MINISTRY OF EDUCATION
GHANA ASSOCATION OF
SCIENCE TEACHERS

Chemistry

for Senior High Schools

Year 2

- Alhassan Kuvidana Mustapha
- Christian Dzikunu
- Delali Robert Akplai
- Yirwelle Santus

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FOREWORD

Ghana's new Senior High School Curriculum aims to ensure that all learners achieve their potential by equipping them with 21st Century skills, knowledge, character qualities and shared Ghanaian values. This will prepare learners to live a responsible adult life, progress to further studies and enter the world of work. This is the first time that Ghana has developed a Senior High School Curriculum which focuses on national values, attempting to educate a generation of Ghanaian youth who are proud of our country and can contribute effectively to its development.

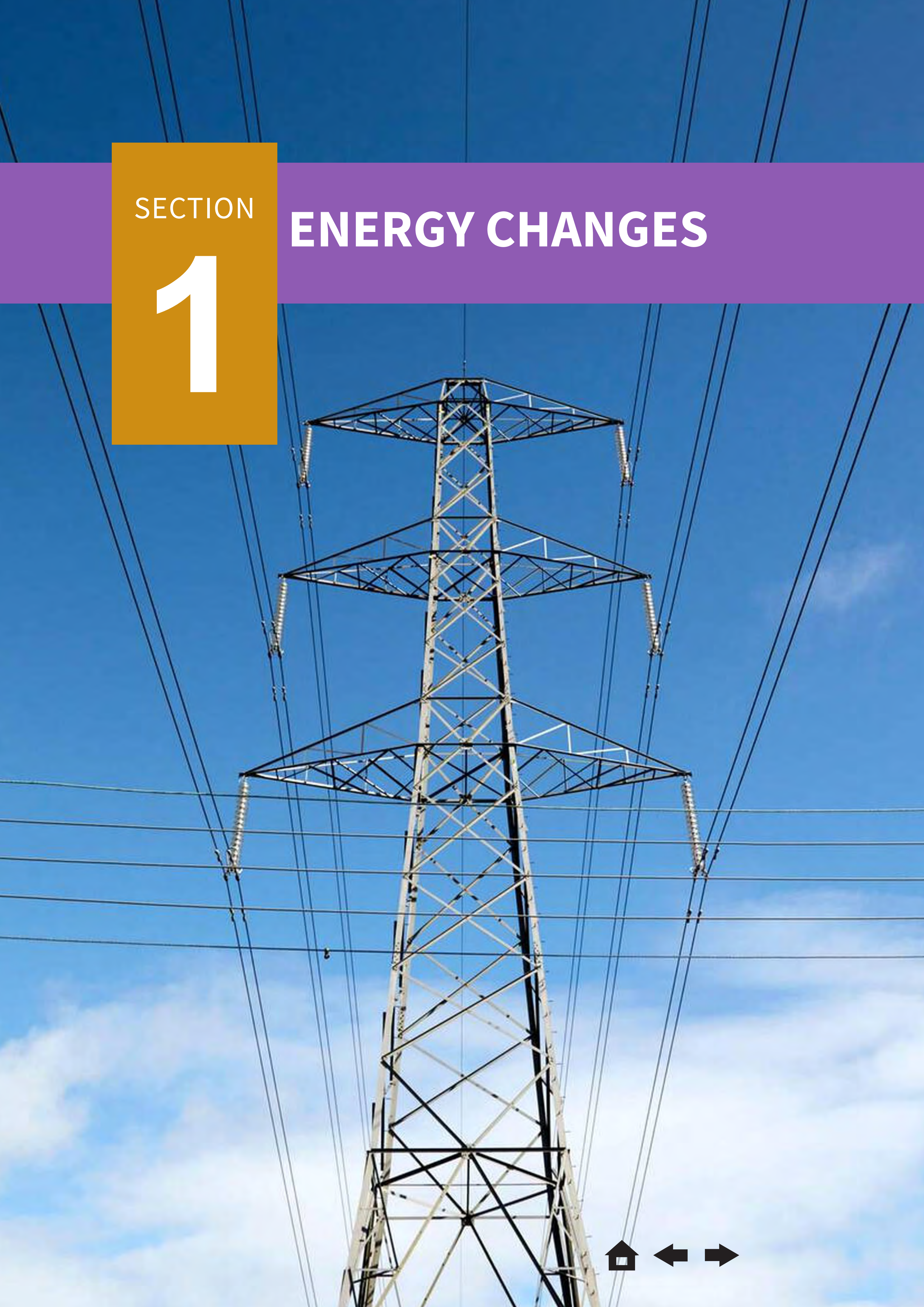
The Ministry of Education is proud to have overseen the production of these Learner Materials which can be used in class and for self-study and revision. These materials have been developed through a partnership between the Ghana Education Service, teacher unions (Ghana National Association of Teachers-GNAT, National Association of Graduate Teacher -NAGRAT and the Coalition of Concerned Teachers- CCT) and National Subject Associations. These materials are informative and of high quality because they have been written by teachers for teachers with the expert backing of each subject association.

I believe that, if used appropriately, these materials will go a long way to transforming our Senior High Schools and developing Ghana so that we become a proud, prosperous and values-driven nation where our people are our greatest national asset.

Haruna Iddrisu MP

Minister for Education

SECTION
1
ENERGY CHANGES



PHYSICAL CHEMISTRY

Matter and Its Properties

INTRODUCTION

In this section, you will learn about how energy changes during chemical reactions (called enthalpy) and how to figure out these changes using experiments. We will also explore Hess's Law, which helps you understand energy in a quantitative way. In addition, you will build important skills like critical thinking, problem solving, and scientific communication.

KEY IDEAS

- **Atomisation energy:** the energy required to break all the bonds in one mole of a substance to form individual atoms in the gas phase under standard conditions.
- **Bond dissociation energy** the energy required to break a specific chemical bond within a compound.
- **Bond enthalpy** is the average energy needed to split one mole of a specific type of covalent bond in a gaseous molecule.
- **Enthalpy change** is the enthalpy change within a reaction system.
- **Enthalpy** is a thermodynamic quantity which describes the energy of a system.
- **Hess's Law of constant heat summation** states that the total enthalpy change of chemical reaction is equal to the sum of all the individual enthalpy changes.
- **Lattice energy** the energy required to separate 1 mole of an ionic solid to gaseous ions.

ENTHALPY CHANGE

A **system** is the part of the Universe that we are focusing on in an experiment. Everything else around it is called the **surroundings**. There are three types of systems:

1. **Open system**: This can exchange both heat and matter with the surroundings. An example is heating water in an open beaker.
2. **Closed system**: This can exchange heat, but not matter, with the surroundings. An example is heating water in a covered beaker.
3. **Isolated system**: This cannot exchange heat or matter with the surroundings. An example is a well-insulated thermos (like a **calorimeter**). It should be noted that no system can be thought of as truly ideal, as some heat or matter will eventually escape to the surroundings. It is a term used to describe systems which will exchange a negligible fraction of heat/matter with their surroundings.

To measure how energy changes during a chemical reaction, scientists use a special quantity called **enthalpy (H)**. Enthalpy is a combination of the system's **internal energy (U)** and the product of its **pressure (P)** and **volume (V)**. Enthalpy has the same units as energy; Joules.

Mathematically, we write this as,

$$\mathbf{H = U + PV}$$

This means that **enthalpy** depends on things like the pressure and volume of the system. When the pressure stays the same, enthalpy is the amount of heat energy that is added or removed during the reaction. This is quite common in chemistry so is worth noting, an "isobaric" reaction.

The **enthalpy change (ΔH)** in a chemical reaction is found by comparing the enthalpy of the products (what we get after the reaction) with the enthalpy of the reactants (what we started with).

Enthalpy change (ΔH) is therefore defined as:

$$\Delta H = H_{products} - H_{reactants}$$

Endothermic and Exothermic Reactions

When we study how energy changes during chemical reactions, we look at what is called the **enthalpy of reaction (ΔH_{rxn})**. This tells us whether a reaction gives off heat or takes in heat.

1. **Exothermic reactions** are reactions that **release heat** to the surroundings, making things around them warmer. When this happens, the enthalpy of the system decreases as the products have less energy/enthalpy than the reactants, and ΔH_{rxn} is **negative**.

   In exothermic reactions the system ends up with a lower heat content than the surrounding.

   An example of an exothermic reaction is the combustion of methane

   $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890.4$ kJ/mol

2. **Endothermic reactions** are reactions that **absorb heat** from the surroundings, making things cooler. In these reactions, the enthalpy of the system increases as the products have more energy/enthalpy than the reactants, and ΔH_{xn} is **positive**.

   In endothermic reactions, the system has more energy than the surroundings because it took heat from them. The difference in energy comes from the heat system absorbs!

   An example of an endothermic reaction is the decomposition of calcium carbonate

   $CaCO_3(s) \rightarrow CaO(s) + CO_2(g) \quad \Delta H = +177.8$ kJ/mol.

   Exothermic and endothermic reactions can be represented by diagrams called Energy profile diagrams. See **Figure 1.1**.

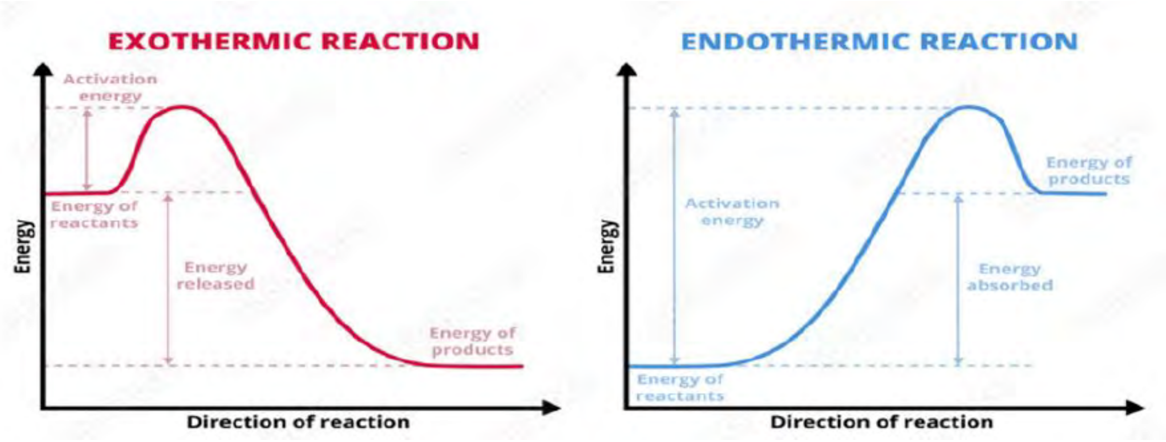


Figure 1.1: Energy profile diagrams for exothermic and endothermic reactions

Activity 1.1 Demonstrating the concepts of chemical systems (Open, Closed, and Isolated)

Materials needed: 3 identical plastic or glass bottles with lids, hot water, thermometers, towel (insulating material), stopwatch or timer, notebook pen and pencil

Steps

1. **Setup for the Demonstration**
   - **a.** **Bottle 1** (Open system): Pour the hot water into a bottle and leave the lid off.
   - **b.** **Bottle 2** (Closed system): Pour the hot water into a bottle and tightly seal it with the lid.
   - **c.** **Bottle 3** (Isolated system): Pour hot water into a bottle, seal it with the lid, and wrap it tightly with towel (insulating material).
2. **Measure Initial Temperature**

   Measure and write down the initial temperature of the hot water in each bottle.
3. **Prediction**
   - **a.** What do you think will happen to the temperature of the water in each bottle over time?
   - **b.** Which bottles will lose heat fast, slow, or not at all?
   - **c.** Why do you think the bottle wrapped in towel stayed warm longer?
4. **Observation**
   - **a.** Leave the bottles for 10–15 minutes. Check and record the temperature at every 5 minutes.
   - **b.** Feel in your palm whether the bottle feels warm or cool and whether heat seems to be escaping.
5. **Explain**
   - **a.** what you observe about the temperature change of the hot water in Bottle 1 (Open System).
   - **b.** what you observe about the temperature change of the hot water in Bottle 2 (closed System).
   - **c.** what you observe about the temperature change of the hot water in Bottle 2 (isolated System).

6. Give examples of real-life examples of open, closed, and isolated systems at home.

Activity 1.2 Exploring Enthalpy and Enthalpy change

Materials Needed: Chart paper and markers for group notes, access to a video on endothermic and exothermic reactions.

Steps

1. Your teacher will divide the class into small group (about 4–5 students per group).
2. Group discussion of some key terms
   - **a.** Activation Energy
   - **b.** Endothermic
   - **c.** Exothermic
   - **d.** Reactants
   - **e.** Products
   - **f.** Enthalpy
   - **g.** Energy
3. Use the link below to watch video on exothermic and endothermic reactions.

   https://www.youtube.com/watch?v=0MBVIXufFbM&t=2s

   After the video:
   - **a.** Give real-life examples of an exothermic reaction.
   - **b.** Give real-life examples of an endothermic reaction.
   - **c.** Explain happens to the surroundings in an endothermic reaction.
   - **d.** Explain happens to the surroundings in an exothermic reaction.
   - **e.** Where does the energy come from in an exothermic reaction?
   - **f.** Where does the energy come from in an endothermic reaction?
   - **g.** Discuss why understanding of energy changes in chemical reactions is important.

Activity 1.3 Understanding how to represent enthalpy profile diagrams

Materials Needed: Blank paper or graph paper, pencils, erasers, and rulers, markers or coloured pencils, reference charts or handouts explaining energy profile diagrams

Steps

1. Introduction to enthalpy profile diagrams
   - **a.** Refer to **Figure 1.1** for energy profile diagram for exothermic and endothermic reactions. Use these as a guide for constructing your own following the instructions below.
2. Draw an energy profile diagram for (Exothermic reaction) combustion of (CH_4), to produce carbon dioxide and water, releasing energy.
   - **a.** Draw and label vertical line representing enthalpy change (vertical 'y' axis).
   - **b.** Draw and label horizontal line representing reaction progress (horizontal 'x' axis).
   - **c.** Draw a horizontal line on the left for the **reactants** ($CH_4 + O_2$), representing a higher energy level.
   - **d.** Draw another horizontal line lower down on the right for the **products** ($CO_2 + H_2O$), showing that the products have less energy.
   - **e.** Show an upward curve, indicating the **activation energy (E_a)** (the energy needed to start the reaction).
   - **f.** Connect the product line with the peak of the activation energy curve.
   - **g.** Label the **enthalpy change (ΔH)** as negative, showing that energy was released.
3. Draw an energy profile diagram for (Endothermic reaction) like photosynthesis.
   - **a.** Draw and label vertical line representing enthalpy change (vertical 'y' axis).
   - **b.** Draw and label horizontal line representing reaction progress (horizontal 'x' axis).
   - **c.** Draw a horizontal line on the left for the **reactants** ($CO_2 + H_2O$) at a lower energy level.
   - **d.** Draw another horizontal line higher up on the right for the **products** (glucose + O_2), showing that the products have more energy.

   - **e.** Show an upward curve to represent the **activation energy (E_a)** required for the reaction to take place.
   - **f.** Connect the E_a curve with the product line.
   - **g.** Label the **enthalpy change (ΔH)** as positive, showing that energy was absorbed.

4. What are the differences between the exothermic and endothermic diagrams?
5. Where does the energy go/come from in each case?
6. Describe how you can see from a reaction profile that a reaction is exothermic or endothermic.

STANDARD ENTHALPY CHANGES

When a chemical reaction happens, energy is either taken in or given off, and we measure this energy change. But the amount of energy can change depending on the reaction conditions, like temperature and pressure. In order to compare processes, scientists use **standard conditions** to measure these energy changes. Standard conditions are used to measure and compare chemical reactions easily.

The Standard conditions are:

1. **Pressure**: 1 atmosphere (1 atm or 101.3 kPa)
2. **Temperature**: 298 K (which is the same as 25°C or room temperature).

When we measure energy changes under these conditions, we call it **standard enthalpy change**. It shows how much energy (heat) is gained or lost during a reaction when everything is at the same pressure and starting temperature. We usually measure this energy in **kilojoules per mole** (kJ/mol), and we use the symbol ΔH° to show standard enthalpy change.

Standard State: This means the most stable form of an element or compound at room temperature and 1 atmosphere of pressure. For example, oxygen's standard state is O_2 gas because that's how it exists in the air. For water, the standard state is liquid because that's its most stable form at room temperature.

For reactions in solutions, scientists also use a **standard concentration**, which is usually:

1.0 mol/dm³ (This means 1 mole of a substance is dissolved in 1 dm^3 of water).

Understanding Enthalpy Changes

1. If the **standard enthalpy change ($\Delta H°$)** is **negative ($\Delta H° < 0$)**, the reaction is **exothermic**, meaning it **releases heat** to the surroundings.
2. If the **standard enthalpy change ($\Delta H°$)** is **positive ($\Delta H° > 0$)**, the reaction is **endothermic**, meaning it **absorbs heat** from the surroundings.

We can calculate the standard enthalpy change by looking at the energy of the **products** and **reactants**. This helps scientists compare how much energy is involved in different reactions and is very important in chemistry and industry!

Types of standard enthalpy changes

1. Standard enthalpy changes of reaction

It is the quantity of heat involved when any valid stoichiometric molar quantities of reactants combine to form products under standard conditions. Any heat change measured under standard conditions is described as a standard enthalpy change of reaction. It is dependent on the stoichiometry of the balanced chemical equation. Changes in the balanced coefficients of reactants and products in any chemical reaction will change the value of its enthalpy changes accordingly. It has a unit of J/mol (or kJ/mol).

For example;

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g) \qquad \Delta H^o_{rxn} = -92.38\ kJmol^{-1}$$

$$\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightarrow 2NH_3(g) \qquad \Delta H^o_{rxn} = -46.19\ kJmol^{-1}$$

$$3N_2(g) + 9H_2(g) \rightarrow 6NH_3(g) \qquad \Delta H^o_{rxn} = -277.14\ kJmol^{-1}$$

Notice here this is the same reaction, the standard enthalpy of reaction changes based upon how much of the substances react. The value is not constrained in any way.

2. Standard enthalpy change of formation

The standard enthalpy change of formation of a compound is the enthalpy change those results when one mole of the compound is formed from its elements under standard conditions (298K and 101.3 kPa), all reactants and products being in their standard states. It is given the symbol ΔH^o_f

E.g., $C(s) + 2H_2(g) \rightarrow CH_4(g),\ \ \Delta H^o_f(CH_4) = -74.8\ kJ\,mol^{-1}$. The standard enthalpy of formation of methane is -74.8 kJmol^{-1}.

The standard enthalpy of formation of all elements in their standard states is **zero**.

Using oxygen as an example, it exists in the most stable form as a diatomic molecule (O_2) in the gaseous form as compared to the other allotropic forms such as ozone (O_3). As such $\Delta H_f^o(\mathrm{O_2}) = 0$ but $\Delta H_f^o(\mathrm{O_3}) = 142.2$ kJ/mol while ΔH_f^o $(O) = 249.4$ kJ/mol.

The most stable form of carbon under standard conditions is graphite, therefore

$\Delta H_f^o(\mathrm{C(graphite)}) = 0$

Table 1.1 describes some elements in their standard state; any other states will have associated enthalpies of formation.

Table 1.1: Elements in their standard state

Substance	ΔH_f^o	Substance	ΔH_f^o
$O_2(g)$	0	$C(graphite)$	0
$N_2(g)$	0	$S(rhombic)$	0
$H_2(g)$	0	$Ag(s)$	0
$F_2(g)$	0	$Ca(s)$	0
$I_2(g)$	0	$Cu(s)$	0
$Cl_2(g)$	0	$Hg(l)$	0

Standard enthalpies of formation values help in calculating the standard enthalpy of reaction ($\Delta H^o_{\mathrm{rxn}}$), defined as the enthalpy of a reaction carried out under standard conditions.

For example, consider the hypothetical reaction

$\mathrm{aA + bB \rightarrow cC + dD}$

where a, b, c, and d are stoichiometric coefficients. For this reaction, $\Delta H^o_{\mathrm{rxn}}$ is given by

$$\Delta H^o_{\mathrm{rxn}} = \left[c\,\Delta H_f^o\left(C\right) + d\,\Delta H_f^o\left(D\right)\right] - \left[a\,\Delta H_f^o\left(A\right) + b\,\Delta H_f^o\left(B\right)\right]$$

We can generalise

$$\Delta H^o_{\mathrm{rxn}} = \Sigma n\,\Delta H_f^o\left(products\right) - \Sigma m\,\Delta H_f^o\left(reactants\right)$$

where m and n are the stoichiometric coefficients for the reactants and products, respectively, and Σ (sigma) means "the sum of."

Standard enthalpy of formation values also gives a measure of the stability of compounds upon their formation. Compounds with negative ΔH_f^o values are more stable than compounds with positive ΔH_f^o values.

3. Standard enthalpy change of combustion

The standard enthalpy of combustion of a substance is defined as the enthalpy change that occurs when one mole of a substance is combusted completely in oxygen under standard conditions. (298K and 101.3 kPa), all reactants and products being in their standard states.

It is given the symbol ΔH_c^o

E.g., $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l), \quad \Delta H_c^o(H_2(g)) = -285.8\ \mathrm{kJ\,mol^{-1}}$.

The standard enthalpy of combustion of hydrogen is $-285.8\ \mathrm{kJ\,mol^{-1}}$.

Notice here that fractional values for stoichiometric co-efficients can be used to balance the equation. This ensures only 1 mole of the substance is combusted.

Knowledge of ΔH_c^o is applied in many fields including nutrition, industries, transportation and research.

Respiration and Combustion

Respiration is like a slow burning process in your body, where the food you eat gets "burned" to release energy that you need for daily activities. The amount of energy you get from food depends on what you eat. The main energy-giving foods are **carbohydrates** and **fats**.

In **nutrition**, the energy from these foods is measured in **Calories**. If you eat more carbohydrates than your body needs, the extra energy gets stored as **fat**. Eating too much can cause your body to store too much fat, which can lead to **obesity**.

In industries, the **enthalpy of combustion** is used to measure how efficiently machines like power plants or engines turn fuel into energy. Fuels with higher energy (higher ΔH_c^o) are better for creating power. It also helps compare different fuels and even different forms of the same material, like allotropes.

4. Standard enthalpy change of neutralisation.

The standard enthalpy change of neutralisation is the enthalpy change when solutions of an acid and an alkali react together under standard conditions to produce 1 mole of water. It is assigned the symbol ΔH_n^o .

When one mole of a monobasic acid reacts with one mole of a monobasic base under standard conditions the amount of energy released is almost the same as the enthalpy of formation of water from its ions.

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l) \quad \Delta H = -\ 57.5\ kJ/mol$

$HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)\ \Delta H = -\ 57.3\ kJ/mol$

Enthalpy of formation of water from its ions is represented by the equation;

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \quad \Delta H = -\ 57.4\ kJ/mol$

Strong acids and strong bases ionise completely in aqueous solution to give off hydrogen ions and hydroxyl ions therefore they tend to have higher enthalpies of neutralisation than the weak acid and weak bases (which do not fully ionise).

5. Standard enthalpy of solution and hydration.

The standard enthalpy of solution is the enthalpy change when one mole of an ionic substance is dissolved in excess water under standard conditions such that no further addition of water produces a further change in energy. It is assigned the symbol ΔH^o_{soln}.

e.g., $NaCl\ (s) + H_2O(l) \rightarrow Na^+\ (aq) + Cl^-\ (aq) \quad \Delta H = -\ 3.9\ kJ/mol$

Standard enthalpy of hydration is the enthalpy change when one mole of gaseous ions is dissolved in water to give one mole of aqueous ions and a solution of infinite dilution. It can also be defined as the energy evolved when one mole of a gaseous ion is surrounded, coordinated and stabilised by a sheath of water molecules under standard conditions.

E.g., $Na^+(g) + H_2O(l) \rightarrow Na^+\ (aq)\ \Delta H = -\ 406\ kJ/mol$

Activity 1.4 Understanding types of standard enthalpy changes

Materials: Computer, pen, calculator and Worksheets with sample calculations,

Step

1. Use the link below to watch the video on types of standard enthalpy changes.

   https://www.nagwa.com/en/videos/298198084514/

   Use the questions below to explore types of standard enthalpy changes.

2. Define and explain the following standard enthalpy changes:

  a. Standard enthalpy change of reaction (ΔH_r^o)

  b. Standard enthalpy change of formation (ΔH_f^o)

  c. Standard enthalpy change of combustion (ΔH_c^o)

  d. Standard enthalpy change of neutralisation (ΔH_n^o)

  e. Standard enthalpy change of solution (ΔH_{soln}^o)

  f. Standard enthalpy change of hydration (ΔH_{hyd}^o)

3. Give an example for each type of enthalpy change

  a. Standard enthalpy change of reaction (ΔH_r^o)

  b. Standard enthalpy change of combustion (ΔH_c^o)

  c. Standard enthalpy change of neutralisation (ΔH_n^o)

  d. Standard enthalpy change of solution (ΔH_{soln}^o)

  e. Standard enthalpy change of hydration (ΔH_{hyd}^o)

  f. Standard enthalpy change of formation (ΔH_f^o)

4. Calculating standard enthalpy change:

  a. Standard enthalpy change of reaction (ΔH_r^o)

  Formula: $\Delta H_r^o = \Sigma \Delta H_f^o(\text{products}) - \Sigma \Delta H_f^o(\text{reactants})$

  Let's calculate the enthalpy change for the reaction: $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

  Given:

  ΔH_f^o of HCl(g) = −92 kJ/mol

  ΔH_f^o of $H_2(g)$ and $Cl_2(g) = 0$ kJ/mol (because they are elements in their standard states).

  Calculation:

  $\Delta H_r^o = \Sigma \Delta H_f^o(\text{products}) - \Sigma \Delta H_f^o(\text{reactants})$

  $= [2(-92)] - [0 + 0] = -184\ kJ/mol$

  The reaction releases 184 kJ of energy.

  b. Standard enthalpy change of combustion (ΔH_c^o)

  Formula: $\Delta H_c^o = \Sigma \Delta H_f^o(\text{products}) - \Sigma \Delta H_f^o(\text{reactants})$

  Calculate standard enthalpy of combustion of methane (CH_4):

  Given:

  $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

ΔH_f^o of $CH_4(g) = -74.8$ kJ mol^{-1}, ΔH_f^o of $CO_2(g) = -393.5$ kJ mol^{-1},

ΔH_f^o of $H_2O(l) = -285.8$ kJ mol^{-1}

Using the equation below, substitute the values into the equation.

$\Delta H_c^o = \Sigma \Delta H_f^o(\text{products}) - \Sigma \Delta H_f^o(\text{reactants})$

5. Discuss the significance of standard enthalpy changes in various fields (e.g., chemistry, physics, engineering) and explore real-life examples and applications (e.g., energy production, materials science).

EXPERIMENTAL DETERMINATION OF ENTHALPY CHANGE

Experimental determination of heat of combustion of alcohol

Materials needed: Alcohol burner (containing ethanol), calorimeter (or a metal can), thermometer, Water (200 ml), balance, retort stand and clamp, heatproof mat.

Procedure

1. Measure initial mass of alcohol burner with the alcohol inside.
2. Add 200 ml of water to the calorimeter (or metal can).
3. Place the calorimeter on a stand above the alcohol burner, ensuring there is space for the flame to heat the water effectively.
4. Record initial temperature of water:
5. Light the alcohol burner and let the flame heat the water.

   Stir the water gently to ensure even heating.
6. Allow the alcohol to burn until the temperature of the water rises by about 15–20 °C.

   Extinguish the flame by capping the burner.
7. Record the final temperature of the water after extinguishing the flame.
8. Weigh the alcohol burner again to find the final mass of alcohol left in the burner.

Data

1. Initial mass of alcohol burner ($m_{initial}$ = x g).

2. Final mass of alcohol burner ($m_{final} = y$ g).
3. Initial temperature of water ($T_{initial} = T_i$).
4. Final temperature of water ($T_{final} = T_f$).
5. Mass of water = z g.

Calculations

1. **Calculate Mass of Alcohol Burned**

   Mass of alcohol burned = $(x - y)\ g$

2. **Calculate Temperature Change**:

   $\Delta T = T_f - T_i$

3. **Calculate Energy Transferred to Water**: Use the formula:

   $Q = m_{water} \cdot c \cdot \Delta T$

   Where,

   Q is the heat absorbed by the water (in joules),

   m_{water} is the mass of the water,

   c is the specific heat capacity of water (4.18 J/g°C),

   ΔT is the temperature change of the water.

4. **Calculate Heat of combustion of the alcohol**

   a. Convert Q from joules to kilojoules (kJ) by dividing by 1,000.

   b. Calculate the heat of combustion per gram of alcohol:

   $$\text{Heat of combustion per gram} = \frac{Q}{\text{mass of alcohol burned}}$$

   c. If you want the heat of combustion per mole, multiply by the molar mass of the alcohol.

   (for ethanol, C_2H_5OH, molar mass 46 g/mol)

Example Calculation

$m_{intial} = x = 120\ g$

$m_{final} = y = 119\ g$

$T_{intial} = T_i = 22°C$

$T_{final} = T_f = 42°C$

Mass of alcohol burned = $(x - y) = 120 - 119 = 1.0\ g$

$\Delta T = T_f - T_i = 42 - 22 = 20°C$

Energy transferred to water:

$$Q = m_{water} \cdot c \cdot \Delta T$$

$$= 200\ g \times 4.18\ \mathrm{J\,g^{-1}{}^{\circ}C^{-1}} \times 20^{\circ}\mathrm{C}$$

$$= 16{,}720\ \mathrm{J} = 16.72\ \mathrm{kJ}$$

$$\text{Heat of combustion per gram} = \frac{Q}{\text{mass of alcohol burned}}$$

$$= \frac{16.72}{1\ g}$$

$$= 16.72\ k\mathrm{J\ g^{-1}}$$

Heat of combustion per mole (for ethanol, C_2H_5OH, molar mass 46 g/mol

Heat of combustion per mole = $16.72\ \mathrm{kJ\,g^{-1}} \times 46\ \mathrm{g\ mol^{-1}} = 769.12\ \mathrm{kJ\ mol^{-1}}$

Experimental Discussion Points

1. Why Stir the Water?

   Stirring the water helps spread the heat evenly so that the thermometer shows a more accurate temperature. Without stirring, some parts might get hotter than others.

2. Comparing Your Results

   Compare the energy value you calculated with the "official" value scientists have found in a lab. Do you notice any differences?

3. Why Are the Values Different?

   The numbers might be different because some heat might have escaped to the surroundings instead of just heating the water.

Calculating % Difference

To see how close you got, calculate the percentage difference between your result and the official value. Here's how to do it:

$$\%\ \text{Difference} = \left(\frac{\text{Official value} - \text{Your value}}{\text{Official Value}}\right) \times 100$$

Improving the Experiment

How could we improve our experiment? Maybe we could insulate the setup better to trap more heat or make sure the flame is steady.

Standard enthalpy of Neutralisation

It is the heat released when one mole of an acid is neutralised by an alkali to produce one mole of water under standard conditions.

Experiment Title: *Determination of the Standard Enthalpy of Neutralisation for a Strong Acid-Base Reaction.*

In this experiment, we will find out how much heat is released when an acid reacts with a base to form water. This is called the **enthalpy of neutralization**. Neutralisation reactions happen between an acid and a base and always release heat.

Materials need: Hydrochloric acid (HCl) solution, sodium hydroxide (NaOH) solution, calorimeter (or a plastic cup in a beaker), thermometer, measuring cylinder, stirrer.

Procedure

1. Pour a specific mass (say, 50 g) of hydrochloric acid (HCl) into the calorimeter. Then, measure the same mass of sodium hydroxide (NaOH).
2. Measure and record the starting temperature of both the acid and base before mixing. They should both be at the same initial temperature.
3. Pour the sodium hydroxide into the hydrochloric acid. Stir the mixture gently record the highest temperature reached.
4. The temperature will rise because this reaction releases heat. The difference between the starting temperature and the highest temperature is the **temperature change**.

Calculating the Enthalpy of Neutralization

1. Find the heat released

   Use the formula:

   $Q = m \times c \times \Delta T$

   Where,

   Q is the heat released,

   m is the total mass of the solution,

   c is the specific heat capacity of water (4.18 J/g°C), here we will approximate the specific heat of the solution to that of water.

   ΔT is the temperature change of the solution.

2. Calculate enthalpy of neutralisation

   Divide the heat released by the number of moles of acid or base that reacted to find the enthalpy of neutralisation in kJ/mol.

Example

If you get a temperature rise of 6°C with 100 g of solution, plug it into the formula:

$Q = 100 \times 4.18 \times 6 = 2508$ J

Divide by the moles of acid or base used to get the enthalpy of neutralisation.

Calculations

Use the calorimeter's heat capacity to adjust for any heat lost or gained from the surroundings.

Find the moles of reactants using their mass, density and concentration ($n = CV$) ($V = m/p$).

Calculate the heat energy released per mole of water produced in the reaction.

Results and Analysis

Find the average enthalpy change for neutralisation from repeated trials.

Compare this value to the known value for strong acid-base reactions (-57 kJ/mol). Calculate the % difference and think about why they differ.

Consider any sources of error and ways to improve the experiment.

Discussion Questions

1. Why is it important to stir the solution regularly?
2. How could we reduce the % difference between our results and the known value?
3. Would a colder room temperature affect the results? Why?

Conclusion: This experiment helps us understand the energy released when a strong acid reacts with a base. By comparing our results with the known value, we can check the accuracy of our experiment and improve it for future trials.

Standard enthalpy of solution

It is the heat absorbed or released when one mole of an ionic substance is dissolved in excess water such that no further addition of water produces further heat change under standard conditions.

Experiment Title: *Determination of the Standard Enthalpy of Dissolution for a Solute*

In this experiment, we will measure the energy change (enthalpy) that occurs when a solid dissolves in water. This process helps us understand how much heat is absorbed or released when a substance mixes with water.

Procedure

1. Pour a known mass of water into a beaker and measure its temperature.
2. Add a known mass of the solid (like salt) to the water.
3. Stir the mixture and watch for temperature changes. Record the final temperature once it stops changing.
4. Subtract the starting temperature from the final temperature to find the temperature change.

Data Needed

Mass of water in the beaker

Mass of the solute (like salt)

Temperature change (final - initial temperature)

Specific heat capacity of water (4.18 J/g°C, a constant that tells us how much heat is needed to change water temperature). Here we will approximate the specific heat capacity of the solution to be that of water, for a more accurate result you could research the specific heat capacity of the concentration of solution you have made.

Calculation

Using the data, we can calculate the heat energy (Q) absorbed or released:

$Q = m \times c \times \Delta T$

Where:

Q is the heat released,

m is the total mass of the solution,

c is the specific heat capacity of water (4.18 J/g°C),

ΔT is the temperature change of the water.

Next, calculate the enthalpy change per mole by dividing Q by the moles of solute added ($n = \frac{m}{RAM}$).

Analysing results

1. Look at the official (theoretical) value for the solute and compare it to what we found our experiment. If they are not the same, that is okay! We will calculate the % *difference* to see how close they are.
2. Sometimes experimental results vary from official values due to small errors in setup or how we measured things.
3. Possible errors and improvements**:**
   - a. Stirring helps spread the heat evenly in the water, giving us a more temperature.
   - b. To improve accuracy, we could use a lid to reduce heat loss or ensure we measure everything carefully.

Conclusion

This experiment shows us the energy change when a substance dissolves in water. Some substances absorb heat (endothermic), while others release heat (exothermic). This energy change, or "enthalpy of dissolution," helps scientists understand how substances behave in solutions.

Activity 1.5 Experimental Determination of enthalpy changes

In small groups, explore and measure different types of energy (enthalpy) changes by carrying out experiments.

Steps

Your teacher will place you into groups.

Each group will do "circus experiments," rotating through stations to explore different enthalpy changes.

Experiments

1. **Enthalpy change of combustion**
   - **a.** Measure the energy released when burning alcohol (e.g., ethanol) and food (e.g., maize or groundnut).
   - **b.** Procedure**:**
     - i. Burn the substance under a container of water.
     - ii. Measure the temperature increase of the water to determine the enthalpy change.

2. **Enthalpy of Neutralisation**
   - **a.** Measure energy changes when an acid (HCl) reacts with a base (NaOH).
   - **b.** Procedure**:**
     - i. Mix the acid and base in a cup and measure the temperature change.
     - ii. Record the temperature before and after the reaction to find the enthalpy change.

3. **Enthalpy of Solution**
   - **a.** Observe the energy absorbed or released when dissolving substances like NH_4Cl and $CaCl_2$ in water.
   - **b.** Procedure:
     - i. Dissolve a measured amount of solute in water.
     - ii. Record the temperature change to determine the enthalpy change.

4. **Calculations**

   Use the formula $\Delta H = mc\Delta T$:

   m = mass of water/solution (in grams)

   c = specific heat capacity (for water, usually 4.18 J/g°C), for more accurate results remember to research the specific heat capacity of the solution you have created.

   ΔT = change in temperature

   Calculate the energy change (ΔH) for each experiment using the recorded temperature changes.

5. **Analysis and Interpretation**
   - **a. Interpret Results:** Discuss if each reaction absorbed or released heat.
   - **b. Presenting Results:** Draw an energy profile diagram for each experiment.
   - **c. Evaluate Experiments:** What challenges did you encounter? How could be improved? How reliable are you results?

Activity 1.6 Digital learning

1. Go online and find a short video explaining how scientists the energy in food (like calories) and fuels (like gasoline or wood).
2. The video should cover
   - a. how burning food or fuel releases energy.
   - b. why measuring energy helps us understand which foods give us more energy activities and which fuels are more powerful.

HESS'S LAW OF CONSTANT HEAT SUMMATION

Some compounds can't be easily made by mixing their basic elements because the reaction is either too slow or creates the wrong products through unwanted side reactions. In these cases, scientists use an indirect method to figure out the energy change, called ΔH_f°. This method is based on Hess's law.

Hess's law states that **if a reaction happens in several steps, the total energy change is the sum of the energy changes of each step**.

Hess's law helps to calculate the energy even if the reaction doesn't happen in one step.

Calorimetry can be used to measure enthalpy changes (ΔH) for many chemical processes, but not all. Consider the hypothetical reaction below:

$A + B \rightarrow D$

Where A and B are the reactants, D is the product

The overall energy change (ΔH) for a chemical reaction remains constant, regardless of the reaction pathway.

$A + B \rightarrow C \qquad \Delta H_1$

$C + B \rightarrow D \qquad \Delta H_2$

Then $A + 2B \rightarrow D\ \Delta H_3$

$\Delta H_3 = \Delta H_1 + \Delta H_2$

Consider the example below, the oxidation of carbon to form carbon monoxide.

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$

When carbon reacts with oxygen, the primary product is CO_2 even with insufficient oxygen is used. As soon as CO is formed, it reacts with O_2 to form CO_2. Because the reaction cannot be carried out in a way that allows CO to be the sole product, it is not possible to measure the change in enthalpy for this reaction by calorimetry.

The enthalpy changes for the reaction between C(s) and O_2 to form CO(g) can be determined indirectly, but, the enthalpy changes for other reactions for which values of ΔH_{rxn} can be determined.

The enthalpy change of the reaction can be calculated using Hess's law, which states that the total enthalpy change of chemical reaction is equal to the sum of all the enthalpy changes.

For example, the oxidation of carbon to carbon (IV) oxide.
This reaction occurs in two steps:
Step 1: the oxidation of C(s) to CO(g) (Equation 1) and

Step 2: the oxidation of CO(g) to CO_2(g) (Equation 2).

Adding the two equations to gives the overall equation for the oxidation of C(s) to CO_2(g) (Equation 3).

Equation 1: $C(s) + 1/2\ O_2(g) \rightarrow CO(g)\ ^\circ\ \Delta H_{rxn1} = ?$
Equation 2: $CO(g) + 1/2\ O_2(g) \rightarrow CO_2(g)\ \Delta H_{rxn2} = -283.0$ kJ/mol
Equation 3: $C(s) + O_2(g) \rightarrow CO_2(g)\ \ \Delta H_{rxn3} = -393.5$ kJ/mol

Hess's law says that the enthalpy change for the overall reaction that is equation 3 is the sum of the enthalpy changes for reactions that equations 1 and 2. these values are then used to calculate the enthalpy change for reaction 1.

$\Delta H_{rxn^\circ 3} = \Delta H_{rxn\ 1} + \Delta H_{rxn\ 2}$
$-393.5\ \text{kJ/mol} = \Delta H_{rxn\ 1} + (-283.0\ \text{kJ/mol})$
$\Delta_{Hrxn\ 1} = -110.5\ \text{kJ/mol}$

Remember that for these calculations it is imperative that the sign of the enthalpy change is preserved correctly during substitution and calculation of values.

Rules for Manipulating Thermochemical Equations

1. The opposite of an exothermic reaction is an endothermic reaction. As a result, the sign of an equation's enthalpy must likewise be reversed or negated when it is stated in the other direction. To illustrate, the reverse of the equation: A + B → C ΔH = +x J/mol is the following equation: C → A + B ΔH = -x J/mol

2. Only when a substance is in the same physical state can it be eliminated from both sides of an equation.
3. If all the coefficients of an equation are multiplied by a factor, the value of its enthalpy must be multiplied by the same factor.
    - i. To illustrate, the factors of an equation: $A + B \rightarrow C \ \Delta H = +x$ J/mol

      When an equation is multiplied by 2: $2A + 2B \rightarrow 2C \ \Delta H = 2\,(+x$ J/mol)
4. When equations are added, the enthalpy changes are also added accordingly with proper signs

   $A + B \rightarrow C \ \Delta H_1$

   $C + B \rightarrow D \ \Delta H_2$

   Then $A + 2B \rightarrow D \ \Delta H3 = \Delta H_1 + \Delta H_2$

Example

Using Hess's law to calculate the enthalpy change (ΔH_{rxn°) of the following reaction

C (graphite) + 2 $H_2(g) \rightarrow CH_4(g)$

Data given from the table below:

Reactions	**Δ*H***
(i) C (graphite) + $O_2(g) \rightarrow CO_2(g)$	ΔH_{f° (i) = – 393.5 kJ mol^{-1}
(ii) H_2 (g) + $\frac{1}{2}$ $O_2(g) \rightarrow H_2O(l)$	ΔH_{f° (ii) = – 285.8 kJ mol^{-1}
(iii) CH_4 (g) + 2 $O_2(g) \rightarrow CO_2(g)$ +2 $H_2O(l)$	ΔH_{rxn° (iii) = – 890.3 kJ mol^{-1}

Solution

C (graphite) + 2 $H_2(g) \rightarrow CH_4(g)$ ΔH_{rxn° (A) =?

$\Delta H_{rxn^\circ} = \Delta H_{f^\circ}$ (products) – ΔH_{f° (reactants)

(ΔH_{rxn°) = -890.3 kJ/mol

ΔH_{f° product = ΔH_{f° (CO_2) + 2 x ΔH_{f° (H_20)

= -393.5 + 2(-285.8)

= - 965.1

$\Delta H_f{}^\circ$ reactants = x + 0

-890.3 = -965.1 – (x)

X = -965.1 + 890.3

= - 74.8 kJ/mol

Alternatively, using the rearrangement method,

Maintain equation (i)

Multiply equation (ii) by 2

Reverse equation (iii)

(iv) $2H_2\,(g) + O_2(g) \rightarrow 2H_2O(l)$ ΔH_{rxn° (iv) = – 2×285.8 kJ mol^{-1} = -571.6 kJ mol^{-1}

(v) $CO_2(g) + H_2O(l) \rightarrow CH_4\,(g) + 2\,O_2(g)$ ΔH_{rxn° (v) = + 890.3 kJ mol^{-1}

equations (i) + (iv) + (v) with corresponding ΔH_{rxn° values, to get equation (vi)

(vi) $[C\,(graphite) + O_2(g)] + [2H_2\,(g) + O_2(g)] + [CO_2(g) + 2\,H_2O(l)] \rightarrow CO_2\,(g) + 2\,H_2O(l) + [CH_4(g) + 2O_2(g)]$

ΔH_{rxn° (vi) = ΔH_{rxn° (iv) + ΔH_{rxn° (v) + ΔH_{rxn° (i)

21 = (– 393.5 + (– 2×285.8 + 890.3) kJ mol–1

= –74.8 kJ mol^{-1}

cancel the common terms on the two sides of equation (vi), we get equation (vii)

(vii) $C\,(graphite) + 2H_2\,(g) \rightarrow CH_4(g)$ ΔH_{rxn° (vii) = –74.8 kJ mol^{-1}

BORN – HABER CYCLES

Hess' law is used to draw Born-Haber cycles, which are energy cycles for ionic molecules. Ionic compounds are created when constituent elements undergo a sequence of events for which, with the exception of the lattice energy, the majority of energy changes are directly measurable. A Born-Haber cycle is used to calculate lattice energy from the enthalpy data. Ionization energies, electron affinities, sublimation energies, bond dissociation energies, and their enthalpy of formation are all related to the lattice energies of ionic compounds.

Consider the reaction between lithium and fluorine:

$Li(s) + 1/2\ F_2(g) \rightarrow LiF(s)$

For this reaction, the standard enthalpy change is -594.1 kJ/mol. Five distinct steps can be used to track the production of LiF from its constituent elements, and the sum of the enthalpy changes for the steps equals the enthalpy change for the entire reaction (-594.1 kJ/mol). Using Hess's law, this pathway aids in the analysis of the energy shifts involved in the synthesis of ionic compounds.

1. Convert solid lithium to lithium vapor (the direct conversion of a solid to a gas is called sublimation):

   $Li(s) \rightarrow Li(g)\ \Delta H° = +155.2$ kJ/mol

   The energy of sublimation for lithium is 155.2 kJ/mol.
2. Dissociate $\frac{1}{2}$mole of F_2 gas into separate gaseous F atoms:

   $\frac{1}{2}F_2(g) \rightarrow F(g)\ \Delta H° = +\ 75.3$ kJ/mol

   The energy needed to break the bonds in 1 mole of F_2 molecules is +150.6 kJ.

   Here we are breaking the bonds in half a mole of F_2, so the enthalpy change is $\frac{150.6}{2}$, or 75.3 kJ.
3. Ionize 1 mole of gaseous Li atoms:

   $Li(g) \rightarrow Li^{+}(g) + e^{-}\ \Delta H° = +\ 520$ kJ/mol

   This process corresponds to the first ionization of lithium.
4. Add 1 mole of electrons to 1 mole of gaseous F atoms.

   $F(g) + e^{-} \rightarrow F(g)^{-}\ \Delta H° = -\ 328$ kJ/mol
5. Combine 1 mole of gaseous Li^{+} and 1 mole of F^{-} to form 1 mole of solid LiF:

   $Li^{+}(g) + F^{-}(g) \rightarrow LiF(s)\ \Delta H° = ?$

   The reverse of step 5,

   $\text{energy} + LiF(s) \rightarrow Li^{+}(g) + F^{-}(g)$

   According to Hess' law

   $\Delta H^{o}(overall) = \Delta H_1^{o} + \Delta H_2^{o} + \Delta H_3^{o} + \Delta H_4^{o} + \Delta H_5^{o}$

   $-594.1 = 155.2 + 75.3 + 520 - 328 + x$

   $x = -594.70 - 422.5$

   $x = -1016.6\ kJ/mol$

BOND ENERGY

One or more pure compounds can be transformed into one or more other chemical substances through chemical reactions. In most cases, chemical bonds are produced in the products and broken in the reactants. These bond-breaking and bond-forming mechanisms are primarily responsible for the enthalpy change for any chemical reaction. A chemical bond must be broken with energy, and when a chemical bond is formed energy is liberated. Bond formation is an exothermic activity that releases energy, whereas bond breaking is an endothermic process that absorbs energy. Consequently, estimating the total enthalpy change in a chemical reaction can be aided by knowledge of the energies of the individual bonds.

The average energy needed to break one mole of a specific type of covalent bond in a molecule is known as the bond enthalpy. The energy absorbed to break one mole of a specific kind of covalent bond in a gaseous molecule into free radicals under standard circumstances is known as the standard bond dissociation energy.

The enthalpy shift that occurs when a compound's bonds are broken and its constituent atoms split apart into single atoms is known as the enthalpy of atomization. It can alternatively be described as the energy shift that takes place when a compound's chemical bonds are broken, producing one mole of gaseous atoms as by-products.

Simple diatomic molecules like H_2, O_2, and Cl_2 often have their bond enthalpies determined spectroscopically; however, complex compounds use Hess's law to calculate bond enthalpies based on thermochemical data.

The experimentally determined bond enthalpy of the diatomic hydrogen molecule, for example, is $H_2(g) \rightarrow H(g) + H(g)\ \Delta H° = +\ 436.4$ kJ/mol

According to this equation, breaking the covalent bonds in 1 mole of gaseous H_2 molecules requires 436.4 kJ of energy.

For the less stable chlorine molecule, $Cl_2(g) \rightarrow Cl(g) + Cl(g)\ \Delta H° = +\ 242.7$ kJ/mol.

Bond enthalpies can also be directly measured for heteronuclear diatomic molecules, such as HCl, $HCl(g) \rightarrow H(g) + Cl(g)\ \Delta H° = +\ 431.9$ kJ/mol

as well as for molecules containing multiple bonds:

$O_2(g) \rightarrow O(g) + O(g)\ \Delta H° = +\ 498.7$ kJ/mol

The table below shows the bond enthalpies of some diatomic molecules

Bond	Bond Enthalpy (kJ/mol)
H – H	+436.4
O = O	+498.7
Cl – Cl	+242.7
H – O	+460

$$\Delta H_{rxn^\circ} = \sum \Delta H^o \, bonds\ broken - \sum \Delta H^o \, bonds\ formed$$

Example

Calculate the enthalpy of reaction when one mole of water is formed from hydrogen gas and oxygen gas.

[Bond enthalpies in kJ/mol are H – H = +436, O = O = +498, H – O = +464]

Solution

$H_2 + \frac{1}{2}O_2 \rightarrow H_2O$

Bonds broken: one mole of (H – H) and half of (O = O)

$= 436 + \frac{1}{2}(498)$

$= +685$

Bonds formed: H – O – H which is twice (H – O)

$= 2(464)$

$= +928$

$\Delta H_{rxn^\circ} = \sum \Delta H^o \, bonds\ broken - \sum \Delta H^o \, bonds\ formed$

$\Delta H_{rxn^\circ} = 685 - 928$

$\Delta H_{rxn^\circ} = -\ 243\ kJ/mol$

Activity 1.7 Hess's Law and its applications

Combustion reaction (combustion of methane)

1. In your own understanding, explain Hess's law of constant heat summation and underlying principles.
2. In groups of five:
   - **a.** Write a balance equation for the reaction between carbon and hydrogen to form methane. ΔH for reaction is -74.8 kJ/mol
   - **b.** Break the methane into its constituent atoms and balanced the equation.
   - **c.** Write a balanced equation for the reaction between the carbon atom and oxygen to form carbon (iv) oxide. ΔH for the reaction is -393.5 kJ/mol
   - **d.** Write a balanced equation for the reaction between the hydrogen atom and oxygen to form water. ΔH for the reaction is -571.6 kJ/mol
   - **e.** Combine equations ii, iii and iv and obtain the enthalpy of reaction for the reaction.

Activity 1.8 Calculating enthalpy change of reaction using Hess law

1. Use the information provided below to draw energy diagram for each reaction:
   - **a.** Reaction between carbon and hydrogen to form methane ΔH = -74.8 kJ/mol
   - **b.** Reaction between carbon and oxygen to form carbon (iv) oxide ΔH = -393.5 kJ/mol
   - **c.** Reaction between hydrogen and oxygen to form water ΔH = -571.6 kJ/mol
2.
   - **a.** Write a balanced reaction equation for the atomisation of nitrogen gas, the enthalpy change for this reaction is -45.9 $kJmol^{-1}$
   - **b.** write a balanced equation for the atomisation of hydrogen gas, the enthalpy change this reaction is -436.4 $kJmol^{-1}$
   - **c.** write a balanced reaction equation for the synthesis of ammonia from nitrogen gas and hydrogen gas, the enthalpy change for the reaction is -389.5 $kJmol^{-1}$

d. calculate the enthalpy change of formation of ammonia from the equations you have formed in iii above using Hess's law.

Activity 1.9 A diagram of Born Haber cycle

1a.

- **i.** Write the atomisation energy of sodium and fluorine
- **ii.** Write the ionisation energy of sodium and electron affinity of fluorine
- **iii.** Find the enthalpy of formation of sodium fluoride
- **iv.** Determine the lattice energy of the sodium fluoride

1b.

- **i.** Write the symbols of sodium and fluorine in their standard states
- **ii.** Draw vertical lines on each of the sodium and fluorine.
- **iii.** At the end of the lines, write the products of the sublimation of sodium and atomisation of fluorine.
- **iv.** Extend the lines and indicate the products of the ionisation of sodium and electron affinity of fluorine
- **v.** Draw lines to show the enthalpy of formation and the lattice energy of sodium fluoride.

2. a. In groups discuss bond energy and its importance in chemical reaction

b. Present your findings to the whole class.

Activity 1.10 Calculations of bond enthalpy of some molecules

Study the table below

Bond	Bond enthalpy (kJ/mol)
H-H	+436.4
N-H	+389
Cl-Cl	+242.7
C-H	+413
C=C	+611
N≡N	+945.4
C-C	+345
C≡C	+837

a. Write a balance equation for the formation of ammonia.

b. Identify the bonds formed and the bonds brokens in the process.

c. From the enthalpy values provided in the table above, calculate the enthalpy of reaction when one mole of ammonia is formed from hydrogen gas and nitrogen gas.

d. Write the equations for the formation of methane, ethene and ethyne.

e. Calculate the bond enthalpy of each of methane, ethene and ethyne.

f. Compare the bond enthalpy of methane, ethene, ethyne and relate it to their bond strength.

EXTENDED READING

1. General, Organic, *and* Biochemistry/ Katherine J. Denniston, Joseph J. Topping, Danaè R. Quirk Dorr, Robert L. Caret. Tenth Edition. (Pg: 230-238)
2. Watch videos or use virtual laboratory to reinforce the measurement of energy content of food and fuels.
    - Investigating The Energy Content of Food | Chemistry Practical - YouTube
    - https://youtu.be/ZrB1VV3ZGqQ
    - https://youtu.be/CkUUj3RGk9A
3. Watch a video of Born-Haber cycle from these link:
    - https://youtu.be/DI-bwx8r3r8
    - https://youtu.be/E6HyNypXR5Q

REVIEW QUESTIONS

Review Questions 1.1

1. Identify the basic difference between exothermic and endothermic reactions in terms of energy transfer.
2. a. What is the standard enthalpy change of a reaction?
   b. Name two types of standard enthalpy changes.
3. Describe a system in thermochemistry and differentiate between an open, closed, and isolated system in terms of energy exchange.
4. If burning 1 mole of methane (CH_4) produces -890 kJ/mol of energy, is this reaction exothermic or endothermic?
5. Draw an energy profile diagram for an exothermic reaction and label reactants, products, activation energy (E_a), and enthalpy change (ΔH).
6. The combustion of propane (C_3H_8) releases −2,220 kJ/mol. How much energy is released when 2 moles of propane are burned?
7. Use the table to answer this question

Reaction	Starting temperature/ °C	Final temperature/°C
X	22	37
Y	25	18
Z	21	25

   a. Decide whether each reaction is endothermic or exothermic, explain how you could tell.
   b. Which reaction has the largest energy change?
8. Why do exothermic reactions feel warm while endothermic reactions feel cold?
9. Calculate the standard enthalpy change of a reaction if the enthalpy of the products is −500 kJ and the enthalpy of the reactants is -200 kJ.
10. a. Design an experiment to determine if a given reaction is exothermic or endothermic and interpret the observed energy changes.

b. When hydrochloric acid reacts with ammonium hydroxide in a beaker, the temperature goes up.

$HCl(aq) + NH_4OH\ (aq) \rightarrow NH_4Cl\ (s) + H_2O\ (l) \Delta H = –53.4 kJ/mol$

Draw the energy profile diagram and state whether the reaction is endothermic or exothermic, explain your answer.

11. Explain why fuels like methane (CH_4) are preferred for energy production over fuels with lower enthalpy of combustion.

12. You are given the following standard enthalpy changes:

a. $\Delta H_f^o(CH_4) = -74.8$ kJ/mol

b. $\Delta H_f^o(CO_2) = -393.5$ kJ/mol

c. $\Delta H_f^o(H_2O) = -241.8$ kJ/mol

The balanced combustion reaction for methane (CH_4) is:

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

a. Using the given enthalpy of formation values, calculate the standard enthalpy change of combustion (ΔH_c^o) for methane.

b. Explain how the result shows whether this reaction is exothermic or endothermic.

c. Discuss why knowing the enthalpy of combustion is important in real-life applications such as energy production.

