VCE CHEMISTRY UNIT 3 & 4
Australian Curriculum

WORKBOOK

Chemistry

TOPICS

UNIT 3: HOW CAN DESIGN AND INNOVATION HELP TO OPTIMISE CHEMICAL PROCESSES?

Area of Study 1: What are the current and future options for supplying energy?

Area of Study 2: How can the rate and yield of chemical reactions be optimised?

UNIT 4: HOW ARE CARBON-BASED COMPOUNDS DESIGNED FOR PURPOSE?

Area of Study 1: How are organic compounds categorised and synthesised?

Area of Study 2: How are organic compounds analysed and used?

SASTA South Australian Science Teachers Association
Supporting Teachers of Science
Advancing Science Education

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Rhys Lewis

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UNIT 3:
How can design and innovation help to optimise chemical processes?

AREA OF STUDY 1: What are the current and future options for supplying energy?

3.1 Carbon-based fuels

3.1 Carbon-Based Fuels

3.1.1 Fuels

Discuss the definition of a fuel, including the distinction between fossil fuels (coal, natural gas, petrol) and biofuels (biogas, bioethanol, biodiesel) with reference to their renewability (ability of a resource to be replaced by natural processes within a relatively short period of time)

A **fuel** is a substance that stores energy in its molecular structure, which can be released in a controlled manner through combustion. Fuels are typically reduced substances, meaning they have a high proportion of hydrogen or carbon atoms that undergo oxidation when burned. Common examples include hydrogen (H_2), carbon (C), methane (CH_4), and octane (C_8H_{18}), all of which react with oxygen during combustion to release energy.

Fossil Fuels

Fossil fuels, including **coal**, **petroleum**, and **natural gas**, are the primary energy sources that have powered human civilisation for centuries. These carbon-based fuels formed over millions of years through the anaerobic decomposition of dead plants, animals, and microorganisms subjected to intense heat and pressure beneath the Earth's surface (see Figure 3.01).

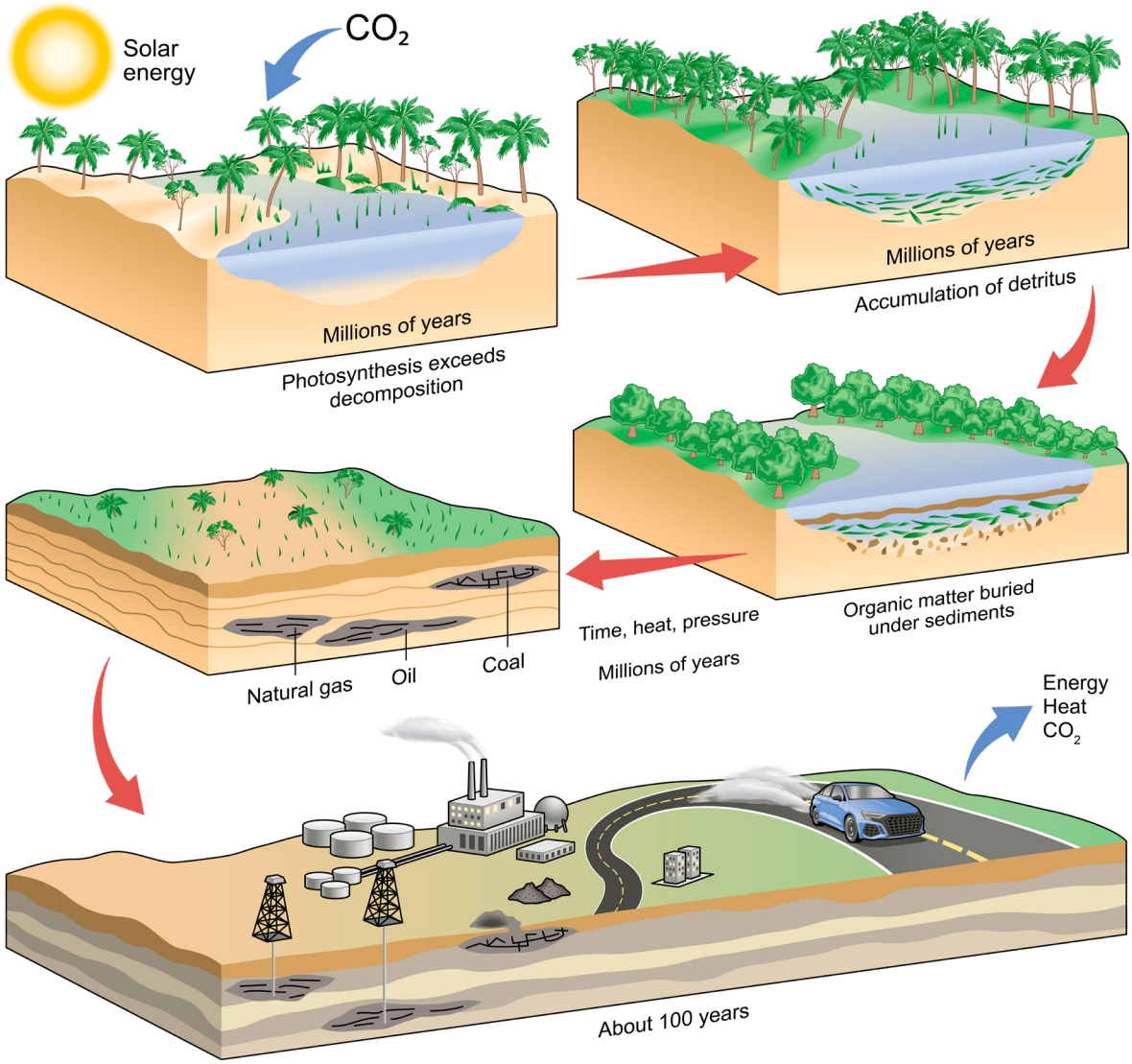


Figure 3.01: Fossil fuel formation and extraction

Question 1

Biofuels are emerging as sustainable alternatives to fossil fuels, contributing to renewable energy production, reduced carbon emissions, and energy security in Australia's transportation and electricity sectors.

Which one of the following is a biofuel?

- ◯ A Methane from natural gas
- ◯ B Octane from petroleum
- ◯ C Hydrogen from coal
- ◯ D Methyl palmitate from plant oils

(1 mark)

Question 2

Methane can be sourced from various origins. It is the primary component of natural gas and exists in coal deposits as coal seam gas, which can be extracted by drilling. In addition, large reserves of methane are stored in ice as methane hydrates. More recently, methane has been generated through the microbial decomposition of organic matter from plants and animals.

Methane is considered a renewable energy source when it is derived from

- ◯ A Coal seam gas
- ◯ B Natural gas
- ◯ C Microbial decomposition
- ◯ D Methane hydrates

(1 mark)

Question 3

A renewable energy resource

- ◯ A is naturally replenished on a short timescale through biological or geological processes.
- ◯ B provides unlimited energy and never requires replenishment.
- ◯ C produces no environmental impact when used for electricity generation.
- ◯ D can be regenerated through natural processes faster than it is consumed.

(1 mark)

Question 4

Consider the following statements about biofuels.

I. The production of biofuels does not damage the environment.
II. The combustion of both biofuels generates greenhouse gases.
III. Biofuels are renewable as they are replaced by natural processes over short timescales.

Which of the statements above are correct?

- ◯ A I and III only
- ◯ B II and III only
- ◯ C I and II only
- ◯ D III only

(1 mark)

3.1.2 Fuel sources for the body

Explore fuel sources for the body, including carbohydrates, proteins and lipids (fats and oils). The energy content of these fuels is measured in kilojoules per gram, kJ g^{-1}

The human body requires a continuous supply of energy to sustain essential functions such as movement, growth, repair, and metabolism. This energy is obtained from three primary macronutrients: **carbohydrates**, **proteins**, and **lipids** (fats and oils). Each of these biological fuel sources plays a unique role in the body, providing energy at different rates and efficiencies.

Carbohydrates

Carbohydrates are naturally occurring organic compounds in the cells and tissues of living things. All living things use carbohydrates for **respiration**, an energy-releasing reaction, and to produce complex macromolecules that store energy. The term "carbohydrate" is used interchangeably with the terms "**saccharide**" and "**sugar**". Most carbohydrates contain carbon, hydrogen, and oxygen in a ratio of 1:2:1, with molecular formulae satisfying the expression $C_n(H_2O)_y$ or $C_xH_{2y}O_y$. Some examples of carbohydrates that satisfy this definition are ribose, $C_5H_{10}O_5$, glucose, $C_6H_{12}O_6$, and sucrose, $C_{12}H_{22}O_{11}$.

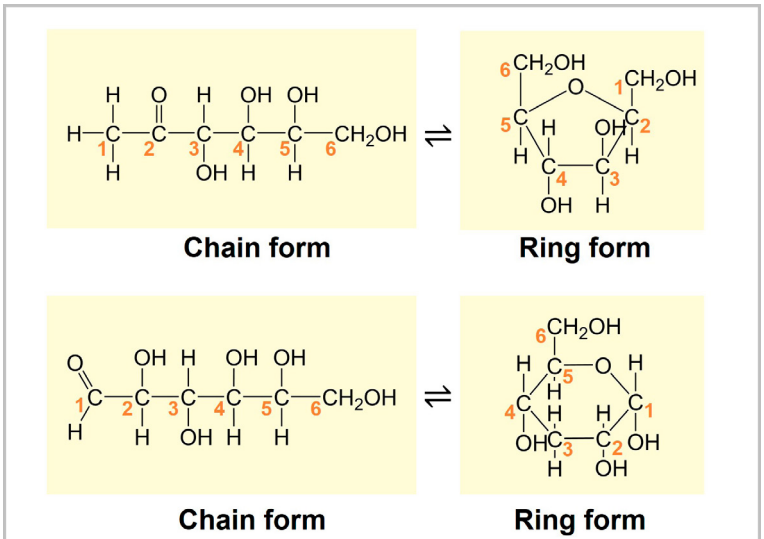


Figure 3.07: Cyclisation of carbohydrates

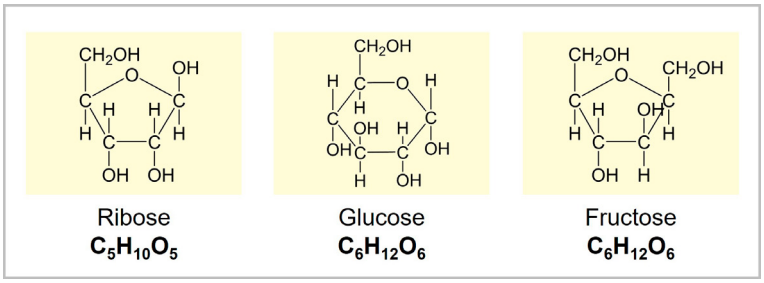


Figure 3.08: Monosaccharides

When carbohydrate molecules with five or six carbon atoms are transferred to water or an aqueous solution such as blood or body tissues, they undergo cyclisation, a type of reaction in which the linear structure changes to a cyclic (ring) structure that no longer contains a carbonyl group. For example, Figure 3.07 shows the cyclisation of fructose (top) and glucose (bottom). An equilibrium is established between the chain and ring forms of these carbohydrates in an aqueous solution. The cyclic forms of carbohydrates can connect in one of several ways to produce chains of varying lengths. These chains are still referred to as carbohydrates and are classified by the number of subunits that comprise them.

Monosaccharides are composed of one subunit in cyclic or ring form that cannot be broken down into simpler compounds. These sugars typically contain between three and nine carbon atoms, with the most common ones containing five or six carbon atoms. Examples of monosaccharides include ribose, glucose, and fructose (see Figure 3.08). The most important carbohydrate for the body is glucose. When a person consumes a carbohydrate-rich meal, the food is digested, and glucose molecules are absorbed into the blood, where they are transported to cells for respiration (see Chapter 3.1.4).

Energy Profile Diagrams

The enthalpy change of a chemical reaction can be illustrated using an **energy profile diagram**, which visually represents the energy changes that occur as reactants are transformed into products. The key feature of an energy profile diagram is the **reaction pathway**, a continuous curve that traces the energy changes throughout the reaction. The pathway begins at a y-coordinate, representing the enthalpy of the reactants and extends horizontally along the x-axis, which represents the reaction's progression. As the reaction proceeds, the pathway rises, indicating an increase in enthalpy as the reactants absorb the **activation energy (E_a)**, the minimum energy required to break chemical bonds. Once the bonds are broken, the pathway descends, reflecting the release of energy as new bonds form in the products. Finally, the reaction pathway stabilises at a new y-coordinate, representing the enthalpy of the products. The ΔH is represented on the diagram as the difference in energy between the reactants and products.

