to answer the questions that follow:

   $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$

   (a) How many moles of water would be obtained from 4 moles of C_2H_6?

   (b) How many moles of C_2H_6 would be needed to produce 3 moles of water?

Review Questions 3.5

1. Write a balanced chemical equation for the reaction between Zn and HCl
2. Zinc metal reacts with Hydrogen Chloride according to the reaction $Zn + 2HCl \rightarrow ZnCl_2 + H_2$.

   Calculate the mass of zinc required to produce 2 moles of hydrogen gas. [A_r: Zn = 65]
3. Calculate the mass of CO_2 produced when 2 moles of propane (C_3H_8) reacts with excess Oxygen. [A_r: $C = 12$, $H = 1$, $O = 16$]
4. Balance the chemical equation $H_2 + O_2 \rightarrow H_2O$ and determine the mole ratio of Hydrogen to Oxygen in the reaction.
5. Balance the chemical equation $Al_{(s)} + Fe_2O_{3(aq)} \rightarrow Al_2O_{3(aq)} + Fe_{(s)}$ and determine the mole ratio of Fe_2O_3 to Fe in the reaction.
6. Consider the reaction, $HCl + NaOH \rightarrow NaCl + H_2O$. If 25cm^3 of 0.25 mol dm^{-3} HCl reacts completely with excess Sodium hydroxide, calculate the mass of Sodium Chloride produced.

   [A_r: **Na = 23, Cl = 35.5**]
7. Consider the reaction, $2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$. Calculate the volume of KOH of concentration 0.10 mol dm^{-3} required to completely neutralise 20 cm^3 of a 0.25 mol dm^{-3} H_2SO_4 solution.

Review Questions 3.6

1. In a reaction, 20 grams of Hydrogen gas (H_2) reacts with 10 grams of Oxygen gas (O_2) to produce water (H_2O). Determine which reactant is the limiting reagent and which is the excess reagent.

2. In the combustion of propane (C_3H_8), 40 grams of propane reacts with 100 grams of Oxygen gas (O_2).

   a. Determine which reactant is the limiting reagent and which the excess reagent is.

   b. Calculate the mass of Carbon dioxide (CO_2) produced.

   c. Calculate the mass of the excess reagent remaining after the reaction is complete.

   $[C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O]$

3. In the synthesis of ammonia (NH_3), 50 grams of Nitrogen gas (N_2) reacts with 20 grams of Hydrogen gas (H_2).

   a. Determine which reactant is the limiting reagent and which the excess reagent is.

   b. Calculate the mass of ammonia produced.

   c. Calculate the volume of nitrogen gas consumed at stp (standard temperature and pressure), assuming the reaction is complete.

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

Review Questions 3.7

1. In a laboratory experiment, 25 grams of Sodium Chloride (NaCl) are reacted with excess Silver nitrate ($AgNO_3$) to produce Silver Chloride (AgCl). If the theoretical yield of AgCl is 30 grams, and the actual yield obtained is 20 grams, calculate the percentage yield of the reaction,

   $[NaCl + AgNO_3 \rightarrow AgCl + NaNO_3]$

2. In the synthesis of aspirin, a student reacts 20 grams of salicylic acid ($C_7H_6O_3$) with 25 grams of acetic anhydride ($C_4H_6O_3$). The theoretical yield of aspirin is 24 grams. If the student obtains 18 grams of aspirin,

   (a) State the actual yield of the reaction.

   (b) Based on the actual yield obtained in (a), calculate the percentage yield of the reaction. $[C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2]$

3. A chemical reaction between Hydrogen gas (H_2) and Nitrogen gas (N_2) produces ammonia (NH_3). If 15 grams of Hydrogen gas reacts with excess Nitrogen gas to produce 25 grams of ammonia,

(a) State the actual yield of the reaction.

(b) The theoretical yield of ammonia in the reaction described in part (a) is 30 grams. Calculate the percentage yield of the reaction based on the actual yield obtained.

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

4. In the production of a pain relief medication, 50 grams of the starting material reacts with excess reagent to produce 70 grams of the medication. However, due to various inefficiencies, the actual yield obtained is 60 grams. Calculate the percentage yield of the reaction. Discuss the potential impacts of low yield on the cost and availability of the medication.

5. In the production of sulphuric acid (H_2SO_4), 100 grams of Sulphur dioxide (SO_2) reacts with excess oxygen to theoretically produce 120 grams of sulphuric acid. The actual yield obtained in the industrial process is 100 grams. Calculate the percentage yield of sulphuric acid. Consider the economic implications of yield on the industrial scale.

6. In a wastewater treatment plant, a reaction is designed to remove 80 grams of a contaminant. The theoretical yield of the process is 90 grams, but due to various factors, only 70 grams are actually removed. Calculate the percentage yield of contaminant removal. Discuss how efficiency in this process affects environmental health and compliance with regulations.

EXTENDED READING

Click on the links below for more information on some topics discussed

1. https://docbrown.info/page04/4_73calcs01ram.htm
2. https://docbrown.info/page04/4_73calcs09mvg.htm

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