Review Questions 1.2

1. What is the enthalpy change of combustion?

2. What is the enthalpy change of neutralisation?

3. What is the enthalpy change of solution?

4. Why is the enthalpy change of neutralisation usually the same for strong acids and bases?

5. Why is it important to know the enthalpy change of combustion for fuels?

6. Why does dissolving ammonium chloride in water cause the solution to cool?

7. How can you experimentally determine the enthalpy change of combustion of ethanol?

8. Describe how you would determine the enthalpy change of neutralisation between HCl and NaOH in a lab setting.
9. Describe how you would determine the enthalpy change of solution for a salt like KCl in the lab.
10. Compare your experimental enthalpy change of combustion for ethanol to the theoretical value. Discuss any differences and possible sources of error.
11. If the experimental enthalpy change of neutralisation differs from the theoretical value, what factors might explain this?
12. Compare your experimental enthalpy change of solution for KCl with the literature value and discuss any discrepancies. How could the experiment be improved?
13. In an experiment to determine the enthalpy change of combustion of ethanol, the initial mass of the ethanol burner (with ethanol in it) was measured as 150.0 g. Next, 100.0 g of water (equivalent to 100 ml) was added to a calorimeter, and the initial temperature of the water was recorded as 25.0 °C. The ethanol burner was placed under the calorimeter and lit. As the ethanol burned, the heat produced caused the temperature of the water to increase. After the flame was extinguished, the final temperature of the water was recorded as 45.0 °C. The final mass of the ethanol burner, after the combustion process, was measured to be 148.50 g. The specific heat capacity of water used in the calculations is 4.18 J/g°C. Determine;
    a. the amount of heat absorbed by the water.
    b. the mass of ethanol burned.
    c. the enthalpy changes of combustion per gram of ethanol.
    d. the molar enthalpy changes of combustion of ethanol (molar mass of ethanol is 46.0 g/mol).
    e. any two precautions you have to take in order to obtain correct result.

Review Questions 1.3

1. State Hess's law of constant heat summation.
2. Given the following information:

$P_4(s) + 3O_2(g) \rightarrow P_4O_6(s)\ \Delta H = -1640.1\ \text{kJ/mol}$

$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)\ \Delta H = -2940.1\ \text{kJ/mol}$

what is the value of $\Delta H_{(rxn)}$ for $P_4O_6(s) + 2O_2(g) \rightarrow P_4O_{10}(s)$?

3. Consider the following data:

$C(s) + O_2(g) \rightarrow CO_2(g)\ \Delta H = -394\ kJmol^{-1}$

$H2(g) + 1/2O_2(g) \rightarrow H_2O(l)\ \Delta H = -286\ kJmol^{-1}$

$C_4H_6(g) + 5/2O_2(g) \rightarrow 4CO_2(g) + 3H_2O(l)\ \Delta H = -254\ kJmol^{-1}$

Determine the standard enthalpy of formation of C_4H_6.

4. **a.** Explain the term lattice energy

b. Use the following data to calculate for the lattice energy of sodium chloride:

$\Delta H_f(NaCl) = -\ 411\ kJmol^{-1}$

$\Delta H_{atom}(Na) = +107\ kJmol^{-1}$

$\Delta H_{atom}(Cl) = +122\ kJmol^{-1}$

$IE(Na) = +\ 494\ kJmol^{-1}$

$E.A.\ (Cl) = -\ 349\ kJmol^{-1}$

Calculate lattice energy of NaCl(s)

5. Define bond enthalpy

6. Distinguish between bond dissociation energy and bond energy.

7. Consider the following bond enthalpies

Bond	BDE (KJ/mol)
H-C	436
C-C	348
C=C	612
N≡N	944
N-H	388

a. Write a balanced chemical equation for the formation one mole of NH_3.

b. Use the bond enthalpy data to calculate for the enthalpy of formation of NH_3.

SECTION

2

CHEMICAL KINETICS

PHYSICAL CHEMISTRY

Matter and Its Properties

INTRODUCTION

In this section, you will explore Chemical Kinetics. You will learn how to measure reaction rates using experiments and understand factors that affect rates. You will also study collision theory and how to use the concept of collisions to explain reaction rates and create rate equations from experiment results. Through problem-solving and analysing graphs, you will learn how to figure out reaction orders and solve rate problems.

KEY IDEAS

- **Average rate of reaction:** this is the average speed of a reaction over a specific time interval. Is calculated by dividing the change in concentration of a reactant or product by the time elapse
- **Catalyst:** a substance that increase the rate of reaction without being consume in the process.
- **Initial rate of a reaction:** the rate of reaction at a very beginning, at time $t = 0$
- **Instantaneous rate of a reaction:** this is the rate of reaction at a specific instant in time.
- **Rate of reaction** is a measure of how quickly a chemical reaction proceeds. This is normally measured by either tracking the reactant being consumed or product being formed

RATE OF REACTION

Chemical kinetics is the area of chemistry that studies how fast chemical reactions take place. This is called the reaction rate, and is important for things like creating medicines, reducing pollution, and processing food. Knowing reaction rates helps make these processes more efficient.

Rate of a chemical reaction is defined as the change in concentration (or moles/mass) of a reactant or product per unit time.

Mathematically expressed as,

$$\text{Rate} = \frac{\Delta[\text{A}]}{\Delta\text{t}},$$

where A is the reactant or product being considered,

the square brackets indicate concentration in mol/dm^3.

the symbol Δ indicates a change in a given quantity.

Consider the hypothetical reaction: $a\text{A} + b\text{B} \rightarrow c\text{C} + d\text{D}$

The rate can be expressed by the following equation

$$\text{rate} = -\frac{1}{a} \times \frac{\Delta[\text{A}]}{\Delta\text{t}} = -\frac{1}{b}\frac{\Delta[\text{B}]}{\Delta\text{t}} = \frac{1}{c}\frac{\Delta[\text{C}]}{\Delta\text{t}} = \frac{1}{d}\frac{\Delta[\text{D}]}{\Delta\text{t}}$$

Notice the negative symbols in front of the reactant stoichiometric coefficients too denote a consumption.

Example 1

Consider the thermal decomposition of calcium carbonate ($CaCO_3$) into calcium oxide (CaO) and carbon dioxide (CO_2) gas through the following equation:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Write expressions for the reaction rate in terms of the rate of change in concentration of the reactant ($CaCO_3$) and the product (CO_2)

Solution

From the balanced chemical equation above shows two moles of $CaCO_3$ decompose to produce 1 mol of CO_2 and 1 mole of CaO. From the equation the mole ratios of CO_2 to $CaCO_3$ and to CaO are 1:1 and 1:1, respectively. This means that the rate of change of $[CaCO_3]$ and $[CO_2]$ must be divided by its stoichiometric coefficient to obtain equivalent expressions for the reaction rate.

$$\text{rate} = \frac{\Delta[CO_2]}{\Delta\text{t}} = \frac{\Delta[CaCO_3]}{\Delta\text{t}} = \frac{\Delta[CaO]}{\Delta\text{t}}$$

Example 2

Consider the thermochemical equation: $2N_2O_5\ (g) \rightarrow 4NO_2(g) + O_2(g)$

a. How is the rate at which N_2O_5 disappears related to the rate at which NO_2 appears in the reaction?

b. If the rate at which NO_2 appears is $6.0 \times 10^{-5} moldm^{-3}$ at a particular instant, at what rate is N_2O_5 disappearing at this same time the rate at which NO_2 appears in the reaction.

Solution

$$2N_2O_5\ (g) \rightarrow 4NO_2(g) + O_2(g)$$

a. $rate = -\frac{\Delta[N_2O_5]}{2 \times \Delta t} = \frac{\Delta[NO_2]}{4 \times \Delta t}$

$$\frac{\Delta[N_2O_5]}{\Delta t} = \frac{2}{4} \times \frac{\Delta[NO_2]}{\Delta t}$$

$$\frac{\Delta[N_2O_5]}{\Delta t} = \frac{1}{2} \times \frac{\Delta[NO_2]}{\Delta t}$$

So, the rate at which N_2O_5 is consumed is equal to half the rate at which NO_2 is created.

b. But the rate of appearance of NO_2 $\left(\frac{\Delta[NO_2]}{\Delta t}\right)$ is $6.0 \times 10^{-5} moldm^{-3}$

= the rate which N_2O_5 disappears

$$\left(\frac{\Delta[N_2O_5]}{\Delta t}\right) = \frac{2}{4} \times 6.0 \times 10^{-5}$$

$$= 3.0 \times 10^{-5}\ mol\,dm^{-3}$$

Exercise

Consider the reaction: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

Write expressions for the reaction rate in terms of the rate of change of the concentration of each species. If the rate of consumption of H_2 in this process is $3 \times 10^{-4}\ mol\,dm^{-3}$ calculate the rate of consumption of O_2 and the rate of appearance of H_2O.

Initial rate of a reaction

When the reagents are first combined, that is the starting rate of reaction. The initial rate can be determined graphically or empirically. It is necessary to combine the chemicals and evaluate the rate of reaction as soon as possible in order to experimentally ascertain the beginning rate.

Measuring the time, it takes for a specific event to occur early in a reaction is a common practice in initial rate experiments. This could be the amount of time needed for a colour shift, the formation of a tiny, detectable amount of precipitate, or the production of a measurable volume of gas.

Average rate of reaction

The change in reactant or product concentration over time period is known as the average rate of reaction for that time period. The average rate of reaction can be determined choosing two points in time say t_1 *and* t_2 and measuring the concentration of the reactants (C_1 *and* C_2) at those points. The average rate of reaction for the reaction (a.v.g rate) $= \frac{\Delta[\text{reactants}]}{\Delta t}$

$$\text{a. v. g rate} = \frac{C_2 - C_1}{t_2 - t_1}$$

The average rate of reaction can also be expressed from a balanced chemical equation taking into consideration the stoichiometric coefficients. Consider chemical reaction below:

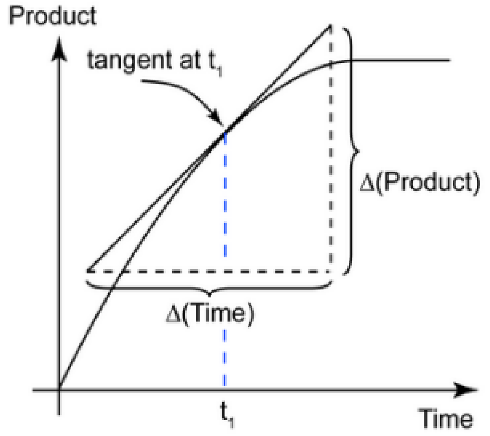
$$aA + bB \rightarrow cC + dD$$

$$\text{a.v.g rate} = -\frac{1}{a}\frac{\Delta[A]}{\Delta t} = -\frac{1}{b}\frac{\Delta[B]}{\Delta t} = \frac{1}{c}\frac{\Delta[C]}{\Delta t} = \frac{1}{d}\frac{\Delta[D]}{\Delta t}$$

Instantaneous rate of a reaction

The instantaneous rate of a reaction is the reaction rate at any given point in time. As the period of time used to calculate an average rate of a reaction becomes shorter and shorter, the average rate approaches the instantaneous rate.

Suppose we have a reaction that forms a product A. A graph of the concentration of A [A] against time t is usually not a straight line but a curve. The instantaneous rate of reaction is the slope of the line (the tangent to the curve) at any time (t).



Instantaneous rate $= \frac{\Delta[P]}{\Delta t}$(where Δt is small)

Measurement of rate of reaction by colour change

Sodium thiosulfate solution reacts with dilute hydrochloric acid according to the reaction below:

$Na_2S_2O_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$

During the reaction, sulphur produces a cloudy yellow-white precipitate. The time it takes for this precipitate to reach a specific level of cloudiness can be used to measure the reaction time. From this time the rate can be determined.

Procedure

SAFETY NOTE: Take care not to inhale any fumes given off in this reaction as $SO_2(g)$ can cause respiratory irritation.

1. Using a measuring cylinder, add 50 cm^3 of dilute sodium thiosulfate solution to a conical flask.
2. Place the conical flask on a piece of white paper with a black cross drawn on it.
3. Using a different measuring cylinder, add 10 cm3 of bench dilute hydrochloric acid to the conical flask. Immediately swirl the flask to mix its contents, and start a stop clock.
4. Look down through the reaction mixture, ensure you observe from a height of more than 20cm to reduce exposure to fumes. When the cross can no longer be seen, record the time on the stop clock.
5. Measure and record the temperature of the reaction mixture, and clean the apparatus.

6. Ensure the still reacting mixture is poured into the alkaline bath to stop the reaction continuing.
7. Repeat steps 1 to 6 with different starting temperatures of sodium thiosulfate solution.
8. Plot a graph of reaction rate on the vertical axis and temperature (°C) on the horizontal axis and draw a curve of best fit.

Measurement of rate of reaction by changes in volume of gas evolved

Calcium carbonate reacts with dilute hydrochloric acid:

$CaCO_3\,(s) + 2HCl\,(aq) \rightarrow CaCl_2\,(aq) + H_2O\,(l) + CO_2\,(g)$.

The volume of carbon dioxide gas produced can be measured using a gas syringe.

Procedure

1. Support a gas syringe with a stand, boss and clamp.
2. Using a measuring cylinder, add 50 cm3 of dilute hydrochloric acid to a conical flask.
3. Add 0.4 g of calcium carbonate to the flask. Immediately connect the gas syringe and start a stop clock.
4. Every 10 seconds, record the volume of gas produced.
5. When the reaction is complete, clean the apparatus.
6. Repeat steps 1 to 5 with different concentrations of hydrochloric acid.
7. Plot a graph of volume of gas (cm3) on the vertical axis and time (s) on the horizontal axis for each concentration of acid. You may wish to plot this data simultaneously on the same axes.
8. Draw a curve of best fit for each concentration.

Activity 2.1 Defining reaction rate and explaining its units

1. Define the term rate of reactions in your own words and explain its units.
2. Share with a friend your definition and idea the rate of reaction.
3. Which definition should be accepted and why?
4. Explain with examples the various ways initial rate, average rate and instantaneous rate can be written.

5. Tell your group the difference between instantaneous and average rates of reactions.
6. What can slow down or hasten the rate of reaction.
7. Share your definitions and ideas with the class and agree on definitions and to express the types of rates of reaction.
8. The examples will guide you on how to express both types of rates.

E → D. the concentration and time for the completion of the reaction are shown the table.

Time (s)	[E]/moldm^{-3}
0	1.00
10	0.95
20	0.90
30	0.85

The instantaneous rate at t= 20s is calculated as:

$\Delta[A] = \Delta[A]_1 - \Delta[A]_2 = 0.90 - 0.95 = -0.05M$

$\Delta t = t_2 - t_1 = 20 - 10 = 10s$

Instantaneous rate = (-0.05)/10s = 0.005 M/s

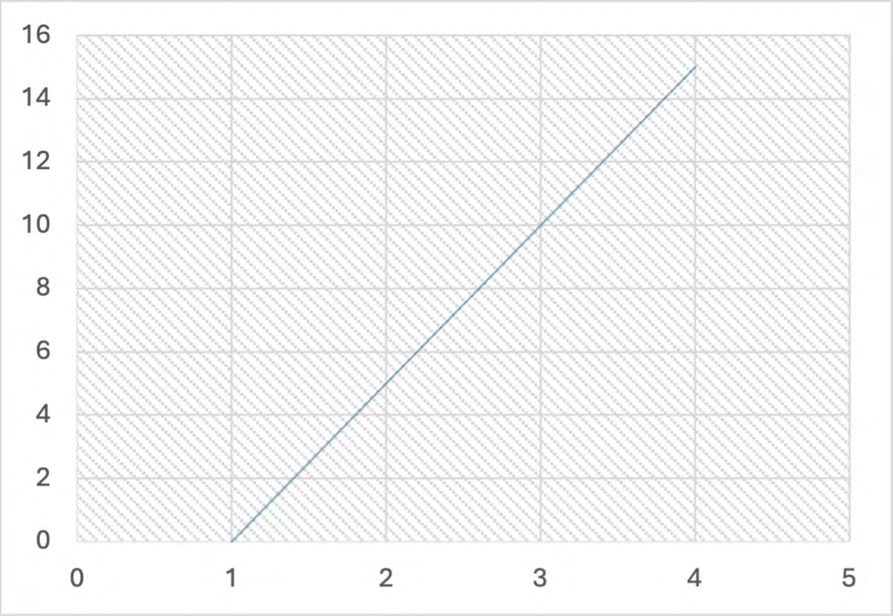
Average rate from concentration-time graph

Data

Time/min	[A]/moldm^{-3}
0	2.0
5	1.8
10	1.6
15	1.4

Average of reaction between 5-10 min.

From graph,



$\Delta[A] = \Delta[A]_2\text{-}\Delta[A]_1 = 1.6\text{-}1.8 =\text{-}\ 0.2\ M$

$\Delta t = t_2\text{-}t_1 = 10\text{-}5 = 5\ \text{min}$

$$\text{Average rate} = \frac{\Delta[A]}{\Delta t} = \frac{-0.2\ M}{5\ min} = 0.04\ M/min$$

Activity 2.2 Measuring reactions rates using specific methods

1. **a.** With the method assigned to your group, plan and perform experiments to calculate the rate of reaction.
    - i. The procedure for measuring rate of reaction.
    - ii. The equipment and materials needed
        - Potential sources of error associated with your method and how to them.
        - Discuss the advantages and limitations of your assigned method
        - How do your results compare with results in the literature?

   b. Watch the video on **CD-ROM encyclopaedia of science** on rate of reaction and discuss,
    - i. the procedures
    - ii. limitations

iii. Potential sources of error.

iv. Each group present your procedure and finding to the class.

v. Compare and discuss your method and the methods that your friends used.

FACTORS THAT AFFECT RATE OF REACTION

Reaction rates can vary greatly over a large range of time. While certain reactions can happen slowly over many years, others can happen at incredibly quick rates.

1. Temperature

All chemical reactions proceed more quickly as the temperature increases. This is because at higher temperature, the average kinetic energy of the particles increases, so more reacting particles have energy equal to or greater than the activation energy of the reaction. Additionally, because the reacting particles move faster at higher temperatures, there are more collisions between them, which raises the frequency of effective collisions. As a result, more products are created per unit of time, increasing the pace of reaction.

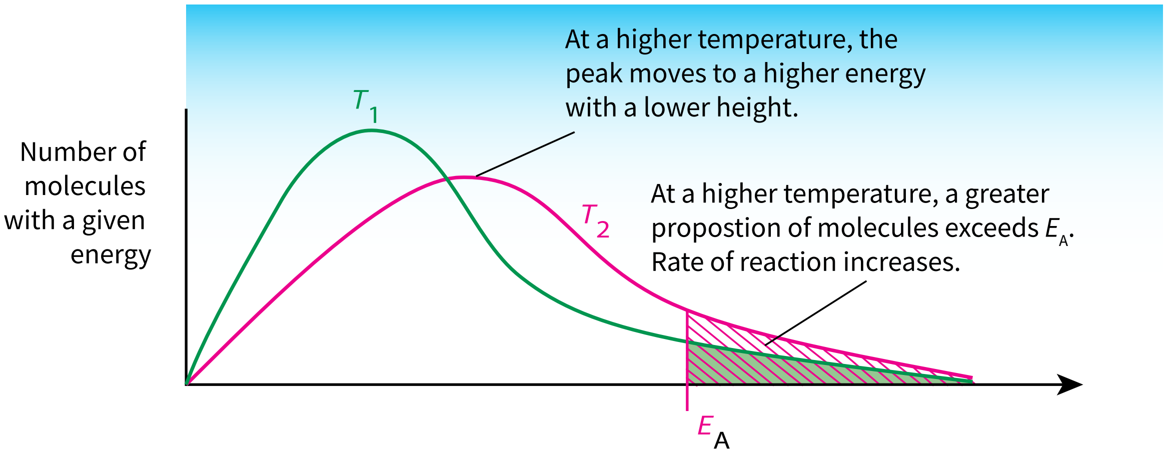


Figure 2.1: Effect of temperature on the kinetic energy distribution of molecules in a sample

This explains why food can spoil quickly when left on the kitchen counter. But that process is slowed down by the refrigerator's lower internal temperature, keeping the same food fresh for days. In the chemistry laboratory, gas burners, hot plates, and ovens are frequently employed to speed up processes that move slowly at room temperature.

2. **Concentration**

   The rate of reaction increases when the concentration of one or more reactants rises. There is a greater likelihood of particle collisions when the reactant concentration rises; more collisions translate into a higher reaction rate.

3. **Available surface area and nature of the reactants**

   The rate of the reaction increases with the amount of surface contact between reactants. The reaction rate decreases with decreasing surface contact.

   Note

   It is also important to note that not all reactions depend on surface area. If both reactants are gases or liquids that mix together, then surface area is not a factor. Surface area is most prevalent when one of the reactants is in the solid phase.

4. **Catalyst**

   A catalyst is a material that increases a chemical reaction's rate without being consumed by the reaction itself. A catalyst functions by offering a different path for the reaction to proceed. The activation energy of this other path is lower. This increases the proportion of reacting particles with energy greater than the activation energy.

   As a result, the frequency of effective collision increases. More products are formed per unit time and the rate of reaction is higher.

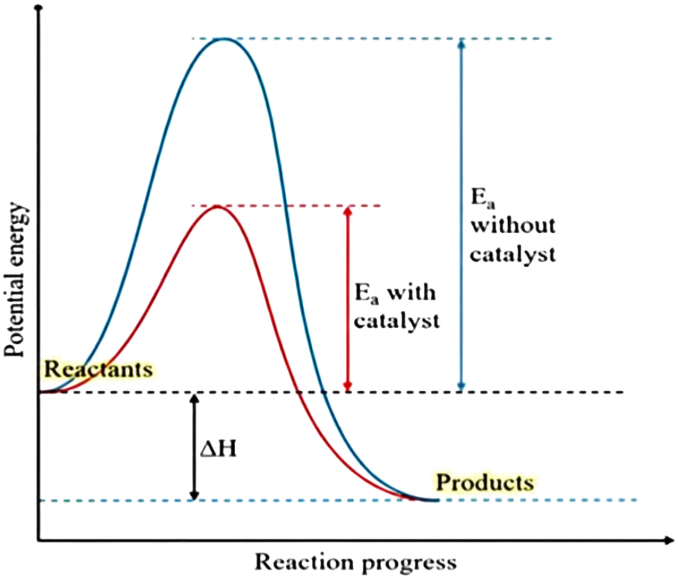


Figure 2.2: Effect of catalyst on the rate of a reaction

5. **Pressure**

   The rate of chemical reactions with reactants and products in their gaseous form increases with increasing pressure. Increasing pressure results in either more particles in a given volume or the same number of gaseous particles being forced into a smaller volume. The particles are now closer together in both situations, which increases the number of collisions per unit time and the reaction rate.

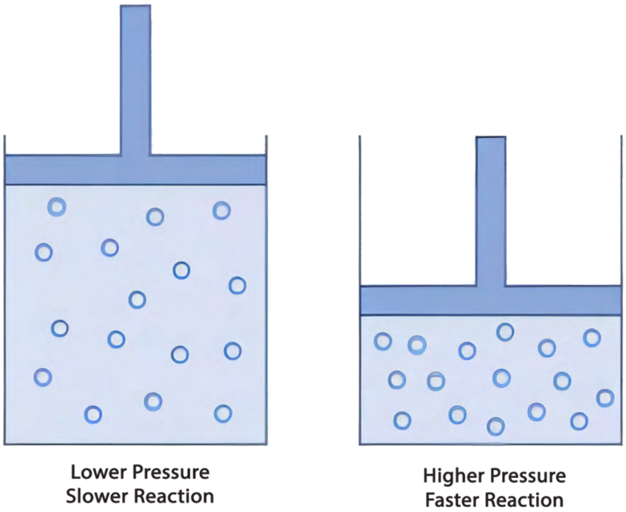


Figure 2.3: *Effect of pressure on the rate of a reaction*

Experiments to investigate the factors that affect rate of reaction

1. Experiment to investigate the effect of temperature on the rate of reaction

Procedure

a. A water bath is used to heat diluted hydrochloric acid to a set temperature.
b. In a conical flask, add the warmed dilute hydrochloric acid.
c. Put a magnesium strip of known mass and surface area in the conical flask and start the stopwatch
d. When the magnesium has completely reacted and disappears, stop the time.
e. Compare the results after repeating at various temperatures, keep the mass of magnesium and available surface area of magnesium constant.

Result

The rate of reaction will increase when the temperature increases. This is because the particles will collide more frequently and effectively due to their greater kinetic energy than the required activation energy, which will increase the rate of reaction.

Why was it important to keep the mass of surface area of the magnesium the same across the different reactions?

2. Experiment to investigate the effect of concentration on rate of reaction

SAFETY NOTE: Take care not to inhale any fumes given off in this reaction as $SO_2(g)$ can cause respiratory irritation.

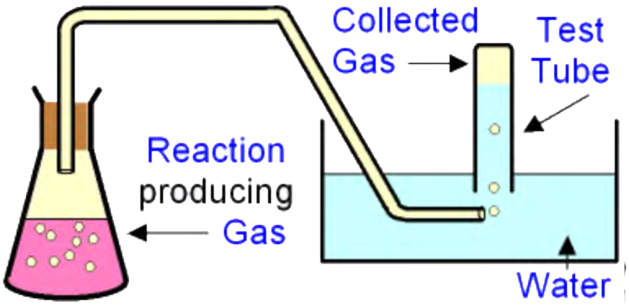
Procedure

- **a.** Fill a conical flask with 50 cm^3 of sodium thiosulfate solution.
- **a.** Fill a measuring cylinder with 5 cm^3 of diluted hydrochloric acid.
- **b.** Place a piece of paper with a cross drawn on it underneath the flask.
- **c.** Start the stopwatch as soon as you add the acid to the flask.
- **d.** Examine the cross from above (ensuring a minimum distance of 20 cm from the top of the flask), and when it is no longer visible, stop the stopwatch.
- **e.** Repeat with varying sodium thiosulfate solution concentrations (dilute with water and mix varying volumes of sodium thiosulfate solution).

Result

When the concentration of a solution increases, the rate of reaction also increases. This is because there will be more reactants in a given volume, allowing more frequent and efficient collisions that speed up the reaction.

3. Experiment to investigate the effect of Surface Area on the rate of reaction



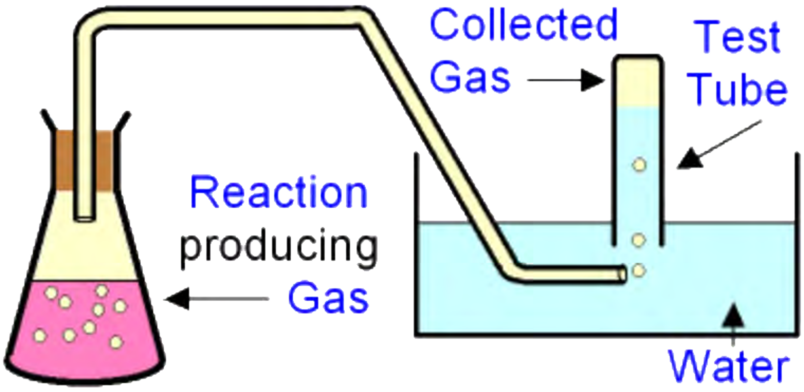
Procedure

a. A conical flask should be filled with diluted hydrochloric acid.
b. This flask should be connected to a measuring cylinder upside down in a bucket of water via a delivery line (downward displacement).
c. Quickly replace the bung into the conical flask after adding magnesium ribbon.
d. Using the measuring cylinder, determine the volume of gas generated i predetermined amount of time.
e. Repeat using pieces of magnesium ribbon of varying sizes (the same mass of magnesium must be used).

Result

Because smaller sections of magnesium ribbon enhance the solid's surface area, the rate of reaction will also rise. This is because a larger particle surface area will be in contact with the other reactant, resulting in more frequent and effective collisions that speed up the reaction.

4. Experiment to investigate the effect of catalyst on the rate of reaction



Procedure

a. Put hydrogen peroxide in a conical flask.
a. This flask should be connected to a measuring cylinder upside down in a tub of water via a delivery line (downward displacement).
b. Quickly insert the bung into the conical flask after adding the catalyst manganese (IV) oxide.
c. Using the measuring cylinder, determine the amount of gas generated i predetermined amount of time.

d. Comparing the findings, repeat the experiment without the manganese (IV) catalyst.

Result

The rate of reaction will be increased by using a catalyst. In order to give more colliding particles, the necessary activation energy to react, the catalyst will offer a different pathway with a lower activation energy. The rate of reaction will increase as a result of more frequent and efficient contacts.

Application of the factors that affect rate of reaction in everyday life

1. **Temperature**

   Food Preservation: The rate of spoiling reactions is slowed down by cold temperatures. Perishables are therefore kept in a refrigerator to prevent bacteria and their enzymes from functioning as well as they could. Fruits and vegetables should be kept in the refrigerator to prolong their freshness.

2. **Concentration**

   Cleaning: Concentrated solutions are frequently found in household cleaning goods. Their efficacy and reaction rates are altered when diluted with water.

3. **Catalysts**

   Chemical Industry

   Catalysts are necessary for effective manufacturing. They expedite responses, enabling producers to satisfy demand while upholding safety regulations.

4. **Surface area**

   Medicine and Pills

   Medication that has been crushed or powdered dissolves faster in the stomach. They are absorbed into the bloodstream more quickly due to their increased surface area.

5. **Pressure**

   Fertiliser production

   The manufacture of ammonia for fertilisers is done in gaseous form so using a high pressure will increase the rate of reaction.

Activity 2.3 Explaining factors that affect reaction rate

1. **a.** In groups, state the factors that affect rate of chemical reaction.
   b. Explain why powdered sugar dissolves faster than sugar cubes in the volume of water?
   c. In the production of nitric acid, ammonia is combusted according to the reaction equation:

   $\mathbf{4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g).}$

   i. Write the equations that relate:

   The rate of consumption of the reactants and the rate of formation of the products.

Activity 2.4 Investigating factors that affect rate of reaction

Spend fifteen (15 minutes) on each activity and move to the next. All the experiments must be performed before leaving the laboratory.

1. In groups, discuss and write down the procedure to investigate the effects of each of the following on reaction rate:
   a. Changing concentrations
   b. Changing surface area
   c. Changing temperature
   d. Introduction of catalyst
2. Perform the experiment following the procedure you outlined.

Activity 2.5 Analysing experimental data and graphs of reactions

Materials needed: Handouts of experimental data and graphs of reaction rates.

Procedure

1. In groups, study and discuss the data and graphs in the handouts and answer the questions below:
   a. What are the variables (dependent and independent) and their units in the experiment?
   b. How do the variables relate to each other?
   c. What do you observe about the steepness of the graph?

d. Which reaction occurred fastest and why?

e. How could you extend this into an investigation?

Activity 2.6 Discussing practical applications of rate of reaction.

a. Discuss the effect of the factors of rate of reaction in the following.

  i. Surface area and safety of grain mill factories.

  ii. Why a glowing splint rekindles when it is put in a bottle of oxygen gas,

  iii. Why smoking is forbidden in areas where bottled oxygen is in use, burning charcoal

  iv. The use of antioxidants as competitive inhibitors to preserve food.

COLLISION THEORY

Collision theory explains how and why chemical reactions happen. It says that for reactants to turn into products, their particles must collide with enough energy and in the right orientation.

Key points of collision theory

1. For particles to react, they need to collide with the correct orientation and have energy equal to or greater than the activation energy (the minimum energy needed for the reaction to occur).
2. The activation energy is the least amount of energy particles need to collide with tom cause a reaction.
3. Different reactions have different activation energy levels.
4. If a collision leads to the reactants turning into products, it is called a successful or effective collision.

Factors Affecting Reaction Rate According to Collision Theory

1. **Temperature**

   When the temperature increases, the particles gain more energy, leading to more collisions with enough energy to cause a reaction. This increases the rate of reaction because a greater fraction of particles will have energy greater than the activation energy.

2. **Surface Area and Reactant Nature**

   A larger surface area allows more particles to collide, increasing the number of successful collisions and increasing the rate of the reaction. More exposed particles mean more frequent collisions, which raises the reaction rate.

3. **Concentration**

   Increasing the concentration of reactants increases the number of particles in a given space, causing more frequent collisions. For gases, increasing the pressure has the same effect as more particles are forced into a smaller space, leading to more collisions and a faster reaction.

4. **Catalyst**

   A catalyst speeds up a reaction by providing an easier pathway with a lower activation energy. This allows more particles to have enough energy to collide effectively, resulting in more successful collisions and a faster reaction.

Maxwell-Boltzmann energy distribution curve

The effect of temperature on the rate of reaction can be shown using a Maxwell-Boltzmann distribution curve. A typical curve looks like this below:

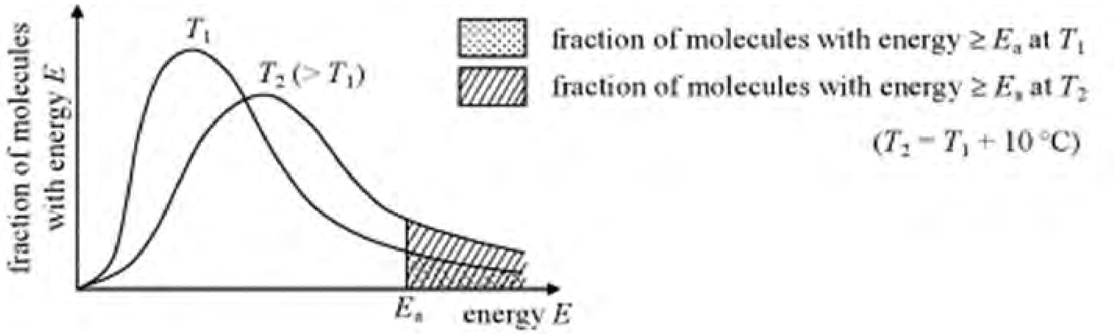


Figure 2.4: Maxwell-Boltzmann energy distribution curve

The area under the curve represents all the molecules, and must remain constant for a given scenario. At a higher temperature (T_2), more particles have energy greater than the activation energy. When drawing the curve at a higher temperature, the curve becomes lower, but the total area under the curve remains the same. This is because no mass can be lost or gained during a chemical reaction.

When the temperature increases, particles gain more kinetic energy, causing more collisions per second. Also, a larger number of particles have energy above the activation energy, meaning more collisions will lead to a reaction. This increases the rate of the reaction because there are more effective collisions.

Activity 2.7

1. Form groups of not more than five. Each group should brainstorm and list all interactions you can think of that might happen in a chemical reaction (e.g., atoms molecules colliding, bonds forming or breaking, energy transfer). Include ideas!

2. Each group will then share their list with the class. Record all interactions on the or on a large piece of paper.

3. Have a whole-class discussion to decide on the most important interactions i chemical reaction.

4. Use the agreed-upon interactions to introduce the collision theory. Think about this theory explains what happens during a chemical reaction.

5. Write the main points of the collision theory on the board as a class.

6. Use examples from the brainstorming session to help explain each key point of collision theory.

7. Each group should make a concept map or diagram that shows the main ideas of collision theory using examples from the discussion.

Activity 2.8 Collision theory- simulating how particles interact and react

1. In the same groups, choose a scenario for each below:

  a. Particles with different energies and orientations.
  b. Particles with various surface areas.
  c. Particles with different concentrations.

2. Act out the collision theory in your groups by pretending to be particles interacting reacting. Discuss how different factors (like temperature, surface area concentration) might affect these interactions.

3. Use your scenarios to think about how each factor could change the rate of a reaction. Draw diagrams and sketches to support your explanations.

4. Draw a sketch of the Maxwell-Boltzmann energy distribution curve and discuss how relates to the collision theory.

5. Your instructor will wrap up with a class discussion to summarize the main ideas key takeaways from the activities.

Activity 2.9 Exploring Reaction Rates Using Collision Theory and Maxwell-B Distribution

1. Understand the basics of collision theory and how it explains reaction rates.
2. Use the Maxwell-Boltzmann energy distribution curve to explain how affects reaction rates.
3. Compare and analyse the effects of concentration and temperature on reaction using collision theory and the Maxwell-Boltzmann distribution.

Materials needed

- Chart or projector displaying the Maxwell-Boltzmann energy distribution curve
- Reaction rate simulation tools (online tools or hands-on demonstrations reacting chemicals)
- Models or diagrams of particle collisions
- Worksheets for group discussions and note-taking
- Whiteboard and markers for visual comparisons

Understand the basics of collision theory and how it explains reaction rates

a. Explain the idea that chemical reactions occur when particles collide with energy (activation energy) and proper orientation after watching the video below.

https://youtu.be/daihzyMQ_X8

b. What happens to the rate of reaction if particles collide more frequently?

c. You will participate in a demonstration where we shake containers filled with numbers of particles. This will show how more frequent collisions lead to reactions.

d. In groups, discuss and explore the factors that can influence the frequency and of collisions (such as concentration, temperature, or surface area).

e. Based on the discussion, write down how collision theory explains why increasing concentration of reactants generally increases the reaction rate.

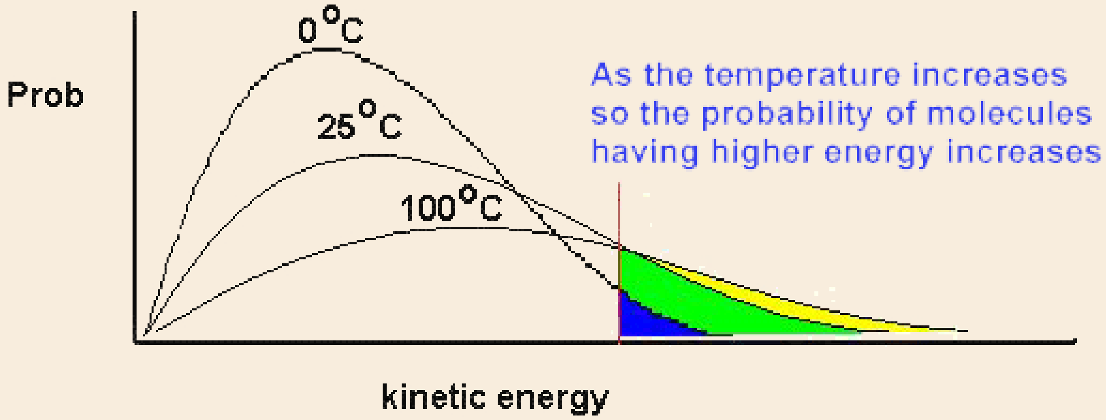
Understanding the Maxwell-Boltzmann Distribution Curve

Aim: Use the Maxwell-Boltzmann energy distribution curve to explain why an increase in temperature increases the rate of a chemical reaction.

a. Watch the video below attentively and carry out the activities that

https://th.bing.com/th?&id=OVP.HyQXE9N3yjrqKjuvCs4l-HAHgFo&w=323&h=181&c=7&pid=2.1&rs=1

b. Look at the Maxwell-Boltzmann energy distribution curve displayed. This curve how the kinetic energies of particles in a gas sample are distributed. Only with energy above the activation energy will react.



c. Compare two curves: one representing a lower temperature and one representing a higher temperature. Observe how the area under the curve changes shape, particularly the number of particles with energy above the activation energy.

d. How does increasing the temperature affect the number of particles with energy to react?

e. Each group will present their findings to the class, explaining how the shift in the at higher temperatures leads to more successful collisions and, therefore, a reaction rate.

Comparing the Effects of Concentration and Temperature

Compare and analyse the effects of concentration and temperature on reaction rates using both collision theory and the Maxwell-Boltzmann energy distribution curve.

a. Working in two teams, you will compare and contrast the effects of increasing concentration and increasing temperature on the rate of a chemical reaction using both collision theory and the Maxwell-Boltzmann energy distribution curve.

b. **Team 1 (Concentration):** Use collision theory to explain how increasing the number of particles (i.e., higher concentration) leads to more frequent collisions and a faster reaction rate.

c. **Team 2 (Temperature):** Use the Maxwell-Boltzmann distribution curve to show that at higher temperatures, more particles have energy greater than the activation energy, leading to a faster reaction rate.

d. After both teams present, engage in a class-wide discussion to compare and contrast the effects of temperature and concentration on reaction rates. Focus on the similarities and differences.

e. Which factor, temperature or concentration, do you think has a greater effect on reaction rates in typical conditions?

Finally, complete the worksheet provided, which asks you to compare and contrast the effects of concentration and temperature using both collision theory and the Maxwell-Boltzmann energy distribution curve.

RATE EQUATIONS AND ORDER OF REACTIONS

1. Rate Equation

The rate equation or rate law for a chemical reaction is an expression that provides a relationship between the rate of the reaction and the concentration of the reactants participating in it.

Consider the hypothetical chemical reaction given by: $aA + bB \rightarrow cC + dD$

Where a, b, c and d are the stoichiometric coefficients of the reactants and products, then the rate equation for the reaction is given by:

$Rate \ \alpha \ [A]^x[B]^y$

$Rate = k[A]^x[B]^y$

Where; **[A]** and **[B]** denotes the concentrations of reactants A and B.

'x' and **'y'** denote the orders with respect to reactants A and B (which may or may not be equal to their stoichiometric coefficients a and b

The proportionality constant '**k**' is the rate constant of the reaction.

2. Order of reaction

Order of a reaction with respect to each reactant is define as the exponent to which the concentration term of that reactant in the rate law is raised.

Consider the general reaction: $aA + bB \rightarrow cC + dD$

suppose the rate expression for this reaction is: $rate = k[A]^x[B]^y$, then the order with respect to A and B are $\boldsymbol{x}$ *and* $\boldsymbol{y}$ respectively.

a. Overall order of a reaction

The overall order of a reaction is determined by adding up the powers (or exponents) of the concentration of all the reactants in the rate equation. To find the overall order, simply add up the individual orders of each reactant involved in the reaction. It is therefore defined as, the sum of the exponents to which the concentrations of all the reactants in the rate law are raised.

Considering the reaction, $rate = k[A]^x[B]^y$ as an example,

the $overall\ order\ of\ the\ reaction = x + y$

b. Zero order reaction

For a chemical reaction in the form $aA \rightarrow product$, if the order with respect to A is a zero order, the rate law for the reaction is: $rate = k[A]^0$

$rate = k(1) = k.$

This means that for a zero-order reaction, the rate is constant. It does not change with concentration. The rate is independent of the concentration of the reactants, meaning an increase in concentration of the reactant will not translate to an increase in rate of the reaction. It is worth noting that a zero-order reaction will obviously change rate when all of the reactant is exhausted (the reaction cannot continue at the constant rate past completion).

c. First-order reaction

For a chemical reaction in the form $aA \rightarrow product$, if the order with respect to A is a first order, then

$rate = -\frac{\Delta[A]}{\Delta t}$

$rate = k[A]$

For a first order reaction, doubling the concentration of a first order reaction causes the rate to double.

d. **Second order reaction**

For a chemical reaction of the form $aA \rightarrow product$, if the order with respect to A is a second order,

$rate = -k\frac{\Delta[A]^2}{\Delta t}$

For a second order reaction, doubling the concentration of a second order reaction causes the rate to quadruple.

e. **Half-life of a reaction**

The time required for a reactant to reduce to half its initial concentration is called the half-life of a reactant and is designated by the symbol $t_{½}$.

For first order reaction,

$t_{\frac{1}{2}} = \frac{0.693}{k}$

For a second order reaction,

$t_{\frac{1}{2}} = \frac{1}{k[A]_o}$

3. Rate determining step

Think of a chemical reaction like an assembly line in a factory.

A whole reaction might be finished in one go, or it might take a series of small steps (each small step is called an 'elementary reaction'). This entire series of steps is known as the 'reaction mechanism'.

The speed of the entire assembly line (the overall rate of the reaction) is always limited by the slowest single worker or machine on that line.

In chemistry, the slowest step in the entire sequence is the most important one. We call this the rate-determining step.

Simply put: We look at the reaction mechanism (the full sequence of steps), and the step that takes the longest is the one that sets the pace for the whole process. That's the rate-determining step.

Graphical Determination of Reaction Order

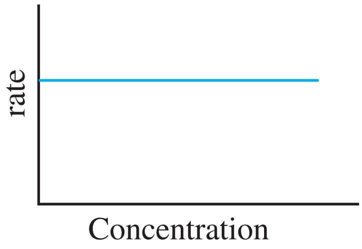
In chemical kinetics, the order of a reaction describes how the rate of the reaction depends on the concentration of reactants. Graphical analysis allows us to determine the order of a reaction with respect to a specific reactant. Here is how it is done:

1. **Experimental Data Collection:** Conduct an investigation where the concentration of the reactant is monitored time. Record data points of concentration at different time intervals.
2. **Plotting the Concentration vs. Time Graph:** Create a graph with concentration (on the y-axis) against time (on the x-axis).

   The resulting curve represents the change in concentration as the reaction progresses.
3. **Drawing Tangents:** Draw tangents (straight lines) to the curve at various points.

   These tangents represent the instantaneous rates of reaction at those concentrations.
4. **Calculating Tangent Gradients:** Calculate the gradient (slope) of each tangent line.

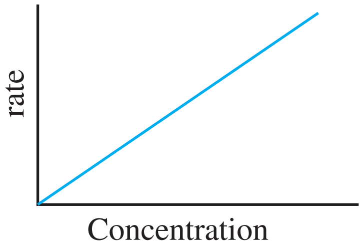
   The gradient at a specific concentration corresponds to the rate of reaction at concentration.
5. **Interpreting the Graph Shape:** The shape of the rate against concentration graph provides clues about the order.
   a. **Zero Order:** A straight horizontal line indicates zero order with respect to the reactant.



$rate = k[A]^0$

$rate = k.$

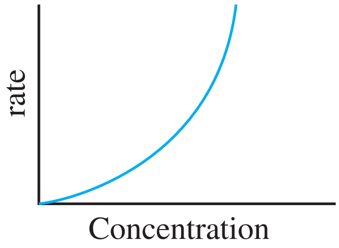
b. First Order: A linear (straight-line) graph suggests first order.



$$rate = -\frac{\Delta[\mathrm{A}]}{\Delta \mathrm{t}}$$

$$rate = k[A]$$

c. Second Order: A curved graph (concave upward) indicates second order.



$$rate = -k\frac{\Delta[\mathrm{A}]^2}{\Delta \mathrm{t}}$$

Calculations involving the rate law expression

Example 1

Consider data for the reaction: $A + B \rightarrow products$

Experiment	[A] / moldm^{-3}	[B] / moldm^{-3}	Rate / moldm^{-3} s^{-1}
1	0.002	0.001	5×10^{-4}
2	0.002	0.002	10×10^{-4}
3	0.004	0.001	20×10^{-4}

Determine,

i. the rate law

ii. the overall order of the reaction

iii. the specific rate constant k

Answer

i. **Determine the exponents separately**

Substituting data from experiment 1 and 2 into the general rate equation,

Rate = $k[A]^x[B]^y$,

$5 \times 10^{-4} = k(0.002)^x(0.001)^y$ … … … … … … . **equation 1.**

$10 \times 10^{-4} = k(0.002)^x(0.002)^y$ … … … … … … . **equation 2.**

Divide equation **2** by**1**

$$\frac{10 \times 10^{-4}}{5 \times 10^{-4}} = \frac{k(0.002)^x(0.002)^y}{k(0.002)^x(0.001)^y}$$

$2 = 2^y$

$y = 1$

Also, substituting data from experiment 1 and 3 into the rate law expression,

$5 \times 10^{-4} = k(0.002)^x(0.001)^y$ … … … … … … . . **equation 3**

$20 \times 10^{-4} = k(0.004)^x(0.001)^y$ … … … … … … … **equation 4**

Divide equation **4** by **3**

$$\frac{20 \times 10^{-4}}{5\times 10^{-4}} = \frac{k(0.004)^x(0.001)^y}{k(0.002)^x(0.001)^y}$$

$4 = 2^x$

$2^2 = 2^x$

$x = 2$

Substituting the exponents into a rate law

Rate = $k[A]^x[B]^y$

Rate = $k[A]^2[B]^1$

ii. The overall order of the reaction is $= y + x$

$= 1 + 2$

$= 3$

Therefore, the reaction is in 3rd order.

iii. Pick any complete set of experimental data, substitute it into the rate equation and calculate the value of the constant k

Substituting data from experiment **three** into the rate equation:

Rate = $k[A]^2[B]$

$5 \times 10^{-4} = k(0.004)^2(0.001)$

$$k = \frac{20\,x\,10^{-4}}{(0.004)^2(0.001)}$$

$$k = 1.25 \times 10^5$$

$$\text{Units of } k = \frac{\text{moldm}^{-3}\text{s}^{-1}}{(\text{moldm}^{-3})\ (\text{moldm}^{-3} \times \text{moldm}^{-3})}$$

$$k = \text{mol}^{-2}\text{dm}^{-6}\text{s}^{-1}$$

Therefore, $k = 1.25 \times 10^5\ \text{mol}^{-2}\text{dm}^{-6}\text{s}^{-1}$

Activity 2.10 Visualising Activation Energy, Collision Theory, and Reaction Rate Influence

1. Reflect on an analogy, such as climbing a hill, to illustrate the concept of energy to initiate a reaction. Write a brief explanation on how this analogy represents concept of activation energy in chemical reactions.
2. Create a table listing various factors (e.g., concentration, surface area, catalysts) that influence reaction rates. Use visual aids, such as particle collision diagrams, to represent how each factor impacts collisions and reaction rates.
3. If possible, use interactive software to manipulate temperature and observe changes in the Maxwell-Boltzmann distribution curve in real time. Record observations on how temperature influences the curve shape and discuss the implications for reaction rates.
4. Practice interpreting concentration vs. time graphs by analysing slopes and trends. Use these graphs to deduce the reaction order and write an explanation of how the graph's shape correlates to different reaction orders.

Activity 2.11 Exploration and Reflection on Rate of Chemical Reactions

1. Use your own experiences to brainstorm definitions for key terms, such as rate constant (k), order of reaction, and rate-determining step (RDS). Write out your ideas and consider how each term applies to chemical reactions you have learned about.
2. After completing the above activities, summarise the key points and insights gained from each. Reflect on any areas where you would like further clarification and note down questions to discuss with your teacher or peers.

Activity 2.12 Experimental Data Analysis on Rate of Chemical Reactions

Given the following experimental data for the reaction A→B, where the concentration of A is measured at different times:

Time (s)	[A] mol/dm^3
0	0.50
10	0.40
20	0.32
30	0.25
40	0.20

a. Use the data provided to deduce the rate expression for the reaction by the reaction order with respect to A.

b. Show your calculations and reasoning, and explain the process used to arrive at rate expression.

c. Sketch the expected concentration vs. time graphs for zero-order, first-order, and second-order reactions. Describe the shape of each graph and how the slope or curve provides insight into the reaction order.

d. Given the half-life data for a first-order reaction with an initial concentration of 0.40 mol/dm^3 and a half-life of 15 seconds, calculate the rate constant and use it to confirm the reaction order.

e. Discuss how this half-life concept is applied in real-life scenarios, such as radioactive decay, and note any insights or questions that arise.

EXTENDED READING

Use **CD-ROM Encyclopaedia** of Science to watch videos on rate of reaction and

- take notes on the procedures
- list any procedural errors
- what modification can be put in place to remove the problems associated with experiment?

- https://youtu.be/QfS1DSOw3Og

- https://youtu.be/plXbgRYW7B8

REVIEW QUESTIONS

Review Questions 2.1

1. Define the term rate of reaction.
2. Compare and contrast the use of changes in volume of gas evolved verses formation of precipitate to measure rate of reaction.
3. Design an experiment to measure the rate of reaction using colour change.
4. State at least three factors that affects rate of chemical reactions
5. Discuss the rationale behind using powdered versus solid reactants.
6. Consider the spoilage of milk, a perishable item that undergoes a chemical reaction involving the enzyme lactase. Explain how refrigeration slows down the spoilage process, using concepts related to rate of reaction, activation energy. Include a graphical representation of the effect of temperature on the rate of reaction.
7. Consider the following reaction in aqueous solution:

   $$5\,Br^-(aq) + BrO_3^-(aq) + 6H^+(aq) \rightarrow 3Br_2(aq) + 3H_2O(l).$$

   If the rate of disappearance of Br^- (aq) at a particular moment during the reaction is $3.5\text{x}10^{-4}\ mol\,dm^{-3}\,s^{-1}$, what is the rate of appearance of Br_2(aq) at that moment?

Review Questions 2.2

1. What is the collision theory, and how does it explain the rate of a chemical reaction?
1. Compare and contrast the effects of increasing concentration and increasing temperature on the rate of a chemical reaction, using collision theory and the Maxwell Boltzmann energy distribution curve to support your answer.
2. Design an experiment to investigate how changing the surface area of reactants affects the rate of reaction and explain how the collision theory supports your hypothesis.
3. Why does increasing the concentration of a solution increase the rate reaction?

4. Why do catalysts increase a rate of reaction?
5. Describe the Maxwell-Boltzmann distribution curve and explain how you can estimate the fraction of particles with sufficient energy to react.

SECTION

3

DYNAMIC EQUILIBRIUM

PHYSICAL CHEMISTRY

Equilibria

INTRODUCTION

In this lesson, we will start by watching a video to learn about two types of reactions: reversible (can go back and forth with products and reactants existing simultaneously) and irreversible (once products are formed, they cannot turn back into reactants). We will see examples of these reactions in everyday life and use simple graphs to understand *dynamic equilibrium*, where reactions reach a state of constant behaviour. Through experiments with anhydrous copper sulphate, we will explore how reactions work. We will also learn to write formulas, called *equilibrium constant expressions*, to characterise these reactions. Finally, we will explore how temperature and/or pressure changes affects equilibrium using *Le Chatelier's principle*, helping chemists to find the best conditions for reactions.

KEY IDEAS

- **Equilibrium constant (K)**: A value that expresses the ratio of the concentrations of products to reactants at equilibrium, specific to a particular reaction at a given temperature or pressure.
- **Equilibrium position**: The relative concentrations of reactants and products in a system at equilibrium, which can shift in response to changes in conditions.
- **Forward reaction**: The reaction that converts reactants into products.
- **Heterogeneous equilibrium**: An equilibrium in which reactants and products are in different phases.
- **Homogeneous equilibrium**: An equilibrium in which all reactants and products are in the same phase.
- **Reverse reaction**: The reaction that converts products back into reactants.

DYNAMIC EQUILIBRIUM

1. Reversible Reactions

Some reactions happen only once. The reactants (starting materials) combine to make a product, and when the reactants are all used up, the reaction stops. In most situations, once a reaction is complete, the new substances (the products) stay as they are. However, with some chemical reactions, the products can actually change back into the original starting materials (the reactants). This means the entire process is happening in two directions at the same time: one process works to form the products, and the other simultaneously works to remake the reactants. This is what we call a reversible reaction.

A reaction is called *reversible* when both forward and backward reactions can happen at the same time under certain conditions like temperature and pressure. We show this with two arrows (⇌) pointing in opposite directions between the reactants and products.

e.g. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

2. Irreversible Reactions

Most chemical reactions only go in one direction and cannot go back to the materials. These are called *irreversible reactions*. For example, when carbon burns air, it forms carbon dioxide (CO_2) and cannot turn back into carbon and oxygen. We show irreversible reactions with a single arrow (→) pointing from the reactants to product, like this: **$\mathrm{C(s) + O2(g) \rightarrow CO2(g)}$**

3. Dynamic Equilibrium

Some reactions do not completely stop but reach a balance called *chemical equilibrium*. In these reactions, both the concentrations of the reactants and products stay constant because they are continuously changing back and forth at the same rate. For each reactant that is changed into a product, a product is changed back into a reactant. Despite there being no change in the overall concentrations of reactants and products the reactions are both still occurring.

Example of Dynamic Equilibrium: When nitrogen dioxide (NO_2) gas reacts to form dinitrogen tetroxide (N_2O_4) in a container, NO_2 changes colour from dark brown to colourless. Eventually, the stops changing, showing that the reaction has reached equilibrium where concentration of NO_2 and N_2O_4 stays constant.

The reaction could better be described with the following equation:

$NO_2(g) + NO_2(g) \rightleftharpoons N_2O_4(g)$

Dynamic Equilibrium means the forward and backward reactions happen at the rate, and the amounts of reactants and products stay constant in a *closed system*.

Characteristics of Dynamic Equilibrium

1. Happens in a closed system (no reactants or products can be added or removed).
2. Both reactants and products are present.
3. Forward and reverse reactions happen at the same rate.
4. Equilibrium can be reached from either direction.
5. Concentrations of reactants and products do not change over time.

Applications of Dynamic Equilibrium

1. Chemical Systems: Rechargeable batteries, like lithium-ion batteries, through dynamic equilibrium, storing and releasing energy.
2. Biological Systems: Our bodies keep glucose levels steady with equilibrium, as the liver stores and releases glucose as needed.
3. Environmental Processes: Dynamic equilibrium helps in carbon cycling photosynthesis, respiration, and decomposition.

Activity 3.1 Reversible and Irreversible Reactions

Go online and watch a brief video that demonstrates both reversible and irreversible reactions. (https://www.youtube.com/watch?v=ty9TczsW5ew here is a possible example)

1. Watch for differences between reactions that can “go back and forth” and those only “go one way.”
2. After the video, define *reversible and irreversible reactions.*
3. Give examples of reversible *and irreversible* reactions
4. Use simple equations to show these forward and reverse reactions.
5. Explain *dynamic equilibrium* as a balance point where the forward and reactions happen at the same speed.

6. Draw a simple graph showing how the amounts of reactants and products constant over time in dynamic equilibrium.
    - a. Label the graph with time on the x-axis and concentration on the y-axis.
    - b. Show how the lines for reactants and products start changing but eventually out to a constant value at equilibrium.

Activity 3.2 Reversible Reaction with Anhydrous Copper (II) Sulfate and Water

Objective: To explore a reversible reaction and understand its relevance through observation, recording, and analysis.