1 Exothermic reaction

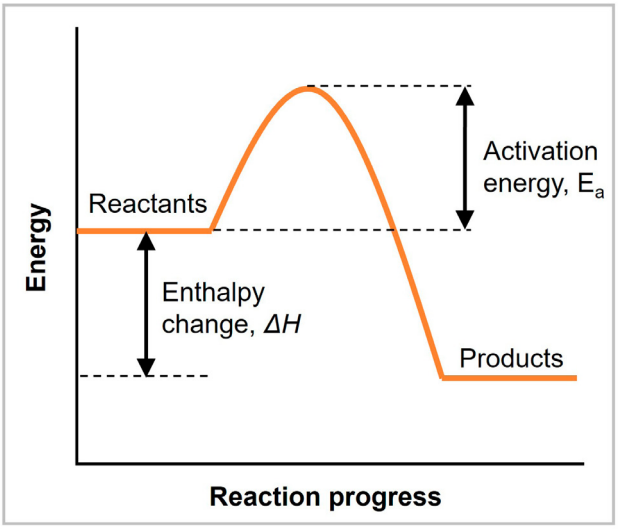


Figure 3.28: Energy profile diagram for an exothermic reaction

2 Endothermic reaction

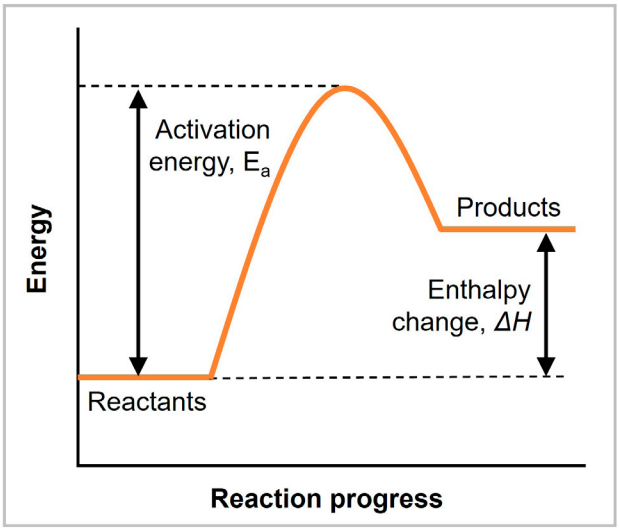


Figure 3.29: Energy profile diagram for an endothermic reaction

Despite sharing key reaction characteristics, such as reaction pathways, activation energies, and enthalpy changes, there are fundamental differences between the energy profile diagrams of exothermic and endothermic reactions. In an exothermic reaction, energy is released to the surroundings, resulting in a negative ΔH (see **Figure 3.28**). This is visually represented by the products having a lower energy level than the reactants, indicating that energy has been released from the system to the surroundings. In contrast, an endothermic reaction absorbs energy from its surroundings, resulting in a positive ΔH (**Figure 3.29**). The products are at a higher energy level than the reactants, reflecting the energy absorbed.

Question 23

Calcium carbonate is decomposed according to the equation below.

$$CaCO_{3\,(s)} \rightarrow CaO_{(s)} + CO_{2\,(g)} \quad \Delta H = +178 \text{ kJ mol}^{-1}$$

Which one of the following is consistent with the information in the equation?

- ○ A More energy is absorbed in bond-making than is released in bond-breaking.
- ○ B More energy is released in bond-making than is absorbed in bond-breaking.
- ○ C Less energy is absorbed in bond-making than is released in bond-breaking.
- ○ D Less energy is released in bond-making than is absorbed in bond-breaking.

(1 mark)

Question 24

Biogas is a fuel mixture containing methane and carbon dioxide.

Consider the following units for the heat of combustion of fuels:

I. kJ mol^{-1}
II. kJ g^{-1}
III. kJ

Which of the units is used to quantify the heat of combustion of biogas?

- ◯ A I and II only
- ◯ B II and III only
- ◯ C II only
- ◯ D III only

(1 mark)

Question 25

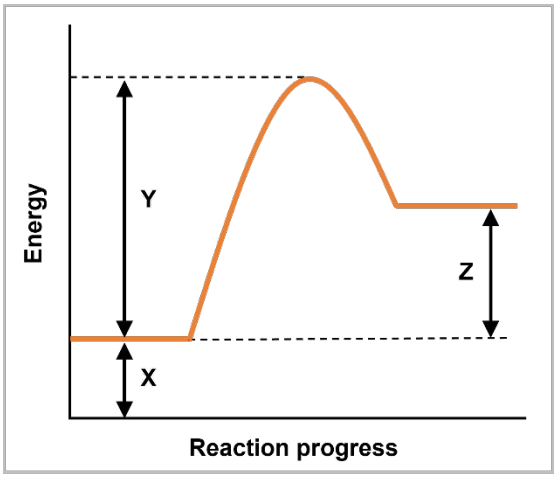
Which one of the following combinations correctly identifies a type of chemical reaction and the sign of the enthalpy change, ΔH?

	Type of reaction	*Enthalpy change*
◯ A	Cellular respiration	Negative
◯ B	Photosynthesis	Negative
◯ C	Fermentation	Positive
◯ D	Combustion	Positive

(1 mark)

Question 26

Refer to the energy profile diagram below:



The activation energy for this reaction is equal to

- ◯ A Y – Z.
- ◯ B X + Z.
- ◯ C Y.
- ◯ D Z.

(1 mark)

3.1.7 Limiting Reagents

Determine the limiting reactants or reagents in chemical reactions.

A chemical reaction involves the transformation of reactants into products. In most cases, the molar quantities of the reactants are unequal, meaning one reactant is present in greater excess. The reactant available in the larger quantity is referred to as the **excess reactant**, while the reactant present in the smaller quantity is known as the **limiting reactant** (or **limiting reagent**). The limiting reactant is completely consumed first, preventing any further reaction from occurring. Once it is exhausted, the reaction stops, even if other reactants remain unreacted.

Understanding limiting reagents is fundamental in stoichiometry, as it enables chemists to predict reaction yields, minimise waste, and optimise chemical efficiency. This concept is particularly crucial in industries such as pharmaceuticals, manufacturing, and environmental chemistry, where precise control of reagents ensures maximum efficiency, cost-effectiveness, and sustainability.

Limiting Reagents

A chemical reaction occurs when reactants combine in fixed mole ratios, as defined by a balanced chemical equation. To determine the limiting reagent, the available mole ratios of reactants in the reaction mixture are compared with the stoichiometric ratios required by the equation.

For example, in the reaction between hydrogen gas (H_2) and oxygen gas (O_2) to form water (H_2O):

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

If five moles of hydrogen (5 mol H_2) and two moles of oxygen (2 mol O_2) are present, the limiting reagent can be determined by analysing the mole ratio. The balanced equation indicates that two moles of H_2 react with one mole of O_2 to form two moles of H_2O. This means that:

- Two moles of H_2 produce two moles of H_2O
- One mole of O_2 produces two moles of H_2O

Using these ratios, the available five moles of H_2 could theoretically produce five moles of H_2O, while the two moles of O_2 can produce only four moles of H_2O. Since oxygen produces the least amount of product, it is completely consumed first, making **O_2 the limiting reagent** in this reaction.

To determine the limiting reagent, we calculate the number of moles of each reactant present and then use the mole ratio to determine the number of moles of product produced by each reactant.

Example 3.01

A 40.0 g sample of calcium carbonate ($CaCO_3$) is reacted with 25.0 g of hydrochloric acid (HCl).

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$$

To determine the limiting reagent, we first calculate the number of moles of each reactant:

$$n(CaCO_3) = \frac{m}{M} = \frac{40.0}{100.09} = 0.400 \text{ mol}$$

$$n(HCl) = \frac{m}{M} = \frac{25.0}{36.458} = 0.686 \text{ mol}$$

Next, we determine the number of moles of one product, such as CO_2, each reactant can produce.

$$n(CO_2) = n(CaCO_3) = 0.400 \text{ mol}$$

$$n(CO_2) = \frac{n(HCl)}{2} = \frac{0.686}{2} = 0.343 \text{ mol}$$

Hence, **HCl is the limiting reagent**, as it produces the fewest moles of the product, CO_2.

Question 28

Hydrazine (N_2H_4) is used as a fuel in spacecraft. It reacts with oxygen as follows:

$$N_2H_4 + O_2 \rightarrow N_2 + 2H_2O$$

A fuel cell contains 50.0 g of hydrazine and 64.0 g of oxygen.

Determine the limiting reagent in this reaction.

(3 marks)

Question 29

In the thermite reaction:

$$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3$$

One thermite mixture contains 30.0 g of aluminium powder and 80.0 g of iron(III) oxide.

Determine the limiting reagent in this reaction.

(3 marks)

Question 30

Ammonia reacts with sulfuric acid to form ammonium sulfate, a common fertiliser:

$$2NH_3 + H_2SO_4 \rightarrow (NH_4)_2SO_4$$

A fertiliser mixture is produced by reacting 20.0 g of ammonia with 50.0 g of sulfuric acid.

Determine the limiting reagent in this reaction.

(3 marks)

Question 31

Aspirin ($C_9H_8O_4$) is synthesised from salicylic acid ($C_7H_6O_3$) and acetic anhydride ($C_4H_6O_3$):

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + CH_3COOH$$

A chemist uses 10.0 g of salicylic acid and 12.0 g of acetic anhydride.

Determine the limiting reagent in this reaction.

(3 marks)

Question 32

The table shows the heat of combustion of some common fuels.

Fuel	Formula	State at STP	Molar enthalpy of combustion (kJ mol^{-1})
hydrogen	H_2	Gas	–286
methane	CH_4	Gas	–890
ethane	C_2H_6	Gas	–1560
propane	C_3H_8	Gas	–2220
butane	C_4H_{10}	Gas	–2880
octane	C_8H_{18}	Liquid	–5470
ethyne	C_2H_2	Gas	–1300
methanol	CH_3OH	Liquid	–726
ethanol	C_2H_5OH	Liquid	–1370

Write thermochemical equations for the complete combustion of:

(a) methane

(2 marks)

(b) ethane

(2 marks)

(c) propane

(2 marks)

(d) butane

(2 marks)

(e) octane

(2 marks)

(f) ethyne

(2 marks)

(g) methanol

(2 marks)

(h) ethanol

(2 marks)

Example 3.07

A 1.25 g sample of methanol, CH_3OH (M = 32.0 g mol^{-1}), was combusted in a laboratory setting, and the heat was transferred to 125 g of liquid water under SLC, increasing its temperature from 22.0°C to 32.8°C.

The quantity of heat absorbed by the water is calculated below.

$$q = mc\Delta T = 125 \times 4.18 \times (32.8 - 22.0) = 5643 \text{ J} \sim 5.64 \text{ kJ}$$

The enthalpy of combustion is calculated below.

$$n(CH_3OH) = \frac{m}{M} = \frac{1.25}{32.0} = 0.0391 \text{ mol}$$

$$\Delta H = \frac{q}{n} = \frac{-5.64}{0.0391} = -145 \text{ kJ mol}^{-1}$$

Notably, the value calculated in **Example 3.07** for the molar enthalpy of combustion of methanol is only 20% of its theoretical value (–726 kJ mol^{-1}), highlighting the significant limitations of calorimetry when carried out using basic laboratory equipment. To better understand these discrepancies, it is important to examine how calorimetry experiments are typically conducted in a laboratory setting.

In a typical laboratory setting, calorimetry is performed using a simple apparatus (see **Figure 3.35**), including a spirit burner, a beaker of water, a thermometer, a tripod, and a gauze mat. The process begins by measuring the initial mass of the spirit burner containing the fuel, as well as recording the initial temperature of a known mass of water placed in the beaker. The beaker is supported on a tripod above the spirit burner, with a thermometer submerged in the water to monitor temperature changes. The fuel is then ignited, and the heat produced by its combustion is transferred to the water above. As the water absorbs this heat, its temperature rises. Once a significant temperature change is observed, the flame is extinguished, and the final temperature of the water is recorded. The spirit burner is reweighed to determine the mass of fuel that was consumed.

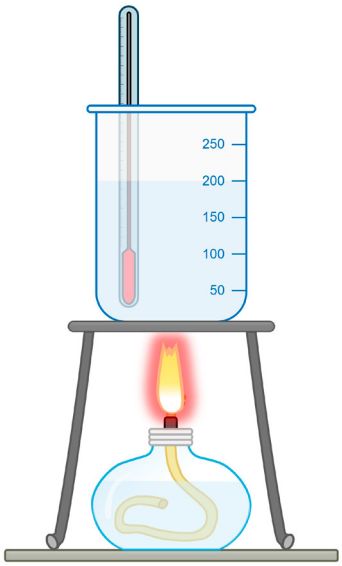


Figure 3.35: Simple calorimetry apparatus

The energy transferred to the water is calculated using $q = mc\Delta T$, and this quantity of heat is then converted to kilojoules and divided by the number of moles of fuel combusted to determine the molar enthalpy of combustion (ΔH).

This method assumes that the water absorbs all of the heat released by the combustion. However, in practice, some heat is inevitably lost to the surrounding air, the beaker, and the apparatus, and combustion may be incomplete as air is used rather than pure oxygen. As a result, the experimental value for the enthalpy change is often significantly lower than the theoretical value, as seen in **Example 3.07**. Despite these limitations, this simple form of calorimetry provides a useful approximation of the energy released during combustion and is commonly used for educational purposes or to compare the energy output of different fuels and the caloric content of various foods.

To improve the accuracy of calorimetry results in the laboratory, several modifications can be made to reduce heat loss and ensure more complete combustion. Using a lid on the beaker helps minimise heat loss to the surrounding air, while placing insulation around the apparatus reduces energy transfer to the environment. A copper calorimeter can replace a glass beaker, as it conducts heat more efficiently and evenly. Ensuring the flame is positioned directly under the beaker improves heat transfer, and using a draft shield around the flame helps prevent heat from being carried away by air currents.

Bomb Calorimetry

Bomb calorimetry is a highly accurate technique used to measure the energy content of a substance, typically a fuel or food, by determining the heat released during complete combustion.

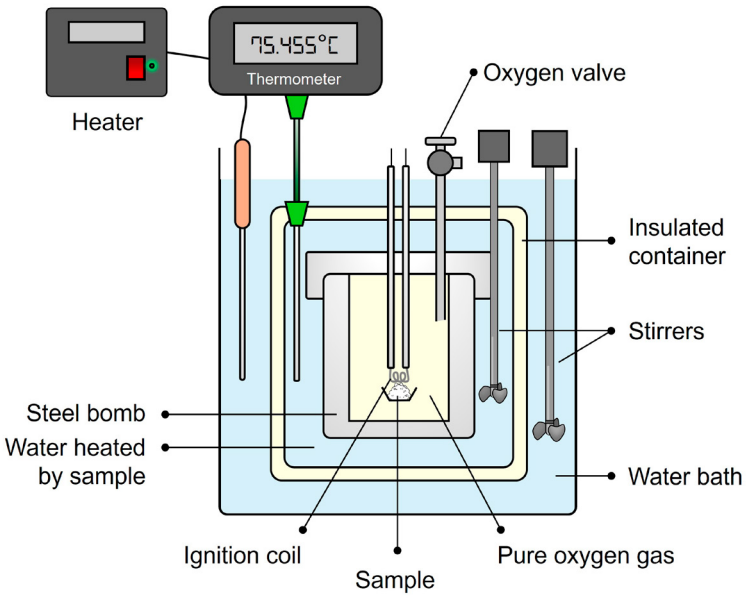


Figure 3.36: Bomb calorimeter

Figure 3.36 shows the components of a typical bomb calorimeter. At the core of the system is the **steel bomb**, a strong, sealed container that holds the sample to be combusted. The bomb is filled with pure oxygen gas through an oxygen valve, ensuring the sample burns completely. An **ignition coil** ignites the sample electrically. As the sample combusts, the heat produced is transferred to the surrounding water bath, raising its temperature. A digital thermometer precisely monitors the temperature change of the water, while stirrers continuously mix the water to ensure uniform temperature distribution. The entire setup is placed within an insulated container to minimise heat loss to the environment, ensuring that nearly all the heat released by the combustion is captured in the water bath. This method significantly reduces heat loss and provides more accurate and reliable energy measurements than simple laboratory calorimetry.

To accurately calculate the heat released by a sample in a bomb calorimeter, a known **calibration factor (CF)** must be determined first. This is done by supplying a known quantity of electrical energy to the calorimeter using a heater. The energy input is calculated using the formula $VI{\cdot}t$, where V is the voltage (in volts), I is the current (in amperes), and t is the time the current is applied (in seconds). The calibration factor is then calculated using:

$$\text{CF} = \frac{VIt}{\Delta T}$$

This value represents the amount of energy required to raise the temperature of the entire calorimeter system by one kelvin (or one degree Celsius). Once the CF is known, the energy released by the combustion of a sample can be calculated by multiplying the calibration factor by the temperature change observed during the reaction:

$$q = \text{CF} \times \Delta T$$

Example 3.08

A 1.50 g sample of potato crisps is placed in the bomb calorimeter and completely combusted. The initial temperature of the water is 23.1°C, and the final temperature is 27.4°C. The calibration factor for the bomb calorimeter is 7.55 kJ $°C^{-1}$.

Calculate the energy content of the potato crisps in kilojoules per gram.

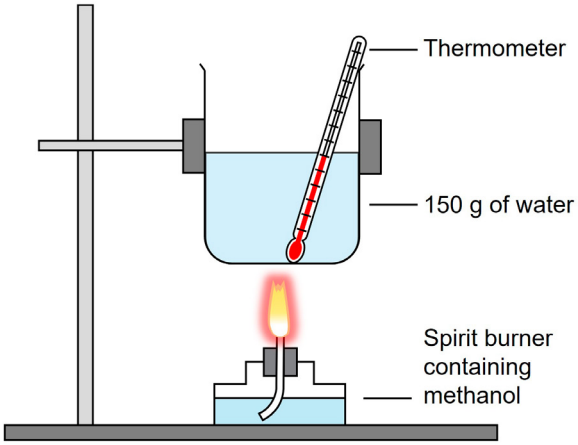
$$q = \text{CF} \times \Delta T = 7.55 \times (27.4 - 23.1) = 32.5 \text{ kJ}$$

$$\frac{q}{m} = \frac{32.5}{1.50} = 21.6 \text{ kJ g}^{-1}$$

Question 72

Methanol, CH_3OH, is a liquid fuel used in cooking.

The molar enthalpy of combustion of methanol, ΔH, was determined in an experiment using the apparatus shown in the diagram below:



The following quantities were measured:

- The mass of water in the calorimeter = 150 g
- The initial temperature of the water = 21.2°C

(a) State two other quantities measured to determine the heat of combustion of methanol.

(2 marks)

(b) The experiment was conducted five times with the same apparatus.

State two factors held constant in all five trials.

(2 marks)

(c) The heat of combustion of methanol is –726 kJ mol^{-1}.

The value determined in the experiment was –182 kJ mol^{-1}.

Incomplete combustion is one reason for the low accuracy of the result.

i. State two other reasons for the low accuracy of the result.

(2 marks)

ii. Suggest two practical improvements to the apparatus above that would increase the accuracy of the result.

(2 marks)

3.2.3 Solution Calorimetry

Investigate the principles of solution calorimetry, including determination of calibration factor and consideration of the effects of heat loss; analysis of temperature-time graphs obtained from solution calorimetry.

Solution calorimetry is a fundamental technique in chemistry used to measure the energy changes that occur during chemical reactions in aqueous solution, such as dissolution, neutralisation, and metal displacement. It is based on the principle of conservation of energy: the heat released or absorbed by the reaction is transferred to the surrounding solution, producing a measurable change in temperature. By accurately recording this temperature change, the energy change of the reaction can be calculated using the formula $q = mc\Delta T$, where q is the quantity of heat transferred, m is the mass of the solution, c is its specific heat capacity (typically assumed to be 4.18 J g^{-1} K^{-1}, the value for water), and ΔT is the temperature change in kelvin or degrees Celsius. This section examines how calorimetry can be utilised to quantify the thermal energy changes associated with reactions occurring in aqueous solutions.

Solution Calorimetry

Solution calorimetry is carried out in the laboratory using a simple yet effective setup designed to minimise heat loss and accurately measure temperature changes during a reaction in solution. The calorimeter shown in **Figure 3.37** consists of two nested foam cups, which provide insulation and reduce heat exchange with the surroundings. A known volume of water or an aqueous solution is placed into the inner cup, where the chemical reaction will occur. A foam lid, which fits loosely over the top, helps limit heat loss to the environment while allowing space for a thermometer and a stirrer.

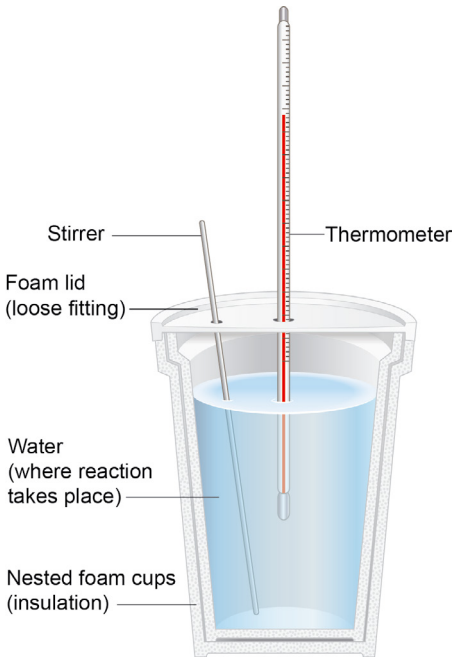


Figure 3.37: Simple calorimeter

Before the reaction begins, the initial temperature of the solution is carefully recorded using the thermometer inserted through the lid. This provides a baseline for calculating the temperature change. The chemical being tested—such as an ionic compound for a dissolution reaction or a second reactant in a neutralisation reaction—is then added quickly to the water through the lid. The solution is stirred continuously to ensure that the reaction mixture is uniform and that heat is evenly distributed throughout the liquid.

As the reaction proceeds, the thermometer is used to monitor the temperature at regular intervals. Depending on whether the process is exothermic or endothermic, the temperature will either rise or fall. Once the temperature stabilises or reaches a clear maximum or minimum, the final reading is measured and recorded. The change in temperature is then used to calculate the quantity of energy transferred using the formula $q = mc\Delta T$, where m is the mass of the solution, c is the specific heat capacity (commonly taken as 4.18 J g^{-1} K^{-1}), and ΔT is the temperature change in kelvin or degrees Celsius. Finally, the enthalpy change, ΔH of the reaction, can be determined by dividing the quantity of energy transferred by the number of moles of the substance involved in the reaction, n. This method offers a practical and accessible approach to measuring energy changes in chemical reactions that occur in solution. The following sections provide specific examples of solution calorimetry.

Dissolution

When a solid ionic compound is added to water, its positive and negative ions are separated and surrounded by water molecules in a process known as **dissolution**. The first stage of this process, known as dissociation, involves the absorption of energy from the surrounding water to overcome the strong electrostatic forces that hold the ions together in the crystal lattice. Water molecules, being polar, play a crucial role in this step: their partially positive hydrogen atoms are attracted to anions, while their partially negative oxygen atoms are drawn to cations. These attractions help pull the ions away from the lattice, allowing them to disperse uniformly throughout the solution. In the second stage, known as **hydration**, the free ions are surrounded by water molecules, forming stabilising electrostatic forces called **ion-dipole interactions**. This step releases energy back into the solution. Overall, the dissolution process may be either exothermic or endothermic, depending on whether more energy is released during hydration than is required to break the ionic bonds during dissociation. The dissolution process is summarised in Figure 3.38.

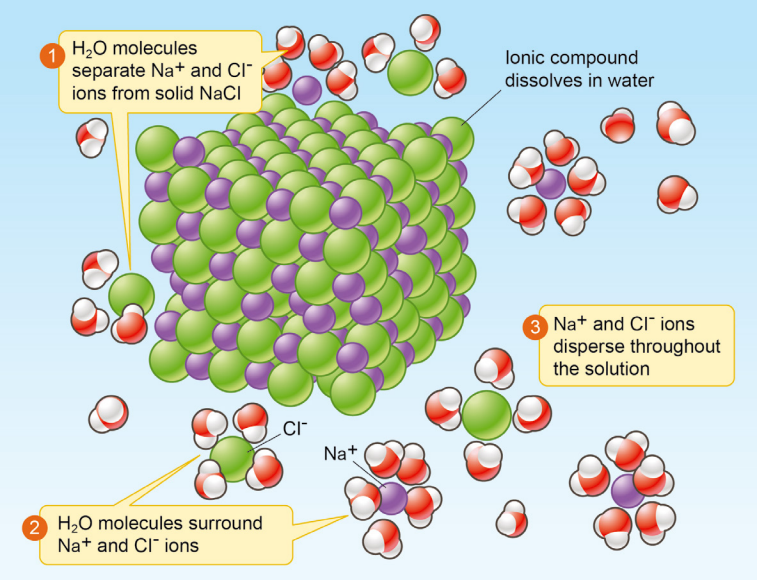


Figure 3.38: Dissolution of an ionic compound

The overall energy change associated with this process is known as the **enthalpy of solution**. It is defined as the enthalpy change when one mole of an ionic compound dissolves completely in water to form a homogeneous aqueous solution of its ions. This value reflects the combined effect of two competing energy changes: the endothermic input required to break apart the ionic lattice during dissociation, and the exothermic release of energy that occurs when water molecules hydrate the separated ions. If the energy released during hydration exceeds the energy required for dissociation, the enthalpy of solution is negative, indicating that the process is exothermic. Conversely, if more energy is required to break the lattice than is recovered through hydration, the enthalpy of solution is positive, and the process is endothermic. The energy change is calculated using the equation $q = mc\Delta T$, where m is the mass of the water, c is its specific heat capacity, and ΔT is the temperature change. Finally, the enthalpy of solution, ΔH, is found by dividing the calculated energy by the number of moles of the ionic compound that dissolved.

Example 3.09

A 23.5 g sample of sodium chloride, NaCl (M = 58.5 g mol^{-1}), was transferred to 175 g of liquid water under SLC, decreasing its temperature from 21.5°C to 19.4°C.

The quantity of heat transferred is calculated below.

$$q = mc\Delta T = 175 \times 4.18 \times (19.4 - 21.5) = -1536 \text{ J} \sim -1.54 \text{ kJ}$$

The energy change is negative because heat was transferred from the surrounding water to the ionic lattice, indicating that energy was absorbed from the solution during the reaction. The enthalpy of solution of sodium chloride is calculated below.

$$n(\text{NaCl}) = \frac{m}{M} = \frac{23.5}{58.5} = 0.402 \text{ mol}$$

$$\Delta H = \frac{q}{n} = \frac{1.54}{0.402} = +3.82 \text{ kJ mol}^{-1}$$

Question 74

50.0 mL of 1.25 mol L^{-1} hydrochloric acid was mixed with 25.0 mL of 2.50 mol L^{-1} sodium hydroxide in a polystyrene cup calorimeter. The initial temperature of both solutions was 25.0°C. After mixing, the final temperature was 33.1°C.

Assume the density of each solution is 1.00 g mL^{-1} and their specific heat capacity is 4.18 J g^{-1} $°C^{-1}$.

What is the enthalpy of neutralisation, in kJ mol^{-1}, for this reaction?