Materials Needed: Anhydrous Copper (II) Sulphate (white powder form), distilled Water, heat source (optional, for drying the copper (II) sulphate back to anhydrous form), spoon or spatula, glass beakers or test tubes, protective gloves and goggles

Steps

Experiment – Conducting the Reaction

1. Observe and describe anhydrous copper (II) sulphate
2. Add a few drops of distilled water to the anhydrous copper (II) sulphate and observe.
3. Note the colour change, texture, and other observations about the substance water is added.
4. Gently heat the blue hydrated copper (II) sulphate. As it dries, it will turn white again, indicating the removal of water.
5. Observe and record any changes during heating. Discuss how the colour returns white as the substance reverts to its original anhydrous state.
6. Compare observations before and after each step.
7. Share what you learned about reversible reactions.

EQUILIBRIUM CONSTANT EXPRESSIONS

1. The Law of Mass Action

The Law of Mass Action explains how substances in a reaction affect each other. It helps us understand how reactions work and what happens when they reac balanced state, called equilibrium.

The **Law of Mass Action** states that the rate of a chemical reaction is proportional to the product of the concentrations of the reactants, each raised t power equal to the coefficient of that reactant in the balanced chemical equation.

Imagine a reaction where: $aA + bB \rightleftharpoons cC + dD$

Here:

A and B are reactants; C and D are products
Forward and Reverse Reaction Rates
The forward rate (reaction moving from reactants to products) is: $r_f = k_f[A]^a[B]^b$ where k_f is a constant specific to the reaction being conducted.

The reverse rate (reaction going back from products to reactants) is:
$r_b = k_b[C]^c[D]^d$ where k_b is a constant specific to the reaction being conducted

Equilibrium Constant (K)

When the reaction reaches equilibrium, the forward and reverse rates are equal:
$k_f[A]^a[B]^b = k_b[C]^c[D]^d$

Rearranging this gives us the equilibrium constant (K), which shows the between products and reactants:

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

This K_c value helps us know the proportions of products and reactants at equilibrium.

2. Equilibrium Constants: K_c and K_p

For a **reversible reaction** like:

$aA + bB \rightleftharpoons cC + dD$

at **equilibrium**, the ratio of concentrations of products and reactants stays the at a certain temperature. This is called the **equilibrium constant** in terms concentration, written as **Kc**:

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Here:

[A] and **[B]** are the concentrations of the reactants.

[C] and **[D]** are the concentrations of the products.

Kc represents the balance point of the reaction.

For gases, we can use their **partial pressures** instead of concentrations. This gives us a new equilibrium constant, called **Kp**:

$$K_P = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

Here, **P** represents the partial pressures of each gas, and **Kp** tells us the balance of pressures at equilibrium for gases.

Example 1

To determine the equilibrium concentration of hydrogen gas (H_2) for the reaction:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

Given:

Equilibrium constant, K_c=13.7
Equilibrium concentration of nitrogen, $[N_2] = 1.88\,\text{mol/dm}^3$
Equilibrium concentration of ammonia, $[NH_3] = 6.62\,\text{mol/dm}^3$

Solution

From the balance equilibrium reaction,

$$K_c = \frac{[NH_3]^2}{[N_2]\,[H_2]^3}$$

Substituting for the known equilibrium concentrations and the Kc, this becomes

$$13.7 = \frac{[6.62]^2}{[1.88]\,[H_2]^3}$$

Solve for $[H_2]^3$

$$[H_2]^3 = \frac{(6.62)^2}{1.88 \times 13.7}$$

$$[H_2] = \sqrt[3]{1.7015} = 1.19 \text{ mol/dm}^3$$

Example 2

Consider the following equilibrium reaction at a certain temperature:

$N_2 + 3H_2 \rightleftharpoons 2NH_3$

At equilibrium, the partial pressures are measured as follows:

$P_{N2} = 0.5$ atm, $P_{H2} = 1.5$ atm, $P_{NH3} = 2.0$ atm

Calculate the equilibrium constant K_p for this reaction.

Solution

$$K_P = \frac{P^2_{NH_3}}{P^3_{H_2} \times P_{N_2}}$$

Substitute the given partial pressures into the expression:

$$K_P = \frac{(2.0)^2}{(1.5)^3 \times 0.5}$$

$= 2.3 \text{ atm}^{-2}$

Notice the units for K_P here. These are determined by the powers in the numerator and denominator of the equilibrium constant equation and will be unique to each reaction. Some will have units of atm^{-2}, atm^{-1}, atm, atm^2 or even no units at all! Atm raised to any power is theoretically possible here.

This principle also applies to the units for a K_c value. Where the units of K_c will be (moldm^{-3}) raised to any theoretical power.

Significance of Equilibrium Constants: Kc and Kp

In chemical reactions, **equilibrium constants** help us understand how much of the reactants turn into products when the reaction reaches a balance, known as **equilibrium**.

a. If K_c or K_p is greater than 1,

The reaction **favours products**, meaning more products are formed than reactants left.

b. If K_c or K_p is less than 1,

The reaction **favours reactants**, meaning more reactants remain than formed.

c. If K_c or K_p equals 1,

The reaction is **balanced**, with similar concentrations of reactants and products.

Equilibrium Constants (Kc and Kp) in Real Life

Equilibrium constants, K_c and K_p, help scientists and engineers predict how chemical reactions behave. Here are some real-world examples:

a. Making Ammonia (Fertilizer Production)

Role of Kc and Kp: These constants help determine the best temperature pressure to produce the most ammonia, which is essential for fertilizers.

b. Airbag Safety

Role of Kc and Kp: Understanding these constants ensures the reaction fast enough to protect passengers.

c. Fuel production (petroleum refining):

Role of Kc and Kp: These constants guide how to adjust conditions to get the mix of fuel products.

By understanding K_c and K_p, we can control chemical reactions to make useful products safely and/or efficiently.

3. Relationship between K_c and K_p using ideal gas equation.

The relationship between K_c and K_p for a particular reaction follows the fact that for an ideal gas, $PV = nRT$, dividing both sides by volume, (V) $PV = nRT$

Suppose the balanced equation for a hypothetical equilibrium reaction has the general form

$aA + bB \rightleftharpoons cC + dD$

$$K_p = \frac{P_C^c \cdot P_D^d}{P_A^a \cdot P_B^b} \quad \text{...equation 1}$$

As number of moles divided by volume can be thought of as concentration the partial pressures can be expressed as shown below

$P_A = C_A RT = [A]RT,\ P_B = C_B RT = [B]RT,$

$P_C = C_C RT = [C]RT,\ P_D = C_D RT = [D]RT$

Substituting P=[ ]RT into equation 1

$$K_p = \frac{[C]^c (RT)^c [D]^d (RT)^d}{[A]^a (RT)^a [B]^b (RT)^b}$$

$$= \frac{[C]^c [D]^d (RT)^{(c+d)}}{[A]^a [B]^b (RT)^{(a+b)}}$$

but K_c expression for the same equilibrium reaction, $K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$

$K_p = K_c (RT)^{(c+d)-(a+b)}$

$\Delta n = (c + d) - (a + b)$

$K_p = K_c (RT)^{\Delta n}$

Δn = change in stoichiometric number of moles of products and reactants

NB: a and b are the coefficients of reactants A and B and c and d are the coefficients of products C and D.

NB: The equilibrium constants K_c and K_p have units unless the units cancel, that is when the sum of the powers in the numerator and the denominator are the same.

Example 3

Write the relationship between K_c *and* K_p for each of the following equilibrium reactions:

a. $A(g) \rightleftharpoons B(g) + C(g)$,

b. $2A\ (g) + B\ (g) \rightleftharpoons 4C\ (g)$

c. $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

Solution

a. $A(g) \rightleftharpoons B(g) + C(g)$,

$K_p = K_c (RT)^{\Delta n}$

$\Delta n = (1 + 1) - (1) = 1$

$K_p = K_c RT$

b. $A\ (g) + B\ (g) \rightleftharpoons 4C\ (g)$

$K_p = K_c (RT)^{\Delta n}$

$\Delta n = 4 - (1 + 1) = 2$

$K_p = K_c RT^2$

c. $2\,SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

$K_p = K_c (RT)^{\Delta n}$

$\Delta n = 2 - (2 + 1) = -1$

$K_p = K_c (RT)^{-1}$

$K_p = \frac{K_c}{RT}$

4. K_c and K_p for homogeneous and heterogeneous equilibrium reactions

Homogeneous equilibrium reactions

A homogeneous equilibrium is simply a reversible reaction where everything involved is in the same physical state. This means that all the initial materials (reactants) and all the final materials (products) are either all gases mixed together, or they're all dissolved in the same solution (liquid). They all share the same phase.

For example, $2\,SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

K_c Uses concentrations (mol/dm^3) of the reactants and products in the expression.

$$K_c = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

K_p Uses the partial pressures of the reactants and products when all species are gases.

$$K_p = \frac{P^2_{SO_3}}{P^2_{SO_3} P_{O_2}}$$

Heterogeneous equilibrium reactions

A heterogeneous equilibrium is a reversible reaction where the materials are not all in the same state. This means the reactants and products are a mix of different phases, such as solids, liquids, and gases all reacting together. They don't share the same physical state.

Equilibrium Constants

In these reactions, pure solids and pure liquids are not included in the expression for K_c or K_p because their concentrations remain constant.

Kc: Includes only the concentrations of gaseous or aqueous species.

Example: For the reaction $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$

$$K_c = \frac{[CaO][CO_2]}{[CaCO_3]}$$

$$K_c = [CO_2]$$

Kp: Includes only the partial pressures of gases.

$$K_p = P_{CO_2}$$

Example 4

Write the K_c and K_p expressions for the following processes:

- **a.** $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$
- **b.** $NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g)$

Solution

a. Here, carbon (C) is a solid, so it is not included in the K_c expression. Only reactants and products are included:

$$K_c = \frac{[CO][H_2]}{[H_2O]}$$

$$K_p = \frac{P_{CO} \times P_{H_2}}{P_{H_2O}}$$

b. Ammonium chloride (NH_4Cl) is a solid, so it is not included in the K_c expression. O the concentrations of the gaseous products, NH_3 and HCl, are included:

$K_c = [HCl][NH3]$

$K_p = P_{HCl} \times P_{NH_3}$

5. Calculations involving K_{sp} of sparingly soluble salts

The solubility product constant (K_{sp}) simply tells us the maximum amount of a substance that can dissolve in water before it starts forming a solid and dropping to the bottom. It measures how much that substance is willing to dissolve. It is used for salts that do not dissolve completely. A higher K_{sp} means the salt dissolves more. If more than 1 gram of a salt can dissolve in 100 cm³ of water, it is considered soluble.

When a salt dissolves in water, it breaks apart into ions and reaches a balance, or *equilibrium*, between the solid salt and the ions in the solution. For example, if a salt is written as A_xB_y, it dissolves like this:

$A_xB_y(s) \rightleftharpoons xA^{y+}(aq) + yB^{x-}(aq)$

The K_{sp} formula for this salt is:

$Ksp = [A^{y+}]^x[B^{x-}]^y$

To find K_{sp}, use the concentrations of the ions, raised to the numbers in the formula. The solid salt itself is not included in the K_{sp} calculation because its concentration doesn't change.

Example 5

The solubility of calcium fluoride (CaF_2) is 0.015 $moldm^{-3}$. Calculate the K_{sp} of calcium fluoride.

Solution

The dissociation of calcium fluoride in water is represented by the equation:

$CaF_2(s) \rightleftharpoons Ca^{2+}(aq) + 2F^-(aq)$

Let the solubility of CaF_2 be $S = 0.015\ moldm^{-3}$.

This means:
$[Ca^{2+}] = S = 0.015\ moldm^{-3}$
$[F^{-}] = 2S = 2 \times 0.015$
$= 0.030\ moldm^{-3}$

Now, the K_{sp} expression is:
$K_{sp} = [Ca^{2+}][F^{-}]^2$

Substitute the values:
$K_{sp} = (0.015)(0.030)^2$

$=0.015 \times 0.0009$

$=1.35\times10^{-5}$

Thus, the K_{sp} of CaF_2 is $1.35 \times 10^{-5}\ mol^3dm^{-9}$

Example 6

The solubility of lead (II) iodide (PbI_2) is $1.1 \times 10^{-4}\ moldm^{-3}$. Calculate the K_{sp} of lead (II) iodide.

Solution

The dissociation of lead(II) iodide (PbI_2) is:

$PbI_2(s) \rightleftharpoons Pb^{2+}(aq) + 2I^{-}(aq)$

Let the solubility of PbI_2 be $S = 1.1 \times 10^{-4}$

This means:
$[Pb^{2+}] = S=1.1 \times 10^{-4}\ moldm^{-3}$
$[I^{-}] = 2S = 2 \times 1.1\times10^{-4}\ moldm^{-3}$

Now, the K_{sp} expression is:
$K_{sp}=[Pb^{2+}][I^{-}]^2$

Substitute the values:
$K_{sp}=(1.1 \times 10^{-4})(2.2 \times 10^{-4})^2$
$= (1.1 \times 10^{-4})(4.84 \times 10^{-8})$
$= 5.32 \times 10^{-12}\ mol^3dm^{-9}$

Example 7

Given that the **Ksp of PbCl2** is 1.6×10^{-5}, calculate the solubility of $PbCl_2$ in moldm^{-3}.

Solution

The dissociation of $PbCl_2$ in water is given by:

$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^{-}(aq)$

Let the solubility of $PbCl_2$ be S moldm^{-3}. At equilibrium:

The concentration of Pb^{2+} will be S.

The concentration of Cl^{-} will be 2S (because each $PbCl_2$ produces 2 Cl^{-} ions).

The expression for K_{sp} is:

$K_{sp} = [Pb^{2+}][Cl^{-}]^2$

Substitute the concentrations in terms of S:

$K_{sp} = (S)(2S)^2 = 4S^3$

Now, substitute the given value of K_{sp}

$1.6 \times 10^{-5} = 4S^3$

Solve for S:

$$S^3 = \frac{1.6 \times 10^{-5}}{4} = 4.0 \times 10^{-6}$$

$$S = \sqrt[3]{4.0 \times 10^{-6}} = 1.58 \times 10^{-2} \text{ mol dm}^{-3}$$

So, the solubility of $PbCl_2$ is 1.58×10^{-2} mol dm^{-3}

Example 8

The **Ksp of BF3** is 3.9×10^{-3}. Calculate the solubility of BF_3 in moldm^{-3}.

Solution

The dissociation of BF_3 in water is:

$BF_3(s) \rightleftharpoons B^{3+}(aq) + 3F^{-}(aq)$

Let the solubility of BF_3 be S moldm^{-3}. At equilibrium:

The concentration of **B3**$^{+}$ will be S.

The concentration of **F**$^{-}$ will be 3S (since three moles of F^{-} are produced for each mole of BF_3).

The expression for **Ksp** is:

$K_{sp} = [B^{3+}][F^{-}]^3$

Substitute the concentrations in terms of S:

$K_{sp} = (S)(3S)^3 = 27S^4$

Now, substitute the given value of K_{sp}:

$3.9 \times 10^{-10} = 27S^4$

Solve for S:

$$S^4 = \frac{3.9 \times 10^{-10}}{27} = 1.44 \times 10^{-11}$$

$$S = \sqrt[4]{1.44 \times 10^{-11}} = 1.94 \times 10^{-3} \text{ mol dm}^{-3}$$

Activity 3.3 Deriving the Expression for the Equilibrium Constant (K_c or K_p)

Steps

1. Using the reversible reactions:
   - **a.** $aA + bB \rightleftharpoons cC + dD$,
   - **b.** $2NH_3 \rightleftharpoons N_2 + 3H_2$
   - **c.** $CO_2 + H_2O \rightleftharpoons H_2CO_3$

   deduce the expression for the equilibrium constant (K_c or K_p) in mixed ability groups.
2. Present your findings and discuss any differences or misunderstandings.

Activity 3.4 Discussing the meaning of the equilibrium constant

Steps

1. Discuss what the value of the equilibrium constant K_c or K_p tells them about a reaction.
2. What does a large value of K mean about the products vs. the reactants?
3. What about a small value of K?

4. How does the value of K influence how we predict the outcome of a reaction?

Activity 3.5 Relationship Between K_p and K_c using the Ideal Gas Equation

Steps

1. Using the ideal gas equation: PV = nRT.

   Discuss how PV = nRT relates to K_p by introducing the formula:

   $K_p = K_c\ (RT)^{\Delta n}$

   where Δn is the change in the number of moles of gas between products and reactants

2. Write the K_p and K_c expressions for the reaction

   $N_2(g) + 2O_2(g) \rightleftarrows 2NO_2(g)$

3. Calculate the Kp using the given value of Kc= 5.21 × 10 −5 moldm^{-3}, temperature (T) = 500 °C, gas constant ® = R = 8.31 kPa dm^3 mol^{-1} K^{-1}

Activity 3.6 Sparingly soluble salts and Calculations Involving K_{sp} based on solubility data

Steps

1. Using $PbSO_4(s)$ as an example of sparingly soluble salts:

   $PbSO_4(s) \rightleftarrows Pb^{2+}(aq) + SO_4^{2-}(aq)$

   Given the equilibrium concentrations: $[SO_4^{2-}] = 1.48 \times 10^{-4}$ moldm^{-3},

   $[Pb^{2+}] = 5.7 \times 10^{-5}$ moldm^{-3}

2. Calculate the K_{sp} of $PbSO_4$.

3. Compare your results with others.

LE CHATELIER’S PRINCIPLE AND FACTORS THAT AFFECT EQUILIBRIUM POSITIONS

Le Chatelier’s Principle

Imagine a see-saw or a balance scale. When it is perfectly balanced, we say it is in “equilibrium.” But if someone pushes down on one side, the see-saw tilts. To balance it again, you need to make changes to bring it back to even.

In a chemical reaction, “equilibrium” means the forward and backward reactions are happening at the same rate, so everything stays balanced. If something changes—like adding more weight (or a reactant/product) or changing how hard someone pushes (like temperature or pressure)—the reaction shifts to balance things out again. This is what

Le Chatelier’s Principle is all about!

Factors That Affect Equilibrium

a. Concentration (Amount of Stuff)

If you add more of one side (reactants or products), the reaction balances it by making more of the opposite side.

Example: Imagine a teeter-totter where you add more toys on one side. To balance it, you move more toys to the other side.

If a reaction is: $\mathbf{A + B \rightleftharpoons C + D}$, and you add more A, the reaction will shift to the right (using up A and making more C and D). If you take away some C, the reaction will also shift to the right to make up for the lost C.

b. Pressure and Volume

Imagine blowing air into a balloon. If you make the balloon smaller by squeezing (reducing the volume), the air inside feels squished, and the pressure goes up. If let the balloon expand (increasing the volume), the air spreads out, and the goes down.

In chemical reactions with gases, the same idea works! How much space the have (volume) and how squished they are (pressure) can change the balance of reaction.

What Happens When Pressure or Volume Changes?

More Pressure (Smaller Volume)

If you squeeze the gases into a smaller space (increase pressure), the reaction will reduce the pressure.

Example: In a reaction like: $xA \rightleftharpoons yB$

If one side of the reaction has fewer gas molecules, the reaction will shift towards side to lower the pressure.

Less Pressure (Bigger Volume)

If you give the gases more room (increase volume), the reaction increases the. pressure.

Example: The reaction will shift to the side with more gas molecules to balance out.

No Effect if the Gases are Equal

If both sides of the reaction have the same number of gas molecules, changing the pressure or volume will not change the balance.

c. Temperature and Equilibrium

Two Types of Reactions

i. Endothermic Reactions (Heat is Absorbed):

Example: $A + B + \text{Heat} \rightleftharpoons C$

If you turn up the temperature (add heat), the reaction shifts make more **C** (products) thereby lowering the temperature.

If you cool it down (remove heat), the reaction shifts backward to make more **A** and **B** (reactants) to produce more heat.

ii. **Exothermic Reactions (Heat is Released)**

Example: $D + E \rightleftharpoons F + \text{Heat}$

If you turn up the temperature (add heat), the reaction gets rid of the extra by shifting backward to make more **D** and **E** (reactants).

If you cool it down (remove heat), the reaction shifts forward to make more **F** (products) because that releases heat.

d. Catalyst Simplified

A **catalyst** speeds up a chemical reaction without getting used up. It makes the reaction happen faster by creating a lower energy pathway for the molecules to react.

Even though the catalyst helps the reaction go faster, it doesn't change the final balance of the reaction (the equilibrium). It doesn't affect the concentrations of reactants and products at the end; it just helps the reaction reach that point more quickly.

Activity 3.7 Exploring factors that affect Chemical Equilibrium

Materials Needed: Flashcards with equilibrium scenarios (e.g., adding reactants, changing pressure, etc.), a simple balance scale or see-saw diagram to represent equilibrium

Step

1. State the factors that can disturb equilibrium.
2. From the given flashcards with specific scenarios, predicting changes when:
   a. Add more reactant.
   b. Increase temperature in an exothermic reaction.
   c. Decrease volume of a gas mixture."
3. For each scenario predict which side reaction will shift to.

Activity 3.8 Applying Le Chatelier's Principle to Optimise Reaction Conditions

Steps

1. Use the simple reaction: $A(g) + B(g) \rightleftharpoons C(g) + 2D(g)$

   Discuss the following conditions:
   a. What happens if more of A is added to the reaction?
   b. What happens if C is removed from the reaction?
   c. If pressure is increased, what will happen to the reaction?
2. Use the simple reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)\ \Delta H = -x\,kJmol - 1$

Discuss the following conditions:

a. What happens when temperature increased?

b. What happens to equilibrium position?

c. What effect will increase in temperature have on K_p?

APPLICATION OF THE CONCEPTS OF EQUILIBRIUM AND RATES TO INDUSTRIAL PROCESSES

The principle of chemical equilibrium and rates of reaction are used by chemists to help make industrial processes more efficient. By understanding how reactions behave and how to control certain factors, industries can increase the volume of desired products. Le Chatelier's Principle helps predict how changes in temperature, pressure, or concentration affect a reaction at equilibrium, allowing chemists to make adjustments that increase production.

1. **The Haber Process (Making Ammonia)**

   The Haber process is used to make ammonia (NH_3), which is important for fertilizers. The reaction is:

   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) \qquad \Delta H = -92\,kJmol^{-1}$

   This reaction is exothermic, meaning it releases heat. To make more ammonia, chemists:

   Lower the temperature (even though it slows the reaction, it favours the production ammonia).

   Increase the pressure (this helps shift the reaction toward ammonia, since there fewer molecules on the product side).

   Use an iron catalyst to speed up the reaction.

2. **The Contact Process (Making Sulfuric Acid)**

   In the contact process, sulphur is used to make sulfuric acid, an important chemical industries. The key reaction in this process is:

   $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) \qquad \Delta H = -196\,kJmol^{-1}$

   To maximize the production of SO_3 (which is used to make sulfuric acid), chemists:

   Increase the pressure (since there are fewer gas molecules on the product side).

   Lower the temperature (this favours the production of SO_3).

Use a vanadium(V) oxide catalyst to speed up the reaction.

3. **Catalytic Converters in Cars**

In cars, harmful gases like carbon monoxide (CO), hydrocarbons (C_xH_y), and oxides (NO and NO_2) are produced when fuel is burned. Catalytic converters help these harmful gases into less harmful ones like carbon dioxide (CO_2), nitrogen (N_2),

and water (H_2O).

The reduction process converts NO to N_2 and O_2:

$$2NO(g) \rightleftharpoons N_2(g) + O_2(g)$$

The oxidation process converts CO to CO_2:

$$2CO(g) + O_2(g) \rightleftharpoons 2CO_2(g)$$

To make these reactions work better, high temperatures are used (because reactions are endothermic, meaning they need heat). Catalysts like platinum (Pt) speed up these reactions. Le Chatelier's principle explains that increasing concentration of harmful gases makes the reaction move forward to create harmful gases.

4. **Petroleum Refining**

In petroleum refining, catalytic cracking and reforming are used to turn long- hydrocarbons into useful fuels like gasoline and diesel.

Catalytic cracking breaks down large molecules into smaller ones (this process heat).

$$C_{12}H_{26}(g) \rightleftharpoons C_6H_{14}(g) + C_7H_{16}(g)$$

Catalytic reforming rearranges molecules to make high-octane gasoline. This also needs heat.

Both of these reactions are controlled by Le Chatelier's principle, and conditions temperature are adjusted to get the right products.

5. **Water Treatment (Ion Exchange)**

In water treatment, hard water (which contains calcium and magnesium ions) can made soft by removing these ions. One way to do this is with ion exchange, calcium ions react with chloride ions:

$$Ca^{2+}(aq) + 2Cl^-(aq) \rightleftharpoons CaCl_2(s)$$

By adding brine (NaCl), which increases the chloride ions, the reaction shifts produce more calcium chloride ($CaCl_2$). This helps remove calcium and ions from the water, making it softer.

Activity 3.9 Mini Research Work

Go online and find other sources about how Le Chatelier's principle is applied in industrial processes.

Activity 3.10 Group Presentations on Economic Reaction Optimization

Objective: In small groups, students work together to learn how chemists use special processes to produce important chemicals like **ammonia** (Haber Process) and **sulphuric acid** (Contact Process) efficiently.

Steps

1. **a.** Research the Haber Process for ammonia production using books, handouts, online resources to gather information.

   b. Research the **Contact Process** for sulphuric acid production using books, handouts, or online resources to gather information.

   Research tasks

   i. Explain the main reaction in each process.

   ii. Describe how Le Chatelier's Principle helps maximise the yield.

   iii. Identify the factors that are controlled in each process (temperature, pressure, catalyst).

   iv. Create a simple diagram or flowchart to show the steps in the process.

2. Create a short presentation for the class.

   The presentation should include:

   a. What the process is used for.

   b. How Le Chatelier's Principle is applied to optimise the reaction.

   c. What conditions (temperature, pressure, catalysts) are used to get the highest yield.

   d. Use charts, drawings, or a simple poster or a PowerPoint slide to help explain.

3. Class reflections:

   a. Why do you think chemists want to control the temperature and pressure in these reactions?

b. How does Le Chatelier’s Principle help in getting the best results in these processes?

c. How does understanding these processes help us in real life?

EXTENDED READING

Refer to the following links for additional reading materials:

- Link 1
- Link 2
- Link 3

REVIEW QUESTIONS

Review Questions 3.1

1. What is a reversible reaction?
2. Give an example of an irreversible reaction.
3. What is meant by "dynamic equilibrium" in a reversible reaction?
4. Describe the visual change observed when water is added to anhydrous copper (II) sulphate.
5. Why doesn't dynamic equilibrium occur in an open system?
6. How is a reversible reaction represented in a chemical equation?
7. Compare reversible and irreversible reactions in terms of product formation energy change.
8. Explain how dynamic equilibrium is crucial in biological systems, providing example.

Review Questions 3.2

1. Given the reaction: $2A(g) + B(g) \rightleftharpoons 3C(g)$

   If the equilibrium concentrations of the species are: $[A] = 0.2$ moldm^{-3}, $[B] = 0.5$ moldm^{-3}, and $[C] = 1.0$ moldm^{-3}, calculate the constant K_c.
2. What does a large value of the equilibrium constant (K_c) indicate about the position equilibrium in a chemical reaction?
3. For the following reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

   At equilibrium, the concentrations of the species are: $[N_2] = 0.1$ moldm^{-3}, $[H_2] = 0.3$ moldm^{-3}, $[NH_3] = 0.2$ moldm^{-3}

   Calculate the equilibrium constant K_c.
4. Given the equilibrium reaction:

   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

   If the equilibrium constant K_c is very small, what can you conclude about concentrations of reactants and products at equilibrium?

5. The following reaction is at equilibrium: $2NO_2(g) \rightleftharpoons N_2O_4(g)$

   At equilibrium, the concentration of NO_2 is 0.3 $moldm^{-3}$ and N_2O_4 is 0.1 $moldm^{-3}$. If initial concentration of NO_2 was 0.5 $moldm^{-3}$, calculate the equilibrium constant K_c. Assume the reaction reached equilibrium and no other substances are involved.

6. For the reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

   If the equilibrium constant K_c is much greater than 1, what would happen to system if the volume of the container is decreased? Justify your answer using Le Chatelier's principle.

7. For the reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

   At a certain temperature, K_p is found to be 2.0×10^3 If the change in the number moles of gas, Δn is 0, what is the relationship between K_p?

8. The solubility of lead(II) iodide (PbI_2) in water is 1.2×10^{-3} mol/dm^{-3}. Calculate the K_{sp} PbI_2. Then, if the solubility is increased to 2.4×10^{-3} $moldm^{-3}$, calculate the new K_{sp} explain why K_{sp} does not change with a change in solubility.

Review Questions 3.3

1. State Le Chatelier's Principle.
2. What is the main product of the Haber Process and what are its uses?
3. If the concentration of a reactant in a system at equilibrium is increased, which direction will the equilibrium shift according to Le Chatelier's Principle?
4. In the Haber Process, what are the key factors that affect the production of ammonia, and how do they influence the reaction?
5. Consider the following equilibrium reaction:

   $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

What will happen to the position of equilibrium if the **temperature** is increased? Assume the reaction is **exothermic**.

6. In the Contact Process for making sulphuric acid, what are the main (temperature, pressure, and catalyst) that affect the reaction, and how are optimised for maximum yield?
7. In the Contact Process, how does the presence of a catalyst influence the rate reaction, and how does it help optimise the production of sulfuric acid?

8. The following reaction is at equilibrium:

   $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

   If the **pressure** is increased by reducing the volume of the container, which direction ill the equilibrium shift? Explain your reasoning based on Le Chatelier's Principle.

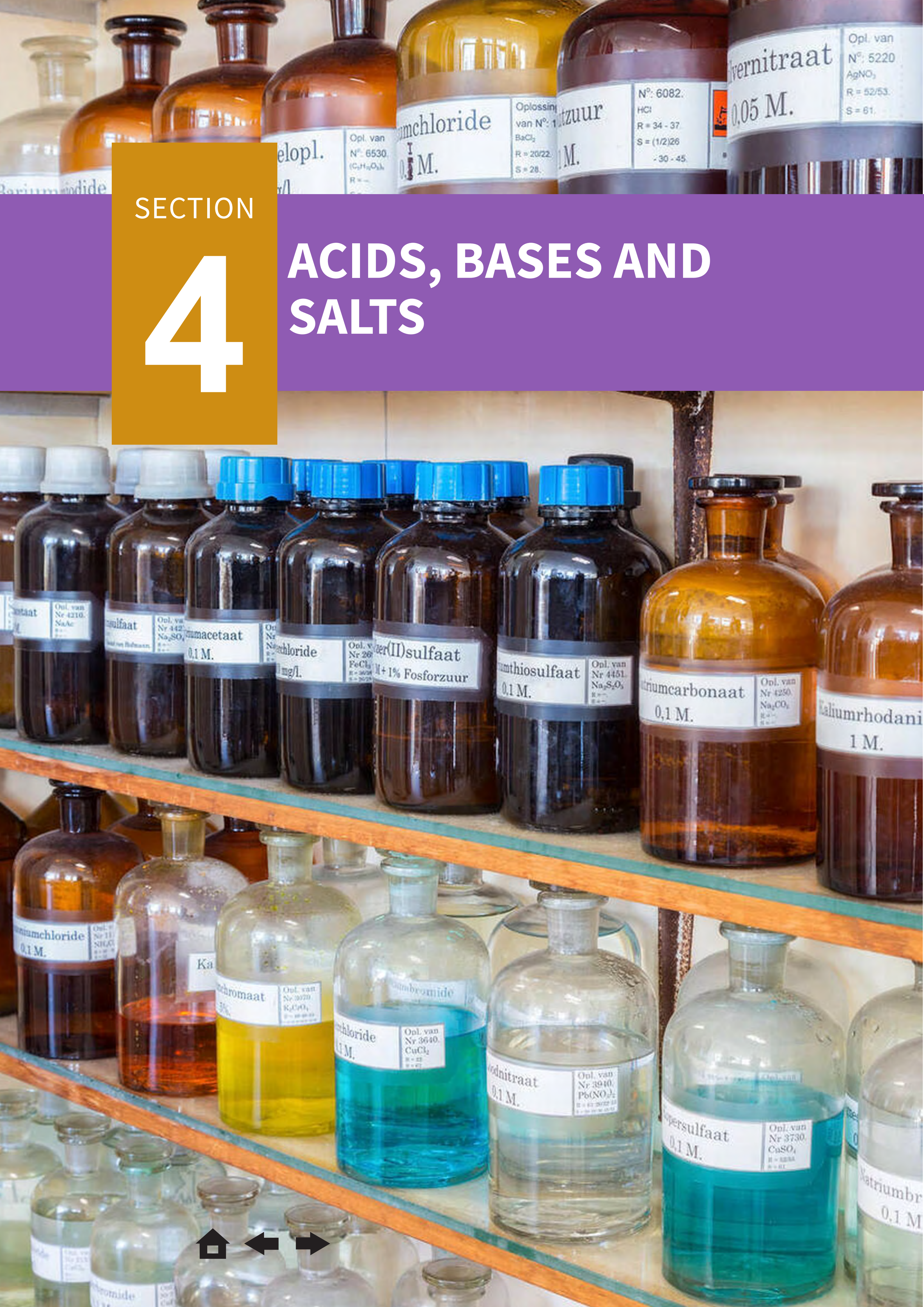
9. Explain how Le Chatelier's Principle applies to the change in the value of equilibrium constant (K) when the temperature is changed.

10. Explain how Le Chatelier's Principle is applied in the Haber Process to ammonia production. What changes in temperature and pressure are made, and why?

SECTION

4

ACIDS, BASES AND SALTS



PHYSICAL CHEMISTRY

Equilibria

INTRODUCTION

In section, we will explore the basics of acids and bases in chemistry. You will learn about different acid-base theories like Arrhenius, Brønsted -Lowry, and Lewis. We will understand the physical and chemical properties of acids and bases, including their everyday uses. This section also covers salts—what they are, how they're made, and their practical uses. Students will practice acid-base titration methods, like simple titration and back titration.

KEY IDEAS

- **Analyte** is a solution of unknown concentration (in the flask).
- **Arrhenius acids** are substances that produce hydrogen ions (H^+) when dissolved in water.
- **Arrhenius bases** are substances that produce hydroxide ions (OH^-) when dissolved in water.
- **Brønsted -Lowry base** a substance that accepts a proton (H^+) from another substance.
- **Brønsted-Lowry acid** a substance that donates a proton (H^+) to another substance.
- **Conjugate acid-base pair** is two species that differ by a single proton (H^+).
- **Endpoint** is a point where the indicator changes colour, showing neutralization.
- **Indicator** is a substance that changes colour to show when a reaction is complete.
- **Lewis acid** a substance that accepts a pair of electrons to form a covalent bond.
- **Lewis base** a substance that donates a pair of electrons to form a covalent bond.

- **Titrant** is a solution of known concentration (in the burette).
- **Titre value** is the volume of titrant used to reach the endpoint.

CONCEPTS OF ACIDS AND BASES

Acids and bases are two important types of chemicals used in everyday items like cleaning products, medicines, and food. Long ago, people could tell them apart by taste and feel: acids have a sour taste, like in vinegar or lemon juice, and bases feel slippery or soapy, like some cleaners.

1. Arrhenius concept of acids and bases

The first scientist to provide a more thorough chemical explanation (as opposed to taste/feel) of the characteristics of bases and acids was Arrhenius. He defines an acid as a substance that produces hydrogen ions or hydronium ions when dissolved in water (aqueous solution). Using H_2SO_4 as an example, in water, the molecules of H_2SO_4 break up (dissociate) to give its ions.

$H_2SO_4\,(aq) \rightarrow 2H^+ + SO_4^{2-}$

$H_2SO_4 + H_2O \rightarrow H_3O^+ + SO_4^{2-}$

Therefore, sulphuric acid contains hydrogen ions or hydronium ions. All other aqueous solutions of acids also contain hydrogen ions and this is what gives them their acidity.

A base, according to Arrhenius, is a compound or substance that produces hydroxide ions when dissolved in water (aqueous) solution.

Using NaOH as an example, in aqueous solution NaOH molecules dissociate to give Na^+ and OH^- ions.

$NaOH\,(aq) \rightarrow Na^+ + OH^-$

Activity 4.1 Exploring Arrhenius Theory of Acids and Bases

Material needed: Computer, and videos to guide their research.

Steps

1. Introduce Arrhenius Theory
   a. Give a simple explanation of Arrhenius acids.

  b. Give 2 examples with diagrams.

2. In small groups, work together to explore and discuss Arrhenius acids and bases.

3. Discuss and share ideas on what makes Arrhenius' theory useful. How does it help explain basic reactions?

4. In your groups, research the strengths (e.g., clear definitions for acids and bases in water) and limitations (e.g., theory only applies to water-based solutions, does not explain all acid-base reactions).

Each group should prepare a short presentation with their findings, using visual aids like charts, chemical reactions, or videos to make it engaging and easy to understand.

Limitation of Arrhenius' concepts

Arrhenius definition of acids and bases are limited because they apply only to aqueous solutions.

2. Brønsted -Lowry concept of acids and bases

Two scientists, **Brønsted** and Lowry, gave a broader definition of acids and bases. They said an acid-base reaction happens when a hydrogen ion is transferred from one compound to another. This can happen even when the reaction is not in water.

A hydrogen ion (H^+) is formed when a hydrogen atom loses its electron, so H^+ is like a proton. **Brønsted** and Lowry defined an acid as something that gives away a proton (H^+) and a base as something that accepts a proton. We call these **Brønsted** -Lowry acids and **Brønsted** -Lowry bases.

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

Here HNO_3 is acting as a **Brønsted** -Lowry acid, and H_2O is acting as a **Brønsted** -Lowry base.

$H_2O + NH_3 \rightarrow NH_4^+ + OH^-$

Here NH_3 is acting as a **Brønsted** -Lowry base, and H_2O is acting as a **Brønsted** -Lowry acid.

This means that in a **Brønsted** – Lowry acid – base reaction, the acid donates protons to the base.

Not all hydrogen atoms in a compound can be given away as ions, so not every compound with hydrogen is a **Brønsted** -Lowry acid. For example, methane (CH_4) is not acidic. Whether hydrogen can act as an acid depends on certain properties like polarity (how the atoms are attracted to each other) and the number of lone electron pairs on the atom it's bonded to. Elements from groups 16 and 17, like oxygen and chlorine, when bonded with hydrogen, can form **Brønsted** -Lowry acids. Common examples of these acids in chemistry labs are HCl, HNO_3, H_2SO_4, CH_3COOH, H_3PO_4, and H_2CO_3.

The basicity of a compound depends on having lone pairs of electrons to accept protons. Elements from group 15 and group 16 often form bases. Common **Brønsted** -Lowry bases are NaOH, KOH, NH_3, C_5H_5N, and Na_2CO_3.

Brønsted -Lowry acids are grouped by how many protons they can donate. **Monoprotic acids** donate one proton (like HCl and HNO_3).

Polyprotic acids donate more than one proton.

Diprotic acids (like H_2SO_4 and H_2CO_3) donate two protons, while

Triprotic acids (like H_3PO_4) donate three.

Brønsted -Lowry bases can also be **monoprotic** (accept one proton, like NaOH and NH_3) or **diprotic** (accept two protons, like CO_3^{2-}). Some compounds can act differently in water, changing based on their surroundings.

When hydrogen chloride gas (HCl) dissolves in water, it gives a hydrogen ion to the water:

$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

In this reaction, HCl is a **Brønsted** -Lowry acid because it donates a hydrogen ion to water, while water acts as a **Brønsted** -Lowry base by accepting the hydrogen ion to form H_3O^+ (hydronium ion).

On the other hand, when ammonia (NH_3) dissolves in water, it accepts hydrogen from water:

$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$

Here, water donates a hydrogen ion to ammonia, making water a **Brønsted** -Lowry acid and ammonia a **Brønsted** -Lowry base.

Since water can act as an acid with NH_3 and as a base with HCl, it is called an **amphiprotic** compound (meaning it can act as both an acid and a base).

Conjugate Acid-Base Pairs

In the **Brønsted** -Lowry theory, acids and bases are connected through “conjugate pairs.” When an acid loses a hydrogen ion (proton), it forms its conjugate base. The acid and its conjugate base differ by one proton. Similarly, when a base gains a proton, it forms its conjugate acid. Each **Brønsted** -Lowry acid has a conjugate base, and each base has a conjugate acid.

For example

1. $H_2O + HCl \rightarrow H_3O^+ + Cl^-$

Here, H_2O (base) and H_3O^+ form one conjugate pair, while HCl and Cl^- form another.

2. $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

NH_3 and NH_4^+ are one pair, and H_2O and OH^- are another.

Activity 4.2 Understanding Brønsted-Lowry Theory of Acids and Bases

Activity Steps

1. Give a simple explanation of Brønsted -Lowry concept of acids and bases.
   a. Use visuals like diagrams, charts, illustrate this idea, showing how protons from acids to bases.
2. In small groups,identify compounds that have *hydrogen atoms*.
   a. Classify the compounds into *simple acids* (e.g., HCl), *oxoacids* (e.g., HNO_3), *hydrated cations* (e.g., $Al(H_2O)_6^{3+}$).
   b. Discuss which compounds could act as Brønsted -Lowry acids and bases.
3. Provide examples of acid-base reactions (e.g., $HCl + NH_3 \rightarrow NH_4^+ + Cl^-$).
4. Demonstrate and discuss with the class how *proton transfer* occurs, which substance is the acid and which is the base in each example.
5. Discuss the strengths of Brønsted -Lowry Theory:
   a. Discuss why this theory is helpful: for example, it explains more types of acid-base reactions than Arrhenius’ theory and works in non-aqueous (non-water) environments.

6. Discuss situations where the theory doesn't fully explain certain reactions (e.g., reactions that don't involve proton transfer).
   a. Identify examples where the theory might not be complete.
7. Present your findings to the class, including examples of Brønsted -Lowry acids and bases, reactions with proton transfer, and key strengths and limitations.

3. Lewis Acids and Bases

The Lewis theory expands the idea by focusing on electron pairs rather than protons. According to Lewis, an acid is a substance that accepts an electron pair, while a base donates an electron pair to form a bond. For example:

$H^+ + NH3 \rightarrow NH4^+$

Here, H^+ (Lewis acid) accepts electrons from NH_3 (Lewis base).

Another example is **$NH3 + BF3 \rightarrow F3B–NH3$**, where NH_3 donates an electron pair to BF_3, creating a bond.

A Lewis acid does not necessarily need a hydrogen, and this is what separates it from a **Brønsted-Lowry acid**.

Activity 4.3 Collaboratively Exploring Lewis Theory of Acids and Bases

Materials needed: Chart papers or whiteboard, markers, handouts with examples Lewis Acids and Bases

Activity Outline

1. In small groups,brainstorm different examples of Lewis acids.

   Starting points, such as:
   a. Metal cations (e.g., Al^{3+})
   b. Electron-deficient molecules like BF_3

   Discuss why each example can act as a Lewis acid, focusing on its ability to accept electron pairs.

2. In small groups,brainstorm different examples of Lewis acids and L bases.

   Starting points, such as:
   a. Anions e.g.O^{2-}

b. Electron-rich molecules like H_2O

Discuss why each example can act as a Lewis bases focusing on its ability to donate electron pairs.

3. Using the chemical reaction provided identify Lewis acids and bases in each reaction:
   a. $Cu^{2+}(aq) + 4NH_3(aq) \rightarrow \left[Cu\left(NH_3\right)_4\right]^{2+}(aq)$
   b. $NH_3 + H^+ \rightarrow NH_4^+$
   c. $BF_3 + NH_3 \rightarrow BF_3NH_3$
4. Draw a diagram to show how ammonia donates an electron pair to the proton.
5. Explain that this is a Lewis acid-base reaction involving a coordinate covalent bond.
6. Brainstorm and discuss the strengths and limitations of the Lewis theory.

Activity 4.5 Identifying Conjugate Acid-Base Pairs

Materials Needed: Acid-base reaction cards (each showing a different acid-base reaction), coloured markers, and a chart with spaces for acids, bases, conjugate acids, and conjugate bases.

1. Organise yourselves into groups of no more than five. Each group will receive a reaction card.
2. Analyse the reaction card, identify the acid and base.
3. Determine the conjugate acid-base pairs.
4. Show the species donating a proton and which is accepting it.

PHYSICAL AND CHEMICAL PROPERTIES OF ACIDS AND BASES

1. Physical properties of acids and bases

a. Physical properties of acids

i. Acids have a sour, sharp taste.

ii. They change blue colour litmus paper to red.

iii. They change the colour of Methyl Orange/Yellow to Pink.

iv. Acidic substances change the colour of Phenolphthalein from deep pink colourless.

v. Have a wet feel.

vi. They are good conductors of electricity.

vii. Acids have a pH value of less than seven.

b. Physical Properties of bases

i. Bases have a bitter taste.

ii. They change red litmus paper to blue.

iii. Basic substances change phenolphthalein from colourless to pink.

iv. Feel slippery between the fingers (they have a soapy feel).

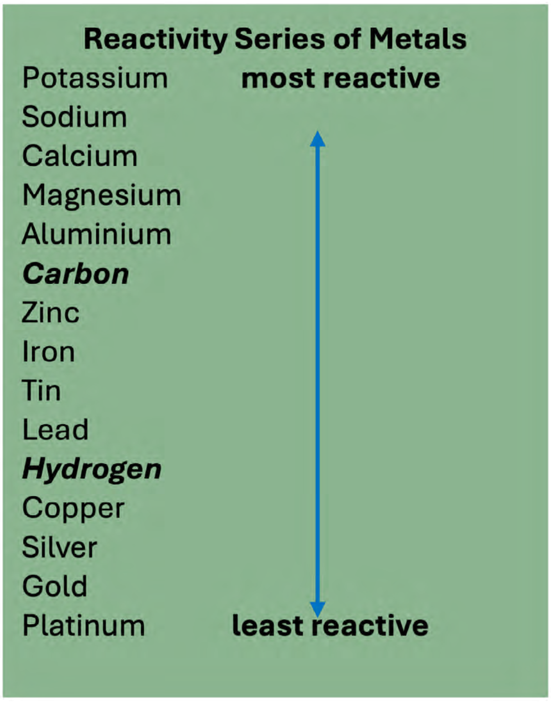
v. Bases conduct electricity in aqueous solution.

vi. Bases have a pH value greater than seven

2. Chemical properties of acids and bases

a. Chemical properties of acids

i. Acids react with some metals, (that are above hydrogen in the series, common examples are shown below) to produce hydrogen gas.



Reactivity Series of Metals

Potassium **most reactive**
Sodium
Calcium
Magnesium
Aluminium
Carbon
Zinc
Iron
Tin
Lead
Hydrogen
Copper
Silver
Gold
Platinum **least reactive**

A typical reaction is that between hydrochloric acid and calcium:

$H_2SO_4\ (aq) + Ca\ (s) \rightarrow CaSO_4\ (aq) + H_2\ (g)$

ii. Acids react with carbonates and bicarbonates, such as Na_2CO_3, $CaCO_3$, $NaHCO_3$, to produce carbon dioxide gas.

$CaCO_3(s) + H_2SO_4\ (aq) \rightarrow CaSO_4\ (aq) + H_2O\ (l) + CO_2\ (g)$

iii. Acids react with bases to produce salt and water. This reaction is neutralisation.

Acid + base → salt + water

$HCl\ (aq) + NaOH\ (aq) \rightarrow NaCl\ (aq) + H_2O\ (l)$

b. Chemical properties of bases

i. The aqueous salt solution of heavy metals reacts with soluble bases (alkalis) form a precipitate (ppt) of the metallic hydroxide.

$3NaOH\ (aq) + AlCl_3\ (aq) \rightarrow A(OH)_3\ (s) + 3NaCl\ (aq)$ White ppt

ii. Bases reacts with warmed ammonium salt to give off ammonia gas

$NaOH\ (aq) + NH_4Cl\ (s) \rightarrow NaCl\ (aq) + H_2O\ (l) + NH_3\ (g)$

iii. Bases reacts with acids to form salt and water

$KOH\ (aq) + HCl\ (aq) \rightarrow KCl\ (s) + H_2O\ (l)$

3. Classification of household substances as acids and bases

Household products are categorized according to their chemical and physical characteristics. The majority of these tests are qualitative and may be conducted outside of a laboratory. The common testing consists of

a. Taste (Do not taste any chemical in the laboratory)

Acids are sour-tasting chemicals. Because they contain more acid than oranges, unripe oranges have a highly sour taste. Conversely, bases have a taste.

b. Litmus paper test

Acidic substances or their solutions turn blue litmus paper red. Bases or their solutions turns red litmus paper blue.

c. The pH scale

Aqueous solutions of acids have pH values less than 7. Whiles aqueous of bases have pH values greater than 7.

d. Universal indicator test

A combination of many indicators that change colour in response to a substance's pH level is known as a universal indicator. By comparing the colour change wit colour code that matches the pH value, one may determine if a chemical is basic, neutral, or acidic.

An acidic solution can become yellow, beige, orange, pink, or red when with the universal indicator. The resultant solution can be dark green, turquoise, pale blue, blue, dark blue, violet, or purple when combined with a basic solution. Universal indicator turns green when using neutral solutions.

4. Application of neutralisation reactions in real life

a. Application of Neutralisation Reactions: Antacids

When we have too much stomach acid, it can cause discomfort, like and acid reflux. **Antacids** are medicines that help by **neutralising the stomach acid**.

Our stomach naturally has hydrochloric acid (**HCl**) to help digest food, but much of it leads to problems. **Antacids** contain basic compounds magnesium hydroxide ($Mg(OH)_2$), aluminium hydroxide ($Al(OH)_3$), carbonate ($CaCO_3$), or sodium bicarbonate ($NaHCO_3$).

These compounds react with the excess hydrochloric acid in a **reaction**.

For example: $Mg(OH)_2 + 2HCl \rightarrow MgCl_2 + 2H_2O$

Here, magnesium hydroxide (from the antacid) reacts with hydrochloric acid form a salt ($MgCl_2$) and water (H_2O). The water reduces the acidity in stomach, providing relief from symptoms like heartburn.

People take antacids as needed, but it's important to follow the correct to avoid side effects. Neutralisation reactions in antacids help to make stomach environment less acidic, bringing temporary relief from acidity.

b. Neutralisation in Wastewater Treatment

Wastewater treatment uses **neutralisation reactions** to adjust the **pH** of water a **neutral level** before it's released into the environment. **Industrial, municipal, and agricultural activities** can create wastewater that's too acidic or basic, which can harm the environment if left untreated.

In **wastewater treatment plants**, acids or bases are added to neutralise water:

i. For acidic wastewater (low pH), a base like calcium hydroxide ($Ca(OH)_2$) added.

ii. For **basic wastewater** (high pH), an **acid** like **sulphuric acid (H2SO4) hydrochloric acid (HCl)** is used.

For example: $2HCl + Ca(OH)_2 \rightarrow CaCl_2 + 2H_2O$

Here, **hydrochloric acid (HCl)** reacts with **calcium hydroxide** to form **calcium chloride (salt)** and **water**. This raises the pH of acidic water to make it neutral.

Neutralising wastewater protects **aquatic life** and keeps **water ecosystems healthy** by preventing harmful effects from untreated acidic or basic water.

c. Neutralisation in Agriculture

In **agriculture**, neutralisation helps improve **soil acidity** for better **plant growth**. **Acidic soil** can limit plant growth and nutrient availability, so farmers **agricultural lime** (calcium carbonate, $CaCO_3$) to **neutralise** the acid in the soil.

The reaction is:

$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + H_2O + CO_2$

Here, **calcium carbonate** reacts with **sulphuric acid** in the soil to form **calcium sulphate**, **water**, and **carbon dioxide**. This reaction **reduces soil acidity** and increases the **pH**, helping plants grow better and boosting **crop yield**.

d. Neutralisation in Cleaning

In **cleaning**, **neutralisation reactions** help remove **acidic or basic residues** from surfaces, making cleaning more effective.

i. When an acidic cleaner (like citric acid) meets alkaline residues (sodium hydroxide), they neutralise each other:

$$\underset{\textit{Citric acid}}{C_6H_8O_7(aq)} + \underset{\textit{sodium hydroxide}}{3NaOH(aq)} \longrightarrow \underset{\textit{water}}{3\,H_2O(l)} + \underset{\textit{Sodium citrate}}{Na_3C_6H_5O_7(aq)}$$

ii. When an **alkaline cleaner** (like **sodium hydroxide**) meets **acidic** (like **hydrochloric acid**), they also neutralise each other:

$$\underset{\textit{sodium hydroxide}}{NaOH(aq)} + \underset{\textit{hydrochloric acid}}{HCl(aq)} \longrightarrow \underset{\textit{water}}{H_2O(l)} + \longrightarrow \underset{\textit{sodium chloride}}{NaCl(aq)}$$

These reactions form **water** and **harmless salts** like **sodium citrate** or **sodium chloride**, which helps remove **unwanted residues** and ensures surfaces are clean.

e. Neutralisation in Food and Beverages

In the **food industry**, **neutralisation reactions** help control **pH** levels to **food safety and quality**, especially in **fermented foods** like yogurt.

During yogurt **fermentation**, bacteria turn **lactose** (milk sugar) into **lactic acid**, making the mixture more acidic. To **stabilise the pH** and prevent too acidity, a base like **calcium hydroxide** (lime water) is used to neutralise lactic acid.

The reaction is:

$$\underset{\textit{lactic acid}}{2C_3H_6O_3(aq)} + \underset{\textit{calcium hydroxide}}{Ca(OH)_2(aq)} \longrightarrow \underset{\textit{water}}{2H_2O(l)} + \underset{\textit{calcium lactate}}{Ca_3C_3H_6O_3(aq)}$$

Here, **lactic acid** reacts with **calcium hydroxide** to form **calcium lactate** (a neutral salt) and **water**. This helps control the acidity in yogurt, ensuring it has the right taste and quality.

Activity 4.6 Investigating physical properties of acids and bases

Conduct the following activity in small groups. Discuss your findings and create a poster or a presentation to share your findings with the entire class.

In your groups, investigate the physical properties of:

a. Sour taste of acids and slippery feel of bases

b. pH Testing

c. Feel Behaviour of Acids and Bases

Materials needed: vinegar (acid) and baking soda, computer, sodium chloride solution, lemon juice, test tube or beaker, test tube or beaker, virtual lab online simulation sites

Step for a:Sour taste of acids and bases

1. Use a virtual lab or online simulation to explore taste of acids and bases.

Steps for b: pH Testing

1. Take about 5 cm^3 the test solution into a test tube or beaker.

2. Using the litmus paper test each substance and record the colour change.

3. Identify which substances are acidic, basic, or neutral.

Steps for c: Exploring the Feel Behaviour of Acids and Base

1. Pour dilute solutions of vinegar (acid) and baking soda (base) into separate beakers.
2. Dip your fingers into each solution rub the fingers and feel the texture of the solution.
3. Based on the feel effect state if the solution is acid or base.

Activity 4.7 Exploring chemical reactions of Acids and Bases

1. Reaction of Acids

a. Reaction of Acids with Metals

Materials needed: Small pieces of **zinc (Zn)**, **iron (Fe)**, **magnesium (Mg)**, and **hydrochloric acid (HCl)** or **vinegar**.

Steps

i. Place each metal in a separate container

ii. Add a small amount of acid and observe.

iii. Observe and record what happens in the reaction.

iv. What is the colour of the gas evolved?

v. Test the gas evolved with blue and red litmus paper.

vi. Test the gas evolved with lighted splint.

b. Reaction of Acids with Bases (Neutralisation)

Materials needed: **Vinegar** (acid) and **baking soda** solution (base).

Steps

i. Measure equal volume of the acid and base into separate beakers

ii. Mix the two solutions and observe.

iii. Observe and record what happen in the reaction.

c. Reaction of Acids with Carbonates

Materials needed: Vinegar or lemon juice and baking soda (sodium bicarbonate) chalk (calcium carbonate).

Steps

i. Add vinegar (25 cm^3) to about 5.0 g baking soda or chalk.

ii. Observe and record what happen in the reaction.

iii. What is the colour of the gas evolved?

iv. Test the gas evolved with blue and red litmus paper.

v. Test the gas evolved with limewater.

d. Reaction of Acids with Basic Oxides

Materials needed: Vinegar and a magnesium oxide powder (or similar basic oxide).

Steps

i. Measure about 25 cm^3 of vinegar into a beaker.

ii. Add about 5.0 g of magnesium oxide to the vinegar solution.

iii. Observe and record what happen in the reaction.

2. Reaction of Bases

a. Reaction of Bases with Acidic Oxides

Materials needed: Limewater (aqueous calcium hydroxide) and carbon dioxide balloon or from exhaled breath.

Steps

i. Pour small quantity of limewater into a beaker.

ii. Fill a balloon with carbon dioxide.

iii. Release the carbon dioxide into the limewater.

iv. Observe and record what happen in the reaction.

b. Reaction of Bases with Ammonium Salts

Materials needed: Ammonium chloride and a strong base, such as sodium hydroxide.

Steps

i. Mix ammonium chloride with sodium hydroxide carefully, in a well-ventilated area.

ii. What is the colour of the gas evolved?

iii. Test the gas evolved with blue and red litmus paper.

Activity 4.8 Classifying Household Items as Acids or Bases

Materials needed: **Litmus paper** or **pH strips**, and a selection of household items (e.g., lemon juice, baking soda solution, vinegar, soap solution, milk, apple juice).