- ◯ A –27.2
- ◯ B –40.6
- ◯ C –54.2
- ◯ D –57.0

(1 mark)

Question 75

85.0 mL of 0.500 mol L^{-1} sulfuric acid (H_2SO_4) was transferred to a highly insulated calorimeter containing 65.0 mL of 1.15 mol L^{-1} potassium hydroxide (KOH) at 22.5°C. The reaction is described in the thermochemical equation below.

$$H_2SO_{4\,(aq)} + 2KOH_{(aq)} \rightarrow K_2SO_{4\,(aq)} + 2H_2O_{(l)} \quad \Delta H = -57.8 \text{ kJ mol}^{-1}$$

The maximum temperature of the solutions in °C was

- ◯ A 26.1
- ◯ B 27.2
- ◯ C 28.3
- ◯ D 29.4

(1 mark)

Question 76

A 1.43 g sample of lithium iodide, LiI, is dissolved in 115 g of water at 22.0°C in a well-insulated calorimeter. The maximum temperature was measured as 23.4°C.

Assume the density of each solution is 1.00 g mL^{-1} and their specific heat capacity is 4.18 J g^{-1} $°C^{-1}$.

What is the enthalpy of solution, in kJ mol^{-1}, for this reaction?

- ◯ A –32.2
- ◯ B –57.2
- ◯ C –63.0
- ◯ D –76.5

(1 mark)

Question 77

What mass of ammonium chloride (NH_4Cl, ΔH = +14.8 kJ mol^{-1}) lowers the temperature of 52.5 g of water by 8.50°C?

- ◯ A 6.75 g
- ◯ B 10.3 g
- ◯ C 14.8 g
- ◯ D 18.9 g

(1 mark)

Using Oxidation Numbers

Redox reactions involve the transfer of electrons from one substance to another in a chemical reaction. It follows that the oxidation number of the substance that transfers electrons will become more positive, reflecting the loss of negatively charged electrons, and the oxidation number of the substance that accepts the electrons will become more negative, reflecting the gain of electrons. By comparing oxidation numbers before and after the reaction, chemists identify which substance was oxidised and which was reduced. For example, in the reaction of magnesium and oxygen, the oxidation number of Mg increases from 0 to +2, reflecting oxidation, and the oxidation number of oxygen decreases from 0 to –2, reflecting reduction.

$2Mg$	+	O_2	→	$2MgO$	
0			→	+2	**oxidation**
		0	→	–2	**reduction**

The substance that loses electrons and causes another to be reduced is called the **reducing agent**, while the substance that gains electrons and causes another to be oxidised is the **oxidising agent**. These agents work together in redox reactions and are identified by changes in oxidation numbers. For example, Mg is the reducing agent because it donates electrons to reduce O_2, while O_2 is the oxidising agent because it accepts electrons and oxidises Mg.

Redox Half-Equations

A redox reaction involves two simultaneous and complementary processes: oxidation and reduction. Each process represents half of the overall reaction and is expressed using a **half-equation**, which shows either the loss or gain of electrons. Together, these half-equations illustrate the full electron transfer that defines a redox reaction. At **neutral pH**, redox half-equations are written by placing the chemical formula of the reactant on the left side of the arrow and the product on the right. The equation is then balanced for both atoms and charge. Electrons are added to the side with the greater overall positive charge to ensure that the electric charge is balanced on both sides of the equation. For example, in the half-equation describing the oxidation of magnesium during its reaction with oxygen, the magnesium atom (Mg) is placed on the left side of the arrow to represent its initial elemental form, while the magnesium ion (Mg^{2+}) is placed on the right to reflect its final state as part of the product, magnesium oxide (MgO). Two electrons (denoted e^-) are written to the right of the arrow to balance the overall electric charge.

$$Mg \rightarrow Mg^{2+} + 2e^-$$

Conversely, the half-equation describing the reduction of oxygen in the reaction is written with the oxygen molecule to the left of the arrow, representing its initial elemental form, and the oxide anion (O^{2-}) to the right to reflect its final state as part of the product, magnesium oxide (MgO). Four electrons are written to the left of the arrow to balance the overall electric charge.

$$O_2 + 4e^- \rightarrow 2O^{2-}$$

The half-equations can be combined to form the overall redox equation. To do this, the number of electrons in each half-equation must first be equal. This typically involves multiplying one or both half-equations by appropriate coefficients to balance the electrons transferred. Once balanced, the half-equations are combined by listing all reactants on the left side of the arrow and all products on the right and cancelling the electrons. For example, the oxidation half-equation for magnesium is multiplied by two, so both half-equations contain four electrons:

$$2Mg \rightarrow 2Mg^{2+} + 4e^-$$

Then each is combined by cancelling the four electrons from both sides of the arrows:

$$2Mg + O_2 \rightarrow 2Mg^{2+} + 2O^{2-}$$

Question 123

Chlorine exhibits a range of oxidation states.

Determine the oxidation number of chlorine (Cl) in the following substances.

(a) Cl_2 ☐ (b) $HClO_4$ ☐ (c) KCl ☐ (d) ClO_3^- ☐

(e) $HClO$ ☐ (f) ClO_2^- ☐ (g) ClO_2 ☐ (h) Cl_2O_7 ☐

(8 marks)

Question 124

Iridium (Ir) exhibits a range of positive and negative oxidation states, including the highest recorded oxidation state for any element, +9.

Determine the oxidation number of iridium in the following substances.

(a) IrO_2 ☐ (b) $IrCl_6^{2-}$ ☐ (c) IrF_6 ☐ (d) $IrBr_3$ ☐

(e) Ir ☐ (f) IrO_4^+ ☐ (g) IrO_4 ☐ (h) Ir_4F_{20} ☐

(8 marks)

Question 125

Determine the oxidation number of manganese (Mn) in the following substances.

(a) MnO_4^- ☐ (b) Na_3MnO_4 ☐ (c) MnO_2 ☐ (d) MnF_3 ☐

(e) Mn_2O_7 ☐ (f) Mn ☐ (g) $MnCl_2$ ☐ (h) Li_2MnO_4 ☐

(8 marks)

Question 126

Determine the oxidation number of phosphorus (P) in the following substances.

(a) H_3PO_4 ☐ (b) $H_4P_2O_6$ ☐ (c) $H_5P_3O_{10}$ ☐ (d) PH_3 ☐

(e) P_4 ☐ (f) P_2H_4 ☐ (g) HPO_2 ☐ (h) H_3PO_2 ☐

(8 marks)

Question 127

Determine the oxidation number of tungsten (W) in the following substances.

(a) W_2C ☐ (b) WO_4^{2-} ☐ (c) WO_3 ☐ (d) WF_6 ☐

(e) WCl_4 ☐ (f) WBr_5 ☐ (g) W_6Cl_{12} ☐ (h) W_6Cl_{18} ☐

(8 marks)

3.3.2 Galvanic Cells

Discuss the common design features and general operating principles of non-rechargeable (primary) galvanic cells converting chemical energy into electrical energy, including electrode polarities and the role of the electrodes (inert and reactive) and electrolyte solutions (details of specific cells not required).

Explore the use and limitations of the electrochemical series in designing galvanic cells. The electrochemical series also serves as a tool for predicting the products of redox reactions, for deducing overall equations from redox half-equations, and for determining maximum cell voltage under standard conditions.

Non-rechargeable galvanic cells, also known as **primary cells**, are devices that use redox reactions to convert chemical energy into electrical energy. These cells are widely used in everyday items such as remote controls, flashlights, and portable electronics. The design and performance of these cells are governed by key principles of electrochemistry, including the use of inert and reactive electrodes, the choice of electrolytes, and the application of the **electrochemical series**. This section examines the structure, operation, and design of non-rechargeable galvanic cells, highlighting their use in harnessing chemical reactions to generate useful electrical energy.

Galvanic Cells

A **galvanic cell** is a device that generates an electric current by capturing the energy released from a spontaneous redox reaction. Its basic structure consists of two **half-cells**, each containing an **electrode** immersed in an **electrolyte** solution (see Figure 3.43). These half-cells separate the oxidation and reduction processes that drive the redox reaction.

At the **anode**, oxidation occurs as electrons are released by a substance undergoing chemical change. These electrons accumulate and flow into the external circuit, giving the anode a **negative polarity.** At the **cathode**, reduction occurs as electrons enter from the external circuit and are accepted by a substance in the electrolyte, resulting in the cathode acquiring a **positive polarity**.

Electrons flow from the anode to the cathode through the external circuit, generating an **electric current**. To complete the internal circuit and maintain electrical neutrality, a **salt bridge** connects the two half-cells, ensuring the system remains electrically neutral. It contains a gel or solution of inert electrolytes that enables ions to migrate—cations toward the cathode and anions toward the anode—allowing the redox reaction to continue without charge build-up.

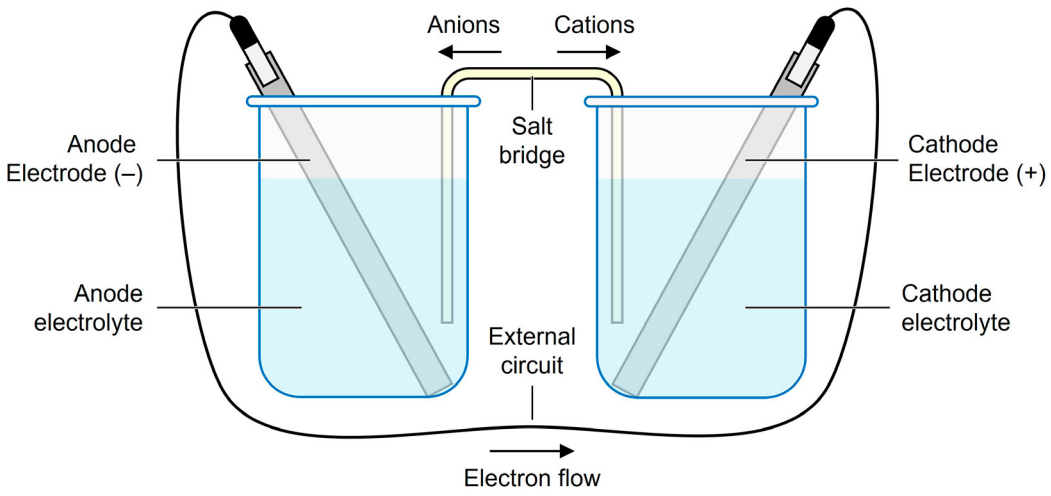
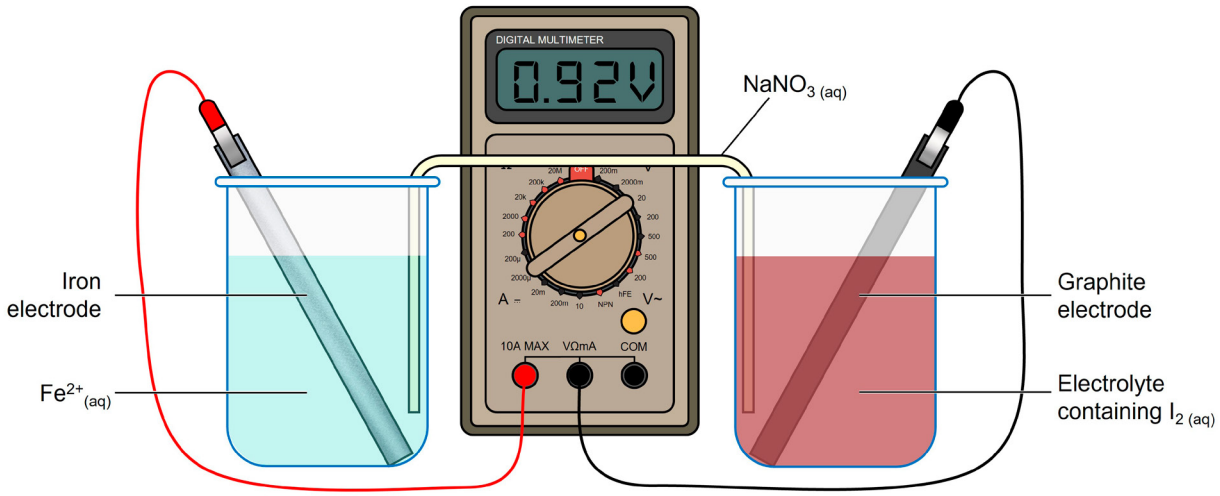


Figure 3.43: The basic features of a non-rechargeable galvanic cell.

Question 165

An investigation was conducted to determine the cell voltage produced in a galvanic cell.



(a) The standard electrode potentials for the half-cell reactions are given below.

Reaction	E^0 (V)
$I_{2\,(s)} + 2e^- \rightleftharpoons 2I^-_{(aq)}$	+0.54
$Fe^{2+}_{(aq)} + 2e^- \rightleftharpoons Fe_{(s)}$	–0.44

i. State whether the iron electrode is the anode or cathode in this cell and give a reason.

(2 marks)

ii. State whether electrons flow towards or away from the graphite electrode in this cell.

(1 mark)

(b) The voltmeter in the diagram reads +0.92 V.

i. Calculate the theoretical cell voltage for this galvanic cell.

(1 mark)

ii. State two reasons why the measured cell voltage differs from the value calculated in (b)

(2 marks)

(c) The salt bridge is composed of a sodium nitrate solution, $NaNO_{3(aq)}$.

i. State the half-cell that nitrate anions, $NO_3^-{}_{(aq)}$ flow into when the cell is operating.

(1 mark)

ii. List two reasons why $NaNO_{3\,(aq)}$ is an appropriate substance for use in the salt bridge.

(2 marks)

Collision Theory

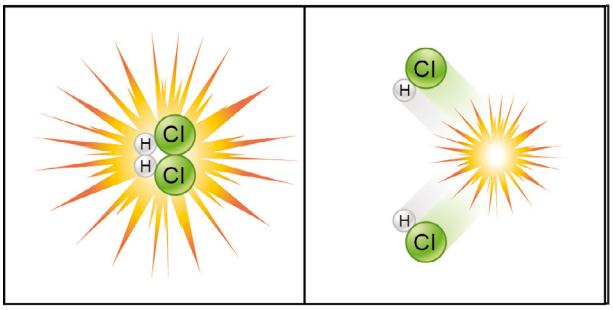
Chemical reactions involve the transformation of reactant particles—such as atoms, ions, or molecules—into one or more new products. From a chemist's perspective, the basic idea behind how a reaction occurs is relatively straightforward. Because all particles possess kinetic energy, they are constantly in motion. In the gaseous and liquid states, this motion causes particles to collide frequently, and some of these collisions can result in a chemical change.

This concept forms the basis of **collision theory**, which proposes that chemical reactions occur when reactant particles collide with enough energy and in the correct orientation. When such a collision happens, existing chemical bonds may break, and the atoms or molecular fragments can rearrange to form new products.

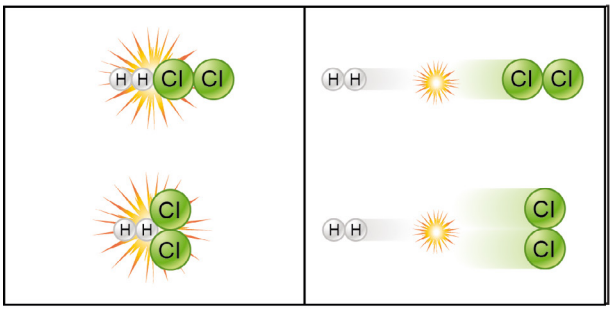
For example, in the reaction between hydrogen gas (H_2) and chlorine gas (Cl_2), the bonds within the H_2 and Cl_2 molecules can break upon collision, allowing new bonds to form and producing hydrogen chloride (HCl) molecules (see Figure 3.51). This process is assumed to occur through a direct collision between a hydrogen molecule and a chlorine molecule.

However, not all collisions result in a reaction. If they did, chemical changes would occur instantaneously due to the enormous number of collisions taking place every second. In reality, only a small fraction of collisions are **successful** because two key conditions must be met for a reaction to occur.

1 Particles collide with sufficient energy in the correct orientation.



2 Particles collide with sufficient energy in the incorrect orientation.



3 Particles collide with insufficient energy in the correct orientation.

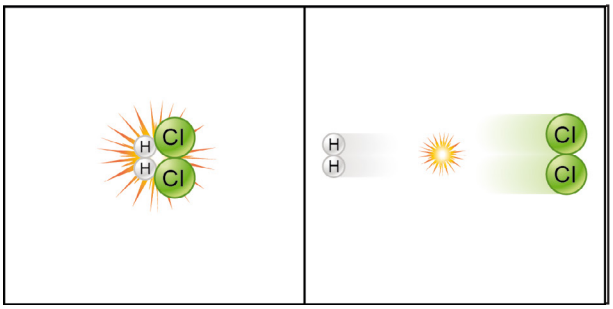


Figure 3.51: Reaction of H_2 and Cl_2

First, reactant particles must collide in a specific geometric arrangement that allows the reactive parts of each molecule to interact with one another. For example, molecules must often align side by side, rather than end to end, to align their atoms correctly for bond breaking and formation. Incorrect orientation results in an ineffective collision.

Second, the particles must collide with enough kinetic energy to break existing bonds and initiate a reaction. This minimum required energy is referred to as the activation energy (denoted as E_a). The value of E_a varies from one reaction to another. Reactions with a low activation energy occur easily, sometimes almost spontaneously, while reactions with a high activation energy may require additional energy input, such as heat or light, to proceed.

If either of these two conditions is not satisfied, the particles bounce off each other unchanged, and no reaction occurs. Next, we investigate the various factors that influence the rate of a chemical reaction, including temperature, surface area, concentration, gas pressure, the presence of a catalyst, activation energy, and particle orientation. Each of these factors affects either the frequency of collisions between reactant particles or the likelihood that a collision will be successful.

Concentration and Reaction Rate

Increasing the rate at which the box of red and blue balls is shaken leads to more frequent collisions between them. However, another effective way to increase collision frequency is to increase the number of balls in the box. As shown in **Figure 3.54**, the top panel displays a box containing four red and blue balls, while the bottom panel shows an increased number of red and blue balls. This increase raises the concentration of blue balls—that is, the number of blue balls per unit volume. By increasing the number of particles of either colour, the likelihood of collisions between red and blue balls also increases.

This concept directly applies to chemical reactions. In a solution, increasing the concentration of one or both reactants results in a greater number of reactant particles moving randomly within a given volume. This leads to more frequent collisions and, therefore, a higher reaction rate. However, the reaction will only proceed faster if the colliding particles possess kinetic energy equal to or greater than the activation energy.

For gas-phase reactions, increasing the pressure has a similar effect. This can be achieved by adding more gas to a fixed-volume container or by reducing the volume of a container with a flexible boundary, such as a gas syringe. Higher pressure increases the concentration of gas molecules, resulting in more frequent collisions. As with solutions, the reaction rate will only increase if the collisions involve sufficient energy to overcome the activation energy barrier.

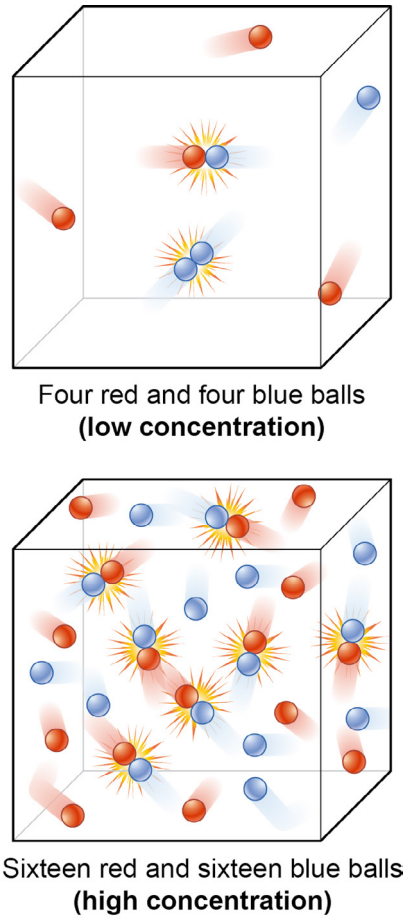


Figure 3.54: Effect of increasing reactant concentration on collision frequency

Surface Area and Reaction Rate

When a solid is involved in a chemical reaction, only the particles located on its surface are available to react, as the inner particles are not directly exposed to the surrounding reactants. The number of reactive surface particles depends on the surface area of the solid. As illustrated in **Figure 3.55**, breaking the solid into smaller pieces increases its total surface area, exposing more particles to potential collisions.

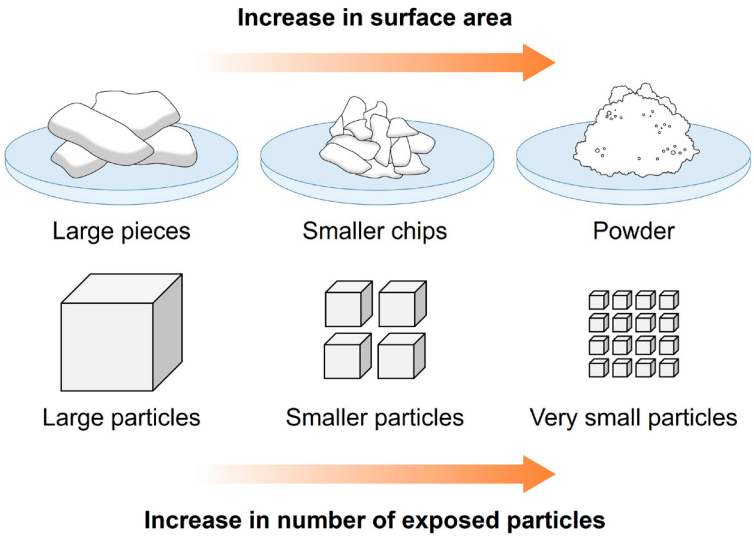


Figure 3.55: Effect of increasing surface area on number of exposed particles

With more surface particles available, the frequency of collisions between the solid and other reactant particles increases. This leads to a faster reaction rate, as a greater number of effective collisions can now occur within a given time. In essence, increasing the surface area of a solid reactant enhances its reactivity by providing more access points for interaction with other substances.

At equilibrium, the concentrations of both gases remain constant over time. However, when additional N_2O_4 is added to the system, its concentration increases, disrupting the equilibrium. In response, the system adjusts by converting some of the added N_2O_4 into NO_2, thereby partially counteracting the change. This results in an increase in the forward reaction rate until a new equilibrium is reached. We describe this as the equilibrium ***shifting to the right*** because, in a chemical equation, the forward reaction (reactants forming products) is written from left to right. A "shift to the right", therefore, indicates an increased formation of products—in this case, more NO_2 being produced from N_2O_4.

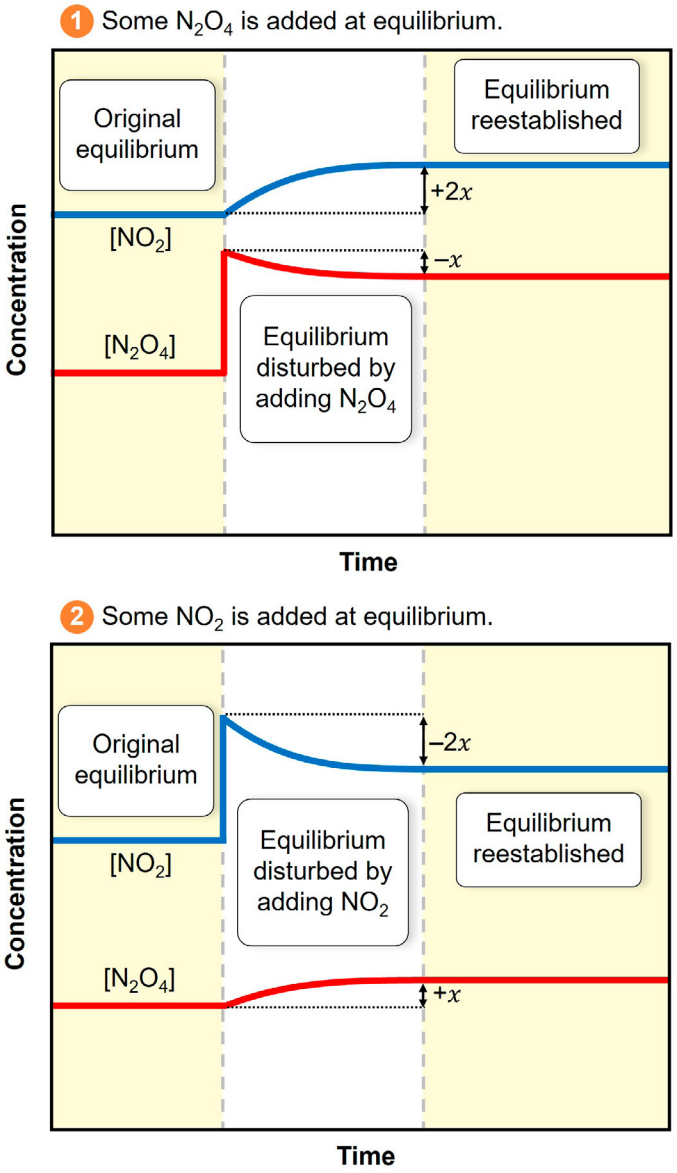


Figure 3.61: Effect of concentration changes at equilibrium.

This shift can be explained using collision theory: a higher concentration of reactant molecules (N_2O_4) results in more frequent collisions, increasing the rate at which NO_2 is formed. As the concentration of NO_2 rises, the rate of the back reaction also increases. Eventually, a new equilibrium is established, this time with a greater concentration of NO_2. Importantly, the ratio of concentrations at this new equilibrium still satisfies the original equilibrium constant (see **Chapter 3.5.4**).

This process is illustrated in **Figure 3.61** 1. The concentration-time graph shows an initial sharp increase in N_2O_4 concentration due to its sudden addition. Following this, the concentration of N_2O_4 declines as the system shifts to form more NO_2. The increase in NO_2 concentration is approximately double the decrease in N_2O_4, consistent with the reaction's stoichiometry.

In contrast, **Figure 3.61** 2 presents a different scenario in which NO_2 is added to the system, increasing its concentration and disturbing the equilibrium. In response, the system shifts to reduce the excess NO_2 by converting some of it back into N_2O_4. This causes the rate of the reverse reaction to increase until a new equilibrium is established. We describe this as a ***shift to the left*** because, in the chemical equation, the back reaction (products converting back into reactants) proceeds from left to right. A "shift to the left", therefore, indicates increased formation of reactants, in this case, more N_2O_4 being produced from NO_2.