Steps

Perform the test on each solution with litmus paper or pH strips and record whether it turns the paper red (acidic) or blue (basic). Record your test and observation in the table below:

Type	Colour observed using		Classify the solution as acid or base
	Blue litmus	Red litmus	
Lemon juice,			
Vinegar			
apple juice			
Baking soda solution,			
Soap solution			

Activity 4.9 Application of neutralization in everyday life

Materials needed: Visual aids (diagrams, images) of an insect bite, an acidic lake, a garden, and a stomach

Steps

1. In small groups, assign each group one scenario:
    - Treating insect bites
    - Helping with stomach indigestion
    - Reducing soil acidity for plants
    - Balancing lake acidity to protect fish
2. What kind of substance (acid or base) is needed to fix this problem?
3. What everyday things can help us neutralize this issue (like baking soda for bites, antacids for indigestion)?
4. How do these substances make things safe or comfortable again?

5. Share your scenario and what you learned with the class.

STRENGTH OF ACIDS AND BASES

Differences between strong acids and weak acids

A strong acid can be distinguished from a weak acid by examining its pH value, electrical conductivity, and reactivity.

1. pH

   pH is the logarithmic value of the inverse of the hydrogen or hydronium ion concentration.

   Consider the reaction: $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$;

   a. Strong acids fully dissociate in water, meaning they release a lot of H^+ ions per mole into solution, results in a low pH (e.g., hydrochloric acid, HCl).

   Equation: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

   b. Strong bases fully dissociate in water, creating a high concentration of OH^- ions per mole in solution, which gives a high pH (e.g., sodium hydroxide, NaOH).

   Equation: $NaOH(s) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq)$

   c. Weak acids only partially dissociate, releasing fewer H^+ ions per mole into solution and resulting in a higher pH compared to strong acids (e.g., acetic acid, CH_3COOH).

   Equation: $CH_3COOH(aq) \rightleftharpoons H^+(aq) + CH_3COO^-(aq)$

   d. Weak bases partially dissociate, resulting in fewer OH^- ions per mole in solution and a lower pH than strong bases (e.g., ammonia, NH_3).

   Equation: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

2. Electrical Conductivity (per mole of acid/base)

   a. Strong acids have high conductivity because they produce more ions in solution.

   b. Strong bases have high conductivity due to the high concentration of OH^- ions.

   c. Weak acids have low conductivity due to fewer ions in solution.

   d. Weak bases have lower conductivity as they produce fewer OH^- ions.

3. Reactivity (per mole of acid/base)
   a. Strong acids are more reactive with metals, releasing more hydrogen gas quickly.
   b. Strong bases react more readily with acids to form water and salt, and they can very corrosive.
   c. Weak acids react more slowly, producing less hydrogen gas over time.
   d. Weak bases are less reactive than strong bases.
4. Enthalpy of Neutralization (ΔH)
   a. Strong acids release more energy per mole when neutralized by a strong base because they fully dissociate.

      Example Equation

      $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)\Delta H = -57.1\,kJmol - 1$
   b. Weak acids release less energy per mole, as some energy is needed to break bonds before neutralization can occur.

      Example Equation

      $CH_3COOH(aq) + OH^-(aq) \rightarrow H_2O(l) + CH_3COO^- \ \Delta H$

      $= -56.1\,kJmol - 1$

Activity 4.10 Introduce acids and bases using visual aids, interactive simulations

Materials needed: Computer, vinegar, vinegar, hydrochloric acid-based cleaning product or drain cleaner, baking soda, stopwatch

Step 1

1. Watch a video online showing basic home substances that are acidic and basic.
2. Use online simulations or videos showing how acids and bases behave in water.

Step 2: Strong vs. Weak Acids and Bases

1. Mix 10 cm^3 vinegar (a weak acid) with 10 g of baking soda and observe the reactions.
2. Mix 10 cm^3 drain cleaner (a strong acid) with baking soda and observe the reactions.

Compare the two reactions. In which reaction are more bubbles produced in the same time interval.

Step 3: **Exploration Stations** (explore the strength of acids and bases)

Organise yourselves into groups of no more than five. Conduct the following exploration stations in your respective groups.

Station 1: Conductivity Measurement

Materials: Provide a small circuit with an LED or bulb and probes to test solutions, beaker, measuring cylinder.

1. Measure 25 cm^3 drain cleaner into a beaker.
2. Insert two probes or electrodes connected with wire into the drain cleaner.
3. Connect bulb between the two probes or electrodes.
4. Observe the level of brightness of the bulb.
5. Measure 25 cm^3 vinegar into a beaker.
6. Repeat steps 2, 3, and 4.
7. Which solutions make the bulb light up more and why?

Station 2: pH Measurement

Materials: pH strips, drain cleaner, vinegar, baking soda, beaker, distilled water

1. Measure 25 cm^3 drain cleaner into a beaker.
2. Dip a pH strip into the drain cleaner for a few seconds.
3. Remove the pH strip and allow the colour to develop.
4. Compare the colour with the pH scale provided to determine the approximate pH the solution.
5. Repeat steps 1, 2, 3, and 4 with vinegar, baking soda, distilled water.

Organise the solutions from most acidic to most basic based on their pH values.

SALTS

A **salt** is a type of chemical compound formed when the hydrogen ions in an acid are replaced by metal or ammonium ions. Salts are important in everyday life and in many industries.

Basic Equation for Salt Formation: Acid + Base → Salt + Water

Types of Salts

There are many types of salts, each with unique properties and uses. Some common types of salts include:

1. **Normal Salts**

   Formed when an acid and base completely neutralize each other.

   Example: $NaOH + HCl \rightarrow NaCl + H_2O$

2. **Acidic Salts**

   Formed when only some hydrogen ions in an acid are replaced.

   Example: $NaOH + H_2SO_4 \rightarrow NaHSO_4 + H_2O$

   Acidic salts often come from reactions between a strong acid and a weak base, ammonium chloride (NH_4Cl).

3. **Basic Salts**

   Formed when some hydroxide ions of a base are replaced.

   Example: $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

4. **Double Salts**

   Made from two salts mixed in equal amounts. They have two types of cations.

   Example: Potash alum, used in dyeing and water purification.

5. **Complex Salts**

   Have a central metal atom bonded to other molecules or ions. They do not fully dissociate into ions in water.

   Example: $[Ni(CO)_4]$ or $[Cr(NH_3)_6]Cl_3$.

6. **Hydrated Salts**

   Crystalline salts that hold water molecules within them.

   Example: Magnesium tetraoxosulphate (VI) heptahydrate (Epsom salt) has trapped inside it and calcium tetraoxosulphate (VI) dihydrate

7. **Deliquescent Salts**

   Salts that absorb water from the air and can even turn into a liquid.

   Example: Calcium chloride and sodium hydroxide.

8. **Hygroscopic Salts**

Salts that absorb water from the air but don't dissolve completely.

Example: Anhydrous copper sulphate and anhydrous cobalt chloride.

9. **Efflorescent Salts**

   Salts that release water into the air, losing their water content.

   Example: Sodium carbonate decahydrate and sodium sulphate decahydrate.

Preparing soluble salts

Preparing Soluble Salts: Two Methods

Method 1: Reacting an Acid with a Solid Metal, Insoluble Base, or Carbonate

Materials needed

Dilute sulfuric acid (H_2SO_4), an insoluble base (like copper (II) oxide) or an insoluble carbonate, beaker, Bunsen burner (or a warm water bath), stirring rod, evaporating dish, glass rod, filter paper

Steps to Prepare the Salt

1. Pour some dilute sulfuric acid into a beaker.
2. Gently heat the beaker using a Bunsen burner or place it in a warm water bath (around 60°C). Be careful not to boil it!
3. Slowly add the copper (II) oxide (or other base) to the warm acid. Stir well until the base stops dissolving and a bit remains at the bottom (this means all the acid has reacted).
4. Pour the mixture through filter paper into an evaporating dish. This removes any leftover solid.
5. Heat the solution in the evaporating dish until some water evaporates and the solution becomes more concentrated. Be careful not to let it boil too much.
6. Dip a cold glass rod into the solution. If crystals form on the rod, the solution is ready.
7. Leave the solution in a warm place and let it sit undisturbed to form crystals.
8. Carefully pour off any leftover solution and let the crystals dry. You can pat them dry with filter paper if needed.

Preparing Soluble Salts: Reacting a Dilute Acid with an Alkali

Materials needed

Alkali (like sodium hydroxide solution), dilute acid (like hydrochloric acid), methyl orange indicator, conical flask, burette, evaporating dish

Steps to Prepare the Salt

1. Measure sodium hydroxide solution into a conical flask using a pipette.
2. Add a few drops of methyl orange indicator to the alkali in the flask. This will help us see when the reaction is complete by changing colour.
3. Pour some hydrochloric acid into the burette and note the starting level.
4. Carefully add the acid from the burette into the flask with the alkali, swirling as you go. Stop when the colour changes from yellow to orange/red. This means the acid has neutralized the alkali.
5. Note how much acid you added by checking the level on the burette. This is the amount you need for the reaction.
6. Now, mix the same amount of acid and alkali in another container without adding any indicator.
7. Heat this mixture in an evaporating dish until some water evaporates, making a concentrated solution. Be careful not to let it boil too much.
8. As the solution cools, crystals will start to form. Leave it in a warm place for more crystals to grow.
9. Pour off any leftover liquid and let the crystals dry completely.

Uses of Salts

1. Medicine and Health
    a. Saline Solution: A mixture of salt (sodium chloride) and water, used hospitals for:
        i. Rehydration: Helps people who are dehydrated get fluids quickly.
        ii. **Wound Cleaning:** Used to clean wounds and prevent infections.
2. Industry
    a. Desalination: Salts are removed from seawater to make fresh water processes like reverse osmosis.
    b. Chemical Production: Table salt (sodium chloride) is used to make chlorine, which is important for treating water and making plastics. Soda ash (carbonate) is used in making glass and detergents.

3. Agriculture
   a. Fertilizers: Some salts provide nutrients for plants. For example, nitrate provides potassium and nitrogen, and ammonium sulphate nitrogen and sulphur.
   b. Animal Nutrition: Livestock are given salt blocks to get important minerals sodium and iodine.
   c. Pest Control: Some salts help control pests and diseases in plants.
4. De-Icing:
   a. In winter, rock salt is spread on roads to melt ice and make surfaces safer walk or drive on. It prevents ice from forming by lowering the freezing point water.

Activity 4.11 Exploring Salt

Step

1. In small groups, brainstorm and think what comes to mind when you hear the word 'Salt'. Use the following questions to guide your discussion:
   a. What is salt made of?
   b. Where do we use salt in everyday life?
   c. How do salts form?
2. Present your findings to the entire class.
3. Using textbooks or the internet, explore the following types of salts, their characteristics and their uses. Present your findings to the class.
   a. normal,
   b. acidic,
   c. complex, or
   d. double salts

Activity 4.12 Preparing Salts and Exploring Their Uses

Part 1: Preparing Soluble Salts

1. Select an acid (like hydrochloric acid) and an insoluble reactant (such as a oxide or carbonate).
2. Pour the dilute acid into a beaker.

3. Gradually add the powdered reactant to the acid, stirring continuously. Watch bubbles (effervescence) as the reaction occurs.
4. Keep adding until no more dissolves and some unreacted powder remains at bottom (this means there's an excess).
5. Pour the mixture through filter paper to separate the solid from the liquid, leavin clear solution of the soluble salt.

Part 2: Preparing Insoluble Salts

1. Mix two solutions of soluble salts (like lead nitrate and potassium iodide) in a beaker.
2. A solid will form (the insoluble salt) and settle at the bottom.
3. Use filter paper to separate the solid precipitate from the liquid.
4. Gently rinse the solid with distilled water to remove any leftover solution.
5. The resulting solution should be handled with care as it contains lead. Ensure a proper disposal route is available.

Part 3: Discuss Everyday Uses of Salts

1. Discuss common salts you know and how they are used, like:
    a. Table Salt (for food)
    b. Baking Soda (in baking and cleaning)
    c. Epsom Salt (for baths and gardening)
2. Share your thoughts on how salts are used at home and in the community.

Part 4: Field Trip or Excursion to a Salt Mine

Objective: To provide a hands-on experience and real-world understanding of salt mining processes.

1. Your teacher will arrange trip to a salt mine or mining facility. Please take safety precautions and wear appropriate attire. During your visit, make sure to observe the following and make notes:
    a. Salt extraction processes
    b. Mining methods (e.g., room and pillar, solution mining)
    c. Safety procedures
    d. Environmental considerations
2. Jot down your observations and share your findings with the class. Engage in a whole class discussion on:
    a. Salt mining processes

  - **b.** Real-world applications of salt
  - **c.** Environmental and economic impacts

3. Connect field trip to classroom learning:
   - **a.** Relate salt mining to geological processes
   - **b.** Discuss chemical properties of salt
   - **c.** Explore industrial uses of salt

ACID - BASE TITRATION

Acid-Base Titration

Definition

Acid-base titration is a method used to find out the concentration of an unknown acid or base by reacting it with a solution of known concentration. Anindicator helps show when the reaction is complete by changing colour.

Apparatus Used in Titration

1. **Conical flask**: Holds the solution with the unknown concentration (called the analyte) and the indicator. It has a narrow neck for easy swirling.
2. **Indicator**: A dye that changes colour to show when the acid has neutralized the (or vice versa). Common indicators are phenolphthalein (turns pink in basic solutions) and methyl orange (turns red in acidic solutions).
3. **White tile**: Placed under the flask to help see the colour change clearly.
4. **Beaker**: It is usually used to hold excess solution or waste during the titration process.
5. **Analyte**: The analyte is the solution we want to learn more about in a experiment. It is usually placed in a conical flask, and an acid-base indicator is to it, which changes colour during the experiment.
6. **Titrant**: It is the solution whose concentration is accurately known and usually in the burette.
7. **Burette**: A long tube that accurately measures and dispenses the solution wit known concentration (called the titrant). It has a tap to control the flow of the titrant.
8. **Pipette**: Measures a specific volume of the analyte and transfers it to the conical flask. See Figure 4.1 for a diagram of pipette filler.

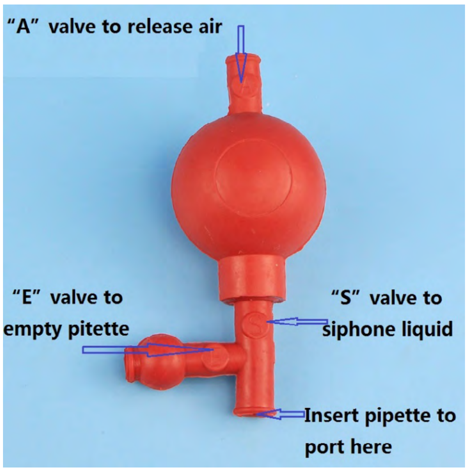


Figure 4.1: Pipette filler

Steps for Using the Pipette Bulb

a. Insert the top (mouth) of the pipette into the bottom of the pipette bulb.

b. Squeeze the button labelled "A" (Air valve) while pressing the bulb to release any air.

c. Place the tip of the pipette into the solution.

d. Squeeze the button labelled "S" (Syphon valve) to draw liquid up into the until it reaches the desired level.

e. To release the liquid, press the button labelled "E" (Empty valve).

9. **Retort stand:** A retort stand is a lab tool with three parts: a base, a rod, and a clamp. It securely holds the burette in place, allowing it to stay directly above the conical for accurate pouring of liquids during experiments. See Figure 4.2 for titration setup.

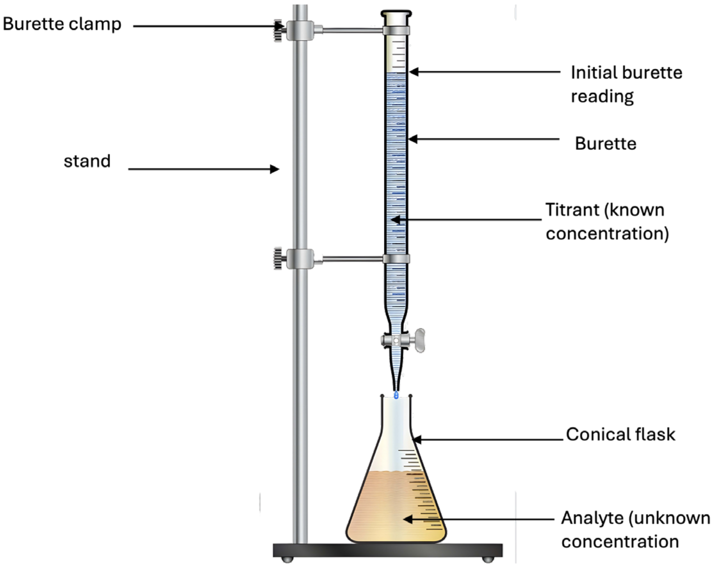


Figure 4.2: Titration set up.

10. Acid-Base Indicator

An **acid-base indicator** is a substance that changes colour to show when a reaction is complete (the endpoint) in a titration. The indicator is added to the solution in the conical flask, and it stays the same colour at first. When the acid and base in the solution balance out, the indicator changes colour to signal that the reaction is done.

a. Phenolphthalein: Colourless in acids, turns pink in a basic solution.

b. Methyl Orange: Red in acidic solutions and yellow in basic solutions.

These indicators help scientists know when an acid and a base have completely reacted with each other.

Most indicators do not change colour at a particular pH. They do so over a range of pH from two to three units. This is called the pH range which is different for various indicators.

Therefore, a pH range is the working range within which an acid-base indicator changes from one colour to another.

The table below gives the pH ranges of some acid-base indicators

Indicator	Colour change (acid to base)	pH range
Methyl orange	Red to orange	3.1 – 4.4
Methyl red	Red to yellow	4.4 – 6.0
Litmus	Red to blue	5.0 – 8.0
Bromothymol blue	Yellow to blue	6.0 – 7.6
Phenolphthalein	Colourless to pink	8.3 – 10.0

Principles of Acid-Base Indicators

1. Phenolphthalein as an Acid-Base Indicator

Phenolphthalein is a weak acid that can change colour depending on the solution's pH.

It exists in two forms in solution: **HIn** (the acidic form with one colour) and **In⁻** (basic form with a different colour).

In an **acidic solution**, there are more H^+ ions, which keeps phenolphthalein in its HIn form, showing the **acid colour**.

In a **basic solution**, there are more OH^- ions, which cause phenolphthalein to to its In^- form, showing the **base colour** (pink in phenolphthalein's case).

2. Other Indicators (Weak Bases)

   Some indicators work as weak bases and also change colour based on pH.

   In an **acidic solution**, H^+ ions from the acid react with OH^- ions from the indicator, causing the indicator to show its **acid colour**.

   In a **basic solution**, there are more OH^- ions, so the indicator shows its **base colour**.

   In simple terms, acid-base indicators change colour because they have one colour acidic conditions and a different colour in basic conditions, helping us tell whethe solution is acidic or basic.

Choosing the Right Indicator for Titration

The choice of indicator in a titration depends on the type of acid and base used. Here is a simple guide:

a. **Strong Acid and Strong Base**

   When both the acid and base are strong, the pH changes a lot near the endpoint, so many indicators will work well.

   Suitable indicators: Litmus, phenolphthalein, methyl orange, or bromothymol blue.

b. **Weak Acid and Strong Base (or Strong Acid and Weak Base)**

   If one is weak and the other is strong, the pH changes less near the endpoint, so you need an indicator that matches the pH at the endpoint.

   Examples: Use methyl orange with strong acids and phenolphthalein with strong bases.

c. **Weak Acid and Weak Base**

   When both are weak, the pH change is very small, so **bromothymol blue** is a good choice.

Key Terms in Titration

1. **Titrant**: The titrant is a solution with a known concentration that is added to the analyte to out its properties. It is placed in the burette during titration.

   Characteristics of a Titrant

   a. Its concentration is known and accurate.

   b. It is as pure as possible.

   c. It reacts completely with the analyte.

   d. The reaction has a clear endpoint, shown by a colour change.

2. **Analyte**: The analyte is the solution with an unknown concentration. We want to learn about by reacting it with the titrant. The analyte is placed in the conical flask.
3. **Equivalence Point**: This is the point where the exact amount of titrant has been added to react with the analyte. It is a theoretical point in the reaction.
4. **Endpoint**: The endpoint is when the colour change happens, showing that the reaction complete. This is observed with the help of an indicator.
5. **Titre Value**: The titre value is the volume of the titrant used to reach the endpoint. It's calculated subtracting the starting reading from the final reading on the burette.
6. **Consistent Titre Values**: These are titre values that are very close to each other (within ± 0.20 cm^3), accuracy in measurements.
7. **Average Titre**: This is the average of the consistent titre values, providing the most measurement of titrant used.
8. **Indicator**: Indicators are substances that change colour based on the pH level, helping us when the titration reaction is complete.

Types of Acid-Base Titration

1. Simple (Direct) Acid-Base Titrations

In a **simple titration**, we directly add a known solution (titrant) to an unknown (analyte) to find its concentration.

This type of titration uses only one acid-base indicator to show the endpoint (change) when the reaction is complete.

Simple titration helps us figure out the exact amount of acid or base in a solution.

Steps for Simple Acid-Base Titration

1. Clamp the burette to a stand, rinse it with distilled water and then with the titrant. Fill with the titrant and record the starting volume.
2. Use a pipette to measure a specific amount of the analyte (usually 20 or 25 cm^3) place it in a conical flask.
3. Add 2-3 drops of an indicator to the analyte in the flask and swirl to mix it.
4. Start Titration: Slowly add the titrant from the burette to the analyte in the flask, swirling until a permanent colour change (endpoint) appears.

5. Record the titre values: Note the initial and final burette readings and calculate volume of titrant used.
6. Repeat the process to get consistent readings and calculate the average titre.
7. Use the balanced chemical equation to find the ratio between the titrant and analyte, and then calculate the analyte's concentration.

Example 1

A solution of 0.2 moldm^{-3} of HCl was titrated against 25 cm^3 of KOH using methyl orange as indicator. The results obtained in the titration experiment are as follows

Titrant: 0.2 moldm^{-3} HCl

Analyte: KOH of unknown concentration

Capacity of Burette: 50 cm^3

Capacity of pipette: 25 cm^3

Indicator used: 2 drops of methyl orange

Colour change: Yellow to red

Titration Table

Burette Reading (cm³)	**1st**	**2nd**
Final	24.60	24.70
Initial	0.00	0.00
Volume of HCl Used	24.60	24.70

Calculate,

i The average titre value

ii The molar concentration of NaOH

iii The mass concentration of NaOH

Solution

i. Average titre $= \frac{24.70 + 24.60}{2} = 24.65$ cm^3

ii. From the titration reaction

$HCl + KOH \rightarrow NaCl + H_2O$

The mole ratio is given by

$$\frac{n(KOH)}{n(HCl)} = \frac{1}{1}$$

$$\frac{C_{KOH}V_{KOH}}{C_{HCl}V_{HCl}} = \frac{1}{1}$$

$$C_{KOH} = \frac{C_{HCl}V_{HCl}}{V_{KOH}} = \frac{0.2\ \text{mol dm}^{-3} \times 24.65\ \text{cm}^3}{25\ \text{cm}^3} = 0.197\ \text{mol dm}^{-3}$$

iii. $M(KOH) = 39 + 16 + 1 = 56\ \text{gmol}^{-1}$

Mass concentration, $\rho = C \times M$

$$= 0.197\ \text{mol dm}^{-3} \times 56\ \text{gmol}^{-1}$$

$$= 11.0\ \text{g dm}^{-3}$$

In addition to solution standardization, direct acid-base titrations have various applications. They can be used to determine:

i. The percentage purity of a substance
ii. The number of moles or ions present in a solution
iii. The molar mass of a substance
iv. The solubility of a substance

2. Back (Indirect) Titration

Back titration is a method used when we cannot directly measure an substance (let's call it "A") by titration. Here's how it works:

a. React Substance A with Excess B

We add more than enough of a known solution (B) to react with all of A, making A is fully neutralized.

b. Measure the Leftover B

After A has reacted with some of B, we titrate the leftover B with another solution find out how much of it is left.

c. Calculate Amount of A

By subtracting the leftover amount of B from the original amount, we can how much B was used to react with A. This tells us how much A was present.

Back titration is useful for many things, including finding out:

a. How much calcium carbonate ($CaCO_3$) is in limestone or eggshells
b. The amount of ammonia in an ammonium salt
c. The amount of calcium hydroxide ($Ca(OH)_2$) in a sample of soil

Example 2

To check if a new brand of aspirin (Ecotrin) contains lower doses of acetylsalicylic acid, a chemist conducted an experiment. They crushed two aspirin tablets and mixed them with 25 cm³ of 1.00 moldm^{-3} NaOH, warming the mixture before transferring it to a 250 cm³ volumetric flask.

0.050 moldm^{-3} H_2SO_4 was titrated against 25 cm³ portions of the sodium hydroxide solution, using phenolphthalein as an indicator.

The following results were recorded in the experiment:

Burette reading/*cm*^{3}	**1**	**2**	**3**
Final reading	10.70	20.40	40.75
Initial reading	0.10	10.30	30.45
Volume of acid used	10.60	10.10	10.30

Calculate

- **a.** The average titre value
- **b.** Moles of H_2SO_4 that reacted
- **c.** Moles of excess NaOH present in the solution.
- **d.** Moles of NaOH that reacted with acetylsalicyclic acid
- **e.** Mass of acetylsalicyclic acid in one tablet.
- **f.** If the mass of each tablet is 2 g, determine the percentage mass of pure acetylsalicyclic acid in each tablet.

  The equations of the reaction involved are:

  $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

  $C_9H_8O_4 + NaOH \rightarrow C_9H_7O_4Na + H_2O$

  $[C = 12, H = 1, O = 16, Na = 23, S = 32]$

Solution

a. Average titre $= \dfrac{10.10 + 10.30}{2} = 10.20\ cm^3$

b. $C_{H_2SO_4} = 0.050\ mol\ dm^{-3};\ V_{H_2SO_4} = 10.20\ cm^3 = 0.0102\ dm^3$

$$n_{H_2SO_4} = C_{H_2SO_4} V_{H_2SO_4}$$
$$= 0.050\ mol\ dm^{-3} \times 0.0102\ dm^3$$
$$= 0.00051\ mol$$

c. From the balanced equation

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

The mole ratio for the titration is

$$\frac{n_{NaOH}}{n_{H_2SO_4}} = \frac{2}{1}$$

$$n_{NaOH} = 2 \times n_{H_2SO_4}$$

$$= 2 \times 0.00051\ mol$$

$$= 0.00102\ \text{mol}$$

Therefore, moles of NaOH that was present in the 25 cm^3 and reacted with the acid is 0.00102 mol.

If 25 cm^3 of NaOH contains 0.00102 mol

Then 250 cm^3 of NaOH will contain $= \frac{250\ cm^3 \times 0.00102\ mol}{25\ cm^3} = 0.0102$ mol

d. From the relation n = CV

Initial concentration of NaOH before it was mixed with the crushed tablets $= 1.0\ mol\ dm^{-3}$.

Volume of the solution that was mixed with the crushed tablets = 25 cm^3 = 0.025 dm^3

$$n_{NaOH} = 1.0\ \text{mol dm}^{-3} \times 0.025$$

$$= 0.025\ \text{mol}$$

If 0.0025 mol of NaOH was mixed with the crushed tablets and 0.0102 mol remained, then,

Mole of NaOH that reacted with the acetylsalicyclic acid = 0.025 – 0.0102

= 0.0148 mol

e. Acetylsalicyclic acid reacts with sodium hydroxide according to the equation

$C_9H_8O_4 + NaOH \rightarrow C_9H_7O_4Na + H_2O$

The mole ratio is given by

$$\frac{n_{aspirin}}{n_{NaOH}} = \frac{1}{1}$$

$$n_{aspirin} = n_{NaOH}$$

$$= 0.0148\ \text{mol}$$

but m = n × M

$$M(C_9H_8O_4) = (12 \times 9) + (1 \times 8) + (16 \times 4)$$

$$= 180\ \text{g dm}^{-3}$$

$$m(C_9H_8O_4) = 0.0148 \times 180$$

$= 2.66\ \text{g}$

2 tablets contain 2.66 g

$\therefore$ 1 tablet contains $\frac{2.66}{2} = 1.33\ \text{g}$

f. If the mass of each tablet is 2 g

$$\%\ (\text{mass}) = \frac{\text{mass of pure acetylsalicyclic acid}}{\text{mass of tablet}} \times 100$$

$$= \frac{1.33}{2} \times 100$$

$$= 66.5\ \%$$

3. Double-Indicator Titration

Double-indicator titration is a method used to find the concentration of substances in a mixture, using two indicators to show separate stages of the reaction.

a. In this method, each part of the mixture reacts differently with the titrant. One forms an acidic product, while the other forms a basic (alkaline) product.

b. Indicators like **methyl orange** and **phenolphthalein** are chosen to match the pH of each stage, changing colour to show when each part of the reaction complete.

Two Methods of Double-Indicator Titration

1. **Continuous Method**: Both stages of the reaction happen one after the other.
2. **Discontinuous Method**: Each stage is carried out separately.

Continuous Method for Double-Indicator Titration

1. Place the acid (titrant) in the burette.
2. Measure 20 or 25 cm³ of the basic mixture and pour it into a conical flask. Add drops of phenolphthalein and swirl to mix.
3. **First Titration (Phenolphthalein Endpoint)**

   Slowly add the acid from the burette until the colour changes. Record the volume acid used as **x cm³**. This step reacts with sodium carbonate (Na_2CO_3) to form bicarbonate ($NaHCO_3$).
4. Add two drops of methyl orange to the solution after the first titration.
5. Continue adding the acid until a second colour change occurs. Record this volume as **y cm³**. Here, the remaining sodium bicarbonate ($NaHCO_3$) reacts with acid to form sodium chloride, water, and carbon dioxide.

6. To find the volume of acid that reacted with the original sodium bicarbonate, the first reading x from the second reading y (**y - x**). These values can then be used calculate the concentration of each component in the mixture.

Example 3

A mixture of Na_2CO_3 *and* $NaHCO_3$ reacted with 0.1 moldm^{-3} HCl.
The results obtained when phenolphthalein indicator was used is

Burette Reading/cm^3	1	2	3
Final	13.00	47.10	13.00
Initial	0.00	34.00	0.00
Volume of acid used	13.00	13.10	13.00

The titration continued with methyl orange being added to the resulting solution. The results obtained at the methyl orange endpoint is

Burette Reading/cm^3	1	2	3
Final	34.00	21.00	34.00
Initial	13.00	0.00	13.00
Volume of acid used	21.00	21.00	21.00

a. Write a balanced chemical equation to represent the reaction that took place at the;
   i. Phenolphthalein endpoint
   ii. Methyl orange endpoint

b. Calculate the concentration in moldm^{-3} of,
   i. Na_2CO_3
   ii. $NaHCO_3$

c. $NaHCO_3$

   In the mixture that was used for the titration

   Determine the mass concentration of;
   i. Na_2CO_3
   ii. $NaHCO_3$

   Determine the percentage composition of Na_2CO_3 in the mixture.

Solution

a. The balanced chemical equation for

i. Phenolphthalein endpoint is represented by:

$$Na_2CO_3(aq) + HCl(aq) \rightarrow NaHCO_3(aq) + NaCl(aq)$$

ii. Methyl orange end point:

$$NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

b. From the titration reaction

i. The average titre value at the phenolphthalein endpoint is given by

$$x = \frac{13.00 + 13.00}{2} = 13.00\ cm^3$$

$V_1 = 13.00\ cm^3$ = volume of HCl required to neutralise half the Na_2CO_3 in the mixture at the end-point.

$$V_1 = \frac{1}{2} = Na_2CO_3$$

$$V_1 = 13.00$$

$$2V_1 = 2 \times 13.00 = 26.00\ cm^3$$

The volume of HCl required to neutralise completely Na_2CO_3 is $2V_1$

The complete reaction of HCl with Na_2CO_3 that requires 26.00 cm^3 of the titrant (HCl) is represented by the chemical equation:

$$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

Mole ratio for the complete reaction is

$$\frac{C_{Na_2CO_3} V_{Na_2CO_3}}{C_{HCl} V_{HCl}} = \frac{1}{2}$$

$$C_{Na_2CO_3} = \frac{C_{HCl} \times V_{HCl}}{2V_{Na_2CO_3}}$$

C_{HCl} = concentration of HCl = 0.100 mol dm^{-3},

V_{HCl} = volume of HCl for complete reaction = 26.00 cm^3

$V_{Na_2CO_3}$ = volume of mixture pipetted for the reaction = 25.00 cm^3

$C_{Na_2CO_3}$ = concentration of Na_2CO_3 =

$$C_{Na_2CO_3} = \frac{C_{HCl} \times V_{HCl}}{2V_{Na_2CO_3}} = \frac{0.100\ mol\ dm^{-3} \times 26\ cm^3}{2 \times 25\ cm^3} = 0.0520\ mol\ dm^{-3}$$

ii. The average titre value at the methyl orange endpoint is given by

$$y = \frac{21.00 + 21.00}{2} = 21.00\ cm^3$$

21.00 cm^3 of HCl reacted with the $NaHCO_3$ that was formed from the Na_2CO_3 and the original $NaHCO_3$ that was present in the mixture.

If half of Na_2CO_3 was neutralised by 13.00 cm^3, then the remaining half that was present in solution will also react with the same volume of HCl.

V_2 = volume of HCl required to neutralise the remaining half of Na_2CO_3 and $NaHCO_3$ in the mixture at the methyl orange end point.

$$V_2 = \frac{1}{2}Na_2CO_3 + NaHCO_3$$

$$V_2 = V_1 + NaHCO_3$$

$$V_2 - V_1 = NaHCO_3$$

$$V_{HCl} = V_2 - V_1 = 21.00 - 13.00 = 8.00\ cm^3$$

$$V_2 - V_1 = \text{Volume acid required to neutralise } NaHCO_3$$

The reaction equation is given by:

$$NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

$$\frac{C_{NaHCO_3}V_{NaHCO_3}}{C_{HCl}V_{HCl}} = \frac{1}{1}$$

$$C_{NaHCO_3} = \frac{C_{HCl}V_{HCl}}{V_{NaHCO_3}}$$

$$C_{NaHCO_3} = \frac{C_{HCl}(V_2 - V_1)}{V_{NaHCO_3}}$$

$$= \frac{0.100\ mol\ dm^{-3} \times 8.00\ cm^3}{25\ cm^3} = 0.032\ mol\ dm^{-3}$$

c. i. $\rho(Na_2CO_3) = c(Na_2CO_3) \times M(Na_2CO_3)$

$$M(Na_2CO_3) = (2 \times 23) + 12 + (3 \times 16)$$

$$= 106\ g\ mol^{-1}$$

$$\rho(Na_2CO_3) = c(Na_2CO_3) \times M(Na_2CO_3)$$

$$= 0.0520\ mol\ dm^{-3} \times 106\ g\ mol^{-1}$$

$$= 5.51$$

$$= 5.51\ g\ dm^{-3}$$

ii. $M(NaHCO_3) = 23 + 1 + 12 + (3x16)$

$$= 84 \text{ g mol}^{-1}$$

$$\rho(NaHCO_3) = c(NaHCO_3) \times M(NaHCO_3)$$

$$= 0.032 \text{ mol dm}^{-3} \times 84 \text{ g mol}^{-1}$$

$$= 2.688$$

$$= 2.69 \text{ g dm}^{-3}$$

d. $\%(Na_2CO_3) = \frac{m(Na_2CO_3)}{\text{Total mass of mixture}} \times 100$

Total mass of mixture $= m(Na_2CO_3) + m(NaHCO_3)$

$$= 5.51 + 2.69$$

$$= 8.2\ g$$

$$\%(Na_2CO_3) = \frac{m(Na_2CO_3)}{\text{Total mass of mixture}} \times 100$$

$$= \frac{5.51}{8.2} \times 100 = 67.2\%$$

Discontinuous Method for Double-Indicator Titration

a. Place the acid (HCl) in the burette.

b. Measure 20 or 25 cm³ of the basic mixture into a conical flask and add 2-3 drops of phenolphthalein.

c. Slowly add HCl from the burette until the colour changes. Record this volume as $\mathbf{V_1}$ **cm³**.

d. This volume of HCl (V_1) converts all sodium carbonate (Na_2CO_3) in the mixture into sodium bicarbonate ($NaHCO_3$).

e. Discard the solution from the first titration.

f. Take another 20 or 25 cm³ of the mixture in a new conical flask, add 2 drops of methyl orange, and swirl to mix.

g. Slowly add HCl from the burette until the colour changes. Record this new volume as V_2 cm³.

h. This time, all sodium carbonate and sodium bicarbonate in the mixture react with HCl, creating sodium chloride, water, and carbon dioxide.

i. To find the amount of original sodium bicarbonate ($NaHCO_3$), use the formula $\mathbf{V3 = V2 - 2V_1}$.

j. This final volume, V_3, helps calculate the amount of sodium bicarbonate initially present in the mixture.

Example 4

A sample of Na_2CO_3 and $NaHCO_3$ were mixed together to form a solution. 25 cm^3 of this solution was titrated against a 0.1 moldm^{-3} HCl solution using phenolphthalein as an indicator.

The results obtained when phenolphthalein indicator was used is

Burette Reading/cm^3	1	2	3
Final	10.60	21.30	32.00
Initial	0.00	10.60	21.30
Volume of acid used	10.60	10.70	10.70

Another 25 cm^3 portions of the same solution was also titrated against 0.1 $moldm^{-3}$ HCl solution using methyl orange as an indicator.

Burette Reading/cm^3	1	2	3
Final	23.70	47.30	27.60
Initial	0.00	23.70	0.00
Volume of acid used	23.70	23.60	27.60

a. Determine the average titre value at the;
  - **i.** Phenolphthalein endpoint
  - **ii.** Methyl orange endpoint

b. Determine the volume of HCl that reacted with;
  - **i.** All of the Na_2CO_3 present
  - **ii.** The original quantity of $NaHCO_3$ present

c. Calculate the concentration in $moldm^{-3}$ of;
  - **i.** Na_2CO_3
  - **ii.** $NaHCO_3$

d. Determine the mass concentration of;
  - **i.** Na_2CO_3
  - **ii.** $NaHCO_3$

e. Determine the percentage composition of $NaHCO_3$ in the mixture

Solution

a. Calculation of the average titre values at the various stages of the titration

i. At the phenolphthalein endpoint, the average titre value is given by;

$$= \frac{10.70 + 10.70}{2} = 10.70 \text{ cm}^3$$

ii. At the methyl orange endpoint, the average titre value is given by;

$$= \frac{23.60 + 23.60}{2} = 23.60 \text{ cm}^3$$

b. To determine the volume of HCl that reacted with the various components of mixture;

i. The equation for the complete reaction of Na_2CO_3 with HCl goes by the equation:

$$Na_2CO_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$$

V_1 = volume of HCl required to neutralise half the Na_2CO_3 in the mixture at phenolphthalein end-point.

$$V_1 = \frac{1}{2} Na_2CO_3$$

$$V_1 = 10.70$$

$$2V_1 = 2 \times 10.70 = 21.40 \text{ cm}^3$$

The volume of HCl required to neutralise completely Na_2CO_3 is $2V_1 = 21.40 \text{ cm}^3$

ii. At the methyl orange endpoint, two sets of reaction occurred where all components of the mixture were neutralised by the HCl. These reactions can represented by the equations;

V_2 = Volume of HCl required to neutralise Na_2CO_3 and $NaHCO_3$ in the mixture at methyl orange end point.

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2 \; (1)$$

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2 \; (2)$$

$$V_2 = Na_2CO_3 + NaHCO_3$$

$$V_2 = 2V_1 + NaHCO_3$$

$$V_2 - 2V_1 = NaHCO_3$$

$V_2 - 2V_1 = 23.60 - 21.40 = 2.20$ = Volume acid required to neutralize $NaHCO_3$

c. **i.** The equation for the reaction of Na_2CO_3 at the phenolphthalein endpoint is given by:

$$Na_2CO_3 + HCl \rightarrow NaHCO_3 + NaCl$$

$$\frac{C_{Na_2CO_3}V_{Na_2CO_3}}{C_{HCl}V_{HCl}} = \frac{1}{1}$$

$$C_{Na_2CO_3} = \frac{C_{HCl}V_{HCl}}{V_{Na_2CO_3}}$$

$$C_{Na_2CO_3} = \frac{0.100 \text{ mol dm}^{-3} \times 10.70 \text{ cm}^3}{25 \text{ cm}^3} = 0.0428 \text{ mol dm}^{-3}$$

ii. The reaction that took place between the original $NaHCO_3$ in the mixture and HCl is represented by:

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$$

This required 2.20 cm^3 of the HCl.

The mole ratio of this reaction is

$$\frac{C_{NaHCO_3}V_{NaHCO_3}}{C_{HCl}V_{HCl}} = \frac{1}{1}$$

$$C_{NaHCO_3} = \frac{C_{HCl}V_{HCl}}{V_{NaHCO_3}}$$

$$= \frac{0.100 \text{ mol dm}^{-3} \times 2.20 \text{ cm}^3}{25 \text{ cm}^3}$$

$$= 0.0088 \text{ mol dm}^{-3}$$

d. **i.** $\rho(Na_2CO_3) = c(Na_2CO_3) \times M(Na_2CO_3)$

$$M(Na_2CO_3) = (2 \times 23) + 12 + (3 \times 16)$$

$$= 106 \text{ g mol}^{-1}$$

$$\rho(Na_2CO_3) = c(Na_2CO_3) \times M(Na_2CO_3)$$

$$= 0.0428 \text{ mol dm}^{-3} \times 106 \text{ g mol}^{-1}$$

$$= 4.54 \text{ g dm}^{-3}$$

ii. $M(NaHCO_3) = 23 + 1 + 12 + (3 \times 16)$

$$= 84 \text{ g mol}^{-1}$$

$$\rho(NaHCO_3) = c(NaHCO_3) \times M(NaHCO_3)$$

$$= 0.0088 \text{ mol dm}^{-3} \times 84 \text{ g mol}^{-1}$$

$$= 0.739 \text{ g dm}^{-3}$$

e. Total mass of mixture $= m(Na_2CO_3) + m(NaHCO_3)$

$$= 4.54 + 0.739$$

$$= 5.28\ g$$

$$\%NaHCO_3 = \frac{m(NaHCO_3)}{\text{Total mass of mixture}} \times 100$$

$$= \frac{0.739}{5.28} \times 100$$

$$= 14.0\%$$

Activity 4.13 Using titration to determine the concentration of an analyte i solution.

Steps

a. Simple Acid-Base Titration

1. Set up the titration apparatus with the acid in the burette and the base in the flask.
2. Add an indicator, like phenolphthalein, to the conical flask.
3. Slowly add the acid from the burette to the base, swirling until the colour shows the endpoint.
4. Record your data and use it to calculate the concentration of the base.
5. Applying the results to determine:
    - **i.** Percentage purity
    - **ii.** Percentage water of crystallization
    - **iii.** Relative atomic mass

A. Determining Percentage Purity

A sample of impure Na_2CO_3 weighing 2.50 g is dissolved in water and made up to 250 cm³. A 25 cm³ portion of this solution is titrated with 0.100 moldm^{-3} HCl. The average volume of HCl required for neutralization is 23.50 cm³. Calculate the percentage purity of the Na_2CO_3 in the sample.

Data given:

Mass of impure Na_2CO_3 = 2.50 g

Total volume of solution = 250 cm³

Volume of solution used in titration = 25 cm³

Concentration of HCl = 0.100 moldm^{-3}

Average volume of HCl used = 23.50 cm³

Steps

1. *Write the balanced equation*

   The reaction between sodium carbonate and hydrochloric acid is:

   $Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$

   From the equation, 1 mole of Na_2CO_3 reacts with 2 moles of HCl.

2. *Calculate the Moles of HCl Used*

   Given:

   Concentration of HCl= $0.100\ \text{moldm}^{-3}$

   Volume of HCl used = $23.50\ \text{cm}^3 = 0.0235\ \text{dm}^3$

   Using the formula:

   $\text{n}(\text{HCl}) = \text{C}_{\text{HCl}} \times \text{V}_{\text{HCl}}$

   $= 0.10 \times 0.0235 = 0.00235\,\text{mol}$

3. *Calculate the Moles of Na_2CO_3 in 25 cm³ of Solution*

   $\frac{\text{n}_{\text{Na}_2\text{CO}_3}}{\text{n}_{\text{HCl}}} = \frac{1}{2}$

   $\text{n}_{\text{Na}_2\text{CO}_3} = \frac{\text{n}_{\text{HCl}}}{2} = \frac{0.00235\,\text{mol}}{2} = 0.00118\,\text{mol}$

4. *Calculate the Moles of Na_2CO_3 in the 250 cm³ Solution*

   Since the titration used only a 25 cm³ portion of the 250 cm³ solution, we scale up:

   $25\ \text{cm}^3$ of Na_2CO_3 contain $0.00118\,\text{mol}$

   $\therefore 250\ \text{cm}^3$ of $Na_2CO_3 = \frac{250\ \text{cm}^3}{25\ \text{cm}^3} \times 0.00118\,\text{mol} = 0.0118\ \text{mol}$

5. *Calculate the Mass of Pure Na_2CO_3*

   Molar mass of Na_2CO_3 = 106 g/mol

   $m(\text{Na}_2\text{CO}_3) = n(\text{Na}_2\text{CO}_3) \times M(\text{Na}_2\text{CO}_3)$

   $= 0.0118 \times 106 = 1.25\ g$

6. *Calculate the percentage purity*

   The sample weighed 2.5 g, so:

   $$\text{percentage purity} = \frac{\text{Mass of pure Na}_2\text{CO}_3}{\text{mass of sample}} \times 100$$

   $$= \frac{1.25}{2.5} \times 100 = 50.0\%$$

B. Determining Percentage of Water of Crystallization

A hydrated sample of magnesium sulphate ($MgSO_4{\cdot}xH_2O$) weighing 4.00 g is dissolved in water, and the solution is titrated with 0.100 moldm^{-33} EDTA solution (a reagent that binds with Mg^{2+}). It requires 35.0 cm³ of EDTA to react with all the Mg^{2+} ions in the sample. Calculate the percentage of water of crystallisation in the hydrated salt.

Note

EDTA binds with Mg^{2+} ions in a 1:1 ratio.

Data

Mass of hydrated $MgSO_4{\cdot}xH_2O$ = 4.00 g

Concentration of EDTA = 0.100 moldm^{-3}

Volume of EDTA used = 35.0 cm³

Molar mass of $MgSO_4$ = 120.4 gmol^{-1}

Molar mass of H_2O = 18 gmol^{-1}

Steps

1. *Calculate the Moles of Mg^{2+} Ions*

   Given:

   Concentration of EDTA (C_{EDTA}) = 0.100 moldm^{-3}

   Volume of EDTA used = 35.0 cm³ = 0.035 dm³

   Since EDTA binds with Mg^{2+} ions in a 1:1 ratio

   $$n(Mg^{2+}) = C_{EDTA} \times V_{EDTA}$$

   $$= 0.100 \times 0.035 = 0.0035 \text{ mol}$$

2. *Calculate the Mass of anhydrous $MgSO_4$*

   The molar mass of anhydrous $MgSO_4$ is 120.4 gmol^{-1}

   $$m(MgSO_4) = n(MgSO_4) \times M(MgSO_4)$$

   $$= 0.0035 \times 120.4 = 0.421\ g$$

3. *Calculate the Mass of Water of Crystallisation*

   The total mass of the hydrated sample is 4.0 g, so:

   Mass of water of crystallisation = Total mass of sample – mass of $MgSO_4$

   $$= 4.0 - 0.421$$

$$= 3.58\,\text{g}$$

4. Calculate the percentage of water of crystallisation

$$\text{Percentage of water of crystallisation} = \frac{\text{Mass of water of crystallisation}}{\text{Total mass of sample}} \times 100$$

$$= \frac{3.58}{4.0} \times 100$$

$$= 89.5\%$$

C. Determining Relative Atomic Mass

A 0.50 g sample of a metal carbonate, MCO_3, is dissolved in water, and the solution is made up to 100 cm³. A 25 cm³ portion of this solution is titrated with 0.100 $moldm^{-3}$ hydrochloric acid. The average volume of HCl required for neutralisation is 24.0 cm³. Calculate the relative atomic mass of the metal M.

Data

Mass of MCO_3 = 0.50 g

Total volume of solution = 100 cm³

Volume of solution used in titration = 25 cm³

Concentration of HCl = 0.100 $moldm^{-3}$

Average volume of HCl used = 24.0 cm³

Steps

1. *Write the Balanced Equation*

The reaction between metal carbonate (MCO_3) and hydrochloric acid HCl) is:

$$MCO_3 + 2HCl \rightarrow MCl_2 + H_2O + CO_2$$

From the equation, 1 mole of MCO_3 reacts with 2 moles of HCl.

2. *Calculate the Moles of HCl Used in the Titration*

Given:

Concentration of HCl = 0.100 $moldm^{-3}$

Volume of HCl used = 24.0 cm³ = 0.024 dm³

Using the formula:

$$n(HCl) = C(HCl) \times V(HCl)$$

$$= 0.100 \times 0.024$$

$= 0.0024$ mol

3. *Calculate the Moles of MCO_3 in 25 cm³ of Solution*

   Since 1 mole of MCO_3 reacts with 2 moles of HCl, we find:

   $$\frac{n(MCO_3)}{n(HCl)} = \frac{1}{2}$$

   $$n(MCO_3) = \frac{n(HCl)}{2} = \frac{0.0024}{2} = 0.0012 \text{ mol}$$

4. *Find the Total Moles of MCO_3 in the 100 cm³ solution*

   25 cm^3 of MCO_3 *contains* 0.0012 mol

   $$\therefore 100 \text{ cm}^3 \text{ of } MCO_3 = \frac{100 \text{ cm}^3}{25 \text{ cm}^3} \times 0.0012 \text{ mol} = 0.0048 \text{ mol}$$

5. *Calculate the Molar Mass of MCO_3*

   The sample of MCO_3 weighs 0.50 g. Using the total moles calculated:

   $$\text{Molar Mass of } MCO_3 = \frac{m(MCO_3)}{n(MCO_3)}$$

   $$= \frac{0.50}{0.0048} = 104.17$$

6. *Calculate the Relative Atomic Mass of Metal M*

   Molar mass of MCO_3 = Relative atomic mass of M + 12 + 48

   $104.17 = M + 60$

   $M = 104.17 - 60 = 44.17$

   The relative atomic mass of metal M is approximately 44.17.

Activity 4.14 Indirect titration: Determining the Calcium Carbonate Content in Limestone

To determine the amount of calcium carbonate ($CaCO_3$) in a 2.5 g sample of limestone, a student adds 50 cm³ of 0.5 moldm^{-3} HCl to the powdered limestone sample. This mixture is left to react until no more gas evolved. The remaining unreacted HCl is then titrated with 0.25 moldm^{-3} NaOH, requiring 20.0 cm³ of NaOH to reach the endpoint. Calculate the percentage of calcium carbonate in the limestone sample.

Steps

1. Calculate the Initial Moles of HCl Added

   Given:

   Volume of HCl = 50 cm^3 = 0.05 dm^3

   Concentration of HCl = 0.5 $moldm^{-3}$

   Using the formula:

   Moles of HCl = Concentration × Volume

   Moles of HCl = 0.5 × 0.05 = 0.025 mol

   This is the initial amount of HCl added to the limestone sample.

2. Calculate the Moles of NaOH Used in the Back Titration

   Given:

   Volume of NaOH used = 20.0 cm^3 = 0.02 dm^3

   Concentration of NaOH = 0.25 $moldm^{-3}$

   Moles of NaOH = concentration × Volume

   Moles of NaOH = 0.25 × 0.020 = 0.0050 mol

   Since HCl and NaOH react in a 1:1 molar ratio, the moles of unreacted HCl in the solution is also 0.005 mol

3. Calculate the moles of HCl that reacted with $CaCO_3$

   Subtract the unreacted HCl from the initial HCl to find the moles of HCl that with $CaCO_3$:

   Moles of HCl that reacted = 0.025−0.005 = 0.020 mol

4. Calculate the Moles of $CaCO_3$ in the Sample

   The reaction between $CaCO_3$ and HCl is:

   $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$

   From the balanced equation, 1 mole of $CaCO_3$ reacts with 2 moles of HCl.

   Therefore, the moles of $CaCO_3$ that reacted is:

   $$\text{Moles of } CaCO_3 = \frac{\text{Moles of HCl that reacted}}{2}$$

   $$= \frac{0.020}{2} = 0.010 \text{ mol}$$

5. Calculate the Mass of $CaCO_3$ in the Sample

Using the molar mass of $CaCO_3 = 100\ gmol^{-1}$

Mass of $CaCO_3$ = Moles of $CaCO_3 \times$ Molar Mass of $CaCO_3$

Mass of $CaCO_3 = 0.01 \times 100 = 1.0\,g$

6. Calculate the Percentage of $CaCO_3$ in the limestone sample

$$\text{Percentage of } CaCO_3 = \frac{\text{Mass of } CaCO_3}{\text{Mass of sample}} \times 100$$

$$= \frac{1.0}{2.5} \times 100 = 40\%$$

Activity 4.15 Titration with pH Monitoring

Materials Needed

pH meter or pH sensor with data logger, burette, pipette, conical flask, beakers, Strong acid (e.g., HCl), strong base (e.g., NaOH), weak acid (e.g., CH_3COOH - acetic acid), weak base (e.g., NH_3 - ammonia), distilled water, white tile, graph paper or computer software for plotting volume vs. pH graphs

Steps

1. Preparation and Setup

- **a.** Set up the burette on a retort stand.
- **b.** Calibrate the pH meter.
- **c.** Rinse and fill the burette with the titrant (acid or base) specific to each titration type.
- **d.** Pipette a measured volume (e.g., 25 cm³) of the analyte into the conical flask.
- **e.** Place the conical flask on a white tile under the burette.

2. For each titration, follow these general steps but with specific combinations for type of acid and base.

- **a.** Strong Acid vs. Strong Base (e.g., HCl and NaOH)
  - i. Place the pH sensor in the conical flask with the analyte (NaOH).
  - ii. Record the initial pH of the solution before adding any HCl.
  - iii. Add the HCl titrant from the burette in small increments (e.g., 1.0 cm³).
  - iv. After each addition, swirl the flask gently and record the pH.

v. As the pH approaches neutral (around pH 7), reduce the increments to 0.20 cm³ for more precise measurements.

vi. Continue until the pH stabilizes after each addition.

vii. Plot the volume of HCl added on the x-axis and pH on the y-axis to determine sharp jump in pH around the equivalence point (expected to be near pH 7 strong acid vs. strong base).

b. Strong Acid vs. Weak Base (e.g., HCl and NH_3)

i. Use NH_3 as the analyte and HCl as the titrant.

ii. Follow the same process, adding HCl in small increments and recording the pH.

iii. This titration will show a more gradual increase in pH until equivalence, which expected to be below pH 7 due to the acidic nature of the solution neutralization.

iv. The volume vs. pH graph should show a less sharp change at the point compared to strong acid vs. strong base.

c. Strong Base vs. Weak Acid (e.g., NaOH and CH_3COOH)

i. Use CH_3COOH as the analyte and NaOH as the titrant.

ii. Add NaOH in small increments, recording the pH after each addition.

iii. The equivalence point will be slightly above pH 7, as the acetate ion (CH_3COO^-) produces a basic solution.

iv. The volume vs. pH graph will have a less steep curve around the point.

d. Weak Base vs. Weak Acid (e.g., NH_3 and CH_3COOH)

i. Use NH_3 as the analyte and CH_3COOH as the titrant.

ii. Follow the same process, adding CH_3COOH in small increments and recording the pH.

iii. The equivalence point will be near neutral, but the pH change will be very gradual, with no sharp increase.

iv. The volume vs. pH graph will have a gentle slope, making it harder to identify a distinct equivalence point.

3. For each titration, plot the volume of titrant added (x-axis) vs. pH (y-axis).

a. Identify the Equivalence Point: Look for the point on each graph where the pH changes most sharply.

b. Compare Graphs: Analyse how the steepness of the curve at the equivalence point differs for each type of titration:

   i. Strong acid vs. strong base: Very sharp pH change around equivalence.
   ii. Strong acid vs. weak base: Moderate pH change with an equivalence below pH 7.
   iii. Strong base vs. weak acid: Moderate pH change with an equivalence above pH 7.
   iv. Weak acid vs. weak base: Very gradual pH change with no sharp point.

4. Discussion and Conclusion

   a. Discuss the pH at the equivalence points for each titration type and why they differ.
   b. Discuss how the shape of each graph relates to the strength of the acid and base in the reaction.

EXTENDED READING

Use the following links below for further reading:

• Link 1	
• Link 2	
• Link 3	

REVIEW QUESTIONS

Review Question 4.1

1. What is the Arrhenius definition of an acid and a base?
2. What is a conjugate acid-base pair in the Brønsted-Lowry theory?
3. What makes a substance a Lewis acid or a Lewis base?
4. How do the Arrhenius, Brønsted-Lowry, and Lewis definitions of acids and differ?
5. Why can the Brønsted-Lowry concept be applied in both water and non-environments, unlike the Arrhenius concept?
6. Give an example of a substance that fits each acid definition (Arrhenius, Brønsted- Lowry, and Lewis).
7. Classify each of the following species in aqueous solution as a Brønsted-Lowry acid base:
   - a. HCl
   - b. NO_2
   - c. CH_3COOH
8. Identify the Brønsted-Lowry acid and its conjugate base in the following reactions
   - a. $H_2SO_4 + HNO_3 \rightarrow H_2NO_3 + HSO_4^-$
   - b. $HCl + H_2O \rightarrow H_3O + + Cl^-$
   - c. $NH_3 + H_2S \rightarrow HS^- + NH_4^+$
9. How would you determine if a given substance is an Arrhenius acid, a Brønsted-L acid, or a Lewis acid? What steps would you take?
10. Explain why the Lewis theory is considered the broadest definition of acids and bases. Can it describe all the situations that the Arrhenius and Brønsted-Lowry definitions cover?
11. Design an experiment that would allow you to identify whether a substance behaves an Arrhenius acid, a Brønsted-Lowry acid, or a Lewis acid.
12. Analyse the limitations of each acid-base theory and discuss scenarios where theory might be more useful than the others.