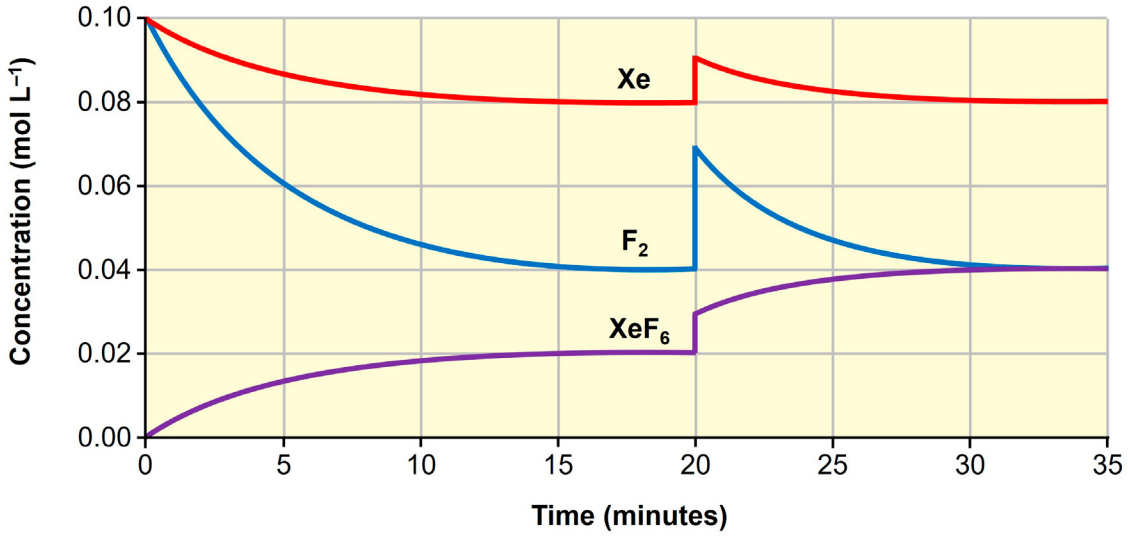
Changing the concentration of a reactant or product is one way to apply stress to a system at equilibrium. Another important factor is pressure, particularly in homogeneous systems of gases. In practice, pressure changes are typically achieved by altering the volume of the reaction vessel. When the pressure increases due to a reduction in volume, the system responds by shifting the equilibrium in the direction that reduces the total number of gas particles, thereby lowering the overall pressure. This happens because gas pressure is directly related to the frequency of particle collisions with the

Question 240

Xenon (Xe) reacts with fluorine according to the equation below.

$$Xe_{(g)} + 3F_{2\,(g)} \rightleftharpoons XeF_{6\,(g)}$$

The graph below shows changes in the concentrations of gases in one mixture.



Describe the change in concentration that occurred at the 20-minute mark and explain the changes in the reaction mixture that occurred during the subsequent 15 minutes.

(4 marks)

Question 241

Tin (Sn) has two allotropes. Grey tin is brittle and less stable than white tin.

The two allotropes are in equilibrium, represented by the equation below:

$$Sn_{\,(grey)} \rightleftharpoons Sn_{\,(white)} \quad \Delta H = +1.92 \text{ kJ mol}^{-1}$$

The great French military leader Napoleon Bonaparte led a disastrous campaign in Russia during the winter of 1812 in which the tin buttons on his infantry's uniforms disintegrated.

Use the information above to suggest how this occurred.

(3 marks)

3.6 Production of Chemicals Using Electrolysis

3.6.1 Electrolysis

Investigate the use and limitations of the electrochemical series to explain or predict the products of the electrolysis of particular chemicals, given their state (molten liquid or in aqueous solution) and the electrode materials used, including the writing of balanced equations (with states) for the reactions occurring at the anode and cathode and the overall redox reaction for the cell.

Discuss the common design features and general operating principles of commercial electrolytic cells (including, where practicable, the removal of products as they form), and the selection of suitable electrode materials, the electrolyte (including its state) and any chemical additives that result in a desired electrolysis product (details of specific cells not required).

Electrolysis is a key process in both laboratory and industrial chemistry, used to drive non-spontaneous redox reactions by applying an external electric current. This chapter explores the design features and operating principles of commercial electrolytic cells, including the electrode reactions, electrode selection, electrolyte properties, and strategies for removing products as they form. Additionally, this chapter explores how the electrochemical series is used—and its limitations—to predict the products of electrolysis, considering the chemical's state (molten or aqueous) and the materials used for the electrodes.

Electrolysis

Chapter 3.3 introduced the galvanic cell, an electrochemical system that converts chemical energy into electrical energy through spontaneous redox reactions occurring at two electrodes: the anode, where oxidation occurs, and the cathode, where reduction takes place. In galvanic cells, the flow of electrons from the anode to the cathode through an external circuit is driven by the natural tendency of certain redox reactions to occur spontaneously. In contrast, this chapter focuses on the **electrolytic cell**, which operates on the opposite principle. Rather than generating electricity, an electrolytic cell consumes electrical energy to drive a non-spontaneous redox reaction. This process, known as **electrolysis**, has a wide range of applications, including the extraction and purification of metals, the industrial production of gases, and electroplating.

The basic structure consists of two electrodes, an anode and a cathode, immersed in an electrolyte, which may be either a molten ionic compound or an aqueous solution. As shown in **Figure 3.68**, the anode is connected to the positive terminal of the power supply, thereby becoming the positive electrode, while the cathode is connected to the negative terminal, thereby becoming the negative electrode. At the anode, oxidation occurs as electrons are drawn away from the substance being oxidised. These electrons then flow through the external circuit toward the cathode, where reduction takes place as another species in the electrolyte accepts them. The electrolyte contains both positive and negative ions. To maintain charge balance, cations (C^+) migrate toward the cathode to gain electrons, while anions (A^-) move toward the anode to lose electrons. Because both electrodes are immersed in a single electrolyte solution, there is no need for a salt bridge—the internal circuit is maintained entirely through ion migration.

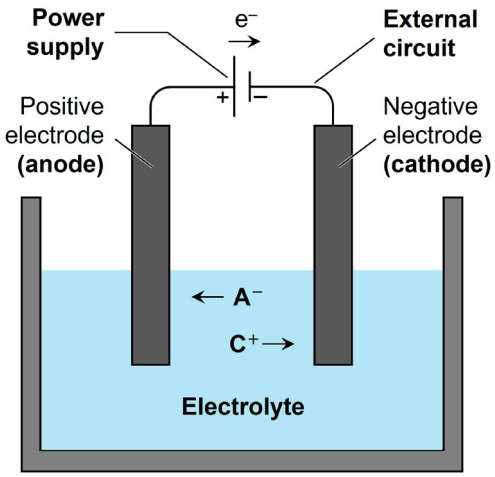
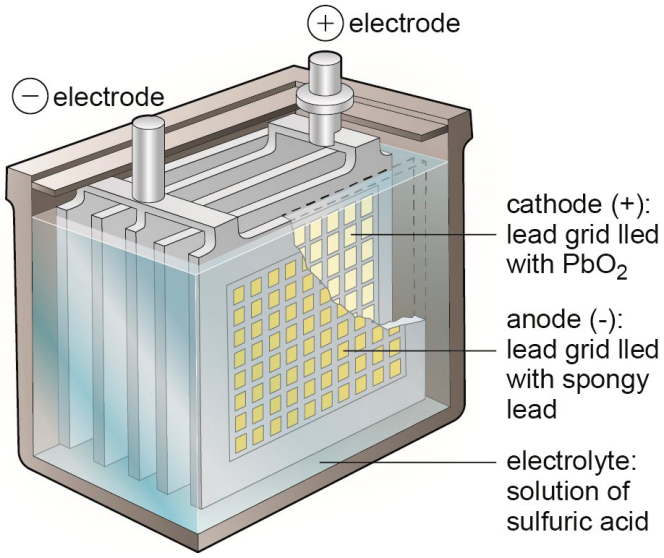


Figure 3.68: Electrolytic cell

Question 315

The diagram below illustrates the primary components of a lead-acid battery commonly used in motor vehicles.



The half-equations below describe the oxidation and reduction reactions when the cell discharges.

Anode (oxidation): $Pb_{(s)} + SO_4^{2-}{}_{(aq)} \rightarrow PbSO_{4\,(s)} + 2e^-$

Cathode (reduction): $PbO_{2\,(s)} + 4H^+{}_{(aq)} + SO_4^{2-}{}_{(aq)} + 2e^- \rightarrow PbSO_{4\,(s)} + 2H_2O_{(l)}$

(a) Write the formula of the reactant that is reduced when the cell is discharging.

(1 mark)

(b) State the charge on the cathode when the cell is discharging.

(1 mark)

(c) Write an equation for the overall reaction when the cell is discharging.

(1 mark)

(d) Suggest why the electrolyte commonly freezes in winter following discharging.

(1 mark)

(e) Lead-acid batteries are rechargeable.

i. Write an equation for the overall reaction when the cell is recharging.

(1 mark)

ii. State the charge on the anode when the cell is recharging.

(1 mark)

Structural Properties

One of the most important properties of carbon is **catenation**, the ability to bond with other carbon atoms to form long chains, branched structures, and rings (see Figure 4.02). This simple property explains why carbon builds an enormous variety of substances, from small molecules like propane to useful plastics such as polyethene and even giant biomolecules like cellulose in plant cell walls.

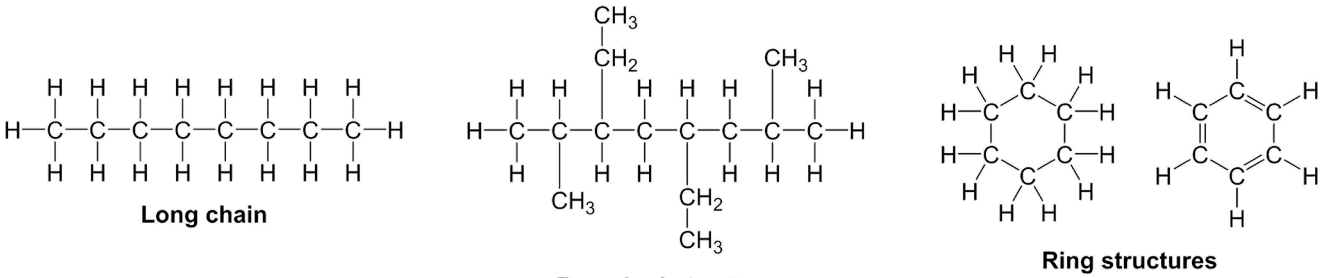


Figure 4.02: Catenation in organic compounds

Another important property is the ability of carbon to form single, double, and triple covalent bonds (see Figure 4.03), and each type behaves differently. Single bonds are flexible because they allow atoms to rotate, while double bonds make parts of a molecule rigid and flat, and triple bonds produce straight, linear sections. These differences affect how molecules react. For example, alkenes and alkynes (with double and triple bonds) are much more reactive than alkanes (with only single bonds), making them especially useful in both industry and living systems.

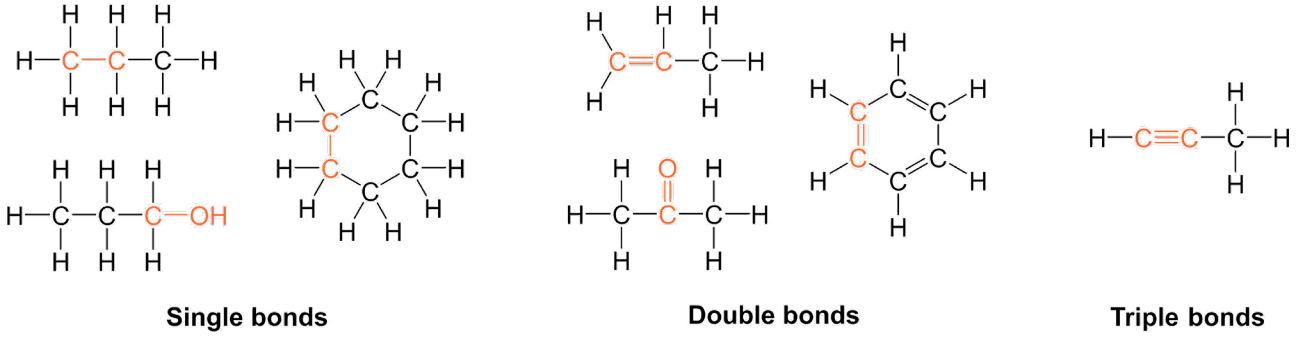


Figure 4.03: Single, double and triple covalent bonds in organic compounds

Catenation and carbon's four bonding electrons also lead to **isomerism**, where compounds have the same formula but different structures. A simple example is C_6H_{14}, which can exist as a straight-chain compound, hexane, or several branched forms, including 2-methylpentane, 3-methylpentane, 2,2-dimethylbutane, and 2,3-dimethylbutane (see Figure 4.04). As molecules get larger, the number of possible isomers rapidly increases, creating millions of compounds from relatively few carbon atoms.

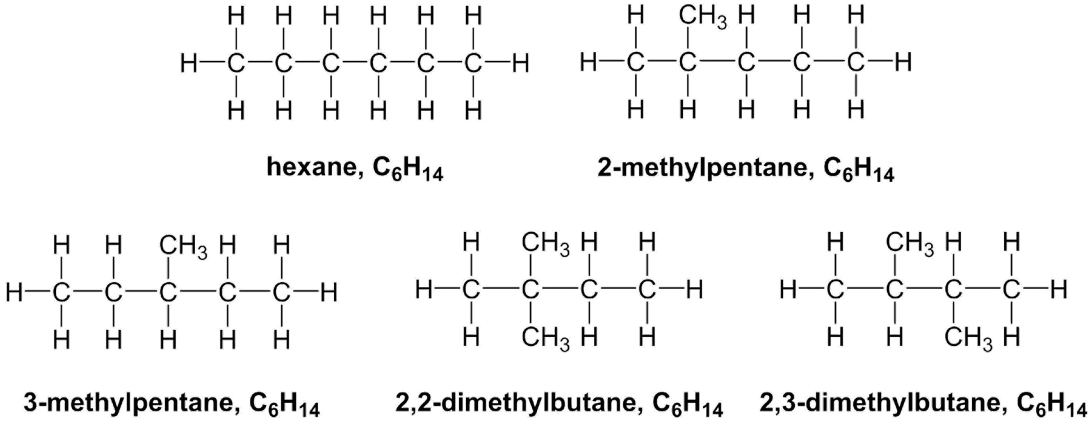


Figure 4.04: Isomerism in C_6H_{14}

Functional Groups

In addition to forming structural isomers, carbon compounds gain even greater complexity through the presence of **functional groups:** specific atoms or groups of atoms (see **Figure 4.05**) that determine how a molecule reacts chemically and interacts physically with other molecules. Each functional group gives a compound its own chemical and physical properties, allowing the classification of millions of carbon compounds into recognisable families.