Review Question 4.2

1. What is a common physical property of acids?
2. What colour does litmus paper turn in the presence of a base?
3. Name a physical property of bases.
4. What is the pH range for acids and bases?
5. Explain why acids are able to conduct electricity.
6. Describe the chemical reaction that occurs when an acid is neutralized by a base.
7. Compare the physical properties of acids and bases based on taste, feel, and on indicators.
8. How can you use pH indicators to differentiate between an acid and a base?
9. A pH meter was used to check the pH level of four solutions labelled as W, X, Y and Z described in the table below

	W	X	Y	Z
pH	7	3	14	1

   a. Which of the solution(s) will taste sour?
   b. Identify the solution that will have the same pH as a dilute Na_2CO_3
   c. Which solution is the most acidic?
   d. Which solution will have no effect on litmus paper?

10. How does the chemical property of neutralisation help in treating an acidic stomach?
11. Analyse how the property of acids and bases as electrolytes is applied in technology.
12. Discuss why alkaline (basic) solutions are often used for cleaning purposes.
13. Evaluate the environmental impact of acid rain on soil and water bodies propose a method to counteract these effects using bases.
14. Propose a solution for farmers facing soil acidity issues by applying the neutralisation and evaluating the pros and cons of different bases that could used.

Review Question 4.3

1. What is a strong acid?
2. What is the pH range for a strong acid compared to a weak acid?
3. How do strong and weak acids differ in terms of electrical conductivity?
4. What does the term “extent of dissociation” mean in acids and bases?
5. Why does a strong acid conduct electricity better per mole than a weak acid?
6. Compare the rate of reaction of strong acids and weak acids with metals.
7. Describe the difference in enthalpy change during neutralisation between a strong and a weak acid.
8. How would you explain the differences in electrical conductivity between solutions hydrochloric acid (strong acid) and acetic acid (weak acid) at the same concentration?
9. Why does the enthalpy change of neutralisation differ between a strong acid like HCl and a weak acid like CH_3COOH?
10. Explain why strong bases like sodium hydroxide have a higher pH than weak bases ammonia, even at the same concentration.
11. Design an experiment to compare the electrical conductivity of strong and weak and bases. What materials would you need, and what results would you expect?
12. Evaluate the enthalpy changes of neutralisation for reactions involving strong weak acids and propose an explanation for any differences in the values.
13. Predict the pH, conductivity, and rate of reaction with magnesium of a 1M solution HCl compared to a 1M solution of CH_3COOH, and explain the reasoning.

Review Question 4.4

1. What is a salt in chemistry?
2. What are the two main components that form a salt?
3. Give an example of a common salt.
4. How is a salt formed during a reaction between an acid and a base?

5. What is the difference between a normal salt and an acidic salt?
6. How can we classify salts based on their properties?
7. What makes a double salt different from a complex salt?
8. Explain why a basic salt might form when a weak acid reacts with a strong base.
9. How would you identify an efflorescent salt, and what conditions affect it?
10. Compare and contrast deliquescent and hygroscopic salts in terms of their behaviour with moisture in the air.
11. Why are hydrated salts important in certain industries?
12. Design an experiment to distinguish between an acidic salt and a basic salt. Explain the steps and expected results.
13. Analyse the environmental impact of efflorescent and deliquescent salts in industrial applications.
14. Predict the effects of using a double salt instead of a normal salt in a specific application, like water treatment or dyeing.

Review Questions 4.5

1. Describe the process of an acid-base titration.
2. Define “water of crystallisation” and give an example.
3. How would you calculate the percentage purity of a sample of impure NaOH if 0.50 g of the sample requires 25 cm^3 of 0.10 M HCl for complete neutralisation?
4. A hydrated salt has a mass of 4.0 g before heating and 2.8 g after heating. C the percentage of water of crystallisation.
5. Explain why phenolphthalein is often used as an indicator in acid-base involving strong bases.
6. Assess potential sources of error in determining percentage purity through titration and suggest methods to minimise them.

SECTION

5

Trends of Chemical and Physical Properties of Elements and their Compounds in the Periodic Table

SYSTEMATIC CHEMISTRY OF THE ELEMENTS

Periodicity

INTRODUCTION

In this section, you will explore the Period 3 elements and their compounds. You will learn about the patterns in their physical and chemical properties, including their hydrides, oxides, hydroxides, and chlorides. By studying atomic number and electron configuration, you will understand what makes each element and compound unique and how they are related.

KEY IDEAS

- Acid strength is a measure of a compound's ability to donate protons (H^+) in solution.
- Amphoteric compounds are compounds that can react with both acids and bases.
- Bond Strength is the energy required to break a bond in a molecule.
- Density is the mass of a substance per unit volume.
- Electrical conductivity is the ability of a material to conduct electricity.
- Hydrolysis is the reaction of a compound with water.
- Melting and boiling points is the temperatures at which a substance changes from solid to liquid (melting) or liquid to gas (boiling).
- Metallic character is a measure of how easily an element can lose electrons.
- Reactivity is how easily an element reacts with others.
- The Periodic Law states that the properties of elements repeat in a regular pattern based on their atomic numbers and electron arrangements.
- Thermal stability is the resistance of a compound to decomposition at high temperatures.
- Volatility is the tendency of a compound to vaporise or form a gas.

PHYSICAL AND CHEMICAL PROPERTIES OF PERIOD 3 ELEMENTS

Modern Periodic Law

The modern Periodic law states that the physical and chemical properties of elements are related to their atomic numbers. This means elements with similar electron arrangements have similar properties. These properties include:

1. Chemical properties (like reactivity)
2. Physical properties (like melting points, boiling points, and density)
3. Chemical bonding (like ionic, covalent, or metallic bonds)

Understanding these properties helps predict how elements react and why they have certain characteristics.

Period 3 Elements

Period 3 elements are the elements in the third row of the periodic table. They include: Sodium (Na), magnesium (Mg), aluminium (Al), silicon (Si), phosphorus (P), sulphur (S), chlorine (Cl) and argon (Ar).

${}^{23.0}_{11}\text{Na}$	${}^{24.3}_{12}\text{Mg}$	${}^{27.0}_{13}\text{Al}$	${}^{28.1}_{14}\text{Si}$	${}^{31.0}_{15}\text{P}$	${}^{32.1}_{16}\text{S}$	${}^{35.5}_{17}\text{Cl}$	${}^{39.9}_{18}\text{Ar}$

These elements have their outer electrons in the third energy level (shell).

1. Metallic Character

Metallic character is how easily an element loses its electrons.

As you move left to right across Period 3:

- Metallic character decreases.
- This happens because, across the period, elements gain electrons more than they lose them in order to exist with a full outer electron shell (increasing electron affinity).
- Ionization energy increases, making it require more energy to lose electrons.

Example: Sodium (Na), magnesium (Mg), and aluminium (Al) are metals with high metallic character.

Silicon (Si) is a metalloid (a mix of metal and nonmetal properties).

Phosphorus (P), sulphur (S), chlorine (Cl), and argon (Ar) are non-metals with low metallic character.

2. Melting and Boiling Points

a. Melting and boiling points of Period 3 elements:
   i. Increase from Na to Si (metals to a metalloid).
   ii. **Decrease** from **Si** to **Ar** (non-metals).

Reason

Metals like **Na**, **Mg**, and **Al** form **metallic bonds**, which require a lot of energy to break.

Silicon (Si), a metalloid, forms **strong covalent bonds**, which also require a lot of energy to break.

Trend in the Melting and Boiling Points of P4, S_8, Cl2, and Ar

The melting and boiling points of **phosphorus (P4), sulphur (S_8), chlorine (Cl2), and argon (Ar)** follow the trend:

$S_8 > P_4 > Cl_2 > Ar$

This trend is primarily influenced by **London dispersion (van der Waals) forces**, the only intermolecular forces present in these **non-polar** substances.

Factors Influencing the Trend

1. Molecular Size: Larger molecules have more electrons, resulting in stronger dispersion forces.
2. Molecular Shape: More complex structures provide a greater surface area for intermolecular interactions.

Among these substances, **S_8** is the largest molecule with the most electrons, leading to the strongest van der Waals forces and the highest melting/boiling points. In contrast, **Ar**, being a monatomic gas, has the weakest intermolecular forces and the lowest melting/boiling points.

Refer to Figure 5.1 for trends in boiling and melting points across Period 3

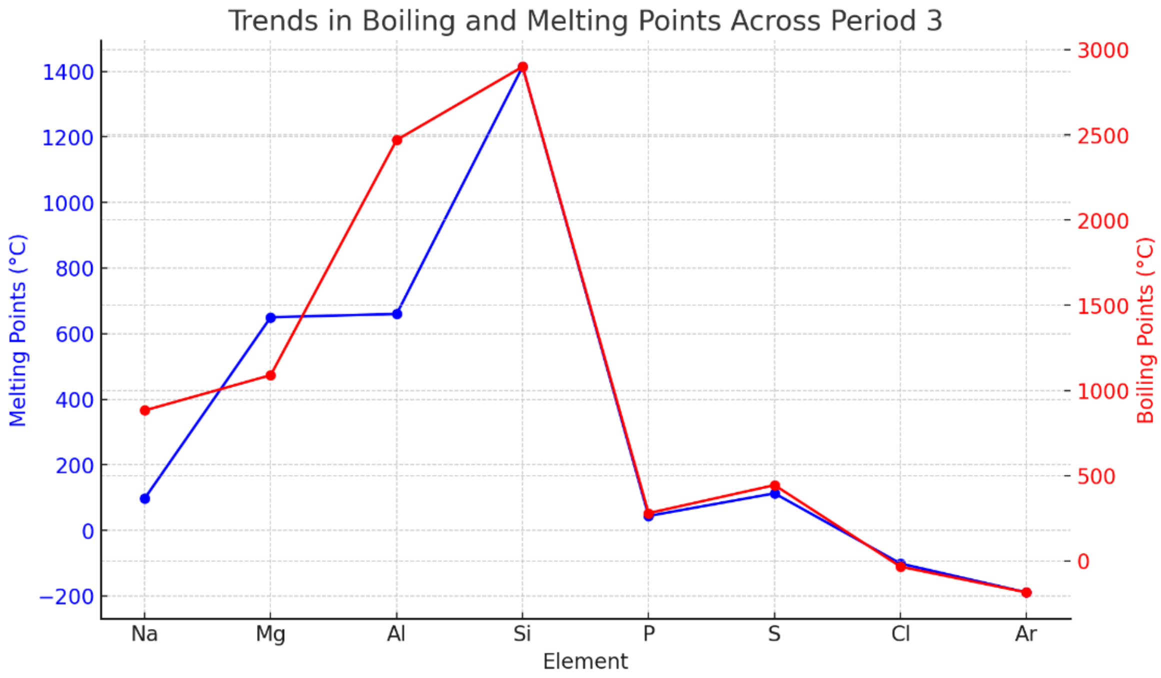


Figure 5.1: Trends in Boiling and Melting Points Across Period 3

The **blue line** represents the **melting points** of Period 3 elements, showing a rise from **Na** (Sodium) to **Si** (Silicon), followed by a decrease for the non-metals (P, S, Cl, Ar).

The **red line** represents the **boiling points**, which follow a similar trend: an increase up to **Si**, followed by a decrease across the non-metals.

3. Density

For Period 3 elements (Na, Mg, Al, Si, P, S, Cl, Ar), the trend in density is influenced by atomic mass, atomic size, and crystal structure.

Density Trend Across Period 3

Refer to Figure 5.2 for density variation in Period 3 Elements

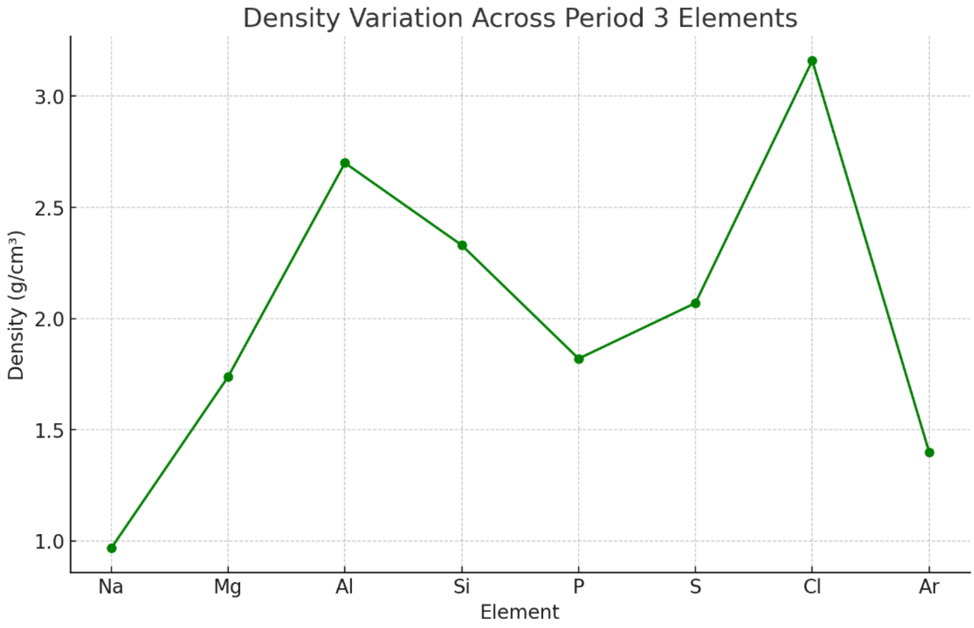


Figure 5.2: *Density Variation in Period 3 Elements*

Explanation of the Graph

The graph illustrates the density trend of Period 3 elements:

1. Increase from Na to Al
    - **a.** Na → Mg → Al: Density generally increases as atomic mass increases and atomic size decreases.
    - **b.** Silicon (Si) has a relatively high density (compared to other covalent molecules) due to its giant covalent structure.
2. Increase from P to Cl
    - **a.** P, S, and Cl have increasing densities due the increase in atomic mass as we go across the period. It should be noted that the density of Phosphorous is variable depending on the allotrope being considered (the most common allotrope is the one quoted here as 1.82 g/cm^3).
    - **b.** Argon (Ar) has a low density as it is a monatomic gas.

3. Electrical Conductivity

Electrical conductivity is the ability of a substance to conduct electricity. It depends on the number of free electrons in the material.

Trend in electrical conductivity across Period 3

The conductivity decreases from left to right.

Na, Mg, and Al are good conductors because they have many free electrons.

Si is a semiconductor: it can conduct electricity under certain conditions, like when it's modified in electronics.

P, S, Cl, and Ar are poor conductors (insulators) because they have few free electrons.

Element	Atomic number	Melting Point (°C)	Boiling Point (°C)	Density (g/cm^3)	Electrical Conductivity	Metallic Character
Sodium (Na)	11	97.8	883	0.968	Good conductor	High
Magnesium (Mg)	12	650	1090	1.738	Good conductor	High
Aluminium (Al)	13	660	2470	2.70	Good conductor	High
Silicon (Si)	14	1414	2900	2.33	Semi conductor	Metalloid
Phosphorus (P)	15	44	280	1.82	Poor conductor	Non-metal
Sulphur (S)	16	113	444.6	2.07	Poor conductor	Non-metal
Chlorine (Cl)	17	-101	-34	3.16	Poor conductor	Non-metal
Argon (Ar)	18	-189	-185	1.40	Poor conductor	Non-metal

Chemical Properties of Compounds of Period 3 Elements

Across Period 3, elements go from **metals** (that easily lose electrons) to **metalloids** (with properties of both metals and nonmetals) to **nonmetals** (that tend to gain electrons). The **reactivity** of these elements decreases as you move from left to right, meaning the metals are more reactive than the nonmetals.

Reaction with Water

Sodium reacts very **violently** with water to form **sodium hydroxide (NaOH)** and hydrogen gas:

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

Magnesium reacts **slowly** with cold water to form **magnesium hydroxide (Mg(OH)2)** and hydrogen gas. It reacts more **vigorously** with hot water (steam) to form **magnesium oxide (MgO)**

$Mg(s) + 2H_2O(l) \rightarrow Mg(OH)_2(s) + H_2(g)$

$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$

Reaction with Oxygen

Most of the elements in Period 3 react with **oxygen**:

Sodium forms **sodium oxide (Na2O)**:

$4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

Magnesium forms **magnesium oxide (MgO)**: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

Aluminium forms **aluminium oxide (Al2O3)**: $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

Silicon forms **silicon dioxide (SiO2)**: $Si(s) + O_2(g) \rightarrow SiO_2(s)$

Phosphorus forms **phosphorus pentoxide (P4O$_{10}$)**: $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(g)$

Sulphur forms **sulphur dioxide (SO2)**: $S_8(s) + 8O_2(g) \rightarrow 8SO_2(g)$

Chlorine does not react with oxygen under normal conditions.

Argon does not react with oxygen because it is a noble gas with a stable electron configuration.

Reaction with Chlorine

Chlorine (Cl2) is a highly reactive gas, and the elements in Period 3 react with it in different ways:

Sodium reacts strongly with chlorine to form **sodium chloride (NaCl)**:

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Magnesium reacts with chlorine to form **magnesium chloride (MgCl2)**:

$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$

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

$2Al(s) + 3Cl_2 (g) \rightarrow 2AlCl_3(s)$

Silicon reacts with chlorine to form **silicon tetrachloride (SiCl4)**:

$Si(s) + 2Cl_2 (g) \rightarrow SiCl_4(s)$

Phosphorus reacts with chlorine to form **phosphorus trichloride (PCl3)**:

$P(s) + 3Cl_2 (g) \rightarrow PCl_3(s)$

Sulphur reacts with chlorine to form **sulphur dichloride (SCl2)**:

$S(s) + Cl_2(g) \rightarrow SCl_2(s)$

Chlorine does not react with itself.

Argon is **inert** and does not react with chlorine.

Activity 5.1 Period 3 Elements

Fill in the table with the corresponding information for each Period 3 element (Sodium, Magnesium, Aluminium, Silicon, Phosphorus, Sulphur, Chlorine, and Argon).

Element	**Atomic number**	**Density (g/cm^3)**	**Melting point (°C)**	**Boiling point (°C)**	**Atomic radius (pm)**	**Phase at room temperature**
Na						
Mg						
Al						
Si						
P						
S						
Cl						
Ar						

Use the periodic table and other reference materials to complete the table with the correct values for each property.

Activity 5.2 Exploring Periodic Trends (physical and chemical properties) in Period 3 Elements

Material needed: Periodic table

Steps

1. Mention what you know about metals and non-metals? ame some metals, metalloids and non-metals from the Periodic Table.

2. **a.** Explain metallic properties.
   a. Discuss how periodic properties such as ionization energy, reactivity are measured.
   b. Using a chart of Period 3 elements explain the trend: Metallic property decreases from left to right.

3. Mention physical properties of period 3 elements.
    - **a.** State how these physical properties measured.
    - **b.** Discuss how density and melting/boiling points vary across the period and down a group.
4. Define chemical properties.
5. Go online and watch videos (or observe these demonstrated in class) illustrating reactions:
    - **a.** Sodium with water (vigorous reaction). https://www.youtube.com/watch?v=dmcfsEEogxs
    - **b.** Magnesium with air (combusting to form MgO). https://www.youtube.com/watch?v=NnFzHt6l4z8
    - **c.** Chlorine with hydrogen (formation of HCl gas). https://www.youtube.com/watch?v=MtygiCwnEzw
6. Discuss how reactivity changes:
    - **a.** Metals (Na, Mg, Al): Reactivity decreases from left to right.
    - **b.** Non-metals (Si, P, S, Cl): Reactivity increases.
7. In mixed-ability groups research using books, internet access, charts, or visual aids, record the element's physical and chemical properties of Period 3 elements.

    Prepare a presentation to share findings with the class.

Activity 5.3 Interpreting trends in the physical properties of Period 3 elements graphs

Material needed: Graph sheet, pencil, ruler

Steps

1. Use the data (e.g., atomic radius, melting point, boiling point, density) generated Activity 5.1 to plot graphs of each property against the atomic number of the elements.

    Example properties to plot:
    - **a.** Atomic radius vs. atomic number.
    - **b.** Melting point vs. atomic number.
    - **c.** Density vs. Atomic number.

2. Examine the graphs and discuss patterns in small groups.

How do the physical properties change as the atomic number increases?

Explain silicon's high melting point compared to neighbouring elements

3. What is the relationship between melting point and atomic number?

Activity 5.4 Reactions of Period 3 Elements

Materials needed: Periodic table for reference, diagrams or pictures of chemical reactions, access to the internet or books for research, paper and pencils for note-taking, (Optional) demonstration materials: small pieces of sodium, magnesium, and other Period 3 elements (for teacher-led demonstration), safety goggles (for demonstrations)

A **chemical reaction** is when one substance changes into a new substance with different properties (e.g., when sodium reacts with water to form sodium hydroxide and hydrogen gas)

Steps

1. In small group research into how period 3 elements (Na, Mg, Al and Si) react with:

- **a.** Water:
  - i. Does it react quickly or slowly?
  - ii. What products are formed?
- **b.** Oxygen:
  - i. Does it form an oxide?
  - ii. What type of oxide is produced?
- **c.** Chlorine
  - i. Does it form a chloride?
  - ii. How reactive is it with chlorine?

2. Use the internet, or books to gather information on:

- **a.** Reaction equations (e.g., $2Na + 2H_2O \rightarrow 2NaOH + H_2$).
- **b.** Type of bond formed (ionic, covalent, etc.).
- **c.** Reaction observations (e.g., gas released, heat produced).

3. In small groups present your findings, showing the reaction equations and explaining:
    - **a.** How your assigned element reacts with water, oxygen, and chlorine.
    - **b.** How the reactivity changes as you move across the period.
    - **c.** If your element is a metal, metalloid, or nonmetal, and how this affects its reactions.
4. Discuss:
    - **a.** Why do metals like sodium and magnesium react more easily with water to non-metals like sulphur and chlorine?
    - **b.** Why does argon not react with other substances?
5. Prepare a multimedia presentation to showcase your findings.

    Slideshows (e.g., Google Slides, PowerPoint), short videos (e.g., using apps like C or Flipgrid), Posters (physical or digital).

    Include visuals (e.g., reaction photos, diagrams), videos of reactions (if online).

PHYSICAL AND CHEMICAL PROPERTIES OF COMPOUNDS OF PERIOD 3 ELEMENTS

Physical and Chemical Properties of Hydrides of Period 3

Period 3 hydrides are compounds made by combining hydrogen with the elements in the third period of the periodic table. Examples include sodium hydride (NaH), magnesium hydride (MgH_2), aluminium hydride (AlH_3), silicon hydride (SiH_4), phosphorus hydride (PH_3), hydrogen sulphide (H_2S), and hydrogen chloride (HCl).

Physical Properties

1. **Bonding and Structure**

    NaH and MgH_2

    These are ionic hydrides where hydrogen forms H^- ions.

    They have high melting and boiling points because of strong ionic bonds.

AlH_3:

A covalent hydride with some ionic character and a polymeric structure.

SiH_4, PH_3, H_2S, HCl

These are covalent hydrides with simple molecular structures.

They have lower melting and boiling points.

2. **Melting and Boiling Points**

   Ionic hydrides (NaH, MgH_2): Have high melting and boiling points due to strong bonds.

   Covalent hydrides (SiH_4, PH_3, H_2S, HCl): Have lower melting and boiling points due to

   weaker intermolecular forces (Van der Waals forces).

3. **Solubility in Water**

   Ionic hydrides (NaH, MgH_2): React with water to form hydrogen gas and hydroxides.

   AlH_3: Reacts with water to form **aluminium hydroxide** and hydrogen gas.

   Covalent hydrides (SiH_4, PH_3, H_2S, HCl): Are generally soluble in water, forming acidic or basic solutions.

4. **Thermal Stability**

   NaH and MgH_2: Thermally stable and decompose only at high temperatures.

   AlH_3: Less stable than NaH and MgH_2.

   Covalent hydrides (SiH_4, PH_3, H_2S, HCl): Decompose at lower temperatures.

Summary of Physical Properties of Period 3 Hydrides

The properties of Period 3 hydrides change as you move from metals to non-metals:

- Metal hydrides (ionic): High melting points and highly reactive with water.
- Non-metal hydrides (covalent): Lower melting points with varied reactivity.
- This shift reflects the change from ionic bonding in metals to covalent bonding in non-metals.

Chemical Properties

1. Reactivity: Reactivity increases from NaH to AlH_3.

   AlH_3 is very reactive and can even catch fire when exposed to air!

2. Acid-Base behaviour: Hydrides become more acidic as you go from NaH to AlH_3.

AlH_3 acts like a strong acid.

3. Redox Reactions: Hydrides become better at giving away electrons (reducing) NaH to AlH_3.

Physical and Chemical Properties of Oxides of Period 3

Period 3 oxides are compounds formed between oxygen and the elements in the third period of the periodic table. Examples include Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10}, SO_3, and Cl_2O_7.

Physical Properties

1. **Structure**

   Ionic oxides: (e.g., Na_2O, MgO) have strong ionic bonds between metal and oxygen ions.

   Giant covalent oxides: (e.g., SiO_2) form a continuous network of covalent bonds.

   Molecular oxides: (e.g., P_4O_{10}, SO_3) consist of small molecules held by weak forces.

2. **Melting and Boiling Points**

   High: In ionic and giant covalent oxides due to strong bonds.

   Lower: In molecular oxides because of weak intermolecular forces.

3. **Electrical Conductivity**

   Ionic oxides: Do not conduct electricity when solid but can conduct when molten because the ions are free to move.

   Covalent and molecular oxides: Do not conduct electricity.

Types of Oxides

i. **Ionic Oxides**

   Examples: Sodium oxide (Na_2O), magnesium oxide (MgO).

   Bonding: Strong ionic bonds between metal cations (e.g., Na^+, Mg^{2+}) and anions (O^{2-}).

   Melting and Boiling Points: High, due to strong attraction between ions.

ii. **Giant Covalent Oxides**

   Example: Silicon dioxide (SiO_2).

Bonding: Each silicon atom is bonded to four oxygen atoms in a 3D network.

Melting and Boiling Points: Very high, because breaking covalent bonds requires lots energy.

iii. **Molecular Oxides**

Examples: Phosphorus pentoxide (P_4O_{10}), sulphur trioxide (SO_3).

Bonding: Made up of small molecules held by weak Van der Waals forces.

Melting and Boiling Points: Low, because less energy is needed to break these forces.

Summary of Trends

Structure: Ionic → Giant covalent → Molecular.

Melting/Boiling Points: High (ionic and covalent) → Low (molecular).

Electrical Conductivity: Only ionic oxides conduct when molten.

Chemical Properties of Period 3 Oxides

1. **Acid-Base Character**

   Basic oxides: Like Na_2O and MgO, react with acids.

   Amphoteric oxides: Like Al_2O_3, can react with both acids and bases.

   Acidic oxides: Like SiO_2, P_4O_{10}, and SO_3, react with bases.

2. **Reactivity with Water**

   Basic oxides: Form hydroxides in water. Example:

   $Na_2O + H_2O \rightarrow 2NaOH$ (sodium hydroxide).

   Acidic oxides: Form acids in water. Example:

   $SO_3 + H_2O \rightarrow H_2SO_4$ (sulphuric acid).

Physical and Chemical Properties of Period 3 Hydroxides

Period 3 hydroxides are compounds formed between elements in Period 3 and hydroxide ions (**OH^-**). Examples include NaOH, $Mg(OH)_2$, $Al(OH)_3$, $Si(OH)_4$, $P(OH)_3$, H_2SO_4, and HClO.

Physical Properties of Period 3 Hydroxides

1. **Solubility in Water**

   NaOH: Highly soluble; forms a strong alkaline solution because it completely in water.

   $Mg(OH)_2$: Sparingly soluble; forms a weak alkaline solution due to its strong energy.

   $Al(OH)_3$: Insoluble in water but dissolves in both acids and bases because it amphoteric.

   $Si(OH)_4$ and $P(OH)_3$: Generally insoluble due to their covalent bonds, which do dissociate easily in water.

2. **Electrical Conductivity**

   NaOH and $Mg(OH)_2$: Conduct electricity in water because they produce free ions.

   $Al(OH)_3$, $Si(OH)_4$, and $P(OH)_3$: Do not conduct electricity as they do not release ions.

3. **Melting and Boiling Points**

   NaOH: High melting point due to strong ionic bonds.

   $Mg(OH)_2$: High melting point but lower than NaOH due to weaker ionic bonds.

   $Al(OH)_3$: Lower melting point; has both ionic and covalent bonds.

   $Si(OH)_4$ and $P(OH)_3$: Very low melting points because they are molecular with weak intermolecular forces.

Summary of Trends

Ionic hydroxides (NaOH, $Mg(OH)_2$): High solubility, high melting points, and good electrical conductivity.

Covalent hydroxides ($Si(OH)_4$, $P(OH)_3$): Low solubility, low melting points, and poor conductivity.

Amphoteric hydroxide ($Al(OH)_3$): Can act like both acids and bases, with intermediate properties.

The properties shift from **ionic compounds** (metals) to **covalent compounds** (non-metals) as you move across Period 3.

Chemical properties of hydroxides of Period 3

1. Acid-Base Behaviour

Metal Hydroxides (Basic Substances)

Metal Hydroxides are substances made by combining metals with hydroxide (OH^-). Most are basic.

Examples

Sodium Hydroxide ($NaOH$): A strong base that fully breaks into its parts in water: $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

It reacts with acids to form salt and water: $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

Magnesium Hydroxide ($Mg(OH)_2$): A weak base that only partly breaks in water: $Mg(OH)_2 \rightleftharpoons Mg^{2+} + 2OH^-$

It also reacts with acids to make salt and water: $Mg(OH)_2\ (aq) + 2HCl(aq) \rightarrow MgCl_2\ (aq) + 2H_2O(l)$

Aluminium Hydroxide ($Al(OH)_3$): This is special because it can act as both an acid and a base (amphoteric).

With acids: $Al(OH)_3\ (aq) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

With bases: $Al(OH)_3\ (aq) + NaOH(aq) \rightarrow NaAl(OH)_4\ (aq)$

Non-Metal Hydroxides (Acidic Substances)

Non-Metal Hydroxides are substances that have hydroxide but act like acids.

Examples

Sulphuric acid (H_2SO_4)

It completely dissociates in water: $H_2SO_4\ (aq) \rightarrow 2H^+\ (aq) + SO_4^{2-}(aq)$

It reacts with bases to form salt and water: $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

Hypochlorous Acid ($HClO$): A weak acid that only partly dissociates in water: $HClO(aq) \rightleftharpoons H^+(aq) + ClO^-(aq)$

It reacts with bases to make salt and water:

$HClO(aq) + NaOH(aq) \rightarrow NaClO(aq) + H_2O(l)$

2. Reactivity

Metal Hydroxides

Sodium Hydroxide (NaOH) and Magnesium Hydroxide ($Mg(OH)_2$) are **basic** and react with acids to form salts and water.

Aluminium Hydroxide ($Al(OH)_3$) is **amphoteric** and reacts with both acids and bases.

Non-Metal Hydroxides

Sulphuric Acid (H_2SO_4) and Hypochlorous Acid (HClO) are **acidic** and react with bases to make salts and water.

3. Thermal Stability

Trend: Thermal stability increases across the period.

NaOH and $Mg(OH)_2$: Decompose at lower temperatures due to weaker bonds.

$Al(OH)_3$ and $Si(OH)_4$: More stable at higher temperatures due to stronger lattice covalent bonds.

Patterns in physical and chemical properties of Period 3 chlorides

The chlorides of Period 3 elements are compounds formed between these elements and chlorine**.** Refer to **Table 5.1** for the Chlorides of period 3 and **Table 5.2** for the Physical Properties of Period 3 Chlorides.

Table 5.1: Chlorides of period 3

Compound	Name of Compound	Type of Chloride	Acidic, Basic, or Neutral
NaCl	Sodium chloride	Ionic	Neutral
$MgCl_2$	Magnesium chloride	Ionic	Neutral
$AlCl_3$	Aluminium chloride	Covalent	Acidic
$SiCl_4$	Silicon tetrachloride	Covalent	Acidic
PCl_5	Phosphorus pentachloride	Covalent	Acidic
$S2Cl_2$	Disulphur dichloride	Covalent	Neutral

Table 5.2: Physical Properties of Period 3 Chlorides

Property	Explanation
Structure	- **Ionic Chlorides** (e.g., NaCl, $MgCl_2$): Have strong, giant ionic lattice structures. - **Covalent Chlorides** (e.g., $AlCl_3$, $SiCl_4$, PCl_5, S_2Cl_2): Have covalent structures. $AlCl_3$ bonds with another molecule to form Al_2Cl_6 in solid form.
Melting & Boiling Points	- **Ionic Chlorides**: NaCl and $MgCl_2$ have high melting/boiling points because of strong ionic bonds. - **Covalent Chlorides**: $SiCl_4$, PCl_5, and S_2Cl_2 have lower melting/boiling points due to weaker forces holding their molecules together.
Electrical Conductivity	- **Ionic Chlorides**: NaCl and $MgCl_2$ can conduct electricity when melted because their ions can move freely. - **Covalent Chlorides**: $AlCl_3$, $SiCl_4$, PCl_5, and S_2Cl_2 don't conduct electricity in any state because they don't have free-moving ions or electrons.

Chemical Properties of Period 3 Chlorides

Refer to **Table 5.3** for chemical properties of Period 3 chlorides.

Table 5.3: Chemical Properties of Period 3 Chlorides

Property	Explanation
Reactivity with Water	- **Ionic Chlorides** (e.g., NaCl, $MgCl_2$): Dissolve in water to form ions. Example: $NaCl(aq) \rightarrow Na^+(aq) + Cl^-(aq)$. - **Covalent Chlorides**: React with water (hydrolysis) to form acids. - **AlCl3**: Forms HCl and Aluminium hydroxide. - **SiCl4**: Forms SiO_2 and HCl. - **PCl_5**: Forms phosphoric acid (H_3PO_4) and HCl.

Property	Explanation
Acid-Base Behaviour	- **Ionic Chlorides**: Neutral in water. - **Covalent Chlorides**: Form acidic solutions when reacting with water.

Thermal Stability of Period 2 and 3 Carbonates

See **Table 5.4** for thermal stability of Period 3 carbonates

Table 5.4: Thermal Stability of Period 3 Carbonates

Property	Explanation
Carbonates	- Thermal Decomposition: Carbonates break down when heated to form metal oxides and CO_2 gas. Example: $MCO_3 \rightarrow MO + CO_2$. - Period 3 Carbonates (e.g., $MgCO_3$): More stable and harder to decompose due to larger cations like Mg^{2+}.

Thermal Stability of Period 3 Nitrates

Refer to **Table 5.5** for thermal stability of Period 3 nitrates

Table 5.5: Thermal Stability of Period 3 Nitrates

Property	Explanation
Nitrates	- **Thermal Decomposition**: Nitrates break down when heated. Example: $2M(NO_3)_2 \rightarrow 2MO + 4NO_2 + O_2$. - **Period 3 Nitrates** (e.g., $NaNO_3$, $Mg(NO_3)_2$): Decompose to form nitrites or oxides and are more stable due to larger cations.

Note

- **Small Cations (like Li^+)**: Make compounds less stable and easier to breakdown with heat.
- **Big Cations (like K^+)**: Make compounds more stable and harder to decompose.
- **Reactions with Water**: Ionic chlorides dissolve; covalent chlorides form acids.

- **Decomposition**: Carbonates and nitrates break down when heated, giving gases and solids.

Activity 5.5 Period 3 Elements and Their Compounds

Objective: Explore the properties, structure, and behaviour of Period 3 elements and their compounds (hydrides, oxides, hydroxides, and chlorides).

1. Use your textbook or provided handouts to learn about Period 3 elements and compounds. Pay attention to:
   a. Identify whether each compound exhibits ionic or covalent bonding.
   b. Note acid-base behaviour (basic, neutral, or slightly acidic).
   c. Understand their behaviour with water and air.
2. Create a table or chart to summarise the properties of each compound type (hydrides, oxides, hydroxides, and chlorides). Use diagrams to illustrate bonding types applicable.
3. Write a brief explanation of the hydrolytic behaviour of chlorides in Period 3 (e.g., $AlCl_3$ is hydrolysed in water while NaCl is not).
4. Write your thoughts on the trends in the acid-base characteristics of Period 3 oxides you move across the period.
5. Complete a quiz or worksheet that includes questions on bonding, behaviour, and reaction equations for at least three compounds.

Activity 5.6 Analysing the Trends in Thermal Stability of Carbonates and Nitrates

Objective: Analyse the trends in thermal stability and their relationship with ionic size and charge density.

1. Read about thermal stability of Group 1 and Group 2 carbonates and nitrates. U additional resources like online simulations or diagrams.
2. Draw a concept map explaining ionic size, charge density, and polarising power, how these factors influence thermal stability.
3. Compare the thermal stability of lithium carbonate, potassium carbonate, carbonate. Write a short paragraph explaining the differences.
4. Write balanced decomposition reactions for $NaNO_3$, and $MgCO_3$, observations such as gas evolution.

5. Complete a short quiz on trends in thermal stability, ionic size, and charge density.

Activity 5.7 Investigation on Thermal Stability of Carbonates and Nitrates

Objective: Conduct and analyse a hands-on experiment to investigate the thermal stability of carbonates.

1. **a.** Review, together with you teacher or laboratory instructor, the safety guidelines heating chemicals.

b. Materials required: test tubes, boiling tubes, delivery tubes, test tubes, holders, Na_2CO_3, K_2CO_3, $CaCO_3$, Bunsen burner, limewater, etc.

2. Heat about 10 g of each solid carbonate in a boiling tube and observe changes. Test CO_2 gas using limewater.

3. Record your observations in a table (e.g., colour change, gas evolved, changes limewater).

4. Write the balanced equations for the decomposition of the carbonates tested and discuss the differences in thermal stability based on observations in relation to size and polarising power.

5. Write a short conclusion on the factors affecting thermal stability in carbonates nitrates.

6. Reflect on the experiment by answering guiding questions:

a. Why does calcium carbonate decompose at a higher temperature compared sodium carbonate?

b. How does the test for CO_2 confirm decomposition

EXTENDED READING

- Properties of the period 3 “hydroxides” – Period 3 Hydroxides
- Hydroxides of Period 3 Elements - Chemistry LibreTexts - Link 1
- Physical Properties of the Period 3 Elements - Link 2

REVIEW QUESTIONS

REVIEW QUESTIONS 5.1

1. What happens to the atomic radius as you move across Period 3?
2. Which Period 3 element has the highest melting point?
3. How does the trend in ionisation energy change across Period 3?
4. Compare the reactivity of sodium (Na) and chlorine (Cl).
5. Explain why the melting points of Period 3 elements first increase and then decrease.
6. Predict how magnesium (Mg) reacts with water compared to sodium (Na) and why.
7. Analyse how the trends in physical properties across Period 3 are related to electron configuration and bonding.
8. Evaluate how the reactivity trends of Period 3 elements influence their applications.

REVIEW QUESTIONS 5.2

1. List the Period 3 elements and their corresponding hydrides. (Use a periodic table for reference)
2. Compare the thermal stability of Na_2CO_3, K_2CO_3, and $CaCO_3$.
3. Having gone through the experiment in this lesson, use the knowledge gained to provide concise answers to the following set of questions as experimental review project.
    - a. List the correct oxides of the Period 3 elements.
    - b. Classify each oxide as acidic, basic, or amphoteric based on their chemical properties. Provide a hint for your reasons for the classification.
    - c. Create a table with the headings like the one below to summarise your findings:

Elements	Oxide(s)	Classification	Explanation

d. Write balanced chemical equations for sodium oxide (Na_2O) and sulphur dioxide (SO_2) reacting with water. the differences in the chemical behaviour of the oxides of sodium (Na_2O) and sulphur (SO_2).

e. Write a short paragraph summarising how bonding and structure influence the acid-base behaviour of these oxides.

SECTION

6

PHYSICAL AND CHEMICAL PROPERTIES OF THE HALOGENS

SYSTEMATIC CHEMISTRY OF THE ELEMENTS

Periodicity

INTRODUCTION

In this section, you will explore halogens, the elements in Group 17 of the periodic table. You will examine their physical and chemical properties, focusing on what makes them unique, how they react, and the patterns they follow as you move down the group.

You will also look at the reactions of halide salts (compounds made with halogens) and explore the strength of hydrogen halide acids, discussing their real-life applications in various industries.

Throughout these lessons, you will use digital tools to enhance your learning and ensure that everyone, regardless of gender or background, has the opportunity to understand how chemistry shapes our world.

KEY IDEAS

- **Displacement reaction** is a reaction where a more reactive halogen displaces a less reactive halide ion from its compound.
- **Halogens** are elements in Group 17 of the periodic table.
- **Ionic halides** are formed when halogens gain one electron (e.g., Cl^-, Br^-).
- **Precipitation reaction of halides** is a reaction where halide ions react with silver ions to form insoluble silver halides
- **Reducing agent** is a substance that donates electrons. Iodide ions are the strongest reducing agents among halide ions.

PHYSICAL AND CHEMICAL PROPERTIES OF HALOGENS (GROUP 17 ELEMENTS)

Halogens are elements in Group 17 of the periodic table. They are special because they exist in all three states of matter at room temperature:

1. Fluorine and Chlorine are gases.
2. Bromine is a liquid.
3. Iodine is a solid.
4. Astatine is radioactive, and not much is known about it.

Halogens are non-metals and usually form diatomic molecules (two atoms of the same element, like Cl_2 or F_2), except for Astatine. They are very reactive, so they are always found in nature combined with other elements, never on their own.

Halogens have a general electron configuration of $ns^2\ np^5$, meaning they almost have a full outer shell of electrons. They only need one more electron to become stable, like noble gases. This makes them the most reactive non-metals. When they gain an electron, they form halide ions (like Cl^-), and they can also form covalent compounds by sharing electrons.

1. Physical Properties of Halogens

Table 6.1 helps you to easily compare halogens' unique features.

Table 6.1: Physical Properties of Halogens

Property	Fluorine (F)	Chlorine (Cl)	Bromine (Br)	Iodine (I)	Astatine (At)
Physical State	Gas	Gas	Liquid	Solid	Solid (Radioactive)
Inter-molecular Forces	Weak (small size)	Weak	Moderate	Strong	Very Strong
Colour	Pale yellow	Yellow green	Reddish-brown	Shiny purple black	Dark (Unknown)
Odour	Sharp and strong	Sharp and strong	Strong and unpleasant	Slightly metallic	Unknown

Property	**Fluorine (F)**	**Chlorine (Cl)**	**Bromine (Br)**	**Iodine (I)**	**Astatine (At)**
Solubility in Water	Slightly soluble	Soluble	Slightly soluble	Not very soluble	Unknown
Density	Very light	Light	Heavier than water	Heavier than water	Very heavy
Melting Point (°C)	-219.6	-101.5	-7.2	113.7	~300
Boiling Point (°C)	-188.1	-34.0	58.8	184.3	~350
Toxicity	Highly poisonous	Highly poisonous	Highly poisonous	Highly poisonous	Highly poisonous
Bond Energy (kJ/mol)	159	243	193	151	Low (Estimated ~120)

Trends in the Table 6.1

Physical State: Changes from gases to liquid to solids as you move down.

Intermolecular Forces: Get stronger due to increasing molecular size.

Colour: Becomes darker down the group.

Density: Increases with atomic mass.

Oxidising Strength: Decreases from fluorine to iodine.

Bond Energy: Decreases due to lone pair repulsion and larger atomic sizes.

Melting/Boiling Points: Increase with stronger intermolecular forces.

Odour & Toxicity: All halogens have strong smells and are highly poisonous.

2. Chemical Properties of Halogens

a. Reactions with Metals

Halogens react with metals to form metal halides, which are ionic compounds.

Example: Sodium + Chlorine → Sodium chloride (table salt).

$Na + ½Cl_2 \rightarrow NaCl$

These compounds have crystal lattice structures.

b. Reactions with non-Metals

Halogens combine with other non-metals to form molecular compounds.

Example: Bromine + Fluorine → Bromine Trifluoride (BrF_3).

$\frac{1}{2}Br_2 + 3F_2 \rightarrow BrF_3$

c. Reactions with hydrogen

Halogens react with hydrogen to form hydrogen halides (e.g., HCl, HF).

Fluorine reacts explosively, while iodine reacts very slowly.

Hydrogen halides dissolve in water to make acids:

HF: Weak acid.

HCl, HBr, HI: Strong acids.

d. Displacement reactions

A more reactive halogen can replace a less reactive one in a compound.

Example: Chlorine replaces bromine in potassium bromide:

$Cl_2 + 2KBr \rightarrow 2KCl + Br_2$.

Reactivity trend: $F_2 > Cl_2 > Br_2 > I_2$.

e. Precipitation reactions

Halogens react with silver ions to form colourful precipitates, helping identify halides:

Chloride (Cl^-): White precipitate (AgCl).

Bromide (Br^-): Pale yellow precipitate (AgBr).

Iodide (I^-): Bright yellow precipitate (AgI).

f. Oxidising Strength

Halogens are strong oxidising agents because they gain electrons easily.

Fluorine is the strongest, followed by chlorine, bromine, and iodine.

Halogens as Strong Oxidising Agents (Simplified Table)

Halogen Pair (X2/2X−)	Standard Reduction Potential (E°)	Oxidizing Strength
$\frac{F_2}{2F^-}$	+2.87	Strongest oxidising agent
$\frac{Cl_2}{2Cl^-}$	+1.36	Can oxidise I^- to I_2

Halogen Pair (X2/2X−)	Standard Reduction Potential (E°)	Oxidizing Strength
$\frac{Br_2}{2Br^-}$	+1.09	Weaker than Cl_2
$\frac{I_2}{2I^-}$	+0.54	Weakest oxidising agent

The higher the **reduction potential (E°)**, the stronger the oxidising power. Fluorine (F_2) is the strongest oxidising agent, while iodine (I_2) is the weakest. Example: Cl_2 can oxidise I^- in solution because it has a higher E° value.

g. Electronegativity and Bonding

Halogens are highly electronegative, meaning they attract electrons easily.

Fluorine is the most electronegative and does not form positive oxidation states.

Heavier halogens (chlorine, bromine, iodine) can show multiple positive oxidation states (e.g., +1, +3, +5, +7).

Remember

1. Halogens react with metals, non-metals, and hydrogen to form important compounds.
2. They are very reactive and strong oxidisers, with fluorine being the strongest.
3. Precipitation reactions with silver ions help identify halides based on their unique colours.
4. Reactivity decreases as you move down the group:

   Fluorine > Chlorine > Bromine > Iodine.

Activity 6.1 Analysing Physical States of Halogens

Table 6.2: Data on physical properties of halogens.

Halogen	Physical state	Boiling point/ °C	Melting point/ °C	Bond energy/ kjmol^{-1}
Fluorine (F)	Gas	-188.1	-219.6	-158
Chlorine (Cl)	Gas	-34.4	-101.5	243
Bromine (Br)	Liquid	58.8	-7.2	193
Iodine (I)	Solid	184.3	113.7	151

Steps

1. **a.** In small groups examine Table 6.2 with data on halogens' physical properties.

   b. Observe the variations in physical states, melting/boiling points, and bond for fluorine, chlorine, bromine, and iodine.

   c. Record how the properties change across the group, focusing on trends.

2. **a.** Discuss why fluorine and chlorine are gases, bromine is a liquid, and iodine is a solid at room temperature. (Relate this to the strength of intermolecular forces (Van der Waals forces) increasing down the group).

   b. Discuss why melting and boiling points increase down the group (from fluorine to iodine). [Connect this trend to the size of the molecules and the strength of intermolecular forces].

   c. Compare the bond energy values for F_2, Cl_2, Br_2, and I_2.

   d. Discuss why fluorine has low bond energy (due to strong repulsion between lone pairs) and why bond energies generally decrease as you move down the group.

3. **a.** Select a property to graph (e.g., melting points, boiling points, or bond energies).

   Plot the graph:
   Horizontal axis: Halogens (F, Cl, Br, I).
   Vertical axis: Selected property (e.g., melting point in °C).

   b. Label the graph clearly and identify trends (e.g., increasing boiling points or decreasing bond energies).

Activity 6.2 Exploring Halogen Properties

1. **a.** Write the electron configuration of halogens (F, Cl, Br, I, At).

   b. Explain where halogens are located in the periodic table and why their outermost electron shell is important for their chemical behaviour.

   - **c.** Discuss how the halogens' electron configurations (with seven valence electrons) allow them to achieve stable configurations through gaining or sharing electrons.
   - **d.** Identify the typical oxidation states for each halogen and how these states vary based on the type of reaction (e.g., -1 in ionic compounds and positive states in covalent interactions).
   - **e.** Highlight common traits like high electronegativity, strong oxidising abilities, and trends in reactivity as you move down the group.

2. **a.** Research and compile data on standard electrode potential values (E°) for fluorine, chlorine, bromine, iodine, and astatine.

   - **b.** Discuss how higher electrode potential values indicate stronger oxidising agents and greater reactivity.
   - **c.** Rank the halogens by their reactivity, based on electrode potentials and other chemical trends, from most to least reactive.

Activity 6.3 Halogen Reactions with Water and Alkalis

1. Watch video on halogen reactions, using the link https://youtu.be/JG8HL9BXN40

2. Discuss and record findings on the:
   - **a.** Reagents
   - **b.** Conditions
   - **c.** products formed
   - **d.** When the halogens react with:
      - i. Water
      - ii. Alkalis

3. Write balanced chemical equations to show how the reactions occur.

4. Write a comprehensive report on halogen reactions following these guidelines:
   - **a.** Introduce halogen reactions with water and alkalis.
   - **b.** Reagents, Conditions, analyse each reaction.
   - **c.** Provide equations for each reaction.

   d. Explain industrial, medical, and environmental applications of halogen reactions.

5. Summarise key points and present it to the class for discussion and feedback.

REACTION OF HALIDE SALTS, DIFFERENCES IN ACID STRENGTH OF THE HYDROGEN HALIDES AND THE USES OF THE HALOGENS

Reactions of Halide Salts

General Behaviour of Halide Salts

1. Halogens (like fluorine, chlorine, bromine, and iodine) are good oxidizing agents, meaning they can gain electrons to form halide ions (F^-, Cl^-, Br^-, I^-).
2. Halide ions (e.g., I^-) can lose electrons to form their halogen gases (e.g., I_2).
   Example: $2I^- \rightarrow I_2 + 2e^-$

Reducing Power of Halide Ions

Iodide (I^-) is the strongest reducing agent, meaning it easily loses electrons.

Fluoride (F^-) is the weakest reducing agent because it holds its electrons tightly.

The **reducing power increases down the group** because larger ions (like I^-) lose electrons more easily.

Reactions with Concentrated Sulfuric Acid (H_2SO_4)

Refer to **Table 6.3** showing how halides react differently based on their reducing power:

Table 6.3: Reducing power of Halides

Halide	Reaction with H_2SO_4	Products
Sodium Fluoride (NaF)	Forms HF (hydrofluoric acid). No further reaction.	$NaF + H_2SO_4 \rightarrow HF + NaHSO_4$
Sodium Chloride (NaCl)	Forms HCl (hydrochloric acid). No further reaction.	$NaCl + H_2SO_4 \rightarrow HCl + NaHSO_4$

Halide	Reaction with H_2SO_4	Products
Sodium Bromide (NaBr)	Forms HBr, which reduces H_2SO_4	$H_2SO_4 + 2H^+ + 2Br^- \rightarrow Br_2 + SO_2 + 2H_2O$
Sodium Iodide (NaI)	Forms HI, which reduces H_2SO_4	$H_2SO_4 + 6H^+ + 6I^- \rightarrow I_2 + S + 4H_2O$

Note

- **Fluoride and Chloride**: Weak reducing agents, only form HF and HCl.
- **Bromide**: Moderate reducing agent, forms bromine gas (Br_2) and sulphur dioxide (SO_2).
- **Iodide**: Strongest reducing agent, forms iodine gas (I_2) and multiple sulphur (SO_2, S, H_2S).

Acid Strengths of Hydrogen Halides

Hydrogen halides are molecules made of hydrogen and halogens: HF (hydrofluoric acid), HCl (hydrochloric acid), HBr (hydrobromic acid), HI (hydroiodic acid)

Acid Strength of Hydrogen Halides

Acid strength depends on how easily the hydrogen halide can break apart (dissociate) in water to release hydrogen ions (H^+).

Order of Acid Strength: HF << HCl < HBr < HI

HF is a **weak acid** because it does not dissociate easily.

HCl, HBr, HI are **strong acids** because they dissociate easily in water.

a. **Relative Bond Strengths of Hydrogen halides (HX)**

The strength of the **H–X bond** affects how easily it breaks:

Strong bonds (like H–F) make it hard to dissociate weak acid.

Weak bonds (like H–I) make it easy to dissociate strong acid.

Refer to **Table 6.4** for relative bond strengths of hydrogen halides.

Table 6.4: Relative Bond Strengths of Hydrogen halides

Hydrogen Halide (H–X)	Bond Strength (kJ/mol)	Acid Strength
H–F	565	Weak acid (HF)
H–Cl	432	Strong acid (HCl)

Hydrogen Halide (H–X)	Bond Strength (kJ/mol)	Acid Strength
H–Br	366	Stronger acid (HBr)
H–I	298	Strongest acid (HI)

As you move **down the group**

Bond strength decreases.

Molecules break apart (dissociate) more easily.

Acids become stronger.

Remember

HF is weak because its bond is very strong and hard to break.

Acid strength increases from HF to HI as bond strength decreases.

HCl, HBr, and HI are strong acids because their bonds are weaker and break easily.

b. Ka Values of Hydrogen Halides

K_a is a number that shows how easily an acid splits into its ions in water.

K_a tells us how strong an acid is.

The bigger the K_a value, the easier the acid breaks apart and the stronger it is.

For hydrogen halides (HX):

$HX(aq) \rightarrow H^+(aq) + X^-(aq)$

Larger K_a values = Stronger acids (dissociate more easily).

Table 6.5 give shows how K_a values relate to the strength of hydrogen halides

Table 6.5: K_a Values of Hydrogen Halides

Hydrogen Halide (HX)	K_a Value	Acid Strength
H–F (Hydrofluoric Acid)	6.6×10^{-4}	Weak acid
H–Cl (Hydrochloric Acid)	1.3×10^{6}	Strong acid
H–Br (Hydrobromic Acid)	5.0×10^{9}	Stronger acid
H–I (Hydroiodic Acid)	5.0×10^{10}	Strongest acid

Reasons why K_a Increase Down the Group

Atomic Size

As the halogen gets larger (from fluorine to iodine), the bond with hydrogen becomes longer and therefore weaker.

Weaker bonds are easier to break, making acids stronger.

Electronegativity

Fluorine is very electronegative and holds electrons tightly, forming a strong H–F bond low K_a (weak acid).

Iodine is less electronegative, forming a weaker H–I bond → high K_a (strong acid).

c. Thermal Stability of Halogen Halides

Thermal stability is how well a compound resists breaking apart (decomposing) when heated.

For halogen halides (HX), it depends on how strong the bond is between hydrogen (H) and the halogen (X).

Trend in Thermal Stability

Thermal stability decreases as you move down the group.

- HF (hydrogen fluoride): Strongest bond → Most stable.
- HI (hydrogen iodide): Weakest bond → Least stable.

Bond Energy and Thermal Stability

See **Table 6.6** for bond energy and thermal stability

Table 6.6: Bond Energy and Thermal Stability

Hydrogen Halide (HX)	Bond Energy (kJ/mol)	Thermal Stability
H–F	565	Most stable (requires more heat to break).
H–Cl	432	Stable.
H–Br	366	Less stable.
H–I	298	Least stable (breaks easily with heat).

Trend in Thermal Stability

Bond Strength

The H–X bond becomes weaker down the group (from HF to HI).

Weaker bonds are easier to break when heated.

Size of Halogen Atom

Larger halogen atoms (like iodine) form longer, weaker bonds with hydrogen.

Everyday Uses of Halogens

Halogens are highly useful in everyday life because of their unique chemical properties. They help in **cleaning, water purification, refrigeration, and sanitation**, improving the quality of life. Here are some key applications:

1. **Purification of Water**

   Chlorine is used in water treatment plants to kill harmful microorganisms and water safe for drinking.

   Chemical Reaction: $Cl_2(g) + H_2O(l) \rightarrow HCl(aq) + HClO(aq)$

   Hypochlorous acid (HClO) is the main disinfectant. It breaks bacteria by damaging heir cell walls and essential components, stopping them from reproducing.

   Chlorine is also used in swimming pools and healthcare settings for disinfection. Refer Figure 6.1 for the water purification steps.