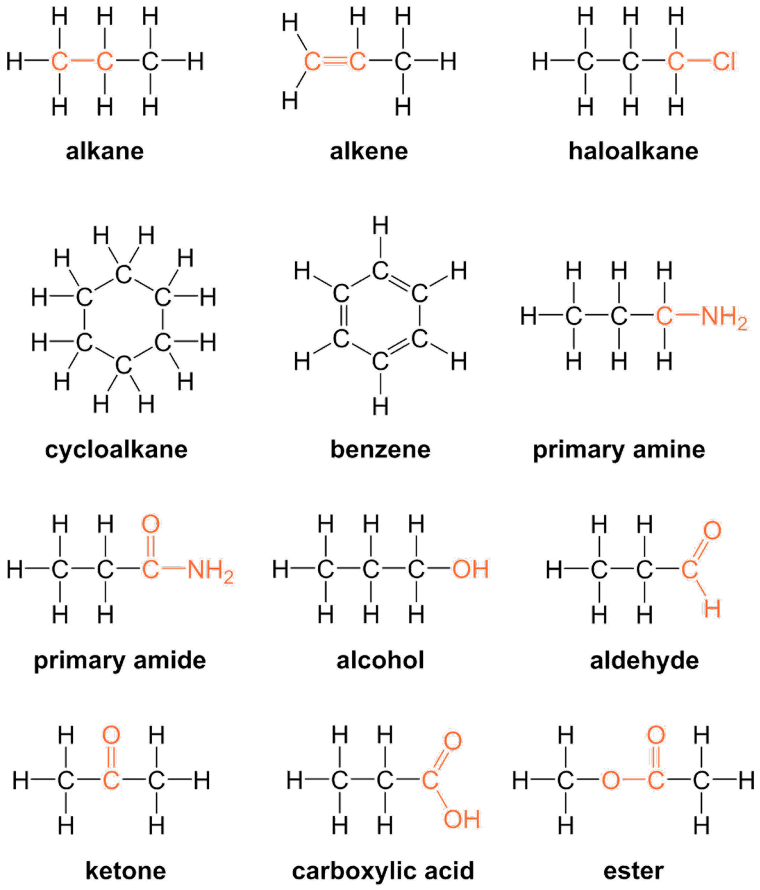


Figure 4.05: Functional groups

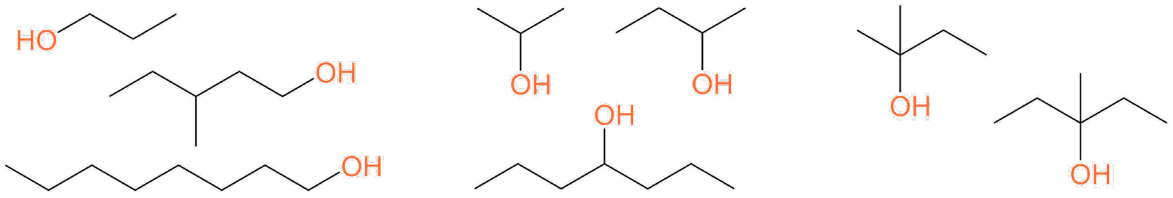
The simplest family of organic compounds are the **alkanes**, which contain only carbon–carbon single bonds. They can exist as straight or branched chains or as rings, such as **cyclohexane**, and are relatively unreactive due to their strong, stable C–C and C–H bonds. Closely related are the **alkenes**, which contain at least one carbon–carbon double bond. This makes them more chemically reactive, especially in addition reactions. A special case of bonding is seen in **benzene**, a cyclic hydrocarbon in which electrons are shared evenly around the ring, giving it unusual stability and characteristic reactions. Adding substituents to these hydrocarbon skeletons introduces further diversity. In **haloalkanes**, one or more hydrogens are replaced with halogen atoms such as chlorine, bromine, or iodine, creating compounds with important industrial and pharmaceutical applications.

Introducing a nitrogen group produces new families: primary **amines**, which contain the $–NH_2$ group, and primary **amides**, which feature a $–CONH_2$ group linking a carbonyl to nitrogen, both important in biological systems and materials. Oxygen-containing groups lead to some of the most significant families in organic chemistry. **Alcohols** are characterised by the hydroxyl group (–OH) attached to a carbon atom and can be classified as primary, secondary, or tertiary depending on how many other carbons are bonded to the carbon bearing the hydroxyl group. The carbonyl group (C=O) defines both **aldehydes**, where it is bonded to at least one hydrogen, and **ketones**, where it is bonded to two carbons. Further oxidation produces **carboxylic acids**, characterised by the –COOH group, which is acidic and plays a central role in metabolism. When a carboxylic acid reacts with an alcohol, the result is an **ester**, a compound with the –COO– linkage that often has distinctive, pleasant fragrances and is widely used in flavourings and polymers.

Together, these families of compounds demonstrate the remarkable versatility of carbon, in which small changes in bonding or substituents yield a vast and diverse chemistry that underpins life and drives countless applications in industry, medicine, and technology. In the chapters that follow, we will examine the structure and systematic naming of each major family of organic compounds, explore their characteristic functional groups and reactivity, and consider the chemical and physical tests used to identify them.

Alcohols

Alcohols are organic compounds in which a **hydroxyl group (–OH)** is bonded to a carbon atom. The presence of this polar functional group significantly alters the molecule's physical and chemical properties compared to hydrocarbons, increasing the boiling point, improving solubility in water, and enabling reactions such as oxidation and substitution. Alcohols are commonly classified as primary, secondary, or tertiary, depending on the number of carbon atoms bonded to the carbon that carries the –OH group. In a **primary alcohol**, the –OH-bearing carbon is attached to only one other carbon atom; in a **secondary alcohol**, it is bonded to two; and in a **tertiary alcohol**, it is bonded to three (see Figure 4.13). Alcohols can also vary in chain length, branching, and position of the hydroxyl group, creating numerous structural isomers. Their versatility makes alcohols fundamental in industry, where they serve as solvents, fuels, and disinfectants.

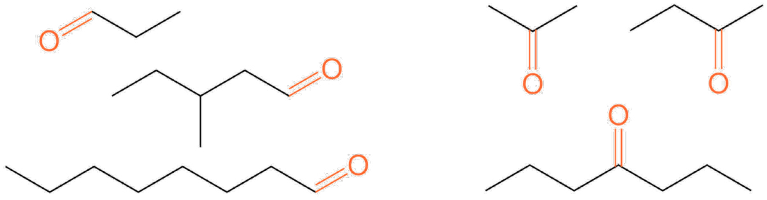


Primary alcohols **Secondary alcohols** **Tertiary alcohols**

Figure 4.13: Alcohols

Aldehydes and Ketones

Aldehydes and **ketones** are organic compounds characterised by the presence of a **carbonyl group** (C=O), a highly reactive functional group that strongly influences their chemical behaviour. In aldehydes, the carbonyl carbon is bonded to at least one hydrogen atom and is always located at the end of a carbon chain (see Figure 4.14). In ketones, the carbonyl carbon is bonded to two other carbon atoms and is therefore positioned within the carbon chain (see Figure 4.14). This structural difference between aldehydes and ketones affects their reactivity: aldehydes are generally more reactive, particularly in oxidation reactions, where they can be readily converted to carboxylic acids, whereas ketones resist further oxidation under normal conditions.



Aldehydes **Ketone**

Figure 4.14: Aldehydes and ketones

Carboxylic acids

Carboxylic acids are organic compounds that contain the carboxyl functional group (–COOH), which combines a carbonyl group (C=O) and a hydroxyl group (–OH) on the same carbon atom. This functional group always occurs at the end of a carbon chain (see Figure 4.15). Their acidity arises from the ability of the –OH group to donate a proton (H^+), forming a carboxylate ion ($R–COO^-$). Structurally, carboxylic acids can vary from simple molecules such as methanoic acid to long-chain fatty acids found in lipids. Their reactivity is also versatile: they can undergo neutralisation to form salts, and condensation reactions to form esters and amides.

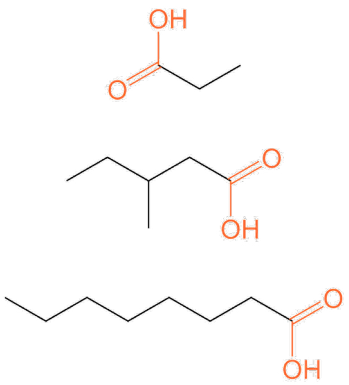


Figure 4.15: Carboxylic acids

Esters

Esters are organic compounds characterised by the functional group –R-COO-R', where the carbonyl carbon (C=O) is directly bonded to an oxygen atom that is further connected to a hydrocarbon group (R) (see **Figure 4.16**). This structure is formed by a condensation reaction between a carboxylic acid and an alcohol. Esters are not terminal functional groups like carboxylic acids; instead, they occur within molecules, bridging two carbon chains (labelled R and R'). Structurally, esters exhibit great diversity, ranging from small molecules, such as methyl ethanoate, to large, complex esters found in fats and oils, such as triglycerides (see **Chapter 3.1.2**). A defining feature of esters is their distinctive and often pleasant odours, which makes them important in flavourings and fragrances. Industrially, they are also vital as solvents, plasticisers, and intermediates in the production of polymers and pharmaceuticals.

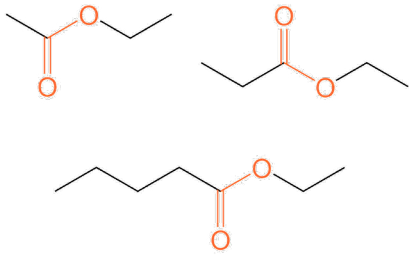


Figure 4.16: Esters

Amines

Amines are organic compounds derived from ammonia (NH_3) in which one or more of the hydrogen atoms are replaced by hydrocarbon groups. In **primary amines**, one hydrogen is substituted ($–NH_2$ group attached to a carbon chain); in **secondary amines**, two hydrogens are replaced, leaving the nitrogen bonded to two carbon groups; and in **tertiary amines**, all three hydrogens are replaced, with the nitrogen bonded to three carbon groups (see **Figure 4.17**). In this course, only the structure and systematic nomenclature of primary amines are considered. This classification has a significant impact on their chemical and physical properties. Amines contain a lone pair of electrons on the nitrogen atom, which allows them to act as bases by accepting protons in chemical reactions. Structurally, amines range from simple molecules, such as methanamine, to large, biologically essential compounds, including amino acids and neurotransmitters. Their strong, often fishy odour is characteristic of many low-mass amines. Amines play a crucial role in both living systems and industry, serving as the building blocks of proteins, pharmaceuticals, dyes, and synthetic materials.

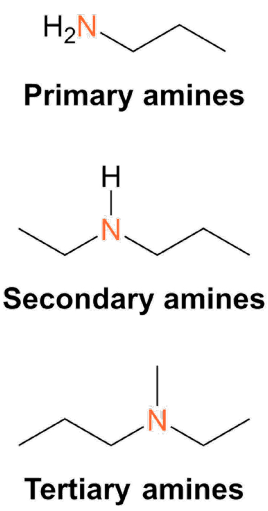


Figure 4.17: Amines

Amides

Amides are organic compounds that contain the functional group –CON, in which a carbonyl group (C=O) is directly bonded to a nitrogen atom (see **Figure 4.18**). They are typically formed through the reaction of carboxylic acids with amines, making them closely related to both carboxylic acids and amines. In **primary amides**, the nitrogen is bonded to one carbonyl carbon and two hydrogens ($–CONH_2$); in **secondary amides**, one of these hydrogens is replaced by a hydrocarbon group (–CONHR); and in tertiary amides, both hydrogens are replaced, leaving the nitrogen bonded to two carbon groups ($–CONR_2$). Structurally, amides play a crucial role in biology, as they form the peptide bonds that link amino acids together in proteins. They also occur in many synthetic materials, such as nylon and Kevlar, highlighting their central role in both living systems and industrial applications.

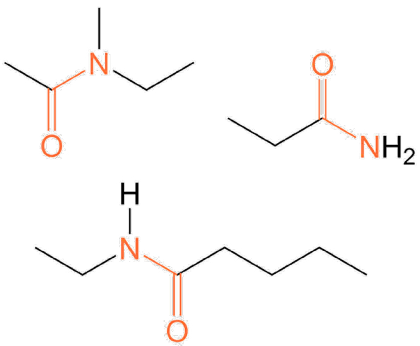


Figure 4.18: Amides

Naming Carboxylic Acids

In IUPAC nomenclature, carboxylic acids are named by replacing the –e of the corresponding alkane with the suffix –oic acid. The carbon of the carboxyl group is always given position 1 in the parent chain, so numbering begins from this carbon, making it unnecessary to include its position in the name (see Figure 4.27). For example, a one-carbon acid is methanoic acid, while a two-carbon acid is ethanoic acid. If substituents are present, their positions are indicated using numbers relative to the carboxyl carbon, and they are listed as prefixes in alphabetical order.