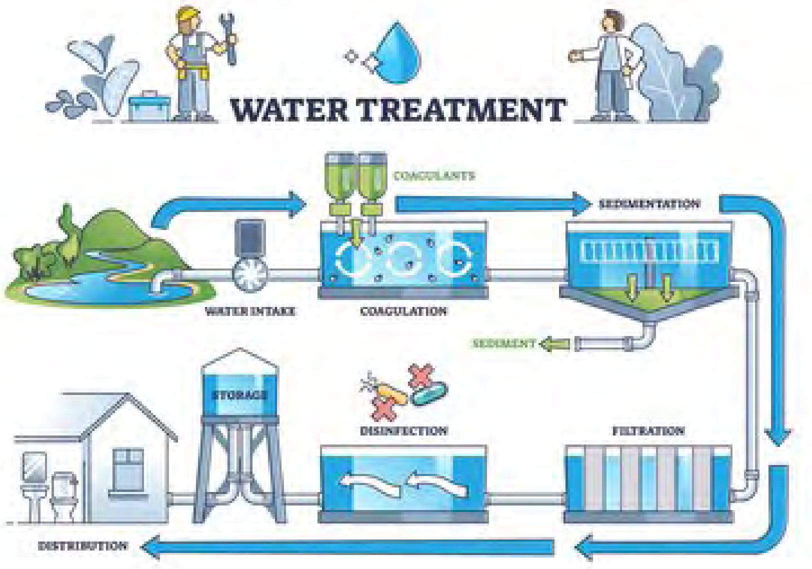


Figure 6.1: Chlorination in water purification

2. **Disinfecting Toilets**

   Chlorine and iodine are strong oxidising agents that kill bacteria and viruses.

   Household bleach (sodium hypochlorite, NaClO is commonly used for toilets.

   When dissolved in water, it forms hypochlorous acid, which removes stains and germs effectively.

Figure 6.2: Disinfecting hospital-bed.

3. **Bleaching**

   Chlorine is used in bleach to remove colours from dyes and stains.

   It oxidises coloured molecules, making them colourless.

4. **Halogenated Hydrocarbons**

   These are compounds containing carbon and one or more halogen atoms (e.g., fluorine, chlorine).

   Uses

   Solvents: Halogenated hydrocarbons like chloroform and dichloromethane fats, oils, and waxes.

   Refrigerants: Compounds like CFCs and HFCs are used in refrigeration because of their ability to absorb and release heat efficiently.

   Aerosols: Halogenated hydrocarbons are used as propellants in spray cans due their stability and low reactivity.

Note

Many halogenated hydrocarbons are now banned due to their role in **global warming** and **ozone layer depletion**.

Activity 6.4 Understanding Reducing Power of Halides

Materials needed: Worksheet containing the key questions

Steps

1. What do you think reducing power means?

   Have you heard the terms chloride, bromide, and iodide? Where might you find these in daily life?
2. How do halides differ in their reducing power?

   Provide the trend.
3. What factors influence reducing power?
4. Present your answers to the class.
5. Discuss as a class:
   - **a.** Why iodide has the greatest reducing power.
   - **b.** How the size and structure of the ions affect their reducing ability.

Activity 6.5 Comparing reducing power of halides

Steps

1. **a.** In small groups, discuss your ideas about the reducing power of halides (Cl^-, Br^-, and I^-).
   - **b.** Which halide do you think is the strongest/weakest reducing agent? Why?
   - **c.** Can you think of any real-life applications of halides' reactions?
2. Write down the differences in reducing power of halides and includes examples.
3. Write and complete the chemical equations for the reactions of halide salts (chlorides, bromides, and iodides) with concentrated tetraoxosulphate (VI) acid.

   For each reaction: balance the equation, identify the products formed and explain the role of the halide ion as a reducing agent and describe any observable changes during the reaction.

   $Cl^- + H_2SO_4 \rightarrow Cl_2 + SO_2 + H_2O$

   $Br^- + H_2SO_4 \rightarrow$

   $I^- + H_2SO_4 \rightarrow$

Activity 6.6 Exploring Halogens

Materials needed: Textbooks, handouts, or supervised online searches

Steps

1. In small groups, investigate the following topics on halogens:
    a. Acid strengths of halogens
    b. Thermal stability of halogen compounds
    c. Everyday uses of halogens
2. **a.** What is the trend in acid strength or thermal stability among halogens?
    b. How are halogens used in daily life?
3. Create visual aids (e.g., tables, graphs, concept maps) to illustrate their findings.
4. Present your findings to the class using your visual aids.
5. Reflect on the following: What did you find most interesting about halogens?

EXTENDED READING

To deepen your understanding of the physical and chemical properties of halogens (Group 17 elements), consider the following resources:

• https://www.britannica.com/science/halogen	
• https://byjus.com/jee/halogens/	
• https://www.thoughtco.com/halogen-elements-and-properties-606650	

- Chemguide – Redox Reactions Involving Halide Ions and Sulphuric Acid
- Chemistry LibreTexts – The Acidity of the Hydrogen Halides
- Chemistry LibreTexts – The Halogens, Group 7A
- ChemicalAid – The Elements of Group 17 (The Halogens).

REVIEW QUESTIONS

Review Questions 6.1

1. What is the physical state of bromine at room temperature?
2. Name the halogen with the highest melting point.
3. Which halogen is the strongest oxidising agent?
4. What type of bond do halogens form when reacting with metals?
5. Why do halogens become less reactive as you move down the group?
6. Explain why fluorine has a lower bond energy than chlorine.
7. Why do halogens have low melting and boiling points compared to other groups?
8. Predict what happens when chlorine gas reacts with potassium bromide solution. Write the equation.
9. Identify which halide ion (Cl^-, Br^-, I^-) forms a yellow precipitate when reacted with silver nitrate.
10. What would you observe if iodine is added to a solution of sodium chloride? Explain why.
11. Compare the trend in melting points with the trend in oxidising strength of halogens.
12. Analyse why HF is a weak acid while HCl, HBr, and HI are strong acids.
13. Why does chlorine exhibit multiple oxidation states (e.g., +1, +3, +5, +7) but fluorine does not?

Review Questions 6.2

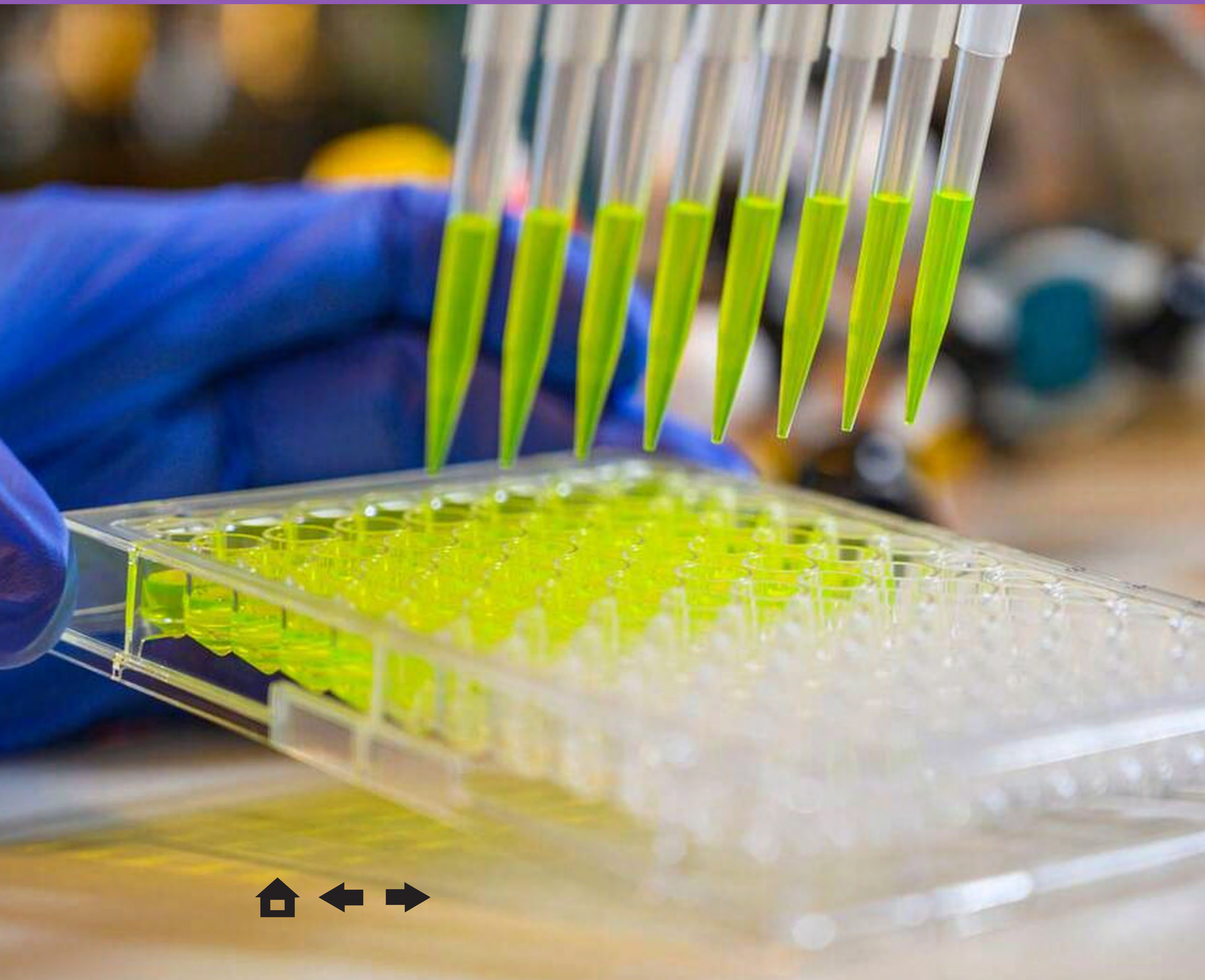
1. Write the balanced equation for the reaction of bromide salts with tetraoxosulphate (VI) acid.
2. Name two common uses of chlorine in daily life.
3. Explain the trend in acid strength of hydrogen halides (HF, HCl, HBr, HI).
4. Describe the observation when iodide salts react with concentrated tetraoxosulphate (VI) acid.

5. Compare the reactions of chloride, bromide, and iodide salts with tetraoxosulphate (VI) acid in terms of reducing power.
6. How does the variation in acid strength of hydrogen halides relate to their structures?
7. Propose an experiment to demonstrate the differences in reactivity of compounds and relate the results to their uses.
8. Evaluate how the acid strength of hydrogen halides affects their industrial uses.

SECTION

7

STRUCTURE, CHEMICAL BONDING AND PROPERTIES OF MOLECULAR COMPOUNDS



BONDING

Systematic Chemistry of the Elements

INTRODUCTION

In this section, you will learn about **hybridisation**, how it shapes molecules, and the types of bonds in them. You will explore how to predict molecule shapes (e.g., straight, bent, or triangular) and measure bond angles. You will also understand the two types of bonds, sigma and pi, and explain how atoms combine their "building blocks" (orbitals) to create different shapes and structures.

KEY IDEAS

- **Valence Shell Electron Pair Repulsion (VSEPR)** Theory is a model used to predict the geometry of molecules based on repulsion between electron pairs.
- **Molecular geometry** is the 3D arrangement of atoms in a molecule.
- **Bond angle**: is the angle formed between two covalent bonds at the atom where they meet.
- **Sigma bond** is a bond formed by the head-on overlap of orbitals, allowing free rotation.
- **Pi (π) bond** is a bond formed by the side-by-side overlap of p orbitals, restricting rotation.
- **Hybridisation** is the mixing of atomic orbitals to form new hybrid orbitals of equivalent energy.

SHAPE AND BOND ANGLES OF MOLECULAR COMPOUNDS AND DIFFERENCES BETWEEN SIGMA AND PI BONDS

Shape and bond angles for molecular compounds

Electronegativity is a way to describe how strongly an atom can pull electrons toward itself when it bonds with another atom. It helps us understand why some atoms share electrons equally while others do not.

Linus Pauling came up with this idea and created a scale called the **Pauling scale** to measure how "electron-hungry" atoms are. The lowest value is 0.7 (like caesium, which does not pull electrons much). The highest value is 4.0 (fluorine, which pulls electrons very strongly).

These numbers do not have units—they just compare how atoms behave.

Electronegativity increases across a row or period because atoms get smaller (due to increased number of protons, which exert a greater electrostatic force of attraction on the surrounding electrons) and pull electrons more strongly.

Electronegativity decreases down a column (group) because atoms get bigger (due to the addition of more electron shells), and their outer electrons are farther from the centre, making it harder to pull more electrons.

Electronegativity helps us predict how atoms bond and how molecules behave in chemical reactions.

Bond Polarity

The difference in electronegativity between two bonded atoms determines the bond type and its polarity:

1. **Non-Polar Covalent Bonds**
   a. Electronegativity Difference: Less than 0.5
   b. Electron Sharing: Equal sharing of electrons.
   c. Charge Distribution: No partial charges, bond is neutral.

   Example: Hydrogen molecule (H_2): Both hydrogen atoms have the same electronegativity, resulting in a non-polar bond.

2. **Polar Covalent Bonds**
    a. Electronegativity Difference: Between 0.5 and 1.7
    b. Electron Sharing: Unequal sharing of electrons.
    c. Charge Distribution:

    The more electronegative atom becomes partially negative (δ−). The electronegative atom becomes partially positive (δ+).

    Example: Water (H_2O): Oxygen is more electronegative than hydrogen, causing a dipole with oxygen as δ− and hydrogen as δ+.

3. **Ionic Bonds**
    a. Electronegativity Difference: Greater than 1.7
    b. Electron Sharing: Electrons are transferred, not shared.
    c. Charge Distribution:

    The more electronegative atom gains electrons, becoming negatively charged (anion).

    The less electronegative atom loses electrons, becoming positively charged (cation).

    Example: Sodium chloride (NaCl): Sodium (Na) loses an electron to become Na^+, and chlorine (Cl) gains an electron to become Cl^-.

Dipole Moments

A **dipole moment** happens in a polar bond when there is a separation of positive and negative charges. It shows how much the electrons are pulled toward one atom.

An arrow points from the positive side to the negative side of the bond to show the direction of the electron pull. See Figure 7.1.

Dipole moments are measured in **Debye units (D)**. It tells us how uneven the charge distribution is in a bond.

Figure 7.1: Dipole Moment

Examples

1. **Hydrogen Chloride (HCl)**

   In hydrogen chloride (HCl):

   Electronegativity Difference: Chlorine (3.0) is more electronegative than hydrogen (2.1), creating a polar covalent bond.

   Chlorine pulls the shared electrons closer, becoming partially negative (δ^-). Hydrogen becomes partially positive (δ^+).

   A dipole moment is created, showing the separation of charges.

   An arrow points from hydrogen (positive) to chlorine (negative), showing the electron pull direction. See Figure 7.2. This dipole moment makes HCl a polar molecule.

Figure 7.2: Polarity of HCl

2. **Methane (CH_4)**

   In methane (CH_4):

   The difference between carbon (2.5) and hydrogen (2.1) is 0.4, forming non-polar covalent bonds.

   The molecule has a tetrahedral shape, which is symmetrical.

   The dipole moments of the C-H bonds cancel out because they point in opposite directions. See Figure 7.3.

   As a result, methane has no net dipole moment, making it a non-polar molecule.

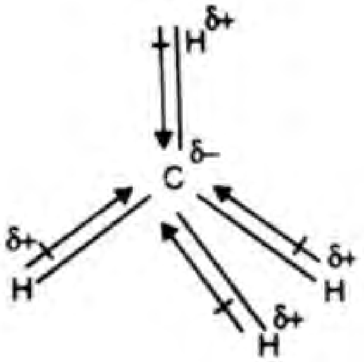


Figure 7.3: Polarity of CH_4

3. **Carbon Dioxide (CO_2)**

   In carbon dioxide (CO_2):

   The difference between carbon (2.5) and oxygen (3.5) is 1.0, forming polar C=O bonds.

   CO_2 has a linear shape, which is symmetrical.

   The dipole moments of the two C=O bonds point in opposite directions and cancel each other out. As a result, CO_2 has no net dipole moment, making it a non-polar molecule. See Figure 7.4.

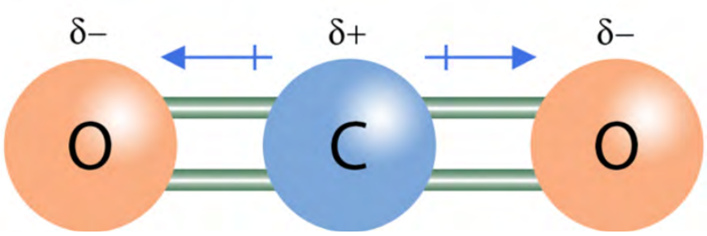


Figure 7.4: Polarity of CO_2

Predicting Molecular Polarity with VSEPR Theory

The **Valence-Shell Electron-Pair Repulsion (VSEPR) theory** helps predict the shape of a molecule by considering how electron pairs repel each other. The goal is to arrange electron pairs around the central atom as far apart as possible to minimize repulsion.

Steps to Use VSEPR Theory

1. **Draw the Lewis structure**

   Show all bonding and lone pairs of electrons.

2. **Count electron pairs around the central atom**

   Include both bonding pairs (BP) and lone pairs (LP).

3. **Account for repulsions**

   Electron repulsions depend on the type of pairs:

   Lone Pair–Lone Pair (LP-LP): Strongest repulsion.

   Lone Pair–Bonding Pair (LP-BP): Medium repulsion.

   Bonding Pair–Bonding Pair (BP-BP): Weakest repulsion.

   Lone pairs take up more space than bonding pairs, so they push other pairs apart.

4. **Adjust Molecular Shape**

   Modify the shape based on repulsions to minimize electron pair interactions.

   For example: Lone pairs cause angles to shrink compared to a shape with only bonding pairs.

   Double bonds occupy more space than single bonds.

1. Species with Four Negative Charge Centres

Example: **Methane (CH_4)**

Structure: Methane (CH_4) consists of one carbon atom bonded to four hydrogen atoms.

Carbon has four valence electrons, and each hydrogen atom has one electron. Together, they form four single covalent bonds.

Electron Groups: The carbon atom has four electron groups (four bonding pairs) around it.

Shape: According to VSEPR theory, these four electron groups repel each equally and arrange themselves in a tetrahedral shape to minimise repulsion.

The bond angles in methane are approximately **109.5°**.

In a 3D structure, the hydrogen atoms form a tetrahedron around the central carbon atom.

See Figure 7.5

Polarity: Methane is **non-polar** because

The bond polarities cancel out due to its symmetrical tetrahedral shape.

This results in no net dipole moment.

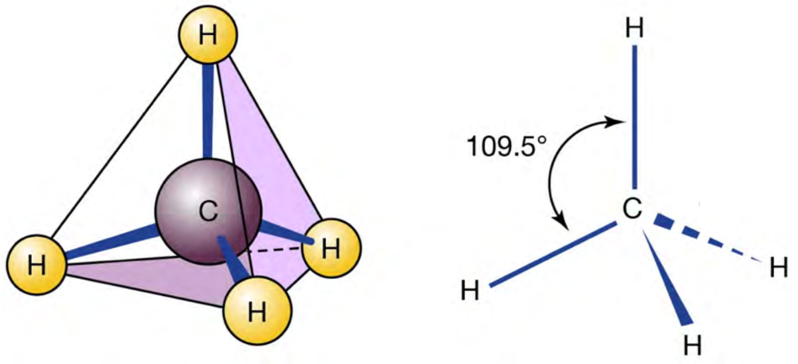


Figure 7.5: Tetrahedral shape of CH_4.

2. Species with 3 Negative Charge Centres

Boron Trifluoride (BF3)

Structure: Boron trifluoride (BF_3) has one boron atom bonded to three fluorine atoms.

Boron has three valence electrons, and each fluorine atom has seven valence electrons.

They form three single covalent bonds.

Electron Groups: The boron atom has three electron groups (three bonding pairs) around it.

Shape: According to **VSEPR theory**, the three electron groups repel each other equally and arrange themselves in a trigonal planar shape to minimise repulsion.

The bond angles are approximately **120°**.

In 3D, the fluorine atoms form an equilateral triangle around the central boron atom.

Polarity: Boron trifluoride is **non-polar** because:

The bond polarities cancel out due to its symmetrical trigonal planar shape. This results in no net dipole moment. See Figure 7.6.

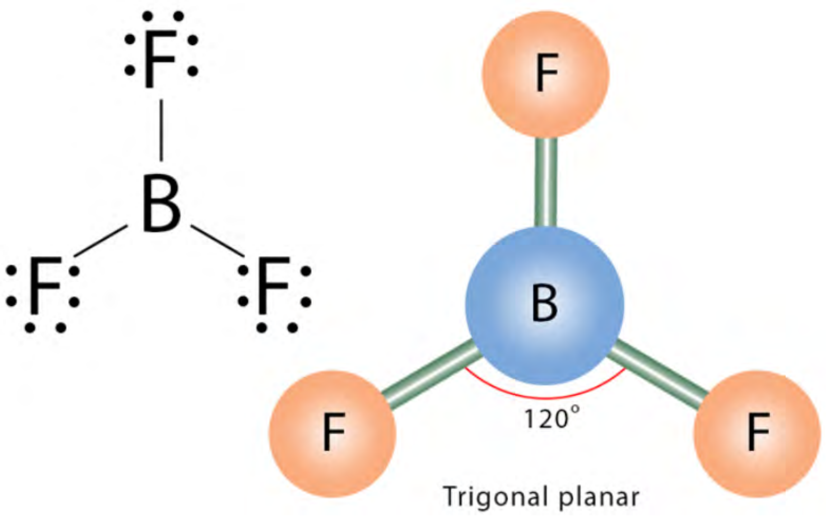


Figure 7.6: Trigonal planar shape of BF_3

3. Species with 2 Negative Charge Centres

Carbon Dioxide (CO2)

Structure: Carbon dioxide (CO_2) has one carbon atom double-bonded to two atoms.

Carbon has **four valence electrons**, and each oxygen has **six valence electrons**.

This forms **two double bonds**.

Electron Groups: The carbon atom has two electron groups (two double bonds) around it.

Shape: According to **VSEPR theory**, the two electron groups repel each other and arrange in a **linear shape** to minimise repulsion. The bond angle is **180°**.

In 3D, the oxygen atoms are aligned in a straight line on opposite sides of the carbon atom. See Figure 7.7.

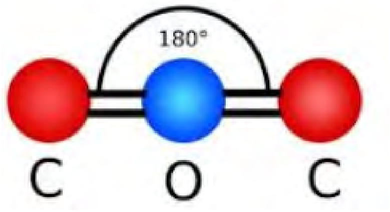


Figure 7.7: Linear shape of CO_2.

Predicting Molecular Polarity

To know if a molecule is **polar**, consider:

- **Bond Polarity**: Check if the bonds are polar (unequal sharing of electrons).
- **Shape of the Molecule**: Look at the molecule's symmetry.
- **Symmetrical Molecules**: If bond dipoles cancel out, the molecule is **non-polar**.
- **Asymmetrical Molecules**: If bond dipoles don't cancel, the molecule is **polar**.

Example

1. **Water (H_2O)**

   Polar Bonds: The O-H bonds are polar because oxygen is more electronegative than hydrogen.

Oxygen has a partial negative charge (δ−), and hydrogen has a partial positive charge (δ+).

Shape: Water has a **bent shape** due to two lone pairs on the oxygen atom.

Bond angle: **104.5°**.

Polarity: The bond dipoles do not cancel because of the bent shape, making water polar. See Figure 7.8. This polarity is responsible for many of water's unique properties, such as its high boiling point and its ability to dissolve many substances.

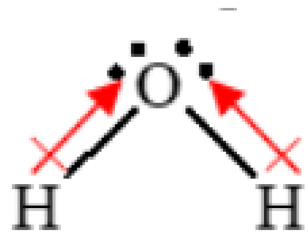


Figure 7.8: Bent shape of H_2O

2. **Carbon Tetrachloride (CCl_4)**

Polar Bonds: The C-Cl bonds are polar because chlorine is more electronegative than carbon.

Shape: CCl_4 has a tetrahedral shape with bond angles of **109.5°**.

Polarity: Even though the bonds are polar, the molecule is symmetrical, so the dipoles cancel out. See Figure 7.9.

Result: CCl_4 is a non-polar molecule.

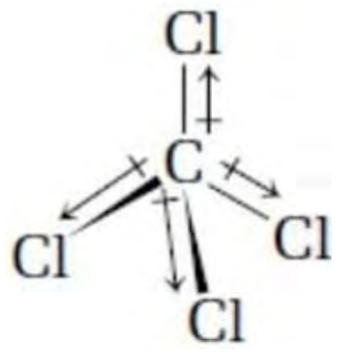


Figure 7.9: Tetrahedral shape of CCl_4

3. **Ammonia (NH_3)**

The N-H bonds are polar because nitrogen (electronegativity 3.04) is more electronegative than hydrogen (2.20). The molecule has a trigonal pyramidal shape with bond angles of about 107°, caused by the lone pair on nitrogen.

Due to its shape, the bond dipoles do not cancel, making ammonia polar with a net dipole moment. This polarity allows ammonia to dissolve well in water and form hydrogen bonds

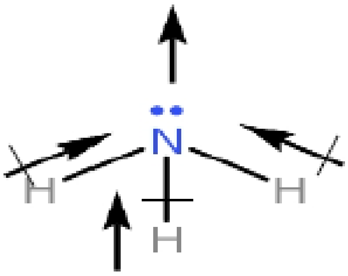


Figure 7.10: Trigonal pyramidal shape of NH_3

4. **Sulphur Dioxide (SO2)**

   The bonds in SO_2 are polar because oxygen pulls electrons more strongly than sulphur. The molecule has a bent shape with bond angles around 119°, caused by a lone pair of electrons on sulphur. Because of this shape, the bond pulls don't cancel out, making SO_2 a polar molecule. This polarity affects how SO_2 behaves and reacts with other substances.

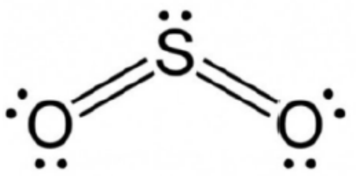


Figure 7.11: *Bent shape of* SO_2

Sigma Bonds (σ- Bonds)

Sigma bonds are formed by the head-on (axial) overlap of atomic orbitals, resulting in electron density concentrated along the line connecting the nuclei of the bonded atoms. This type of overlap allows for maximum orbital interaction, making σ-bonds stronger than π\piπ-bonds.

Types of Overlap in Sigma Bonds

1. s–s Overlap: Two s-orbitals overlap directly.

   Example: H_2 molecule, where two hydrogen atoms form a sigma bond.

2. s−p Overlap: An s-orbital overlaps with a p-orbital.

   Example: In CH_4, the carbon atom's sp^3-hybrid orbitals overlap with the s-orbitals hydrogen atoms.
3. p−p Overlap: Two p-orbitals overlap along their axes.

   Example: In F_2, two fluorine atoms form a sigma bond through p−p overlap.

Sigma bonds are always present in single bonds and in the first bond of double or triple bonds.

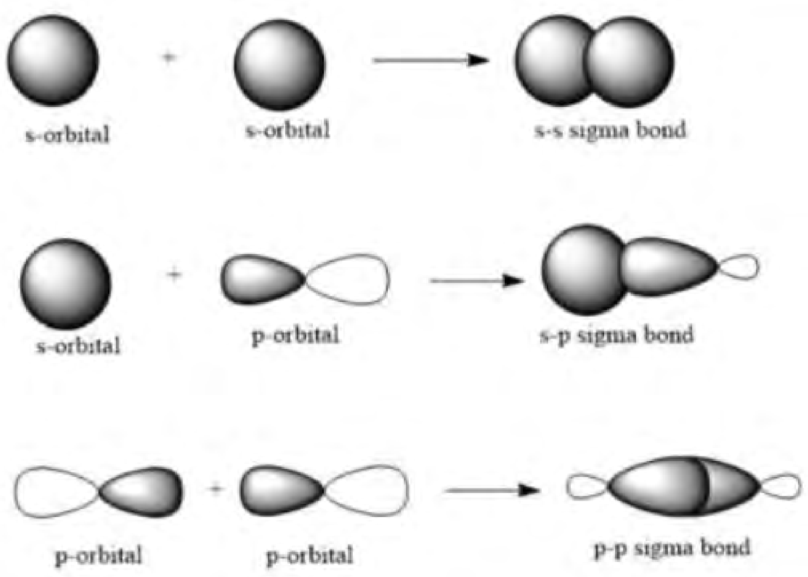


Figure 7.12: Three types of sigma bonds between s-s, s-p, and p-p atomic orbitals

Characteristics of Sigma Bonds

1. Sigma (σ) bonds are formed by the direct, head-on (axial) overlap of atomic orbitals.
2. They are always the first bond formed between two atoms and are found in all single bonds.
3. Sigma bonds are generally stronger than pi (π\piπ) bonds due to the greater overlap of orbitals.
4. The electron density in a sigma bond is concentrated along the internuclear axis (the line connecting the nuclei of the bonded atoms).
5. Sigma bonds have cylindrical symmetry around the bond axis, meaning the electron density is evenly distributed in a circular fashion.
6. Atoms connected by a sigma bond can freely rotate around the bond axis without breaking the bond.

7. They can be formed through s–s, s–p, or p–p orbital overlaps, and also through hybrid orbital overlaps (e.g., sp^3–s).

Pi Bonds (π Bonds)

A **pi (π) bond** is a type of covalent bond formed by the sideways (lateral) overlap of atomic orbitals. It typically occurs in addition to a sigma (σ) bond in double or triple bonds.

Characteristics of Pi Bonds

1. Formed by the sideways overlap of unhybridized p-orbitals.
2. Found in double bonds (1 π-bond and 1 σ-bond) and triple bonds (2 π-bonds and 1 σ-bond).
3. The electron density in a π-bond is concentrated above and below the bond axis, not along it.
4. Weaker than sigma bonds due to the less effective overlap of orbitals.
5. π-bonds restrict rotation around the bond axis because breaking the sideways overlap requires energy.

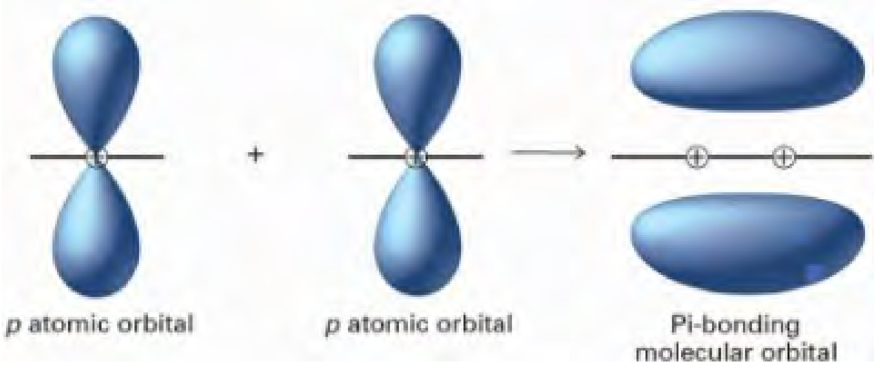


Figure 7.13: pi bond between p-p atomic orbitals

Differences Between Sigma (σ) and Pi (π) Bonds

Table 7.1: Differences Between σ and π Bonds

Sigma (σ) Bond	Pi (π) Bond
Formed by the head-on (axial) overlap of atomic orbitals.	Formed by the sideways (lateral) overlap of p-orbitals.
Electron density is concentrated along the internuclear axis.	Electron density is above and below the internuclear axis.

Sigma (σ) Bond	Pi (π) Bond
Stronger due to greater orbital overlap.	Weaker due to less effective orbital overlap.
Allows free rotation of bonded atoms around the bond axis.	Restricts rotation due to the rigidity of the bond.
Found in single, double, and triple bonds (always one σ).	Found only in double and triple bonds (accompanies a σ).
Can involve s–s, s–p, or p–p overlaps.	Involves p–p overlaps only.

Activity 7.1 Exploring Molecular Shapes, Polarity, and Bonding

Materials needed: A table of electronegativity values, clay or playdough, toothpicks a list of element pairs (e.g., H-F, C-H, Na-Cl, CH_4, NH_3, H_2O, CO_2, SO_2)

Steps

1. Calculate electronegativity differences for each pair.

H-F, C-H, Na-Cl, CH_4, NH_3, H_2O

Classify the bonds as: Nonpolar covalent, polar covalent and ionic

2. **a.** Use VSEPR theory to predict shapes and bond angles.

H-F, C-H, Na-Cl, CH_4, NH_3, H_2O

b. Build 3D models representing molecular shapes (linear, trigonal planar, tetrahedral).

3. Predicting bond polarity

- **a.** Identify bond polarities based on electronegativity differences: $CH_{4,}$ CO_2, SO_2
- **b.** Predict molecular shapes using VSEPR theory.
- **c.** Determine if the molecule is polar or nonpolar.

4. Analyse how molecular shape influences overall polarity.

Give examples of real-life implications

Activity 7.2 Exploring Molecular Orbitals and Bond Formation

Materials needed: Diagrams of s- and p-orbitals.

Steps

1. Brainstorm what is meant by sigma and pi bonds. Make notes for each of the definitions.
2. Draw diagrams to illustrate the overlaps. Make sure to included s-p, p-p and and s-s.
3. Use examples to explain sigma bonds (e.g., single bonds in H_2) and pi bonds (e.g., double bond in C_2H_4). Consider why each bond would apply to each scenario that you choose.
4. Compare sigma and pi bonds, ensure that you include mention of how the electrons are contained within these bond structures.

HYBRIDISATION, STRUCTURE, AND BONDING IN ORGANIC COMPOUNDS

Hybridisation is a key concept in organic chemistry that explains the bonding and geometry of carbon-containing compounds. Carbon atoms form stable covalent bonds with other carbon atoms and elements, creating a wide variety of molecular structures. Hybridisation involves mixing atomic orbitals to create new hybrid orbitals, enabling carbon to form these diverse structures.

The three main types of hybridisation in organic compounds are:

- **sp^3**: Tetrahedral geometry
- **sp^2**: Trigonal planar geometry
- **sp**: Linear geometry.

Each type determines the **molecular geometry** and **bonding patterns**, influencing the **stability** and **reactivity** of organic molecules. Hybridisation is essential for understanding organic bonding.

Types of Hybridisation and Bonding

1. sp^3 Hybridisation

sp^3 hybridisation occurs in alkanes, where each carbon atom forms four sigma (σ) bonds, resulting in a tetrahedral geometry with bond angles of approximately 109.5°.

Examples:

Methane (CH_4)

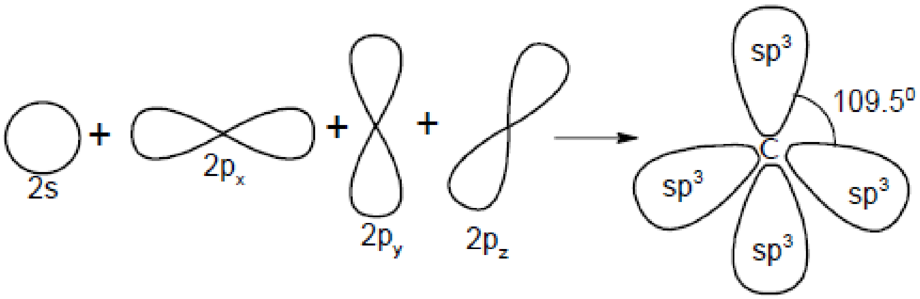
Carbon forms four sp^3-hybrid orbitals, each overlapping with hydrogen's 1s-orbital to form four σ-bonds.

Ethane (C_2H_6)

Each carbon atom starts with $1s^2 2s^2 2p^2$, having only two unpaired electrons in the 2p-orbitals.

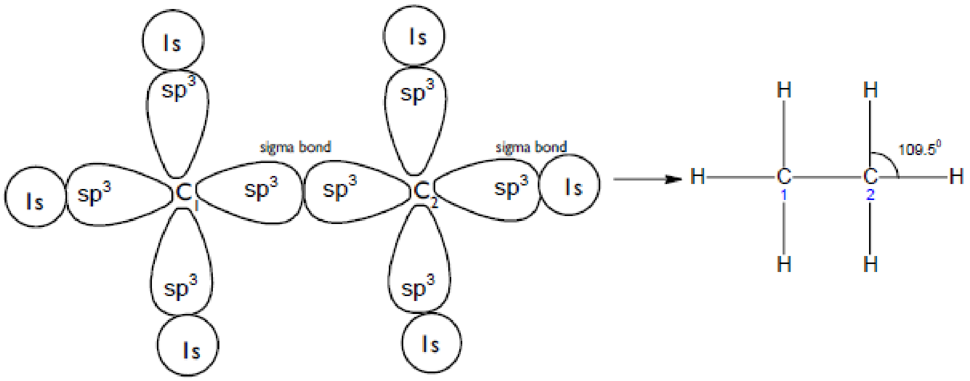
One electron from the 2s-orbital moves to the $2p_z$-orbital, creating four unpaired electrons with the configuration $1s^2 2s^1 2p^3$.

The 2s-orbital mixes with the $2p_x$, $2p_y$, and $2p_z$-orbitals, forming four equivalent sp^3-hybrid orbitals, each with 75% p character and 25% s character.



Each carbon uses one sp^3-orbital to overlap with the other carbon's sp^3-orbital, forming a single σ-bond.

The remaining three sp^3-orbitals on each carbon overlap with hydrogen's $1s^1$-orbitals, forming six σ-bonds (three per carbon).



The four sp^3-orbitals arrange in a tetrahedral shape to minimise repulsion, with bond angles close to 109.5°.

2. sp^2 Hybridisation

sp^2 hybridisation occurs in alkenes like ethene (C_2H_4).

Each carbon atom forms three sigma (σ) bonds and one pi (π) bond.

The sp^2 hybrid orbitals lie in a plane with bond angles of approximately 120°.

The unhybridised $2p_z$ orbital forms the π-bond.

Each carbon atom has the electron configuration $1s^2 2s^2 2p^2$.

There are two unpaired electrons in the 2p orbitals—insufficient for forming four bonds.

One electron from the 2s orbital is promoted to the empty 2p orbital.

New configuration: $1s^2 2s^1 2p_x^{\,1} 2p_y^{\,1} 2p_z^{\,1}$.

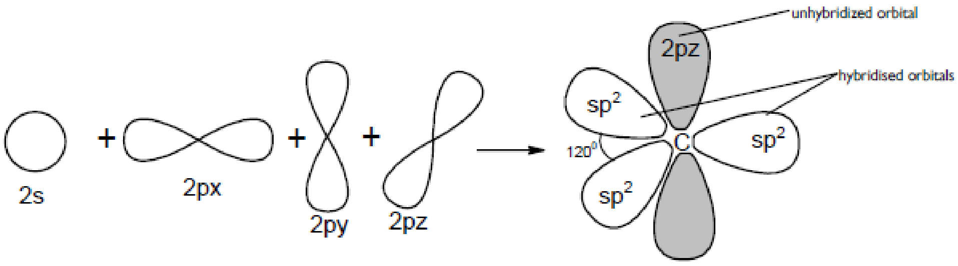
This results in four unpaired electrons.

The 2s orbital mixes with two 2p orbitals ($2p_x$ and $2p_y$).

Forms three equivalent sp^2 hybrid orbitals.

One $2p_z$ orbital remains unhybridised.

Each sp^2 hybrid orbital has 33% s character and 67% p character.



The three sp^2 hybrid orbitals arrange themselves 120° apart in a plane to minimize electron repulsion.

Each carbon atom uses its three sp^2 hybrid orbitals to form σ-bonds:

Two σ-bonds with hydrogen atoms.

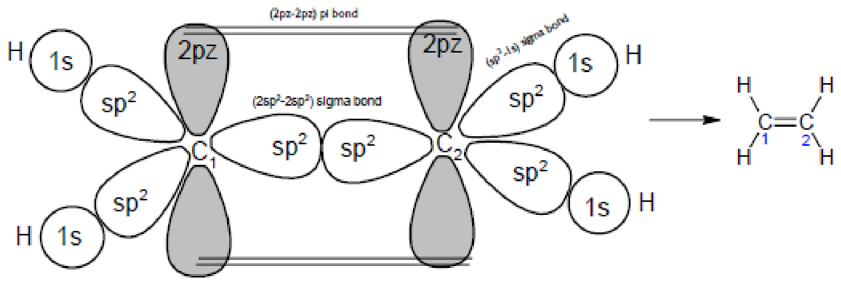
One σ-bond with the other carbon atom.

Pi (π) Bond Formation: The unhybridised $2p_z$ orbitals on each carbon overlap side-by-side. This forms a π-bond above and below the plane of the molecule.

Resulting double bond: The double bond between the carbon atoms consists of:

One σ-bond (from sp^2-sp^2 overlap).

One π-bond (from $2p_z$ overlap).



3. sp Hybridisation

sp hybridisation occurs in alkynes like ethyne (C_2H_2).

Each carbon atom forms two sigma (σ) bonds and two pi (π) bonds.

The sp hybrid orbitals are linearly arranged with bond angles of 180°.

Electron Configuration and Hybridisation:

Each carbon atom starts with $1s^2 2s^2 2p_x^1 2p_y^1$.

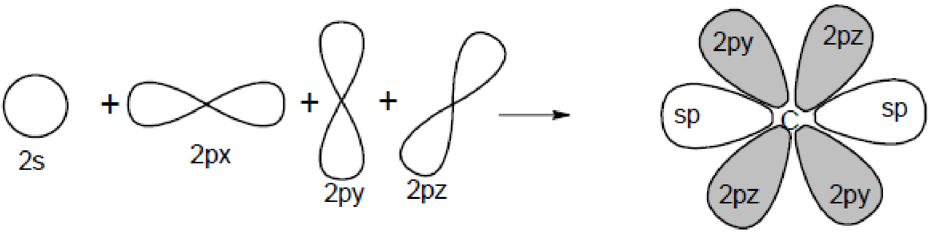
There are two unpaired electrons—insufficient for forming four bonds.

One electron from the 2s-orbital moves to the empty $2p_z^2$-orbital.

New configuration: $1s^2 2s^1 2p_x^1 2p_y^1 2p_z^1$.

The 2s-orbital mixes with one $2p_x$-orbital to form two sp hybrid orbitals.

50% s character and 50% p character.

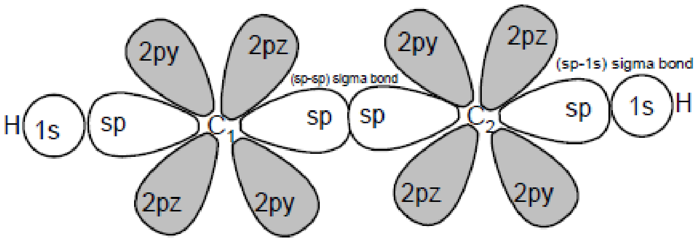


The remaining $2p_y$-orbitals remain unhybridised.

The two sp hybrid orbitals arrange themselves 180° apart to minimize repulsion.

One sp orbital from each carbon overlaps to form a σ-bond between the carbons.

The other sp orbitals overlap with hydrogen's 1s-orbitals to form two sigma bonds (one per carbon).

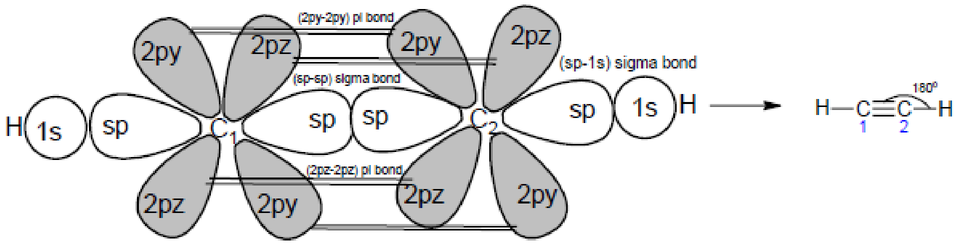


The unhybridised $2p_y$ and $2p_z$-orbitals on each carbon overlap side-by-side to form two π-bonds.

The triple bond between the two carbon atoms consists of:

One σ-bond (head-on sp-sp overlap).

Two π-bonds (sideways overlap of $2p_y$ and $2p_z$).



Activity 7.3 Building Carbon-Carbon Double and Triple Bonds

Material needed: Molecular model kits

Organise yourselves into groups of no more than five for this activity.

Steps

1. Building a carbon-carbon double bond
   - **a.** Identify the following components:
     2 carbon atoms, 4 hydrogen atoms, and 2 connectors for the double bond.
   - **b.** Connect two carbon atoms using two connectors to represent the double bond.
   - **c.** Attach two hydrogen atoms to each carbon atom using single connectors.
   - **d.** Observe the planar geometry with bond angles of 120°.
2. Building a Carbon-Carbon Triple Bond (10 minutes)
   - **a.** Identify the following components:
     2 carbon atoms, 2 hydrogen atoms, and 3 connectors for the triple bond.
   - **b.** Connect two carbon atoms using three connectors to represent the triple bond.
   - **c.** Attach one hydrogen atom to each carbon atom using single connectors.
   - **d.** Observe the linear geometry with bond angles of 180°.

3. Present your model to the class, explaining:
    - **a.** How you built the double and triple bonds.
    - **b.** The bond types and geometry of each structure.
    - **c.** What is the difference between double and triple bonds?
    - **d.** How does geometry affect molecular properties?

Activity 7.4 Exploring sigma and pi bonds in unsaturated hydrocarbons

Organise yourselves into groups of no more than five for this activity

Materials needed: Molecular model kits

Videos, charts, and pictures illustrating:

Formation of sigma (σ) and pi (π) bonds.

Orbital overlaps in double and triple bonds.

Examples of unsaturated hydrocarbons such as ethene (C_2H_4) and ethyne (C_2H_2).

Steps

1. Watch videos showing the formation of σ- and π-bonds. https://www.youtube.com/watch?v=i2WY_JElXlU
    - **a.** Study charts and pictures, analyse diagrams, focusing on:
        - i. Sigma bonds.
        - ii. Pi bonds
    - **b.** Discuss your observations with the class.
2. Build the following molecular model of ethene and ethyne using the following steps:

    Ethene (C_2H_4)
    - **a.** Two sigma bonds per carbon (one with hydrogen, one with the other carbon).
    - **b.** One pi bond between the carbons.
    - **c.** Observe planar geometry with 120° bond angles.

    Ethyne (C_2H_2)
    - **a.** One sigma bond and two pi bonds between carbons.
    - **b.** One sigma bond between each carbon and hydrogen.
    - **c.** Observe linear geometry with 180° bond angles.

3. Using the model you have created, compare

  a. the geometry of the molecules,

  b. how this geometry is influenced by the bonding within the molecules.

EXTENDED READING

For an in-depth understanding of molecular shapes, bond angles, sigma (σ) and pi (π) bonds, and hybridisation in organic compounds, consider exploring the following resources:

- Predicting Molecular Shapes and Bond Angles; Distinguishing Between Sigma and Pi Bonds:
    - Chemistry LibreTexts: "Sigma and Pi Bonds"
      This article explains the formation and characteristics of σ and π bonds, including their roles in molecular geometry.

      LibreTexts Chemistry
    - Chemistry LibreTexts: "Structure and Bonding in Ethene: The π Bond"
      This resource discusses the bonding in ethene, highlighting the formation of π bonds and their impact on molecular structure.

      LibreTexts Chemistry
- Understanding Hybridisation and Its Relation to Molecular Structures; Application to Sigma and Pi Bonds in Organic Compounds:
    - Leah4Sci: "Sp^3, Sp^2 and Sp Hybridisation, Geometry and Bond Angles"
      This article provides a detailed overview of sp^3, sp^2, and sp hybridisation, correlating each type with molecular geometry and bond angles.

      Leah4Sci
    - Master Organic Chemistry: "Hybrid Orbitals and Hybridisation"
      This resource delves into the concept of hybrid orbitals, explaining how different hybridisations (sp^3, sp^2, sp) influence molecular shapes and bonding in organic molecules.

      Master Organic Chemistry

- Visual Learning Resource: YouTube Video: "Hybridisation of Atomic Orbitals - Sigma & Pi Bonds - Sp Sp^2 Sp^3"

REVIEW QUESITONS

Review Questions 7.1

1. What is a sigma bond, and how is it different from a pi bond?
2. Identify the shape and bond angle of CH_4.
3. Predict the shape of NH_3 and explain how the lone pair affects the bond angles.
4. Distinguish between the types of bonds in C_2H_4 (ethylene).
5. Predict the molecular shape and bond angles of SO_2 using VSEPR theory and determine if it is polar or nonpolar.
6. Explain how sigma and pi bonds contribute to the rigidity of C_2H_4.
7. Analyse the bonding in C_2H_2 (ethyne) and predict its geometry, bond angles, and the type of bonds present.
8. Compare the molecular shapes of BF_3, NH_3, and H_2O using VSEPR theory and discuss the role of lone pairs.

Review Questions 7.2

1. What is hybridisation?
2. Identify the three main types of hybridisation in organic compounds.
3. How does hybridisation explain the structure of methane (CH_4)?
4. Explain the difference between sigma (σ) and pi (π) bonds.
5. Apply hybridisation to explain the bonding in ethene (C_2H_4).
6. How does hybridisation explain the structure and bonding in ethyne (C_2H_2)?
7. Compare the geometry and bonding in methane (CH_4), ethene (C_2H_4), and ethyne (C_2H_2) based on hybridisation.
8. Evaluate how the presence of pi bonds affects the rigidity of molecules.

SECTION

8

ORGANIC COMPOUNDS

CHEMISTRY OF CARBON COMPOUNDS

Organic Functional Groups

INTRODUCTION

This section introduces you to the fascinating world of organic chemistry, focusing on the structure, bonding, properties, and everyday uses of key organic compounds such as: alkanes, alkenes, alkynes, benzene, alkanols, alkanoic acids.

KEY IDEAS

- **Alkanes** are saturated hydrocarbons with single sigma (σ) bonds; general formula C_nH_{2n+2}
- **Alkanoic acids** are organic compounds that contain a carboxyl group (−COOH); general formula $C_nH_{2n+1}COOH$.
- **Alkanols** are hydrocarbons with one or more hydroxyl (−OH) groups attached.
- **Alkenes** are hydrocarbons that contain at least one double bond; general formula C_nH_{2n}.
- **Alkynes** are hydrocarbons that contain at least one triple bond; general formula C_nH_{2n-2}.
- **Benzene** a cyclic hydrocarbon with alternating double bonds.

ALKANES

Alkanes are hydrocarbons with carbon atoms connected by single bonds only. They are saturated hydrocarbons with the general formula C_nH_{2n+2}. One common use of alkanes is as fuel.

Naming a Straight-Chain Hydrocarbon

1. Determine the parent name by identifying the longest carbon chain in the compound. Refer to Tables 8.1 and 8.2 to determine the parent name. Notice that these names are made up of a prefix related to the number of carbons

in the chain and the suffix *-ane,* indicating that the molecule is an alkane (Table 8.2).

Table 8.1: Names and Formulae of the First Ten Straight-Chain Alkanes

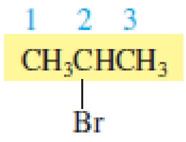
Name	Molecular Formula	Condensed Formula
Alkanes	C_nH_{2n+2}	
Methane	CH_4	CH_4
Ethane	C_2H_6	CH_3CH_3
Propane	C_3H_8	$CH_3CH_2CH_3$
Butane	C_4H_{10}	$CH_3CH_2CH_2CH_3$
Pentane	C_5H_{12}	$CH_3CH_2CH_2CH_2CH_3$
Hexane	C_6H_{14}	$CH_3CH_2CH_2CH_2CH_2CH_3$
Heptane	C_7H_{16}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_3$
Octane	C_8H_{18}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
Nonane	C_9H_{20}	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$
Decane	$C_{10}H_{22}$	$CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_2CH_3$

Table 8.2: Carbon Chain Length and Prefixes Used in the IUPAC Nomenclature System

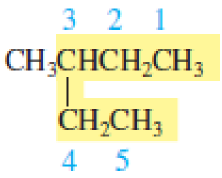
Carbon Chain Length	Prefix	Alkane Name
1	Meth-	Meth*ane*
2	Eth-	Eth*ane*
3	Prop-	Prop*ane*
4	But-	But*ane*
5	Pent-	Pent*ane*
6	Hex-	Hex*ane*
7	Hept-	Hept*ane*
8	Oct-	Oct*ane*
9	Non-	Non*ane*
10	Dec-	Dec*ane*

2. Number the parent chain to assign the lowest number to the first substituent encountered.
3. Name and number each substituent on the parent chain, indicating its type and position. F example, it may be one of the halogens

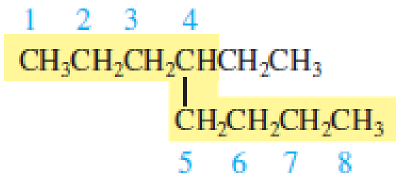
   [F—(fluoro), Cl—(chloro), Br—(bromo), and I—(iodo)] or an alkyl group (Tables 8.3). In the following examples, the parent chain is highlighted in yellow:



$CH_3\underset{\underset{\displaystyle Br}{|}}{C}HCH_3$



$CH_3\underset{\underset{\displaystyle CH_2CH_3}{|}}{C}HCH_2CH_3$



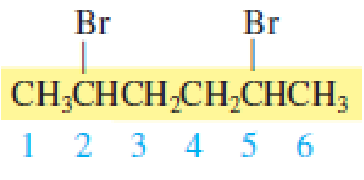
$CH_3CH_2CH_2\underset{\underset{\displaystyle CH_2CH_2CH_2CH_3}{|}}{C}HCH_2CH_3$

Substituent:	2-Bromo	3-Methyl	4-Ethyl
IUPAC name:	2-Bromopropane	3-Methylpentane	4-Ethyloctane

Table 8.3: Names and formulae of the first five continuous-chain alkyl groups

Alkyl Group Structure	Name
$-CH_3$	Methyl
$-CH_2CH_3$	Ethyl
$-CH_2CH_2CH_3$	Propyl
$-CH_2CH_2CH_2CH_3$	Butyl
$-CH_2CH_2CH_2CH_2CH_3$	pentyl

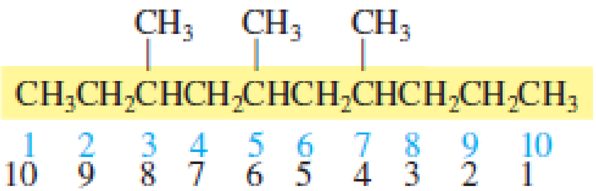
4. Use prefixes (di-, tri-, tetra-, etc.) and assign separate position numbers for occurrences of the same substituent:



$CH_3\overset{\overset{\displaystyle Br}{|}}{C}HCH_2CH_2\overset{\overset{\displaystyle Br}{|}}{C}HCH_3$

2,5-Dibromo

2,5-Dibromohexane

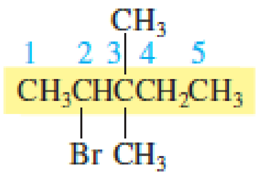


$CH_3CH_2\overset{\overset{\displaystyle CH_3}{|}}{C}HCH_2\overset{\overset{\displaystyle CH_3}{|}}{C}HCH_2\overset{\overset{\displaystyle CH_3}{|}}{C}HCH_2CH_2CH_3$

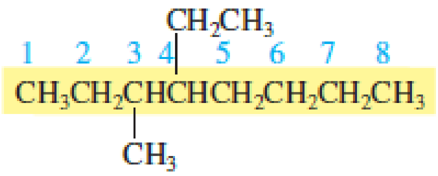
3,5,7-Trimethyldecane

Not 4,6,8-Trimethyldecane

5. List substituents in alphabetical order before the parent compound name, separating numbers with commas and numbers from names with hyphens; halogens are listed before alkyl groups.



$CH_3CHCCH_2CH_3$



$CH_3CH_2CHCHCH_2CH_2CH_2CH_3$

2-Bromo-3,3-

Not 3,3-Dimethyl-2-bromopentane

4-Ethyl-3-methyloctane

Not 3-Methyl-4-ethyloctane

Drawing the Structure of Alkanes

1. Determine the Number of Carbon Atoms (Parent Chain)

   Use the name of the alkane to determine the number of carbons in the longest chain: Methane: 1 carbon, Ethane: 2 carbons, Propane: 3 carbons, Butane: 4 carbons, and so on for longer chains.
2. Draw a straight line of carbon atoms to represent the parent chain.
3. Attach hydrogen atoms to each carbon until it has four bonds.
4. For branched alkanes, identify the substituents from the name and attach them to appropriate carbon atom.

Types of organic reactions

Reactivity of Alkanes

Alkanes are **not very reactive** and do not react easily with acids, bases, oxidants, or reductants under normal conditions.

The unreactive of alkanes is due to the following reasons;

1. Alkanes have saturated bonds, leaving no room for additional bonding.
2. The primary bonds in alkanes are strong, non-polar sigma bonds, making them difficult to break.
3. There is little electronegativity difference between carbon and hydrogen, so alkanes do not attract polar substances.
4. Alkanes lack lone pairs or empty orbitals, which are often needed to initiate reactions.

However, under certain conditions, alkanes can undergo combustion, halogenation and cracking.

Halogenation

Halogenation is a substitution reaction where a halogen atom (F, Cl, Br, or I) replaces a hydrogen atom in an alkane.

Conditions Required: Alkanes need **light**, **heat**, or a **catalyst** to react with halogens.

Alkanes do not react in the **cold or dark** because low temperatures reduce molecular energy, making it hard to overcome the high activation energy.

Light energy breaks halogen molecules into halogen radicals, initiating the reaction.

General Reaction: $R{-}H + X_2 \rightarrow R{-}X + HX$

(Where R is the alkane and X is the halogen).

Chlorination

Reaction Conditions

Chlorine (Cl_2) and methane (CH_4) do not react at room temperature in the absence of light.

Reaction begins when the mixture is heated above **100°C** or exposed to **intense light**, releasing heat.

Products

Initially forms chloromethane (CH_3Cl) and hydrogen chloride (HCl).

With more chlorine and continued reaction, additional products are formed:

Dichloromethane (CH_2Cl_2)

Trichloromethane ($CHCl_3$, chloroform)

Tetrachloromethane (CCl_4, carbon tetrachloride)

Initiation Step

The covalent bond in Cl_2 is broken, forming **chlorine free radicals** (atoms with unpaired electrons).

These radicals start the chain reaction of chlorination.

$$Cl - Cl \xrightarrow{\text{heat or light}} Cl^{\cdot} + Cl^{\cdot}$$

The chlorine radical is highly reactive and removes a hydrogen atom from methane to become stable. This reaction produces HCl and a methyl radical ($\cdot CH_3$).

$$\cdot Cl + CH_4 \rightarrow HCl + \cdot CH_3$$

The methyl radical is also reactive enough to abstract a chlorine atom from the excess Cl_2 if present.

$$H_3C\bullet + Cl{-}Cl \rightarrow H_3C{-}Cl + Cl$$

This is part of the **propagation phase** of the chlorination chain reaction, which continues until the reactants are consumed or two radicals combine to terminate the reaction.

Fluorine (F_2) reacts with alkanes but releases a lot of energy, making the reaction hard to control. Iodine (I_2) is rarely used because it needs too much energy to start, making it inefficient and difficult to manage.

Combustion of Alkanes

The combustion of alkanes is a chemical reaction in which alkanes react with oxygen (O_2) to produce carbon dioxide (CO_2) and water (H_2O), along with the release of energy in the form of heat and light.

Types of Combustion

1. Complete Combustion

   Occurs when there is enough oxygen available. Produces carbon dioxide and water.

   General equation:

   $$C_nH_{2n+2} + \left(\frac{3n+1}{2}\right) O_2 \rightarrow nCO_2 + (n+1)H_2O + \text{heat}$$

   where n = number of carbon atoms

   n = 1, 2, 3, 4 etc

   Examples

   $$CH_4(g) + 2O_2(g) \rightarrow 2H_2O(g) + CO_2(g) + \text{heat}$$

   $$C_2H_6(g) + 7\ \_\ 2\ O_2(g) \rightarrow 3H_2O(g) + 2CO_2(g) + \text{heat}$$

2. Incomplete Combustion

   Happens when there is insufficient oxygen. Produces **carbon monoxide (CO)** or **soot (C)** along with water.

General equation:

$$C_nH_{2n+2} + \left(\frac{2n+1}{2}\right) O_2 \rightarrow nCO + \left(n+1\right) H_2O + \text{heat}$$

Example

$$CH_4(g) + \frac{3}{2}O_2(g) \rightarrow 2H_2O(g) + CO(g) + \text{heat} \;;$$

$$CH_4(g) + O_2(g) \rightarrow 2H_2O(g) + C(s) + \text{heat}$$

Carbon monoxide is poisonous because it binds strongly to haemoglobin, preventing oxygen transport to tissues. Alkanes are excellent energy sources because they release a lot of energy when burned.

Cracking of Alkanes

Cracking is a chemical process where **long-chain alkanes** are broken down into **shorter-chain alkanes** and **alkenes**. This is done to make more useful products, such as fuels and chemicals used in industry.

Types of Cracking

1. **Thermal Cracking**

   Uses **high temperatures (400–900°C)** and **high pressure**.

   Produces a mixture of shorter alkanes and alkenes.

   Example

   $$C_{10}H_{22}\ \text{h}\xrightarrow{\text{eat}} C_5H_{12} + C_5H_{10}$$

2. **Catalytic Cracking**

   Uses a **catalyst** (e.g., zeolites) at **lower temperatures (450–750°C)**.

   Produces branched alkanes and aromatic compounds, which are useful in fuels like gasoline.

   Example: Heptane is reformed into toluene and hydrogen.

   $$C_{12}H_{22}\ \text{c}\xrightarrow{\text{atalyst}} C_8H_{18} + C_4H_8$$

Reformation

Reformation in chemistry, refers to a process where straight-chain hydrocarbons (alkanes) are converted into branched-chain hydrocarbons, cycloalkanes, or aromatic compounds. This process improves the quality of fuels, such as gasoline, by increasing their octane number.

Catalytic Reformation

Uses a **catalyst** such as **platinum**, **rhenium**, or **zeolites**.

Operates at **high temperature (500–600°C)** and **moderate pressure**.

Example reaction:

$$C_7H_{16}\ \mathrm{c}\xrightarrow{\text{atalyst, heat}} C_7H_8 + 4H_2$$

Specific terminology

Electrophile

An electrophile is a species (atom, ion, or molecule) that is electron-deficient and seeks to accept electrons. Electrophiles are attracted to electron-rich areas in a reaction.

Electrophile typically have a positive charge or partial positive charge ($\delta+$) or molecule with empty orbitals on the central atom and they act as Lewis acids (electron-pair acceptors).

Examples

H^+ (proton), NO_2^+ (nitronium ion), CH_3^+, BCl_3.

Nucleophile

A nucleophile is a species that is **electron-rich** and donates electrons to form a bond. Nucleophiles are attracted to electron-deficient areas in a reaction.

Features of nucleophile

Have a negative charge or lone pair of electrons.

Act as Lewis bases (electron-pair donors).

Examples: OH^- (hydroxide ion), NH_3 (ammonia), Cl^- (chloride ion)

Free Radicals

A free radical is a highly reactive species with an unpaired electron. They are formed by the homolytic cleavage of bonds.

Features of free radical are;

Represented with a dot (·) next to the element or group.

Highly unstable and reactive because of the unpaired electron.

Can initiate chain reactions in processes like halogenation of alkanes.

Examples

$\cdot Cl$ (chlorine radical)

$\cdot CH_3$ (methyl radical)

Formation

Free radicals are formed by heat, UV light, or peroxides, which cause bond homolysis.

Example

$$Cl_2 \xrightarrow{\text{UV light}} \cdot Cl + \cdot Cl$$

Isomerism

Isomerism is a phenomenon where two or more compounds have the same molecular formula but different structures or arrangements of atoms, leading to distinct chemical and physical properties. There are two types of Isomerism: structural Isomerism and stereoisomerism.

Isomerism in Alkanes

Alkanes, exhibit structural isomerism. This means that alkanes with the same molecular formula can have different arrangements of their carbon atoms.

Steps to determine the number of structural isomers in alkanes

1. Use the general formula for alkanes, C_nH_{2n+2}, to determine the number of carbons (n) and hydrogens (H).
2. Start by drawing the straight-chain (normal) alkane structure. This is the simplest isomer, with all carbons connected in a single continuous chain.
3. Create shorter chains by branching the carbon atoms. Follow these rules:
    a. Move one carbon at a time from the main chain to create a side branch.
    b. Avoid repeating structures (e.g., moving the branch from the first carbon to the last produces the same structure).
    c. Consider the symmetry of the molecule to avoid duplicates.
4. Assign systematic IUPAC names to the structures to ensure they are unique.

For example:

For C_5H_{12} (Pentane):

$CH_3{-}CH_2{-}CH_2{-}CH_2{-}CH_3$

$CH_3{-}CH(CH_3){-}CH_2{-}CH_3$

$CH_3{-}C(CH_3)_2{-}CH_3$

Physical properties of alkanes

Physical Properties of Alkanes (Simplified and Summarized)

1. **State at Room Temperature**

   Alkanes are non-polar compounds that can exist as:

   Gases: $C_1$to C_4 (weak Van der Waals forces).

   Liquids: C_5 to C_{17}.

   Solids: C_{18} and above (stronger Van der Waals forces).

2. **Melting and Boiling Points**

   Increase with longer carbon chains due to stronger Van der Waals forces (e.g., butane: −1.0°C, pentane: 36.1°C).

   Branching decreases boiling points because it reduces surface area and weakens Van der Waals forces (e.g., pentane: 36.1°C, 2-methylbutane: 27.8°C).

3. **Solubility**

   Alkanes are insoluble in water (non-polar) but dissolve in non-polar solvents like alkanes.

   Water's hydrogen bonds exclude alkanes, preventing solubility.

4. **Density**

   Alkanes have densities between 0.62 g/ml and 0.79 g/ml, which is less than water (1.0 g/ml).

   Low density is due to:

   Light atoms (hydrogen and carbon) and the tetrahedral molecular structure, increases volume relative to mass.

Effect of Combustion on the Environment

Complete Combustion: Produces carbon dioxide (CO_2).

Incomplete Combustion: Produces carbon monoxide (CO) in low oxygen conditions.

Impact on Global Warming: CO_2 contributes to the greenhouse effect, trapping heat and leading to global warming.

Greenhouse gases absorb and re-radiate infrared radiation, increasing Earth's temperature.

Uses of Alkanes (Simplified)

1. Gaseous alkanes (C_1 to C_4):

   Main component of natural gas, used for cooking, heating, and industrial processes. Found biogas from organic matter decomposition.

   Propane and butane: Components of Liquefied Petroleum Gas (LPG), used as heating and cooking fuels due to their high energy yield and easy transport in liquid form.

2. Liquid alkanes (C_5 to C_1):

   Gasoline (Pentane to Octane): Used in cars due to high volatility and energy release.

   Kerosene, diesel, and jet fuel (C_9 to C_{17}):

   Diesel: High energy density for heavy-duty vehicles.

   Jet fuel: Designed to remain liquid at low temperatures and provide thrust efficiently.

3. Waxy solids (C_{18} and Above):

   Paraffins: Used in wax coatings for fruits and vegetables to retain moisture and mould.

   Vaseline: A semisolid hydrocarbon mix for ointments, cosmetics, and lubrication.

4. Specialized uses:

   Motor oil: Lubricates engine components.

   Mineral oil: Used as a laxative and lubricant due to its insolubility in water.

Activity 8.1 Molecular models and Alkanes

Materials needed: Whiteboard, markers, molecular model kits (optional, for hands-on building of alkanes), worksheets with examples of alkanes for naming practice

Steps

1. Using the formula C_nH_{2n+2} calculate the molecular formulas for the first ten alkanes.
2. Write the molecular formula and draw the structures of these alkanes on a chart worksheet.
3. Guide to naming alkanes:
   - **a.** Identify the longest continuous chain of carbon atoms (parent chain).
   - **b.** Number the chain starting from the end nearest to a substituent or branch.
   - **c.** Name the substituents and place them in alphabetical order with the appropriate numbers.
4. Practice naming alkanes

   Name the following:

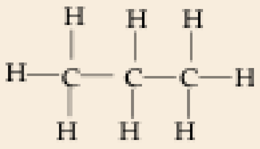
   i. $CH_3CH_2CH_2CH_3$
   ii. $CH_3CH_2CH_2CH_2CH_3$
   iii. $CH_3CH(CH_3)CH_3$
   iv. $CH_3CH_2CH(CH_3)CH_2CH_2CH_3$
   v. $CH_3CHClCH_3$

Activity 8.2

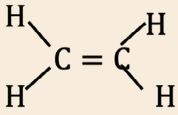
Materials Needed: Molecular model kits or 3D printed structures, charts or drawings illustrating electrophiles, nucleophiles, free radicals, types of reactions, bond cleavage, and isomerism, Flashcards with examples for discussions.

Steps

1. Engage in a whole class discussion and identify which of the following structures (a), (b) and (c) is saturated and unsaturated hydrocarbons.



(a)



(b)

$H-C\equiv C-H$

(c)

2. Explain why your chosen structure is saturated or unsaturated.

3. Name compound A.

4. Identify electrophile or nucleophile or free radicals from the species below:

$C_2H_5O^-$, $(CH_3)_3N:$, Cl^+, $H_3\dot{C}$, $CH_3-C^+=O$, N^+O_2

5. Identify the type of reaction and complete reaction:

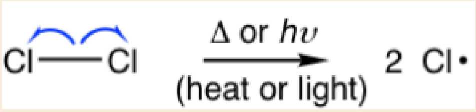
a. $CH_4 + 2O_2 \longrightarrow$

b. $CH_4 + Cl_2 \xrightarrow{\text{UV light}}$

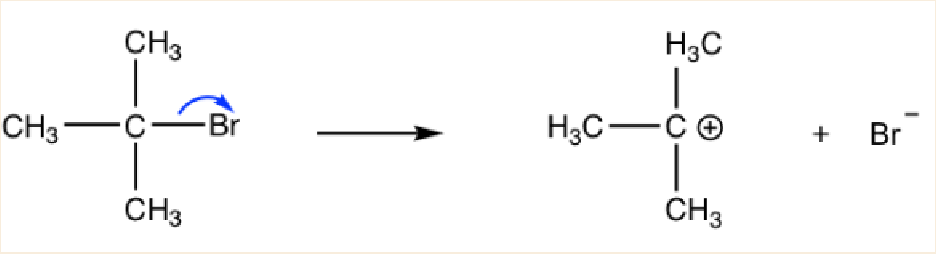
c. $C_{10}H_{22} \longrightarrow C_6H_{14} +$

6. Why are alkanes less reactive under normal conditions?

7. Diagram (a) and (b) illustrate homolytic and heterolytic cleavage respectively



(a)



(b)

a. What would methane produce in homolytic cleavage?

b. What would methane produce in heterolytic cleavage?

8. How does adding a functional group change the properties of an alkane?

9. Use model kits to construct different isomers of C_4H_{10} (butane) and C_5H_{12} (pentane).

   Identify and name the isomers.

Activity 8.3 Exploring Isomers and Naming Alkanes

Materials Needed: Molecular model kits (or paper cutouts for atoms and bonds), whiteboard and markers, handouts or flashcards of alkane examples, worksheets for drawing and naming exercises.

Steps

1. Use model kits or cut-outs to build and identify the isomers of butane and pentane.

   Compare their structures and write the molecular formulas for each.

2. Draw the structural formula and write the molecular formula for names like:
   - **a.** 2-Methylpropane
   - **b.** 3-Methylpentane

Activity 8.4 Investigate Physical Properties of Alkanes

Steps

1. Review the chart below of melting points, boiling points, and densities for different alkanes.

Table 8.4: Melting Points, Boiling Points, and Densities of Alkanes

Alkane	Formula	Melting Point (°C)	Boiling Point (°C)	Densities of Alkanes
Methane	CH_4	-182.5	-161.5	0.000656
Ethane	C_2H_6	-183.3	-88.6	0.00126
Propane	C_3H_8	-187.6	-42.1	0.002
Butane	C_4H_{10}	-138.4	-0.5	0.00258
Pentane	C_5H_{12}	-129.7	36.1	0.626
Hexane	C_6H_{14}	-95.3	68.7	0.66
Heptane	C_7H_{16}	-90.6	98.4	0.684
Octane	C_8H_{18}	-56.8	125.6	0.703
Nonane	C_9H_{20}	-51.3	150.8	0.718
Decane	$C_{10}H_{22}$	-29.7	174.1	0.73

2. Analyse the trends in boiling points, melting points and densities and explain using Van der Waals forces.

3. Write the balanced equations for

- **a.** Complete combustion of hexane
- **b.** Incomplete combustion of hexane

4. Engage in a whole class discussion to:

- **a.** Discuss the difference in energy release between complete and incomplete combustion.
- **b.** Discuss the harmful effects of CO from incomplete combustion.

Activity 8.5 Exploring Alkanes and Their Reactions

Material needed: computer, internet, other chemistry textbooks

Steps

1. Organise yourselves into groups of no more than five. In your groups:
   - a. Brainstorm solutions to reduce acid rain.
   - b. Create posters showing the impact of CFCs on the ozone layer
   - c. Draw flow diagrams illustrating the mechanism.
   - d. Research the uses of assigned fractions of petroleum.
2. Present this to your class for discussion and feedback.

Activity 8.6 Mini reports

1. Organise yourselves into groups of no more than five. In your groups, produce a mini report on the following:
2. Outline the reasons why some alkanes will undergo cracking, make sure to include examples and chemical reactions.
3. Outlines the reasons why some alkanes will under reforming, make sure to include examples and chemical reactions.

ALKENES

Alkenes are hydrocarbons with carbon-carbon double bonds. They are found in plant extracts like citrus fruits (limonene), steroids (cholesterol), and insect pheromones (muscalure). Industrially, alkenes are produced by cracking hydrocarbons.

Physical Properties of Alkenes

1. **Boiling Points**

   Alkenes have low boiling points.

   Boiling points increase with longer carbon chains and decrease with branching.
2. **Density**

   Alkenes have lower densities than alkanes because their double bond reduces molecular packing.

They are less dense than water.

3. **Solubility**

   Alkenes are insoluble in water (non-polar) but dissolve in organic solvents.

4. **State**

   Alkenes with fewer than five carbons are colourless gases.

   Alkenes with longer chains are volatile liquids.

5. **General Formula**

   Alkenes follow the formula C_nH_{2n} and form straight or branched homologous series.

Isomerism in Alkenes

Alkenes show **structural isomerism** and **geometric isomerism**:

1. **Structural Isomerism**: Compounds with the same molecular formula but different atom arrangements.

   Types structural isomerism in alkenes:

   Position isomerism: The double bond is in different positions.

   Chain isomerism: The carbon chain has different arrangements (straight or branched).

2. **Geometric isomerism**: Arises from restricted rotation around the double bond, resulting in **cis** (same side) and **trans** (opposite side) forms.

Position Isomerism

Position isomerism happens when the location of the carbon-carbon double bond changes in the molecule.

Example

The double bond is between the first and second carbon atoms.

$CH_2{=}CH{-}CH_2{-}CH_3$ 1-Butene

The double bond is between the second and third carbon atoms.

$CH_3{-}CH{=}CH{-}CH_3$ 2-Butene

Chain Isomerism in Alkenes

Chain isomerism occurs when alkenes with the same molecular formula have different arrangements of carbon atoms. The carbon chain can be straight or branched, while the number of carbon and hydrogen atoms stays the same.

Example

For C_4H_8

1. But-1-ene: A straight chain with the double bond at the first carbon. $CH_2{=}CH{-}CH_2{-}CH_3$
2. 2-Methylpropene: A branched chain with the double bond in a different arrangement. $CH_2{=}C(CH_3){-}CH_3$

IUPAC Nomenclature of Alkenes

To name alkenes (up to C), follow these steps

1. Find the longest carbon chain containing the double bond.
   Replace the suffix "-ane" (from alkanes) with "-ene."
   For multiple double bonds, use prefixes like **di**, **tri**, or **tetra** (e.g., butadiene).
2. Start numbering from the end closest to the double bond to give the lowest possible numbers to the double bond.
3. If the double bond is equally distant from both ends, start numbering from the end closest to the first branch.
4. Use the lower-numbered carbon of the double bond to specify its position in the chain (e.g.,1-butene, 2-pentene).

Synthesis of Alkenes

Alkenes are important in organic chemistry because they are reactive and useful intermediates. They can be made from:

1. **Alcohols**

   Alkenes are produced by dehydrating alcohols using a catalyst like concentrated H_2SO_4 heat.

   Example

   $$CH_3CH_2OH \xrightarrow{H_2SO_4,\ heat} CH_2 = CH_2 + H_2O$$

2. Synthesis of Alkenes from Alkyl Halides

Alkenes are made from alkyl halides through elimination reactions, where a hydrogen halide (HX) is removed.

A double bond forms when;

a. The halogen is removed.

b. The hydrogen is taken from the carbon with fewer hydrogen atoms (following Zaitsev's rule).

Example

$$CH_3CH_2Br \xrightarrow{\text{base}} CH_2 = CH_2 + HBr$$

Reactions of symmetrical and unsymmetrical alkenes

Reaction of Alkenes with Hydrogen

Hydrogen gas (H_2) reacts with alkenes under special conditions to produce alkanes. This reaction is called hydrogenation.

Reagent: H_2 gas.

Conditions:

Temperature: Above 130°C.

Catalyst: Nickel, platinum, or palladium.

Reaction:

Two hydrogen atoms are added to the carbon-carbon double bond (C=C).

The double bond breaks, forming a saturated alkane.

This reaction is highly exothermic.

For a symmetrical alkene:

$$CH_3 = CH_2 + H_2\ h\xrightarrow{\text{eat, catalyst}} CH_3 - CH_3$$

For an unsymmetrical alkene:

$$CH_3 - CH = CH_2 + H_2 \xrightarrow{\text{heat, catalyst}} CH_3 - CH_2 - CH_3$$

This reaction is very useful in margarine production from vegetable oils which are soft and have a high proportion of cis-alkenes.

a. Reaction of Alkenes with Halogens

Halogens (Cl_2, Br_2) are symmetrical reagents that react with alkenes to form dihaloalkanes. This reaction is an example of an **addition reaction**.

Reagent: Halogen (Cl_2, Br_2).

Conditions

Room temperature.

Dark environment.

Polar solvent.

Reaction

The halogen molecule adds across the double bond, breaking it and forming a saturated dihaloalkane.

Reaction Representation

For the simplest alkene (ethene): $CH_2{=}CH_2 + Cl_2 \rightarrow CH_2Cl{-}CH_2Cl$

(Ethene reacts with chlorine to form 1,2-dichloroethane)

For an unsymmetrical alkene (propene)

$CH_3{-}CH{=}CH_2 + Br_2 \rightarrow CH_3{-}CH(Br){-}CH_2(Br)$

b. Reaction of Alkenes with Steam (Hydration)

The reaction of alkenes with steam is called **acid-catalysed hydration**, where water (H_2O) adds across the double bond to form an alcohol.

Reagent: Steam (H_2O).

Catalyst: Acid (e.g., H_2SO_4 or H_3PO_4).

The acid helps dissociate water molecules, which are otherwise too weak to react.

Reaction

Water adds across the double bond of the alkene, breaking the double bond and forming an alcohol.

$$CH_2 = CH_2 + H_2O \xrightarrow{\text{acid catalyst}} CH_3 - CH_2OH$$

For an unsymmetrical alkene (propene)

$$CH_3 - CH = CH_2 + H_2O \xrightarrow{\text{acid catalyst}} CH_3CHOH - CH_3$$

Propene reacts with steam to form propan-2-ol, following **Markovnikov's rule**: the OH group attaches to the more substituted carbon atom.

c. Reaction of Alkenes with Hydrogen Halides

Hydrogen halides (HX, where X is Cl, Br, or I) are unsymmetrical reagents that react with alkenes to form **haloalkanes**. In this reaction, the double bond becomes saturated by adding hydrogen and the halogen atom.

Reagent: Hydrogen halide (HX, e.g., HCl, HBr, HI).

Reaction: The alkene reacts with the hydrogen halide, breaking the double bond and forming a single-bonded haloalkane.

Markovnikov's Rule (for unsymmetrical alkenes)

The hydrogen atom from HX adds to the carbon with more hydrogen atoms, and the halogen adds to the carbon with fewer hydrogen atoms.

For a symmetrical alkene (ethene): $CH_2{=}CH_2 + HCl \rightarrow CH_3{-}CH_2Cl$

For an unsymmetrical alkene (propene):

$CH_3{-}CH{=}CH_2 + HBr \rightarrow CH_3{-}CH(Br){-}CH_3$

Propene reacts with hydrogen bromide to form 2-bromopropane, following Markovnikov's rule.

d. Reaction of $KMnO_4$ with Ethene

Aqueous $KMnO_4$ reacts with ethene at room temperature:

Product: 1,2-ethanediol ($HOCH_2CH_2OH$), a diol.

Oxidation: Manganese's oxidation state decreases from +7 (in $KMnO_4$) to +4.

Observation: The solution changes colour from purple (permanganate) to brown (manganese dioxide, MnO_2).

$$CH_2 = CH_2 + [O] \xrightarrow{KMnO_4} HOCH_2CH_2OH$$

Distinguishing Alkanes and Alkenes Using Bromine Water

Alkanes (saturated hydrocarbons): Do not react with bromine water under normal conditions because they lack a double bond. The reddish-brown colour of bromine remains unchanged.

Alkenes (unsaturated hydrocarbons): React with bromine water, adding bromine across the double bond. The reaction forms a dibromoalkane and decolourizes the reddish-brown bromine water.

Example: $CH_2{=}CH_2 + Br_2 \rightarrow CH_2Br{-}CH_2Br$

Importance of Alkene Reactions

1. Alkenes are raw materials for making alcohols, polymers, and other organic compounds.
2. Used to create plastics and synthetic materials.
3. Alkenes like butene and propene enhance gasoline combustion and performance.
4. Serve as intermediates in making detergents.
5. Used to produce ethylene glycol, an antifreeze chemical for car radiators.

Activity 8.6 Understanding Alkenes: General, Molecular, and Structural Formulae

Steps

1. Recall and share the general formula of alkenes (Think-Pair-Share).
2. State the molecular formulae of the alkenes with n=3,4,5 number of carbons.
3. In pairs or small groups, use molecular model kits or paper cutouts to create the structure ethene, propene and butene.
4. Share your model or drawing with the class.

Activity 8.7 Exploring Isomerism, Properties, and Reactions of Alkenes

Materials needed: Molecular model kits

Steps

1. Write the molecular formula of butene
2. Build models or draw the structural formulae of 1-Butene and 2-Butene.
3. Name the structures $CH_3{-}CH{=}CH{-}CH_3$ and $CH_3{-}CH_2{-}CH{=}CH{-}CH_2{-}CH_3$.
4. Create a mind map of alkenes showing:

   Sources of alkenes

   Laboratory Preparation: Dehydration of alcohols, dehalogenation.

   Physical Properties: Low boiling points, insolubility in water, reactivity due to double bonds.

5. Complete and balance equations for reactions with:

a. Hydrogen:

$$CH_3 - CH_2 - CH = CH_2 + H_2 \xrightarrow{\text{Ni, catalyst}}$$

b. Halogens:

$$CH_3 - CH_2 - CH = CH_2 + Br_2 \rightarrow$$

c. Steam:

$$CH_3 - CH_2 - CH = CH_2 + H_2O \xrightarrow{H_2SO_4}$$

d. Hydrogen halides:

$$CH_3 - CH_2 - CH = CH_2 + HCl \longrightarrow$$

e. $KMnO_4$

$$CH_3 - CH_2 - CH = CH_2 + [O] \xrightarrow{\mathbf{KMnO_4}}$$

Activity8.8 Distinguishing Between Alkanes and Alkenes Using Bromine Water

Materials Needed: Test tubes and test tube rack, bromine water or bromine dissolved in tetrachloromethane (reddish-brown solution), samples of an alkane (e.g., hexane) and an alkene (e.g., hexene), dropper or pipette, gloves and safety goggles.

Steps

1. Label two test tubes: one for the alkane and one for the alkene.
2. Pour a small amount of bromine water or bromine in tetrachloromethane into each test tube.
3. Add a few drops of hexane (alkane) to one test tube and hexene (alkene) to the other.
4. Gently shake or swirl both test tubes to mix the solutions.
5. Observe and record the colour change (or lack thereof) in each test tube.
6. Engage in a class discussion to discuss why the alkene reacted whereas the alkane did not.
7. Write down the reactions that occur.

Activity 8.9 Understanding the economic importance of alkene reactions

Steps

1. Identify the products you use that might involve margarine or alcohol.
2. Organise yourselves into groups of no more than five. In your groups, discuss why it is important to produce these products economically.
3. Use the internet and watch videos on:
   - **a.** Hardening vegetable oils (margarine) through hydrogenation
   - **b.** Hydration process for making alcohol (ethanol)
4. After the Video
   - **a.** Discuss what you learned about
     - i. How margarine is made.
     - ii. Why is ethanol important in our daily lives
5. Engage in a class discussion to share your insights.

ALKYNES

Introduction to Alkynes

Alkynes are hydrocarbons with at least one carbon-carbon triple bond, making them distinct from alkanes (single bonds) and alkenes (double bonds). Their general formula is $\mathbf{C_nH_{2n-2}}$ ($n \geq 2$), showing they are unsaturated hydrocarbons with fewer hydrogen atoms than alkanes.

Examples include:

Ethyne (C_2H_2) for $n = 2$

Propyne (C_3H_4) for $n = 3$

Butyne (C_4H_6) for $n = 4$

Alkynes are less common than alkenes in nature and labs. The triple bond includes one sigma and two pi bonds, making it shorter, stronger, and resulting in a linear structure with a bond angle of 180°.

Alkynes are highly reactive due to the electron density in the triple bond, which readily undergoes addition reactions. These reactions are vital in organic synthesis for producing complex molecules.

Structure of Alkynes

In alkynes, the carbon atoms in the triple bond are sp-hybridised. This involves the mixing of one s orbital and one p orbital from each carbon atom, forming two sp hybrid orbitals. These orbitals create a linear structure with a bond angle of 180°, giving alkynes a straight-line geometry. See Figure 8.1.

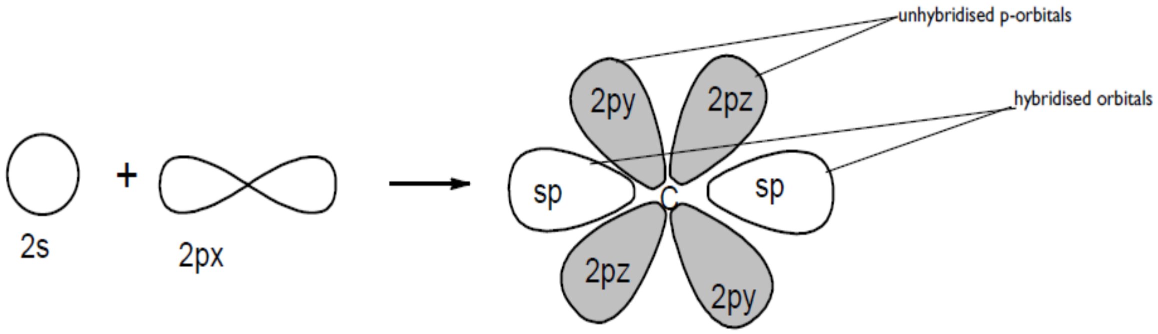


Figure 8.1: sp hybridisation in alkyne

The triple bond in alkynes consists of one sigma bond and two pi bonds:

The sigma bond is formed by the head-on overlap of sp orbitals from each carbon atom.

The pi bonds are created by the side-by-side overlap of the unhybridised p orbitals on each carbon atom. Refer to Figure 8.2.

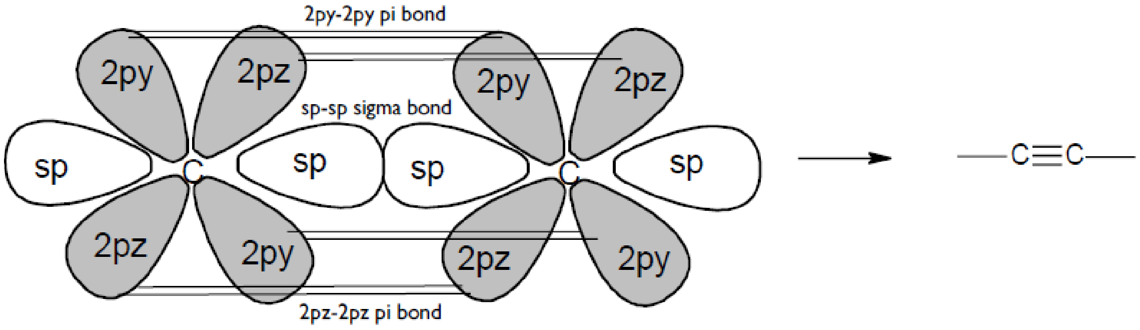


Figure 8.2: sigma and pi bonds formation in alkyne

The triple bond in alkynes is shorter and stronger than single or double bonds, making alkynes highly reactive. Their linear structure influences their properties:

Alkynes are generally non-polar and insoluble in water, but they dissolve in organic solvents.

The triple bond increases reactivity, enabling reactions like hydrogenation and halogenation.

Nomenclature of Alkynes

Alkyne name is derived from the names of the alkanes of the same carbon chain length. The same prefixes are used (*meth-*, *eth-*, etc.), but the suffix is different

(-*yne* for alkynes). To determine the name of an alkyne using the IUPAC Nomenclature System, use the following simple rules:

1. Count the number of carbon atoms in the longest continuous carbon chain containing the triple bond (alkynes). Name the alkane with the same number of carbon atoms. This is the parent compound.
2. Replace the -*ane* ending of the alkane with the ending for an -*yne* ending for an alkyne. F example:

   $CH \equiv CH$ — *ethyne* , $CH \equiv C - CH_3$ — *propyne*

3. Number the parent chain to give the double or triple bond the lowest number.

   1 2 3 4 5

   $CH{\equiv}CCH_2CH_2CH_3$

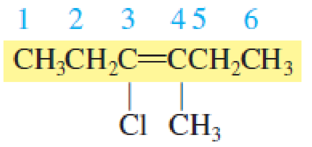
   1-pentyne

4. Determine the name and carbon number of each group bonded to the parent alkyne, and place the name and number in front of the name of the parent compound. Remember that with alkynes the triple bond takes precedence over a halogen or alkyl group, as shown in the following examples:

1 2 3 4 5 6

$CH_3CHBrC{\equiv}CCH_2CH_3$

2-Bromo-3-hexyne



$CH_3CH_2C(Cl){=}C(CH_3)CH_2CH_3$

3-Chloro-4-methyl-3-hexene

Physical Properties of Alkynes

1. **State and Appearance**:

   The first three alkynes (ethyne, propyne, butyne) are gases at room temperature.

   Higher alkynes are liquids or solids.

   Alkynes are colourless, and ethyne has a garlic-like odour.

2. **Polarity and Solubility**:

   Alkynes are non-polar.

   Slightly soluble in water but dissolve well in organic solvents like benzene, acetone, and chloroform.

3. **Boiling and Melting Points:**

   Increase with molecular weight.

   Slightly higher than those of alkanes and alkenes due to stronger London dispersion forces from the triple bond.

4. Density:

   Less dense than water.

   Density increases with molecular weight but remains lower than corresponding alkanes and alkenes.

Chemical properties of alkynes

Alkynes are highly reactive due to the presence of a carbon-carbon **triple bond**, which makes them susceptible to **addition reactions** and other transformations. Below are the key chemical properties of alkynes:

1. **Addition Reactions**

   Alkynes undergo **addition reactions** where reactants add across the triple bond, reducing it to a double or single bond.

   a. **With Hydrogen (Hydrogenation)**

   Alkynes react with hydrogen gas (H_2) in the presence of a catalyst (e.g., nickel, palladium, or platinum) to form alkenes or alkanes.

   Example:

   $$CH \equiv CH + H_2 \xrightarrow{catalyst} CH_2 = CH_2 + H_2 \xrightarrow{catalyst} CH_3 - CH_3$$

   b. **With Halogens**

   Alkynes react with halogens (Cl_2, Br_2) to form dihaloalkenes or tetrahaloalkanes.

   Example:

   $CH{\equiv}CH + Br_2 \rightarrow CHBr{=}CHBr$ (dibromoethene)

   c. **With Hydrogen Halides**

   Hydrogen halides (HCl, HBr) add across the triple bond, following **Markovnikov's rule**.

Example:

$CH\equiv CH + HCl \rightarrow CH_2{=}CHCl + HCl \rightarrow CH_3{-}CHCl_2$

d. Hydration of Alkynes

Alkynes react with water (H_2O) in the presence of a catalyst, like mercuric sulphate and sulphuric acid, to form ketones or aldehydes. This reaction follows Markovnikov's rule.

a. Ethyne (acetylene): Forms ethanal (an aldehyde) upon hydration.

$$CH \equiv CH + H_2O \xrightarrow{H_2SO_4,\ HgSO_4} CH_3CHO$$

b. Other Alkynes: Produce ketones

Example: Prop-1-yne forms propan-2-one (acetone)

$$CH \equiv C - CH_3 + H_2O \xrightarrow{H_2SO_4,\ HgSO_4} CH_3COCH_3$$

e. Combustion

Alkynes burn in the presence of oxygen to produce carbon dioxide and water.

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

f. Polymerization

Alkynes can undergo polymerization to form polymers. For example:

Formation of Benzene: Three ethyne (C_2H_2) molecules polymerize to form benzene (C_6H_6)

$$3CH \equiv CH \xrightarrow{heat,\ catalyst} C_6H_6$$

g. Oxidation

Alkynes can be oxidized with strong oxidizing agents like potassium permanganate ($KMnO_4$) to form carboxylic acids.

Example: $CH\equiv CH + [O] \rightarrow 2HCOOH$

h. Acidic Nature of Terminal Alkynes

Terminal alkynes (with a triple bond at the end of the chain) are acidic. They react with strong bases like sodium amide ($NaNH_2$) to form alkynide salts.

Example: $CH\equiv CH + NaNH_2 \rightarrow HC\equiv CNa + NH_3$

Laboratory preparation of alkynes

1. **Elimination reactions of dihalides to form alkynes**

   Alkynes can be made from **dihalides** through a **double elimination reaction**, where two molecules of hydrogen halide (HX) are removed.

   Types of Dihalides

   Vicinal dihalides: Halogens are on adjacent carbons.

   Geminal dihalides: Halogens are on the same carbon.

   Mechanism

   A strong base like sodium amide ($NaNH_2$) in liquid ammonia (NH_3) removes a proton next to a halogen.

   This creates a double bond (alkene).

   The base removes another proton from the alkene, forming a triple bond (alkyne).

   Example: $CH_2Br - CH_2Br \xrightarrow{NaNH_2/NH_3} CH \equiv CH + 2HBr$

2. **Dehydrohalogenation of Alkyl Dihalides**

   Alkynes can be prepared by removing hydrogen halides (HX) from vicinal or geminal dihalides. This process requires a strong base and high temperatures.

   Mechanism

   The base removes a proton (H^+) from a carbon adjacent to the halogen, forming a double bond (alkene).

   The base removes another proton, resulting in the formation of a triple bond (alkyne).

   Reactants: Vicinal or geminal dihalides.

   Conditions: Strong base (e.g., $NaNH_2$), high temperature.

   Product: Alkyne.

   Example: $CH_2Br - CH_2Br \xrightarrow{NaNH_2,\ heat} CH \equiv CH + 2HBr$

3. **Preparation of Alkynes from Alkenes**

   Alkynes can be prepared from alkenes through **halogenation** followed by **dehydrohalogenation**.

   Halogenation: Halogens (e.g., Br_2) are added across the double bond of an alkene, forming a dihaloalkane.

Dehydrohalogenation: A strong base removes two hydrogen halides (HX) from the dihaloalkane, forming a **triple bond** (alkyne).

Mechanism

Step 1: $CH_2{=}CH_2 + Br_2 \rightarrow CH_2Br{-}CH_2Br$

Step 2: $CH_2Br - CH_2Br \xrightarrow{NaNH_2,\ heat} CH \equiv CH + 2HBr$

Preparation of Alkynes from Calcium Carbide

Alkynes, especially acetylene (ethyne), can be produced from calcium carbide in an industrial process.

Limestone ($CaCO_3$) is heated to produce calcium oxide (CaO):

$$CaCO_3 \xrightarrow{heat} CaO + CO_2$$

Calcium oxide reacts with carbon (coke) at high temperatures to form calcium carbide (CaC_2):

$$CaO + 3C \xrightarrow{heat} CaC_2 + CO$$

Calcium carbide reacts with water to produce acetylene (C_2H_2):

$$CaC_2 + 2H_2O \rightarrow C_2H_2 + Ca(OH)_2$$

Everyday Uses of Alkynes (Simplified)

Alkynes are highly useful in various fields due to their reactivity and versatility. Some common applications include:

1. **Fuel and Energy**

   Acetylene (C_2H_2) is used as a fuel for welding and cutting metals due to its high temperature.

   It is also used in portable lighting, like miner's lamps.

2. **Chemical Synthesis**

   Alkynes are intermediates in making pharmaceuticals, agrochemicals, and dyes, aiding in the production of complex molecules.

3. **Polymer Production**

   Acetylene is used to make polyvinyl chloride (PVC), commonly used in pipes, cables, and clothing.

   Other alkynes help in producing plastics like polyethylene.

4. **Artificial Ripening**

   Ethyne is used to ripen fruits artificially, ensuring optimal ripening for the market.

5. **Solvents and Additives**

   Alkynes are used to create solvents and improve the performance of fuels, paints, coatings.

6. **Pharmaceuticals**

   Alkynes are used in synthesising active ingredients for drugs, enabling the development of targeted therapies.

Activity 8.10 Understanding Alkynes

Material needed: Chemistry textbooks, diagrams, or internet access

Steps

1. Organise yourselves into groups of not more than five. In your groups, brainstorm what you know about alkynes.
   a. What makes alkynes unique?
   b. How might alkynes be useful in everyday life?
2. Name the following structures:
   a. $CH\equiv CH$
   b. $CH\equiv C\text{-}CH_2\text{-}CH_3$
   c. $CH_3\text{-}CH_2\text{-}C\equiv C\text{-}CH_2\text{-}CH_3$
3. Discuss the structure, acidity, or reactions of alkynes.
4. a. Write the reaction for converting ethyne into ethanol.
   b. Identify isomers of $C_4H_{6.}$
5. Discuss your findings with the class.
6. Use the internet to watch video on alkynes (e.g., their reactions or real-world uses). In your groups, discuss your insights and thoughts about the video.
7. Use the thoughts and insights from question 5 to write an essay on the role of alkynes in modern technology and medicine.

Activity 8.11 Comparing Alkanes, Alkenes and Alkynes

Organise yourselves into groups of no more than five. In your groups, complete the following and discuss your findings with the class for feedback.

1. Create a Venn diagram comparing and contrasting alkynes, alkanes, and alkenes.
2. Compare the bond angles and lengths in alkynes with those in alkanes and alkenes, and alkanes and explain the differences.
3. Explain the difference is physical properties between the alkanes, alkenes and alkynes. Focus specifically on boiling/melting points and reactivity.

BENZENE

Introduction to Benzene

Benzene is a fundamental organic compound with unique structure and properties, making it essential in organic chemistry.

Natural Sources: Benzene occurs naturally in volcanoes, forest fires, plants, and animals.

Industrial Production: It is produced from coal and oil for commercial use.

Applications: Benzene is found in pharmaceuticals (e.g., aspirin, paracetamol), pesticides, polymers, and dyes.

Discovery of Benzene

1825: Michael Faraday discovered benzene while studying illuminating gas.

1834: Eilhardt Mitscherlich produced benzene by heating benzoic acid with lime.

1845: A.W. von Hofmann isolated benzene from coal tar.

Structure of Benzene

Benzene (C_6H_6) consists of six carbon and six hydrogen atoms arranged in a ring. Kekulé's Model (1865): August Kekulé proposed that benzene has a ring of six carbon atoms with alternating single and double bonds.

This structure suggested benzene would behave like alkenes, but it does not.

Refined Understanding

The double bonds are not fixed but resonate between positions, creating a stable structure. Refer to Figure 8.3. This phenomenon, called resonance, explains benzene's unique stability.

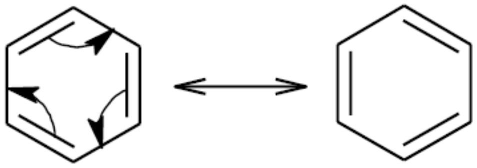


Figure 8.3: Kekule structures of benzene

Bonding in Benzene

Benzene's structure is explained by the delocalisation of electrons using modern bonding theories.

Hybridization

Each carbon atom in benzene undergoes sp² hybridisation. This combines one s orbital and two p orbitals to form three sp² hybrid orbitals.

Structure

The sp² hybrid orbitals form a trigonal planar arrangement around each carbon atom, with bond angles of 120°. This creates a flat, hexagonal ring structure.

Delocalization

The remaining unhybridized p orbitals on each carbon overlap sideways, forming a pi (π) electron cloud above and below the ring.

This delocalization gives benzene its unique stability and aromatic properties.

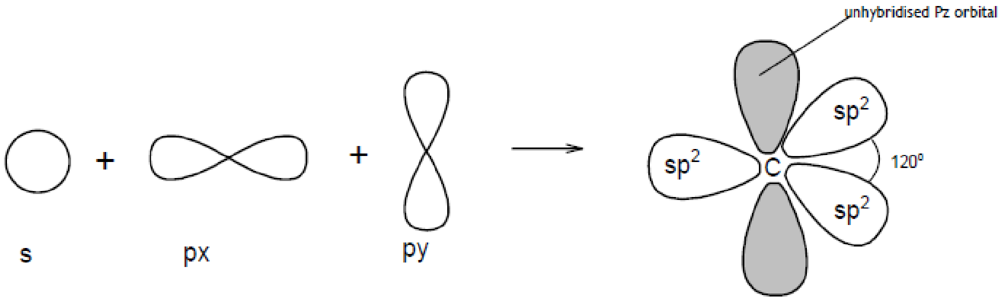


Figure 8.4: sp^2 hybridisation in benzene

Each carbon atom in benzene uses two of its sp^2 hybrid orbitals to form σ-bonds with two neighbouring carbon atoms, creating a continuous ring.

The third sp^2 hybrid orbital forms a σ-bond with a hydrogen atom.

Each carbon atom has one unhybridised p-orbital, perpendicular to the plane of the ring.

These p-orbitals overlap sideways with adjacent p-orbitals, forming a π-electron cloud above and below the ring.

The overlapping π-orbitals result in delocalised electrons, giving benzene its stability and a hexagonal ring structure.

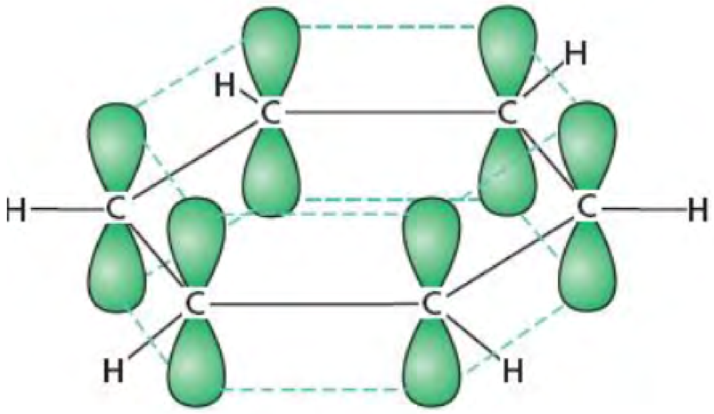


Figure 8.5: Sideways overlap of unhybridised p orbitals to form pi bonds

Delocalised π-electrons: The π-electrons in benzene are not confined to specific carbon atoms but are spread evenly across the entire ring.

This creates a continuous π-system, shared equally among all six carbon atoms.

Aromatic stabilization: The delocalisation of π-electrons gives benzene extra stability, known as **aromatic stabilisation**.

Representation: Benzene is often shown with a circle inside the hexagonal ring, symbolising the equal distribution of π-electrons.

It can also be represented by two resonance structures where the double bonds alternate, but the actual structure is a resonance hybrid.

Bond lengths: Delocalisation results in all carbon-carbon bonds in benzene having the same length, intermediate between single and double bonds.

Implications of Benzene's bonding

1. Chemical reactivity

Benzene's delocalised π-electrons make it less reactive in addition reactions, as these would disrupt its aromatic stability.

Instead, benzene undergoes electrophilic aromatic substitution, where a hydrogen atom is replaced by another group without affecting the π-system.

2. Aromaticity

Benzene is a classic example of an aromatic compound due to its cyclic, planar structure and delocalised π-electrons, which provide extra stability.

Nomenclature of Benzene

Naming benzene derivatives can be complex because of the variety of possible substituents. Key points include:

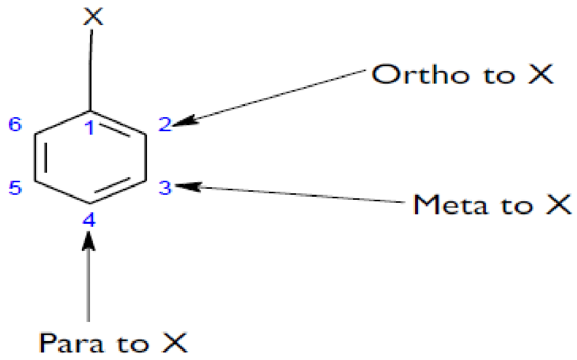
1. **Simple Derivatives**

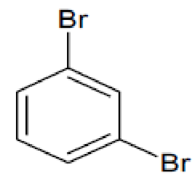
   If there is only one substituent, the compound is named by placing the substituent's name before "benzene" (e.g., **chlorobenzene**, **nitrobenzene**).

2. **Multiple Substituents**

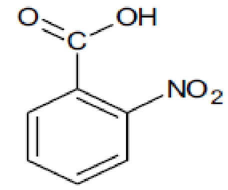
   When there are two or more substituents, their positions on the ring are indicated by numbers or prefixes:

   - **Ortho (o-)**: Substituents on adjacent carbons.
   - **Meta (m-)**: Substituents separated by one carbon.
   - **Para (p-)**: Substituents opposite each other.

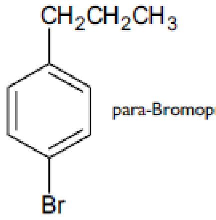




meta-Dibromobenzene



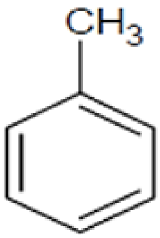
ortho-Nitrobenzoic acid



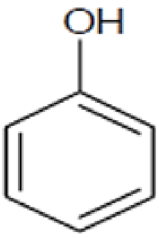
para-Bromopropylbenzene

3. **Special Names**

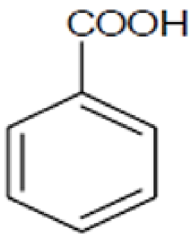
   Some derivatives have common names, such as **toluene** (methylbenzene), **phenol** (hydroxybenzene), and **aniline** (aminobenzene).



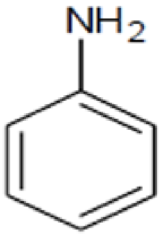
toluene



phenol



benzoic acid



aniline

4. When a benzene ring is attached to a larger molecule as a substituent, it is called a "phenyl group" (C_6H_5).

   Example

   Phenylalanine: An amino acid that contains a phenyl group attached to its structure. See Figure 8.6.

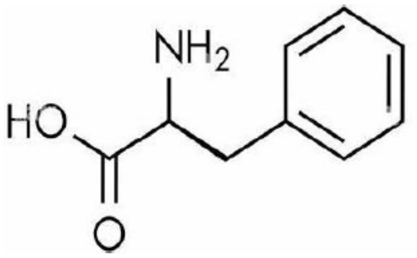


Figure 8.6: Structure of phenylalanine

Properties of Benzene

Physical Properties of Benzene (Simplified)

1. **Appearance**

   Benzene is a clear, colourless liquid with a sweet odour, noticeable even at concentrations.

2. **Density**

   Its density is 0.8765 g/cm³ at 20°C, making it less dense than water, so it floats on water.

3. **Melting and Boiling Points**

   Melting Point: 5.53°C (just above water's freezing point).

   Boiling Point: 80.1°C (low, making it easy to vaporise).

4. **Solubility**

   Benzene is insoluble in water due to its non-polar nature.

   It dissolves well in organic solvents like alcohol, ether, chloroform, and acetone.

5. **Flammability**

   Benzene is highly flammable and burns with a sooty flame due to incomplete combustion, requiring careful handling.

6. **Refractive Index**

   Its refractive index is 1.5011 at 20°C, useful for optical applications and checking purity.

Chemical Properties of Benzene

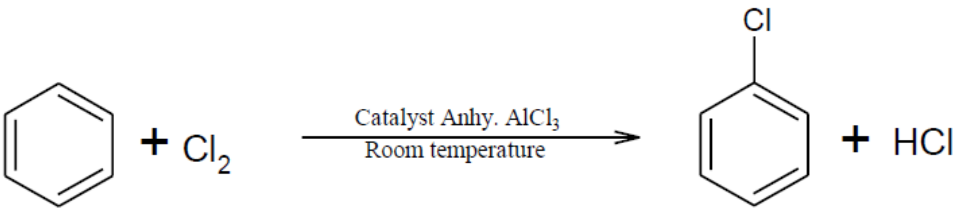
Benzene mainly undergoes electrophilic substitution reactions instead of addition reactions to preserve its stable aromatic ring. Key reactions include:

1. **Halogenation**

   Benzene reacts with halogens (e.g., Cl_2 or Br_2) in the presence of a catalyst like $FeBr_3$ or $AlCl_3$.

   The catalyst generates the electrophile X^+ (halogen atom).

   $$C_6H_6 + Cl_2\ C\xrightarrow{\text{atalyst Anhy. } AlCl_3\text{ , room temperature}} C_6H_5Cl + HCl$$



2. **Friedel-Crafts Reactions**

   Alkylation: A hydrogen atom in benzene is replaced with an alkyl group.

   Acylation: A hydrogen atom is replaced with an acyl group.

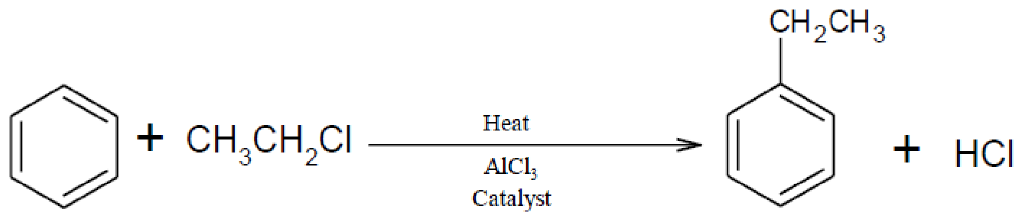
   Both reactions require a catalyst like $AlCl_3$.

$AlCl_3 + Cl_2 \rightarrow AlCl_4 + Cl$ and

$FeBr_3 + Br_2 \rightarrow FeBr_4 + Br$

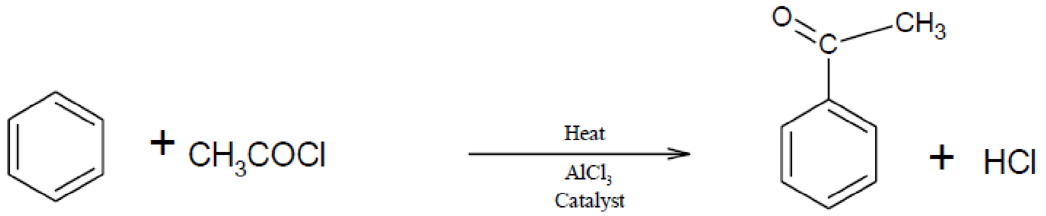
Friedel-Crafts alkylation

$$C_6H_6 + CH_3CH_2Cl\ C\xrightarrow{\text{atalyst Anhy. }AlCl_3\text{ , heat}} C_6H_5CH_2CH_3 + HCl$$



Friedel-Crafts acylation

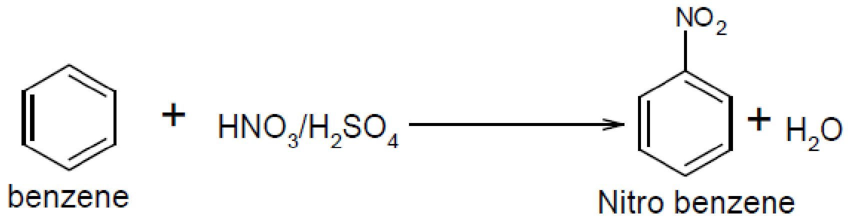
$$C_6H_6 + CH_3COCl\ C\xrightarrow{\text{atalyst Anhy. }AlCl_3\text{ , heat}} C_6H_5COCH_3 + HCl$$



3. **Nitration**

Benzene reacts with a mixture of concentrated nitric acid (HNO_3) and sulphuric acid (H_2SO_4) at 25-60°C to form nitrobenzene.

$$C_6H_6 + HNO_3/H_2SO_4 \longrightarrow C_6H_5NO_2 + H_2O$$



4. Combustion (Reaction with Oxygen)

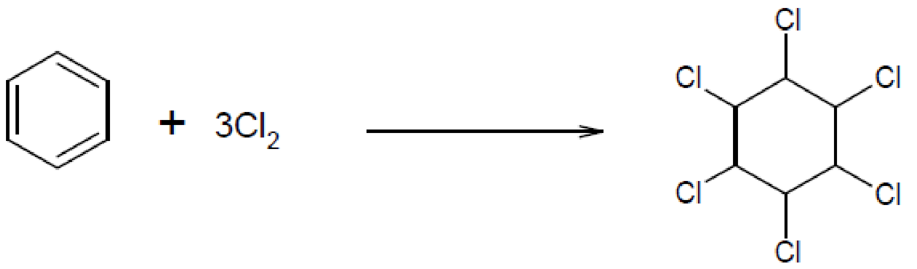
Benzene burns in oxygen to produce carbon dioxide and water, but with a sooty flame due to its high carbon content.

$$2\ C_6H_6 + 15\ O_2 \longrightarrow 12\ CO_2 + 6\ H_2O$$

5. Halogen Addition (Hexachloride Formation)

In the presence of sunlight, benzene reacts with chlorine to form benzene hexachloride (C_6H_6).

$$C_6H_6 + 3\ Cl_2 \xrightarrow{\text{UV light}} C_6H_6Cl_6$$

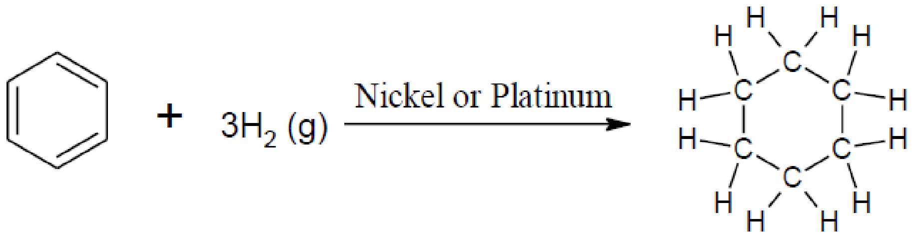


Benzene hexachloride

6. **Hydrogenation**

   Benzene reacts with hydrogen under high pressure and in the presence of a catalyst (e.g., nickel) to form cyclohexane.

   $$C_6H_6 + 3\,H_2 \xrightarrow{\text{Ni, heat, pressure}} C_6H_{12}$$



Uses of Benzene

1. **As a Solvent**

   Industrial Use: Used in making rubber, lubricants, dyes, detergents, drugs, and pesticides due to its ability to dissolve substances.

   Laboratory Use: Used as a solvent for chemical reactions and extractions in research.

2. **Chemical Production**

   Plastics and Resins: Benzene is used to produce **styrene** (for polystyrene plastics) and **phenol** (for epoxy resins).

   Synthetic Fibers: Used to make nylon and other fibres for textiles.

3. **Printing Industry**

   Found in printing inks and cleaning agents, helping to maintain equipment and improve ink quality.

4. **Fuel Additive**

   Added to gasoline to **increase octane rating**, improving engine performance and reducing knocking (its use is now regulated due to toxicity).

5. **Pharmaceuticals**

   Used as an intermediate in the production of drugs and complex molecules.

6. **Other Uses**

   Explosives: Used in manufacturing certain explosives.

   Pesticides: Serves as a starting material for making effective pest control agents.

Activity 8.12 Benzene Structure and Stability

Material needed: Computer and internet, visual aids (videos, handouts, diagrams), molecular model kits or online simulations, paper, and colour-coding supplies

Steps

1. Use the internet to watch a video about the discovery of benzene and its importance in organic chemistry or use the link below:

   https://www.youtube.com/watch?v=xD7Z7SHix-w
2. Organise yourselves into small groups. In your groups, undertake the following. Make notes in your:
3. Discuss the Kekulé structure of benzene (alternating single and bonds).
4. Draw the Kekulé structure of benzene showing alternating single and double bonds.
   a. How many carbon-carbon single and double bonds are shown in this structure?
5. Draw two resonance structures of benzene using colour-coded double bonds.
   a. Why are these two structures not "real" but rather a representation?
6. Explain why benzene is more stable than expected from Kekulé's model.
7. Draw the molecular orbital diagram of benzene showing the arrangement of the delocalized π-electrons.
   a. How does electron delocalization contribute to benzene's stability?
   b. Benzene does not undergo typical addition reactions like alkenes. Instead, it prefers substitution reactions. Why does benzene resist addition reactions despite having π-bonds?

8. Using the concept of resonance energy, explain why benzene is more stable than cyclohexatriene. (Hint: Resonance energy is the difference between the actual energy of benzene and the theoretical energy of its Kekulé structure).
9. Compare the enthalpy of hydrogenation of benzene (208 kJ/mol) with cyclohexene (120 kJ/mol for one double bond). Why is the enthalpy of benzene not three times that of cyclohexene? What does this reveal about its stability?
10. Aromaticity is a concept that explains benzene's structure. According to Hückel's Rule, a molecule is aromatic if it has $4n + 2$ π-electrons. Show how benzene satisfies Hückel's rule. Why is this important for its stability?

Activity 8.13 Exploring Benzene's Substitution Reactions

Steps

1. Use the link below to watch video explaining substitution reactions of benzene:

   https://www.youtube.com/watch?v=9IgmnYooHOQ

2. Use the link below to observe a virtual laboratory platform to simulate benzene reactions with:
   - **a.** Halogens (e.g., chlorine, bromine)
   - **b.** Concentrated HNO_3 (nitration)
   - **c.** Concentrated H_2SO_4 (sulfonation)
   - **d.** Alkyl halides (Friedel-Crafts alkylation)
   - **e.** Acyl halides (Friedel-Crafts acylation).

   Link to virtual laboratory platform: Electrophilic aromatic substitution Section

3. In small groups, share your findings from the virtual lab and videos.

   Focus on questions like:
   - **a.** What happens to benzene during each reaction?
   - **b.** How does the reaction mechanism work?
   - **c.** What patterns do you notice?

d. What is common in all substitution reactions?

e. Why does benzene prefer substitution over addition reactions?

Activity 8.14 Addition Reactions

Materials Needed: Videos or articles explaining benzene's addition reactions, access to virtual lab simulations, laptops.

Steps

1. Use the link https://www.youtube.com/watch?v=pnr9ylSu64c to watch video on addition reactions of benzene with:
   - **a.** Hydrogen in the presence of nickel or platinum catalyst to form cyclohexane.
   - **b.** Halogens like chlorine or bromine in sunlight to form hexahalocyclohexanes.
2. Use a virtual lab to simulate benzene's addition reactions. Link: Electrophilic aromatic substitution Section
3. Create concept maps to visually organise and summarise benzene's reactions, including both addition and substitution reactions.
4. Share with your class for discussion and feedback. Refine it if needed, based on the feedback you receive. Research on the real-world uses of benzene. Exchange your findings with a peer.

ALKANOLS

Introduction and Structure of Alkanols

Alkanols are compounds with a **hydroxyl group (-OH)** attached to an **sp^3-hybridised carbon atom**. They can be viewed in two ways:

1. As alkanes where one hydrogen atom is replaced by a hydroxyl group.
2. As water molecules where one hydrogen atom is replaced by an alkyl group.

Alkanols follow the **general formula $C_nH_{2n+1}OH$** ($n \geq 1$). Examples include:

n = 1: CH_3OH (methanol)

n = 2: C_2H_5OH (ethanol).

Nomenclature of Alkanols

The naming of alkanols follows **IUPAC rules** for clarity and consistency.

Basic Steps for Naming Alkanols

1. Find the longest continuous carbon chain containing the hydroxyl group (-OH).

   Replace the suffix **"-ane"** of the parent alkane with **"-anol"**.
2. Start numbering the chain from the end nearest to the hydroxyl group.

   This gives the hydroxyl group the **lowest possible number**.
3. Identify any substituents (e.g., alkyl groups, halogens).

   Name and number them based on their position on the chain.

   List substituents in alphabetical order.
4. Combine the position number of the hydroxyl group, the parent chain name, and names/positions of substituents.

   Use the suffix **"-anol"** to indicate the presence of the hydroxyl group.

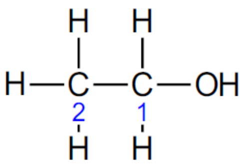
Example

Give the IUPAC name for each compound

- **a.** CH_3CH_2OH
- **b.** $CH_3CH(OH)CH_3$
- **c.** $CH_3C(OH)(CH_3)CH_3$
- **d.** $(CH_3)_2C(OH)CH_2CH_2CH(Br)CH_3$

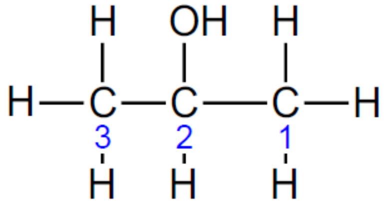
Solution

1. CH_3CH_2OH



- **a.** Parent hydrocarbon: Ethane
- **b.** Replace '-ane' with '-anol': Ethanol
- **c.** The hydroxyl group is on the first carbon, so no number is needed.

2. $CH_3CH(OH)CH_3$

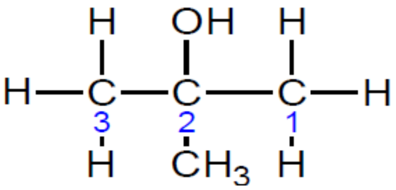


a. Parent hydrocarbon: Propane

b. Number the chain: The hydroxyl group is on the second carbon.

c. Replace '-ane' with '-anol' and indicate the position: Propan-2-ol.

3. $CH_3C(OH)(CH_3)CH_3$



a. Parent hydrocarbon: Propane

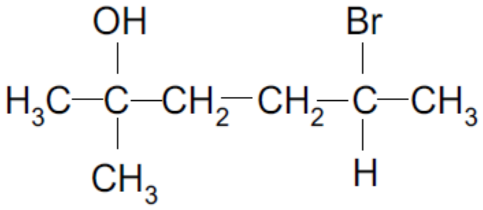
b. Identify and number the longest chain: The hydroxyl group is on the second carbon.

c. Name the substituent: A methyl group is attached to the second carbon.

d. Replace '-ane' with '-anol' and indicate the position: Propan-2-ol.

e. Combine the name of the substituent and the parent name as one name: 2-Methylpropan-2-ol.

4. $(CH_3)_2C(OH)CH_2CH_2CH(Br)CH_3$



a. Parent hydrocarbon: Hexane

b. Identify and number the longest chain: The hydroxyl group is on the second carbon.

c. Name the substituents: A methyl group is attached to the second carbon and a bromine on the fifth carbon.

d. Replace '-ane' with '-anol' and indicate the position: Hexan-2-ol.

e. Combine the names of the substituent in alphabetical order and the parent name -bromo-2-methylhexan-2-ol

Naming of Polyhydroxy Alkanols

For alkanols with more than one hydroxyl group:

Use suffixes like **"-diol"** (2 OH groups) or **"-triol"** (3 OH groups) to show the number of hydroxyl groups.

Indicate the **positions** of the hydroxyl groups using numbers.

Example

Ethane-1,2-diol → Two OH groups on carbons 1 and 2.

Propane-1,2,3-triol → Three OH groups on carbons 1, 2, and 3.

Example

1. 1,2-ethanediol $HOCH_2CH_2OH$

   Parent hydrocarbon: Ethane

   Two hydroxyl groups: '-diol'

   Indicate positions: 1,2-ethanediol.

2. Propan-1,2,3-triol $CH_2(OH)CH(OH)CH_2(OH)$
   - **a.** Parent hydrocarbon: Propane
   - **b.** Three hydroxyl groups: '-triol'
   - **c.** Indicate positions: 1,2,3-propantriol.

Classification of Alkanols

Alkanols are classified based on the number of carbon atoms attached to the carbon bearing the hydroxyl group (-OH). This affects their reactivity and behaviour, especially in oxidation reactions.

1. Primary (1°) Alkanols

   The carbon with the -OH group is attached to one other carbon or none (e.g., methanol).
   The -OH group is at the end of the carbon chain.
   They are the most reactive in oxidation reactions.
   Examples:

Methanol (CH_3OH): No carbon attached to the -OH carbon.

Ethanol (CH_3CH_2OH): The -OH carbon is bonded to one other carbon.

2. Secondary (2°) Alkanols

   The carbon with the -OH group is attached to two other carbons.

   Less reactive in oxidation than primary alkanols.

   Examples:

   Propan-2-ol ($CH_3CH(OH)CH_3$): -OH group on the second carbon in a 3-carbon chain.

   Butan-2-ol ($CH_3CH_2CH(OH)CH_3$): -OH group on the second carbon of a 4-carbon chain.

3. Tertiary (3°) Alkanols

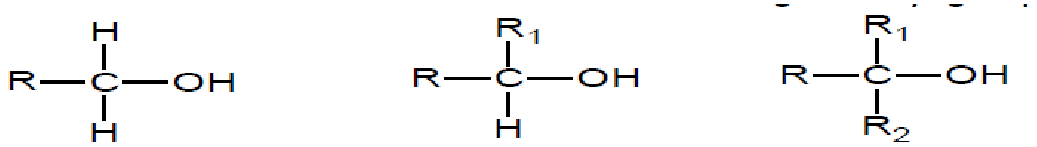
   The carbon with the -OH group is attached to three other carbons.

   They are the least reactive in oxidation because there is no hydrogen on the -OH carbon.

   Examples:

   Tert-butanol ($(CH_3)_3COH$): -OH carbon bonded to three methyl groups.

   2-Methylpropan-2-ol ($(CH_3)_2C(OH)CH_3$): -OH group on a carbon attached to two groups and one ethyl group



primary (1°) secondary (2°) tertiary (3°)

where R = alkyl group

Properties of alkanols

Physical Properties of Alkanols

The hydroxyl group (-OH) in alkanols influences their physical properties, including boiling points, solubility, viscosity, and odour.

1. **Boiling Points**

   Alkanols have higher boiling points than alkanes or ethers of similar molecular mass due to hydrogen bonding between hydroxyl groups.

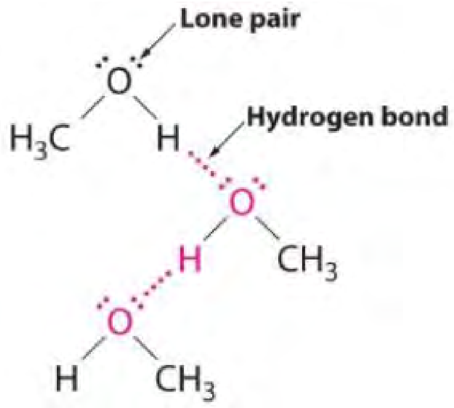


Figure 8.7: hydrogen bonding in methanol

Example: Methanol (CH_3OH) boils at 65°C, while methane (CH_4), with a similar molecular mass, boils at -164°C.

2. **Solubility**

   Short-chain alkanols (e.g., methanol, ethanol) are highly soluble in water because they form hydrogen bonds with water molecules.

   Solubility decreases as the carbon chain length increases because of the hydrophobic nature of the alkyl group.

   Examples:

   Methanol, Ethanol, Propanol → Completely soluble in water.

   Octanol ($C_8H_{17}OH$) → Limited solubility due to a long hydrophobic chain.

3. **Viscosity**

   Viscosity (resistance to flow) increases as the carbon chain length and molecular weight increase.

   Reasons

   a. Molecular Weight: Larger molecules have greater surface areas, leading to intermolecular interactions.

   b. Intermolecular Forces: Hydrogen bonding between hydroxyl groups increases viscosity.

   Trend

   a. Shorter alkanols → Lower viscosity.

   b. Longer alkanols → Higher viscosity.

4. **Odour**

   Alkanols have distinct odours depending on their molecular structure:

Methanol → Sweet, pungent smell.

Ethanol → Alcoholic smell (found in beverages).

Isopropanol → Sharp, medicinal odour.

Higher alkanols → Fruity, heavier odours (e.g., butanol, pentanol).

a. The hydroxyl group influences volatility and sensory perception.

b. Larger molecules are less volatile and produce heavier odours.

Chemical Properties of Alkanols

Alkanols react due to the presence of the hydroxyl group (-OH), which influences their chemical behaviour.

Acidity of Alkanols

Alkanols are weakly acidic because the hydroxyl group (-OH) can donate a proton (H^+) to form an alkoxide ion (RO^-).

Factors Affecting Acidity

1. **Polarity of the O-H Bond**

   The O-H bond is polar because oxygen is more electronegative than hydrogen. This polarity helps release the proton.

2. **Inductive Effects**

   Electron-withdrawing groups (e.g., halogens) increase acidity by stabilizing the negative charge on the alkoxide ion.

   Electron-donating groups (e.g., alkyl groups) reduce acidity by increasing electron density on oxygen, making proton release harder.

3. **Resonance Stabilization**

   Alkanols like phenol (aromatic alcohols) are more acidic because the conjugate base is stabilized by resonance.

   Example: Phenol (pKa ~10) is much more acidic than aliphatic alkanols.

   The order of acidity for alkanols generally follows:

$$\underset{\xrightarrow[\text{Increasing acidity}]{\hspace{12em}}}{1^{0}\text{ alkanols} < 2^{0}\text{ alkanols} < 3^{0}\text{ alkanols}}$$

Order of Acidity

Primary alkanols > Secondary alkanols > Tertiary alkanols

Fewer electron-donating alkyl groups in primary alkanols increase acidity.

As the carbon chain length increases, acidity decreases due to the stronger electron-donating effect of alkyl groups.

Example:

Methanol (CH_3OH) > Ethanol (CH_3CH_2OH) > Propanol ($CH_3CH_2CH_2OH$)

Amphoteric Nature of Alkanols

Alkanols can show both **acidic** and **basic** behaviour in specific reactions.

1. **Reaction with Bases**

   Alkanols react with strong bases like sodium hydroxide (NaOH) under specific conditions to form alkoxides and water.

   This occurs when the hydroxyl group (-OH) donates a proton (H^+).

   Example: $R{-}OH + NaOH \rightarrow R{-}O^-Na^+ + H_2O$

2. **Reaction with Carbonates and Bicarbonates**

   Alkanols are much weaker acids than carboxylic acids and typically do **not react** with carbonates (e.g., Na_2CO_3) or bicarbonates (e.g., $NaHCO_3$) under normal conditions.

   However, carboxylic acids do react with these compounds to produce:

   Carbon dioxide (CO_2), Water (H_2O), and salts (alkanoates)

   Example: $2R{-}COOH + Na_2CO_3 \rightarrow 2R{-}COONa + H_2O + CO_2$

Reaction with Sodium Metal

Alkanols react with sodium metal to form alkoxide ions and release hydrogen gas. This reaction demonstrates the weak acidity of alkanols.

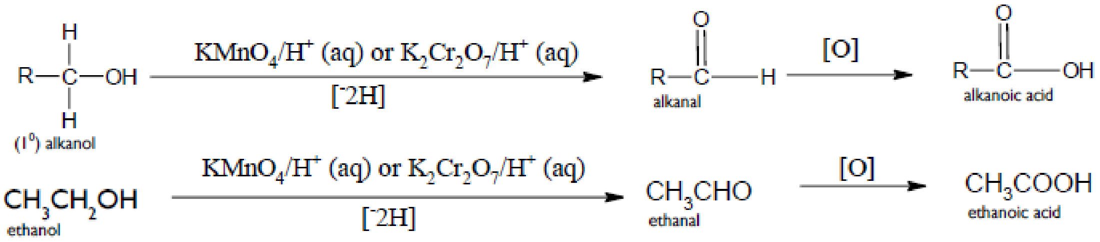
Reaction: $2ROH + 2Na \rightarrow 2RONa + H_2$

The hydroxyl group (-OH) donates a proton (H^+) to sodium, forming an alkoxide (RONa) and hydrogen gas (H_2).

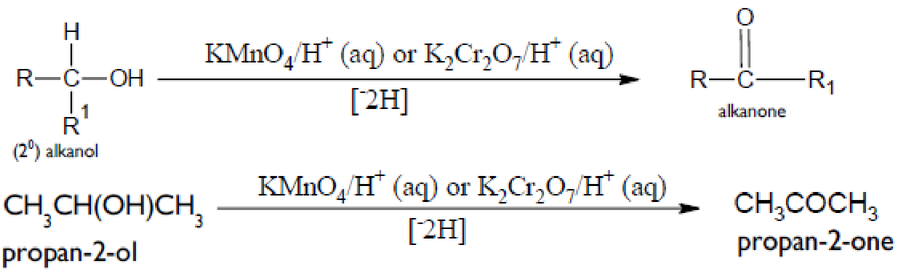
Oxidation Reactions of Alkanols

The oxidation of alkanols depends on their type (primary, secondary, or tertiary) and is carried out using oxidizing agents like **potassium dichromate (K2Cr2O7)** or **potassium permanganate (KMnO4)** in the presence of an acid (e.g., **H2SO4**).

Primary alkanols → Oxidize to **aldehydes** → Further oxidize to **carboxylic acids**.



Secondary alkanols → Oxidize to **ketones**.



Tertiary alkanols → Do not undergo oxidation because there is no hydrogen atom on the carbon bearing the -OH group.

Oxidation is an important property used to identify and distinguish between different types of alkanols.

Combustion Reactions of Alkanols

Alkanols burn in oxygen to produce **carbon dioxide (CO2)**, **water (H2O)**, and release **heat**. This exothermic reaction makes alkanols useful as fuels.

Complete Combustion

Occurs when there is **sufficient oxygen**.

General equation: $C_nH_{2n+1}OH + (n + \frac{1}{2})O_2 \rightarrow n\,CO_2 + (n + 1)H_2O$

Example: Ethanol burns with a blue flame, releasing a large amount of heat.

Incomplete Combustion

Happens when there is **insufficient oxygen**.

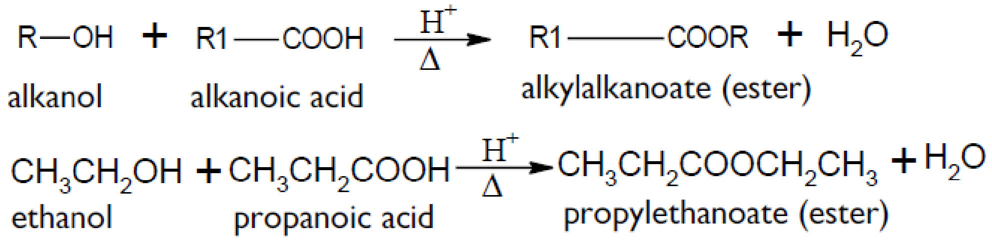
Produces **carbon monoxide (CO)**, **carbon (soot)**, and water instead of CO_2.

It is **less efficient** and releases less energy.

Reaction with Alkanoic Acids

Alkanols react with alkanoic acids to form esters (alkyl alkanoates) and water. This process is called esterification.

General Reaction: Alkanol + Alkanoic Acid → Ester + Water



$$R{-}OH + R1{-}COOH \xrightarrow[\Delta]{H^+} R1{-}COOR + H_2O$$

$$CH_3CH_2OH + CH_3CH_2COOH \xrightarrow[\Delta]{H^+} CH_3CH_2COOCH_2CH_3 + H_2O$$

Catalyst: The reaction requires concentrated sulphuric acid (H_2SO_4) as a catalyst. The acid speeds up the reaction.

It also helps remove water, which pushes the reaction forward to form more ester. Importance: Esterification is used to produce esters, which are important in:

- Fragrances
- Flavourings
- Solvents

Preparation of Alkanols in the Laboratory

1. Hydrolysis of Halogenoalkanes:

   Halogenoalkanes (alkyl halides) are alkanes with one or more halogen atoms.

   Reaction: Heated with aqueous alkali (e.g., NaOH or KOH) under reflux.

   Process:

   Hydroxide ion (OH^-) replaces the halogen atom (nucleophilic substitution).

   Reflux ensures the reaction completes.

   The alkanol (alcohol) is then distilled.

   Example: $R{-}X + NaOH \rightarrow R{-}OH + NaX$

2. Hydration of Alkenes

   Alkenes react with water to form alkanols.

   Conditions: Water is added in the presence of an acid catalyst (e.g., H_3PO_4, phosphoric acid).

   Process

   A carbocation intermediate form.

   Water is added to produce the alkanol.

   Example: $C_2H_4 + H_2O \rightarrow C_2H_5OH$

   Ethene → Ethanol

3. Fermentation of Sugars

   Sugars are converted into ethanol and carbon dioxide using yeast.

   Process:

   Enzyme (zymase) in yeast breaks down sugars.

   Reaction: $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$

   Ethanol is separated from the mixture by distillation.

Test for Alkanols (Alcohols)

1. **Lucas Test**

   The Lucas test is used to differentiate primary, secondary, and tertiary alcohols based on their reactivity with Lucas reagent (a mixture of $ZnCl_2$ and concentrated HCl). The reaction produces alkyl chlorides, which cause turbidity (cloudiness).

   Procedure

   Add 2-3 mL of Lucas reagent to a test tube.

   Add 1-2 drops of the alcohol to be tested.

   Shake the test tube and observe how quickly turbidity forms.

   Observations

   Primary Alcohols:

   React very slowly or not at all at room temperature.

   No turbidity is seen immediately.

   $R{-}CH_2OH + HCl \rightarrow R{-}CH_2Cl + H_2O$

   Secondary Alcohols

   React in 5-10 minutes.

Turbidity forms after some time.

$R_2CH{-}OH + HCl \rightarrow R_2CH{-}Cl + H_2O$

Tertiary Alcohols

React immediately.

Turbidity forms instantly.

$R_3C{-}OH + HCl \rightarrow R_3C{-}Cl + H_2O$

2. **Iodoform Test for Alkanols**

The Iodoform test is used to identify alcohols with the structure R-CH(OH)-CH_3. This includes ethanol and secondary alcohols with a methyl group next to the hydroxyl group.

Reagents: Iodine (I_2) and sodium hydroxide (NaOH)

Procedure

1. Add a few drops of the alcohol to a test tube.
2. Add a few drops of iodine solution.
3. Add sodium hydroxide dropwise until the iodine colour disappears.
4. Warm the mixture gently if needed.

Observation

A yellow precipitate of iodoform (CHI_3) forms, with a distinctive antiseptic smell.

Reaction

$R{-}CH(OH){-}CH_3 + 3I_2 + 4NaOH \rightarrow CHI_3 + R{-}COONa + 3NaI + 3H_2O$

3. **Ester Test**

To confirm the presence of an alcoholic group in a compound.

Procedure

Add 1 ml of the organic liquid to a clean, dry test tube.

Add 1 ml of glacial acetic acid and 2-3 drops of concentrated sulphuric acid.

Heat the mixture in a water bath for 10 minutes.

Pour the hot mixture into a beaker containing cold water.

Smell the water in the beaker.

Observation

A fruity smell confirms the presence of an alcoholic group due to the formation of an ester.

Everyday Uses of Alkanols

1. Household Uses
   - a. Cleaning Agents: Ethanol and isopropanol dissolve grease and stains, disinfect surfaces, and are used in products like window cleaners and multi-surface cleaners.
   - b. Personal Care Products: Ethanol is found in hand sanitizers, perfumes, and hairsprays due to its antimicrobial properties, ensuring hygiene and freshness.
2. Medical and Pharmaceutical Uses
   - a. Antiseptics and Disinfectants: Ethanol and isopropanol are used to clean wounds, disinfect skin, and sterilize instruments.
   - b. Pharmaceutical Solvents: Alkanols dissolve and extract active ingredients, aiding in the formulation of medications like pills and tinctures.
   - c. Preservatives: Ethanol ensures the stability and longevity of pharmaceutical products.
3. Industrial Uses
   - a. Solvents: Used in paints, varnishes, and coatings for smooth application.
   - b. Fuel Additives: Ethanol improves combustion efficiency and reduces emissions in gasoline. It's also used as bioethanol, a renewable fuel source.
   - c. Chemical Synthesis: Alkanols are intermediates in producing esters, ethers, and aldehydes.
4. Food and Beverage Industry
   - a. Alcoholic Beverages: Ethanol is the main alcohol in beer, wine, and spirits, produced by fermenting sugars with yeast.
   - b. Flavouring Extracts: Alkanols extract flavours from natural sources for food and beverages.
5. Laboratory and Research Uses
   - a. Solvents: Useful for dissolving organic compounds in chemical reactions.
   - b. Cleaning and Sterilization: Ethanol cleans and sterilizes equipment, killing microorganisms and evaporating quickly.

  c. Sample Preparation: Helps prepare samples for analysis by dissolving organic compounds.

  d. Biomolecule Precipitation: Ethanol precipitates DNA and RNA, aiding in biochemical research.

6. Automotive Uses

  a. Antifreeze: Methanol and ethanol lower the freezing point of radiator coolant.

  b. Fuel: Ethanol serves as a high-energy, clean-burning fuel for racing cars and rockets.

Activity 8.15 Exploring Alkanols: Formula, Classification, Isomerism, and Naming.

Steps

1. **a.** Write down the general formula for alkanols.

   b. Write down the formulae for methanol, ethanol, propanol, butanol, pentanol, and hexanol.

2. In pairs compare and discuss your answers.

3. Classify the following alkanols as primary, secondary, or tertiary, based on the carbon atom bonded to the hydroxyl group:

   CH_3OH, CH_3CH_2OH, $CH_3CH_2CH_2OH$, $CH_3CH(OH)CH_3$, $CH_3CH_2CH(OH)CH_3$, $(CH_3)_3COH$, $(CH_3)_2C(OH)CH_3$

4. In pairs compare and discuss your examples of primary, secondary, and tertiary alkanols.

5. Draw structural isomers of alkanols for each class.

6. Explain the differences between primary, secondary, and tertiary alkanols, showcasing structural isomers.

7. Review the IUPAC naming rules for alkanols, focusing on:

   a. Identifying the longest carbon chain.

   b. Numbering the chain to give the hydroxyl group the lowest number.

8. In pairs, practice naming alkanols: $CH_3CH_2CH_2OH$, $CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2OH$, $CH_3\text{-}CH(OH)\text{-}\ CH_2\text{-}CH_3$, $(CH_3)_2CH\text{-}CH_2OH$.

9. Draw out the structural formula of each of the alcohols in question 8.

10. In pairs compare and discuss your answers.

11. Write a short essay on the differences between primary, secondary, and tertiary alkanols, including examples and their structural differences.

Activity 8.16 Exploring the Preparation and reactions of Alkanols

Materials needed: laptop and internet

Steps

1. Use the internet and watch a video on the preparation of ethanol or use the link: https://www.youtube.com/watch?v=vWBf9s56H7E

2. In small groups, discuss ethanol preparation using both fermentation and hydration methods.

3. Design an experiment to demonstrate the esterification reaction of an alkanol with a carboxylic acid. Conduct the experiment, record your observations, and present your findings to the class.

4. Write a report on the oxidation reactions of primary, secondary, and tertiary alkanols. Include balanced chemical equations and explain the products formed.

5. In your groups, compare and contrast the fermentation and hydration methods, considering:

  a. Reactants and conditions.
  b. Yield and efficiency.
  c. Environmental and industrial impacts.

6. In your groups, explore the physical properties of alkanols, such as:

  a. Boiling points (increasing with chain length).
  b. Solubility (in water, decreasing with chain length).
  c. Density

Activity 8.17 Alcohol Breath Analysers and Biofuels

Material needed: laptop and internet

Steps

1. List at least three common uses of ethanol in everyday life. Include pictures or diagrams to illustrate each use.
2. Go online and watch video on breath analysers or use the link:

   https://www.youtube.com/watch?v=3z5nMtqvGks
3. Organise yourselves into groups of no more than five. In your groups, research the mechanism of alcohol breath analysers, focusing on;
   - **a.** The redox reaction involving potassium dichromate and ethanol.
   - **b.** The chemical changes and colour shift in the reaction.
4. Brainstorm and sketch your own alcohol breath analyser using potassium dichromate.
5. Research the advantages and disadvantages of biofuels, such as ethanol and biodiesel.

   Focus on;
   - **a.** Energy efficiency compared to fossil fuels.
   - **b.** Environmental impact and sustainability.
6. Share your group's findings and insights with the class in a discussion.

ALKANOIC ACIDS

Introduction to Alkanoic Acids

Alkanoic acids, also called carboxylic acids, are organic compounds with a carboxyl group

(R-COOH). The carboxyl group contains:

1. A carbonyl group (C=O).
2. A hydroxyl group (-OH) attached to the same carbon atom.

General Formula

The homologous series of alkanoic acids follows the formula: $CnH_{2n+1}COOH$, where $n \geq 0$.

Examples

n=0: HCOOH (methanoic acid)

n=1: CH_3COOH (ethanoic acid)

n=2: C_2H_5COOH (propanoic acid)

n=3: C_3H_7COOH (butanoic acid)

Properties

Weak Acids

Alkanoic acids partially dissociate in water, meaning they do not fully release hydrogen ions (H^+).

This explains their moderate conductivity in solution and reactivity with bases to form salts and water.

Acidity Factors

The presence of electron-withdrawing groups (e.g., halogens) increases acidity by stabilising the conjugate base.

Nomenclature of Alkanoic Acids

The naming of alkanoic acids follows IUPAC guidelines.

Rule

1. Start with the name of the corresponding alkane.
   Replace the suffix "-e" with "-oic acid".
2. Priority of the Carboxyl Group:
   The carboxyl group (-COOH) is always assigned the highest priority.
   It is always located at the end of the carbon chain and designated as carbon number 1.

Examples

Methane → Methanoic acid (HCOOH)

Ethane → Ethanoic acid (CH_3COOH)

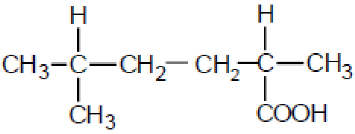
Propane → Propanoic acid (C_2H_5COOH)

$CH_3CH(Cl)CH_2COOH$

3-Chlorobutanoic acid

$CH_3CH{=}CHCH_2COOH$

3-pentenoic acid



$CH_3-\underset{CH_3}{\overset{H}{C}}-CH_2-CH_2-\underset{COOH}{\overset{H}{C}}-CH_3$

2,5-dimethylhexanoic acid

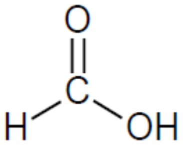
Cl—COOH

Chloromethanoic acid

Common Names of Alkanoic Acids

The common names of alkanoic acids often come from their natural sources:

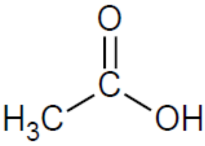
Formic acid: Named after “formica” (Latin for ant), as it is found in ant venom.



formic acid

Acetic acid: Comes from **“acetum”** (Latin for vinegar), as it is a key component of vinegar.

These common names are widely used in both industry and daily life.

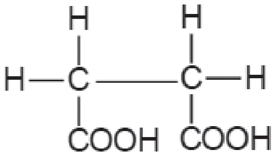


acetic acid

Naming Alkanoic Acids with Multiple Carboxyl Groups

For alkanoic acids with two carboxyl groups, the suffix “-dioic acid” is used.

Example: A compound with two carboxyl groups on an ethane chain is called ethanedioic acid.



$H-\underset{COOH}{\overset{H}{C}}-\underset{COOH}{\overset{H}{C}}-H$

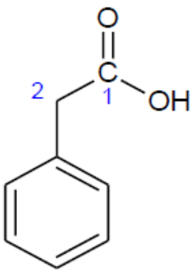
Ethanedioic acid

HOOC—CH_2—COOH

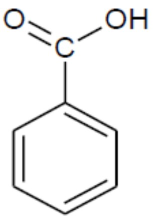
1,3-propanedioic acid

Naming Alkanoic Acids Attached to Rings

When an alkanoic acid group is attached to a ring, the suffix **"-alkanoic acid"** is used. The carbon with the carboxyl group is always numbered as **carbon 1**.



2-phenylethanoic acid



Phenylmethanoic acid (benzoic acid)

Structure of Alkanoic Acids

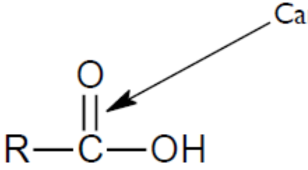
1. **Carboxyl Group (-COOH)**

   Contains a carbon atom

   Double-bonded to an oxygen atom (C=O).

   Single-bonded to a hydroxyl group (-OH).

   General formula: R-COOH, where R can be an alkyl group, hydrogen, or a phenyl group.



Carbonyl group (containing the carbonyl carbon and the carbonyl oxygen)

$$R-\overset{\overset{\displaystyle O}{\|}}{C}-OH$$

2. **Resonance and Stability**

   The carboxyl group is planar and exhibits resonance.

   Resonance involves electron delocalization between the carbonyl and hydroxyl groups, stabilizing the molecule.

   This stability enhances the acidity of alkanoic acids by stabilizing the conjugate base after proton loss.

3. **Molecular Geometry**

   The carbon in the carboxyl group is sp^2 hybridized, creating a trigonal planar structure.

   Bond angles around the carbonyl carbon are approximately 120°.

4. **Hydrogen Bonding**

   The hydroxyl group (-OH) forms hydrogen bonds, which affect:

   Boiling points: Higher due to strong intermolecular forces.

   Solubility: High solubility in water for smaller alkanoic acids.

Properties of Alkanoic Acids

The physical properties of alkanoic acids depend on the length of their carbon chain and the presence of the carboxyl group.

1. **Boiling and Melting Points**

   Alkanoic acids have higher boiling and melting points compared to other organic compounds of similar molecular weight due to strong hydrogen bonding.

   Hydrogen Bonding:

   a. Alkanoic acids form dimers through two hydrogen bonds between:

      The hydrogen atom of the hydroxyl group (-OH).

      The oxygen atom of the carbonyl group (C=O) of another molecule.

      This dual bonding leads to higher boiling and melting points compared to alkanols of similar molecular weight.

   b. Chain Length Effect

      Boiling and melting points increase as the carbon chain lengthens due to stronger Van Waals forces.

      Example

      Ethanoic acid (118°C) has a significantly higher boiling point than ethanol (78°C) despite their similar molecular weights.

2. **Solubility**

   Lower alkanoic acids, like methanoic acid and ethanoic acid, are highly soluble in water due to hydrogen bonding with water molecules.

   As the carbon chain length increases, the hydrophobic alkyl group reduces solubility.

   Example: Butanoic acid is less soluble than ethanoic acid.

3. **Density**

   Alkanoic acids are generally denser than water.

Density decreases as the carbon chain length increases.

Example: Ethanoic acid has a density of ~1.049 g/cm³, while hexanoic acid is ~0.927 g/cm³.

4. **Physical State**

   Lower members (e.g., methanoic acid, ethanoic acid) are liquids at room temperature.

   Higher members (e.g., octanoic acid, decanoic acid) are waxy solids due to stronger Van der Waals forces.

5. **Odour**

   Alkanoic acids have strong, pungent odours.

   Example

   Ethanoic acid: Vinegar-like smell.

   Butanoic acid**:** Rancid butter smell.

   Odour becomes less intense as molecular weight increases.

Chemical Properties of Alkanoic Acids

Alkanoic acids are highly reactive due to their carboxyl group (-COOH), which can:

Donate a proton (acidic behaviour).

Undergo nucleophilic attack at the carbonyl carbon.

1. **Acid-Base Reactions**

   Alkanoic acids partially dissociate in water, forming hydronium ions (H_3O^+) and carboxylate ions ($R\text{-}COO^-$).

   React with strong bases (e.g., NaOH) to produce salts and water in neutralization reactions.

   Example: $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

   Importance: Used in soap production, where long-chain alkanoic acids react with bases to form soap molecules.

2. **Esterification Reactions**

   Alkanoic acids react with alcohols in the presence of an acid catalyst to form esters and water.

   Example: $CH_3COOH + C_2H_5OH \rightarrow CH_3COOC_2H_5 + H_2O$

   Widely used in producing fragrances, flavourings, and plastics.

Equilibrium Reaction: Yield can be increased by removing water or using an excess of one reactant.

3. **Decarboxylation Reactions**

   Alkanoic acids lose a carbon dioxide molecule (CO_2) to form hydrocarbons when heated, often with a catalyst.

   Example: $CH_3COOH \rightarrow CH_4 + CO_2$

   Significance: Crucial in metabolic pathways in biochemistry.

4. **Formation of Amides (Amidation)**

   Alkanoic acids react with amines to form amides, often requiring heat and a dehydrating agent.

   Example: $CH_3COOH + NH_3 \rightarrow CH_3CONH_2 + H_2O$

   Uses: Important in synthesising pharmaceuticals and polymers.

Tests for Alkanoic Acids

1. **Litmus Test**

   Procedure: Dip blue litmus paper into the sample solution.

   Observation: Blue litmus turns red, indicating the presence of an acid.

2. **Reaction with Sodium Bicarbonate**

   Procedure: Add sodium bicarbonate ($NaHCO_3$) to the sample solution.

   Observation: Effervescence (bubbling) indicates carbon dioxide (CO_2) release, confirming a carboxylic acid.

   Reaction:$RCOOH + NaHCO_3 \rightarrow RCOONa + H_2O + CO_2$

3. **Esterification Test**

   Procedure: Mix the sample with ethanol and a few drops of concentrated sulphuric acid (H_2SO_4), then gently heat.

   Observation: A sweet, fruity smell confirms the formation of an ester, indicating a carboxylic acid.

   Reaction:$RCOOH + C_2H_5OH \rightarrow RCOOC_2H_5 + H_2O$

4. **Reaction with Magnesium**

   Procedure: Add a small piece of magnesium ribbon to the sample.

   Observation: Effervescence indicates hydrogen gas (H_2) release, confirming a carboxylic acid.

   Reaction:$2RCOOH + Mg \rightarrow (RCOO)_2 Mg + H_2$

5. **Formation of a Salt**

   Procedure: Add the sample to sodium hydroxide (NaOH).

   Observation: Formation of a salt and water confirms a carboxylic acid.

   Reaction:$RCOOH + NaOH \rightarrow RCOONa + H_2O$

Preparations of Alkanoic Acids

1. **Oxidation of Primary Alcohols**

   Reaction: Primary alcohols are oxidised to carboxylic acids using strong oxidising agents like $KMnO_4$ or $K_2Cr_2O_7$ in an acidic medium.

   Equation: $R{-}CH_2OH + 2[O] \rightarrow R{-}COOH + H_2O$

   Example: $CH_3CH_2OH + 2[O] \rightarrow CH_3COOH + H_2O$

   Conditions: Acidic medium (e.g., dilute H_2SO_4) and heating under reflux.

2. **Oxidation of Aldehydes**

   Reaction: Aldehydes are oxidised to carboxylic acids using similar oxidising agents.

   Equation: $R{-}CHO + [O] \rightarrow R{-}COOH$

   Example: $CH_3CHO + [O] \rightarrow CH_3COOH$

   Conditions: Acidic medium and heating under reflux.

3. **Hydrolysis of Nitriles**

   Reaction: Nitriles are hydrolysed to carboxylic acids in the presence of an acid or base.

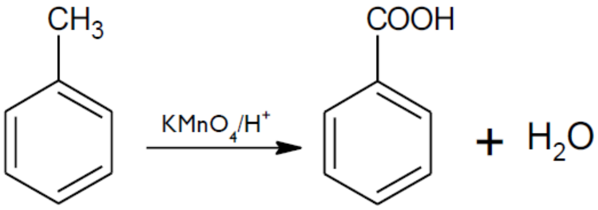
   Equation: $R{-}CN + 2H_2O \rightarrow R{-}COOH + NH_3$

   Example: $CH_3CN + 2H_2O \rightarrow CH_3COOH + NH_3$

   Conditions: Acidic or basic medium and heating.

4. **Oxidation of Alkylbenzenes**

   Reaction: Alkylbenzenes are oxidised to aromatic alkanoic acids using strong oxidising agents like **KMnO4**.



5. **Hydrolysis of Esters**

   Reaction: Esters are hydrolysed to carboxylic acids in the presence of an acid or base.

   Equation: $R{-}COOR^1 + H_2O \rightarrow R{-}COOH + R^1{-}OH$

   Example: $CH_3COOCH_3 + H_2O \rightarrow CH_3COOH + CH_3OHCH_3$

   Conditions: Acidic or basic medium and heating.

Uses of Alkanoic Acids

1. **Preservatives**: Benzoic acid is used in pharmaceuticals to prevent microbial growth.
2. **Flavouring Agents:** Citric acid and lactic acid enhance tartness and flavour in foods and drinks.
3. **Soaps and Detergents**: Fatty acids like stearic acid are used to make soaps and detergents.
4. **Medicinal Compounds**: Salicylic acid is used to produce aspirin and other anti-inflammatory medications.
5. **Fragrances**: Alkanoic acids are used to synthesize esters, key components in perfumes.
6. **Animal Feed**: Propionic acid acts as a preservative to prevent mould in animal feed.

Activity 8.18 Exploring Alkanoic Acids

Steps

1. Organise yourselves into small groups. In your groups, research on preparation, properties and uses alkanoic acid.
2. Create a visual map illustrating the production methods of alkanoic acids (e.g., oxidation, hydrolysis).
3. Compare the physical properties (e.g., boiling points, solubility) of different alkanoic acids.
4. Present your work to the class for discussion and feedback

Activity 8.19 Investigating Reactions of Alkanoic Acids

Materials needed: Test tubes, alkanoic acids (e.g., ethanoic acid, propanoic acid), NaOH, $NaHCO_3$, Na_2CO_3, NH_3, ROH, $LiAlH_4$

Steps

1. In small groups, research the expected reaction between alkanoic acid and the following substances: NaOH, $NaHCO_3$, ROH.
2. Design a simple test tube experiment observing and recording results (e.g., colour changes, gas evolution, or formation of precipitates).
3. Analyse your results and determine whether they match the expected reaction.
4. Draw conclusions about the chemical behaviour of alkanoic acids.
5. Present your research, experimental setup, observations, and conclusions to the class for discussion and feedback

Activity 8.20 Testing for Alkanoic Acids

Materials needed: Test tubes, ethanoic acid (or other alkanoic acid), ethanol (alkanol), concentrated sulphuric acid (catalyst), water bath

Steps

Test for alkanoic acids by reacting them with alkanols to produce esters.

Procedure

- **a.** Add a small amount of ethanoic acid and ethanol to a test tube.
- **b.** Add 2-3 drops of concentrated sulphuric acid.
- **c.** Heat the mixture gently in a water bath for 5-10 minutes.
- **d.** Observe the reaction and note any changes (e.g., smell of a fruity ester).
- **e.** Record observations in a provided template, including:
  - **i.** Initial setup details.
  - **ii.** Observations during heating.
  - **iii.** Final product (e.g., ester smell).

2. Discussion: Other Tests for Alkanoic Acids
   - **a.** Litmus Test
   - **b.** Reaction with Sodium Bicarbonate
   - **c.** Reaction with Magnesium

Activity 8.21 Exploring Alkanoic Acids: Discussion, Reflection, and Differentiated Learning

Steps

1. Organise yourselves into groups. In your groups, share what you already know about alkanoic acids.
2. Using visual aids (e.g., reaction diagrams, flowcharts), describe how alkanoic acids are made (e.g., oxidation of alcohols, hydrolysis of nitriles).
3. Using real-life examples and multimedia resources (e.g., videos, pictures), highlight everyday applications of alkanoic acids. (e.g., preservatives, soaps, medicines).

EXTENDED READING

• https://byjus.com/chemistry/alkanes	
• https://testbook.com/chemistry/alkanes	
• https://chemistnotes.com/organic/alkenes/	
• https://en.wikipedia.org/wiki/Organic_chemistry	

REVIEW QUESTIONS

Review Questions 8.1

1. What is the general formula for alkanes?
2. Name one use of methane.
3. Write the molecular formula for 2-methylpropane.
4. Why do alkanes with longer carbon chains have higher boiling points?
5. Which alkane fractions from crude oil are used as fuel for vehicles?
6. Explain how the structure of branched alkanes affects their boiling points compared straight-chain alkanes.
7. Explain why alkanes are insoluble in water but soluble in non-polar solvents.
8. Explain why propane and butane are suitable for use in Liquefied Petroleum Gas (LPG).
9. Analyse the trends in melting and boiling points of alkanes from C1 to C10 and relate them to their molecular structure.
10. Evaluate the environmental impact of using alkanes as fuels and suggest alternatives to minimise these effects.

Review Questions 8.2

1. What type of bond is present in alkenes and alkynes?
2. Are alkenes and alkynes soluble in water?
3. Draw the structure of 2-pentyne.
4. Why do alkenes and alkynes have low boiling points compared to alcohols?
5. Explain the geometry around the double bond in alkenes and the triple bond in alkynes.
6. Compare the reactivity of alkanes, alkenes, and alkynes.
7. Compare the bond lengths and strengths of C–C, C=C, and C≡C.
8. Analyse how branching affects the boiling points of alkenes and alkynes.

Review Questions 8.3

1. a. What is the molecular formula of benzene?
   b. Name the type of bonding found in benzene.
2. a. What is the functional group of alkanols?
   b. Name the simplest alkanol.
3. Explain why benzene is more stable than expected based on its structure.
4. Compare the boiling points of benzene and water.
5. Write the IUPAC name for the compound $CH_3CH_2CH(OH)CH_3$.
6. Compare the solubility of ethanol and hexanol in water.
7. Predict how benzene's structure affects its reactivity in electrophilic substitution reactions.
8. Analyse the environmental impact of benzene's use in industrial applications.
9. Explain how the Lucas test distinguishes between primary, secondary, and tertiary alkanols.
10. Analyse the environmental and economic benefits of using bioethanol as a fuel.
11. Design an experiment to differentiate benzene from alkanes in terms of reactivity.
12. Design an experiment to compare the boiling points of methanol, ethanol, and propanol.
13. Evaluate the effectiveness of ethanol as a solvent in pharmaceutical preparations compared to water.

Review Questions 8.4

1. What functional group is present in alkanoic acids?
2. Name one method used to prepare alkanoic acids.
3. Are alkanoic acids strong or weak acids?
4. What gas is produced when alkanoic acids react with sodium bicarbonate?
5. Explain the structure of the carboxyl group in alkanoic acids.
6. Write the reaction for the test of ethanoic acid with sodium bicarbonate.

7. Why does the carboxyl group exhibit resonance?
8. Compare the hydrolysis of nitriles and the oxidation of aldehydes for preparing alkanoic acids.
9. How does chain length affect the solubility of alkanoic acids in water?
10. Compare the use of benzoic acid and citric acid in the food industry.
11. Explain why effervescence is observed when alkanoic acids react with sodium bicarbonate.
12. Predict how the resonance in the carboxyl group affects the acidity of alkanoic acids.
13. Analyse the advantages and disadvantages of using oxidation versus hydrolysis in industrial alkanoic acid preparation.
14. Predict the effect of adding an electron-withdrawing group (e.g., chlorine) to a carboxylic acid.
15. Propose a method to confirm the presence of CO_2 in the reaction with sodium bicarbonate.

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GLOSSARY

Acetylene: A simple alkyne with a triple bond, used in welding and synthesis.

Acid salt: a salt that still contains replaceable hydrogen ions

Acid strength: is the tendency of a hydrogen halide to donate a proton (H^+) in an aqueous solution.

Acidic oxides: these are oxides found in non-metals like silicon dioxide (SiO_2), phosphorus pentoxide (P_4O_{10}), and sulphur dioxide (SO_2); react with water to form acids.

Addition reactions: Reactions where atoms add across the double bond of alkenes.

Amphoteric oxides: these are found in elements like aluminium oxide (Al_2O_3); react with both acids and bases.

Analyte: the solution of unknown concentration in a titration

Antiseptics: Alcohols used to kill microorganisms and sanitise surfaces.

Aromatic hydrocarbon: A hydrocarbon with a stable ring structure and delocalised electrons.

Atomic radius: the distance from the nucleus to the outermost electron shell.

Basic oxides: are found in metals like sodium oxide (Na_2O) and magnesium oxide (MgO); react with water to form hydroxides.

Basic salt: a salt formed from the partial neutralisation of a polybasic base, where one or more hydroxyl groups remain

Bitter taste: a characteristic of bases (e.g., soap)

Bond angle: The angle between two adjacent bonds in a molecule.

Born-haber cycle: a series of hypothetical steps that describes the formation of an ionic compound from its constituent elements in their standard states.

Carboxyl group (-COOH): The functional group that defines carboxylic acids.

Chlorides: these are compounds formed when elements combine with chlorine.

Complex salt: a salt containing complex ions, typically involving coordination compounds

Covalent bond: A chemical bond formed by the sharing of electrons between two atoms.

Covalent chlorides: compounds found in non-metals like silicon tetrachloride ($SiCl_4$), phosphorus pentachloride (PCl_5), and sulphur dichloride (S_2Cl_2); hydrolyse in water to form acidic solutions.

Covalent hydrides: contain non-metals and hydrogen, such as phosphine (PH_3). Typically, gaseous.

Delocalised π-electrons: Electrons spread over the ring structure, stabilizing benzene.

Double bond: A bond involving one sigma bond and one pi bond. **Double salt**: a salt that contains two different cations or anions

Electrical conductivity: the ability of an element to conduct electricity.

endothermic reaction: a reaction where heat is absorbed from the surroundings

Endpoint: the point in a titration where the reaction is complete, often indicated by a colour change from an indicator

enthalpy change: the difference in enthalpy between the products and reactants of a reaction

Enthalpy: a thermodynamic quantity which describes the energy of a system

Equivalence point: the point at which equal moles of acid and base have reacted in a titration

Esters: Compounds formed from the reaction of carboxylic acids with alcohols

exothermic reaction: a reaction where heat is released into the surroundings

Fermentation: A process that produces alcohols like ethanol from sugars using yeast.

First-order reaction: a reaction where the rate is directly proportional to the concentration of one reactant

Fuels: Used as energy sources, e.g., methane and propane.

Halides: these are compounds formed when halogens (group 17 elements) bond with metals, e.g., NaCl, KBr, KI.

Hybridisation: The process by which atomic orbitals mix to form new hybrid orbitals with equivalent energy. This explains bonding and geometry in molecules.

Hydrides are compounds formed when hydrogen combines with another element.

Hydrogen halides: are compounds where a halogen is bonded to hydrogen, e.g., HF, HCl, HBr, HI.

Hydroxides: are compounds that include the hydroxide ion (OH^-), formed when oxides react with water.

Indicator: a chemical substance that changes colour at a specific ph level, used to signal the endpoint of a titration

Ionic chlorides: compounds found in metals like sodium chloride (NaCl) and magnesium chloride ($MgCl_2$); dissolve in water to form neutral solutions.

Litmus test: a test to determine acidity or basicity; acids turn blue litmus paper red, and bases turn red litmus paper blue

Melting and boiling points: temperature at which an element changes state from solid to liquid (melting) or liquid to gas (boiling).

Metallic hydrides: these contain metal and hydrogen, such as magnesium hydride (MgH_2). Often solid and ionic.

Molecular geometry: The three-dimensional shape of a molecule, determined by the number of bonding and lone pairs of electrons around the central atom. **Neutralisation**: a reaction where an acid and a base react to form water and a salt, typically resulting in a neutral solution.

Normal salt: a salt formed when all acidic protons have been replaced by metal ions or ammonium ions

Order of reaction: the power or exponent to which a reactant concentration is raised in the rate law

Oxides: compounds formed when an element combines with oxygen.

Pi (π) bond: A type of covalent bond formed by the sideways (lateral) overlap of p-orbitals.

Plastics: Synthetic materials made from polymerized alkenes.

Polymers: Large molecules formed from repeating alkene units.

Rate law: a mathematical expression that shows how the rate of a reaction depends on the concentration of reactants

Rate-determining step: the slowest step in a reaction sequence that controls the overall reaction rate

Reactivity the tendency of an element to participate in chemical reactions.

Reducing power: It is the ability of a substance to donate electrons in a chemical reaction. In halides, reducing power increases down the group.

Resonance: Stabilization of benzene by the movement of electrons between structures.

Salt: an ionic compound formed from the neutralisation reaction between an acid and a **Base**: usually consisting of the cation from the base and the anion from the acid.

Saturated hydrocarbon: Hydrocarbons with only single bonds between carbon atoms.

Second-order reaction: a reaction where the rate is proportional to the square of one reactant's concentration

Sigma (σ) bond: A type of covalent bond formed by the direct, head-on overlap of orbitals. Found in all single bonds and as the first bond in double and triple bonds.

Sour taste: a characteristic of acids (e.g., lemon juice)

sp^3 hybridisation: Mixing of one s-orbital and three p-orbitals, forming four equivalent hybrid orbitals.

sp^2 hybridisation: Mixing of one s-orbital and two p-orbitals, forming three equivalent hybrid orbitals.

sp hybridisation: Mixing of one s-orbital and one p-orbital, forming two equivalent hybrid orbitals.

Orbital overlap: The interaction of atomic orbitals to form bonds.

Standard enthalpy change of combustion: the heat released when one mole of a substance is completely burned in oxygen under standard conditions

Standard enthalpy change of formation: the enthalpy change when one mole of a compound is formed from its elements under standard conditions

Standard enthalpy change of hydration: the heat released when one mole of gaseous ions dissolves in water to form a solution under standard conditions

Standard enthalpy change of neutralisation: the heat change when one mole of water is formed from the reaction of an acid and a base under standard conditions

Standard enthalpy change of reaction: heat change for a chemical reaction measured under standard conditions

Standard enthalpy change of solution: the enthalpy change when one mole of solute dissolves in a solvent under standard conditions

Standard solution: a solution of known concentration used in titrations

Strong acid: an acid that completely ionizes in solution, producing a high concentration of H^+ ions

Strong base: a base that completely dissociates in solution, producing a high concentration of OH^- ions

Substitution reactions: Reactions where one atom or group in benzene is replaced by another.

Thermal stability: decreases across the period as covalent character increases.

Titrant: the solution of known concentration that is added to the analyte in a titration

Triple bond: A bond involving one sigma bond and two pi bonds. Example: In ethyne (C_2H_2), the triple bond between carbons consists of one sigma and two pi bonds.

Unhybridised Orbital: A p-orbital that does not participate in hybridisation and forms pi bonds.

Unsaturated hydrocarbon: Hydrocarbons containing at least one carbon-carbon double bond.

VSEPR Theory (Valence Shell Electron Pair Repulsion): A model used to predict molecular shapes by minimising repulsion between electron pairs around the central atom.

Weak acid: an acid that partially ionises in solution, producing a lower concentration of H^+ ions

Weak base: a base that partially dissociates in solution, producing a lower concentration of OH^- ions

Weak bases/acids: hydroxides of nonmetals like phosphorus and sulphur are unstable and often acidic.

Zero-order reaction: a reaction where the rate does not depend on the concentration of the reactant

This book is intended to be used for the Year Two Chemistry Senior High School (SHS) Curriculum. It contains information and activities to support teachers to deliver the curriculum in the classroom as well as additional exercises to support learners' selfstudy and revision. Learners can use the review questions to assess their understanding and explore concepts and additional content in their own time using the extended reading list provided.

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