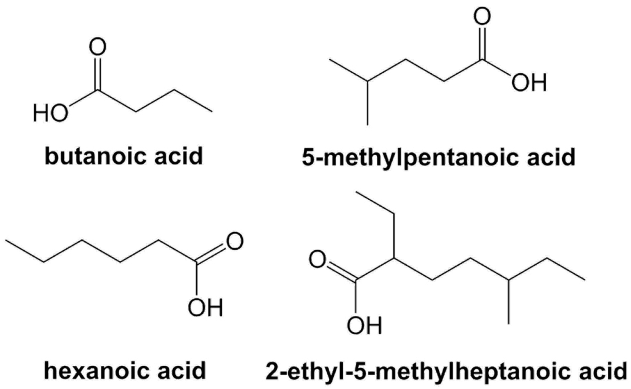


Figure 4.27: Carboxylic acids

Naming Esters

Esters are organic compounds formed from the reaction of a carboxylic acid with an alcohol. In IUPAC nomenclature, esters are named in two parts. The first part comes from the alcohol and is written as an alkyl group, indicating the length of the alcohol chain (e.g., methanol → methyl, ethanol → ethyl, propan-1-ol → propyl). The second part comes from the carboxylic acid, with its name modified to end in the suffix –oate (e.g., methanoic acid → methanoate, ethanoic acid → ethanoate, propanoic acid → propanoate). Combining the two parts gives the ester's name. For instance, the ester formed from methanol and ethanoic acid is named methyl ethanoate (see Figure 4.28). In simple, non-branched esters, no numbering is required because the structure is straightforward: one chain comes from the acid and the other from the alcohol.

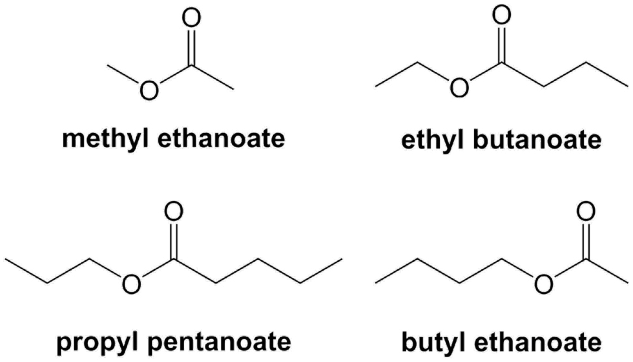


Figure 4.28: Esters

Naming Primary Amines

Primary amines are organic compounds in which a nitrogen atom is bonded to a carbon atom by a single covalent bond. Since nitrogen forms three covalent bonds in total, the remaining two are with hydrogen atoms, giving rise to the functional group $–NH_2$ attached to a carbon chain. In IUPAC nomenclature, they are named by adding the suffix –amine to the name of the parent alkane. For example, methanamine (CH_3NH_2) is the simplest primary amine. The carbon chain is numbered so that the $–NH_2$ group is given the lowest possible position number (see Figure 4.29). If the amine group is not the main functional group, it is treated as a substituent and named with the prefix amino–. For instance, 4-aminohexanoic acid contains both an amino group and a carboxylic acid group.

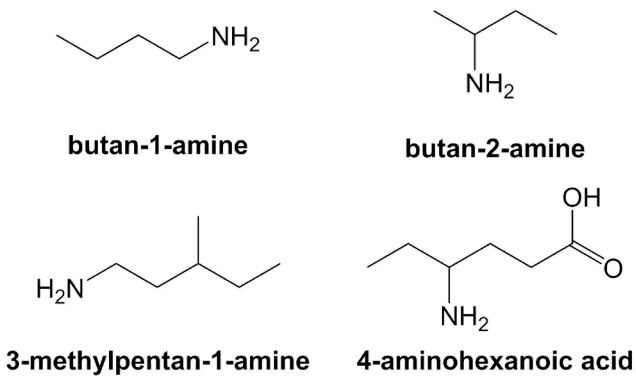
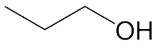
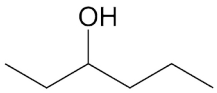


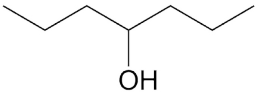
Figure 4.29: Primary amines

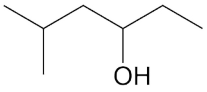
Question 360

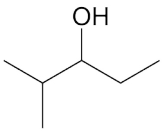
Write the systematic name of the following alcohols.

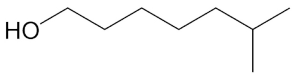
(a) 

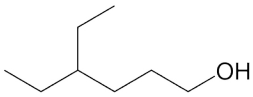
(b) 

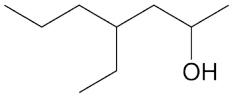
(c) 

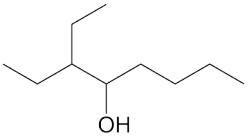
(d) 

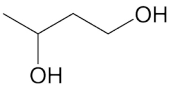
(e) 

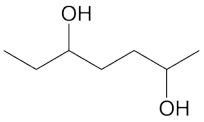
(f) 

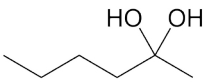
(g) 

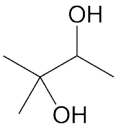
(h) 

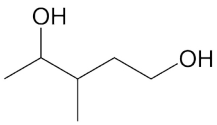
(i) 

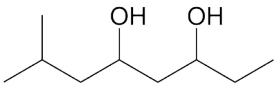
(j) 

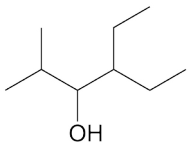
(k) 

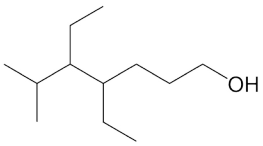
(l) 

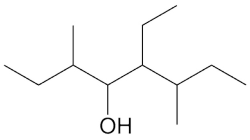
(m) 

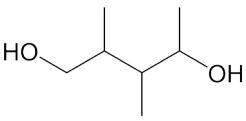
(n) 

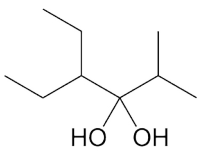
(o) 

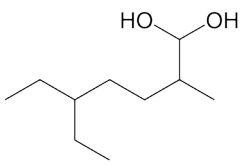
(p) 

(q) 

(r) 

(s) 

(t) 

(u) 

(21 marks)

4.3: Analysis of Organic Compounds

4.3.1 Qualitative Tests for Functional Groups

Investigate the qualitative tests for the presence of carbon-carbon double bonds, hydroxyl and carboxyl functional groups.

Functional groups are the defining features of organic compounds, determining their chemical reactivity and physical properties. Detecting these groups is an essential skill in organic chemistry, allowing chemists to identify and classify unknown substances. Simple qualitative tests provide clear visual evidence of particular functional groups through observable changes such as colour shifts, effervescence, or the release of characteristic odours. This chapter investigates three key tests: the use of bromine water to detect carbon–carbon double bonds, the use of acidified dichromate solution to confirm the presence of hydroxyl groups, and procedures for identifying carboxyl groups. Together, these tests illustrate practical applications of structure–property relationships.

Testing For Functional groups

The carbon–carbon double bond (C=C), characteristic of alkenes and unsaturated organic compounds, can be detected using bromine water. In this test, a few drops of bromine (Br_2 (l)) are shaken with the sample. If a double bond is present, the bromine reacts across it in an addition reaction, rapidly decolourising the solution from **brown to colourless** (Figure 4.48). However, if bromine water ($Br_{2(aq)}$) is used, the colour change is from **orange to colourless**.

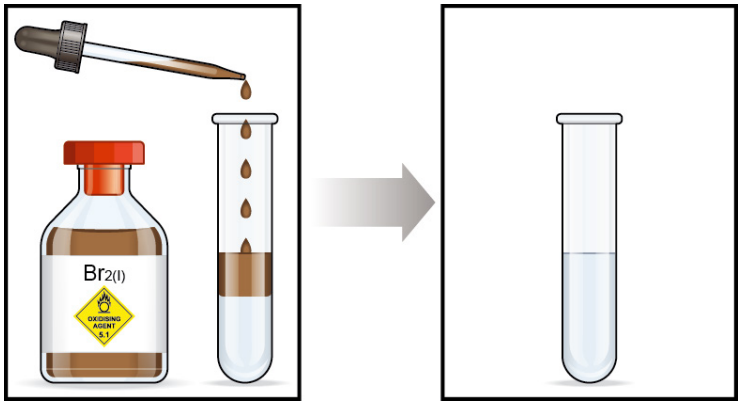


Figure 4.48: Bromine test for C=C

The hydroxyl group (–OH) is a defining feature of alcohols, which are classified as primary, secondary, or tertiary depending on the number of carbon atoms attached to the carbon bearing the –OH group. In primary and secondary alcohols, this carbon is bonded to at least one hydrogen atom. The presence of this hydrogen allows these alcohols to undergo oxidation when treated with an oxidising agent such as acidified potassium dichromate. During the reaction, electrons are transferred to the dichromate ion ($Cr_2O_7^{2-}$), reducing it to chromium(III) ions (Cr^{3+}), resulting in a colour change from **orange to green**, making the test a useful indicator of oxidation (see Figure 4.49). In contrast, tertiary alcohols lack a hydrogen atom on the carbon bonded to the –OH group. As a result, they cannot be oxidised under normal conditions with acidified dichromate, allowing them to be readily distinguished from primary and secondary alcohols.

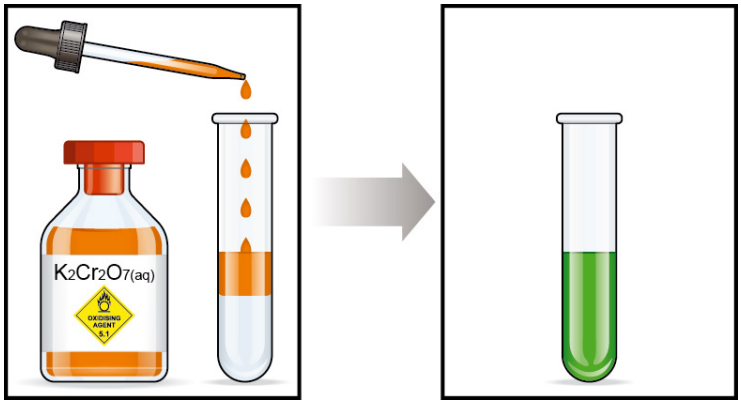


Figure 4.49: Dichromate test for –OH

Fractional distillation, by contrast, is used when the boiling points of the components are closer together, usually less than 25°C apart. A fractionating column filled with glass beads or plates provides repeated cycles of condensation and vaporisation, improving separation (see Figure 4.52). This technique is especially useful for analysing and purifying mixtures such as liquid hydrocarbons in petroleum. In laboratory analysis, fractional distillation provides more precise separation, enabling accurate determination of composition and purity.

Figure 4.52: Fractional distillation

Measuring the Degree of Unsaturation Using Iodine

The degree of unsaturation of an organic compound reflects the number of carbon–carbon double or triple bonds in its structure. A common method for determining this is the **iodine value test**, which measures the amount of iodine a compound can absorb. Iodine undergoes addition reactions with carbon–carbon multiple bonds, with each double or triple bond consuming one molecule of iodine (I_2). The greater the unsaturation, the more iodine is absorbed. Iodine solutions are typically brown when dissolved in organic solvents such as cyclohexane or ethanol. As iodine reacts with unsaturated bonds, it is consumed, causing the brown colour to fade and eventually disappear, resulting in a colourless solution (see Figure 4.53). This decolourisation provides a clear and simple indication of unsaturation, allowing both qualitative identification of double or triple bonds and quantitative assessment of the degree of unsaturation.

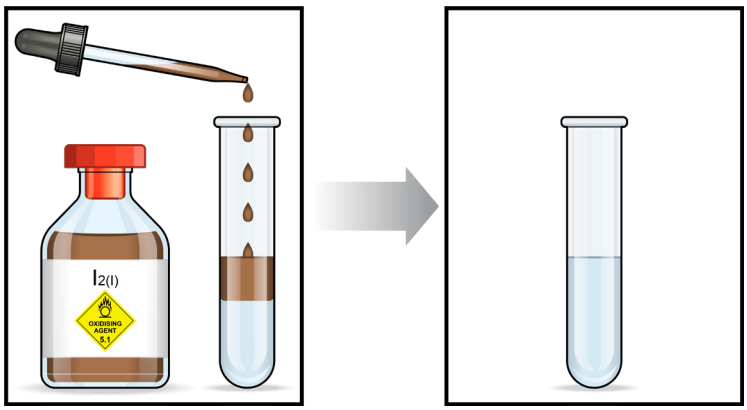


Figure 4.53: Iodine test for C=C

In practice, the compound, often an alkene, oil or fat, is treated with a known excess of iodine solution. The unreacted iodine is then determined by titration with sodium thiosulfate solution, using starch as an indicator. The amount of iodine that reacted with the compound is calculated by difference, yielding the **iodine value**, which is usually expressed as grams of iodine absorbed per 100 grams of substance. A higher iodine number indicates a greater degree of unsaturation, as seen in polyunsaturated oils compared to saturated fats.

This method is particularly useful for analysing edible oils and fats. For example, polyunsaturated oils, such as linseed oil, have a much higher iodine value than saturated fats, like butter. Measuring the degree of unsaturation is crucial for evaluating the stability, nutritional properties, and potential industrial applications of organic compounds.

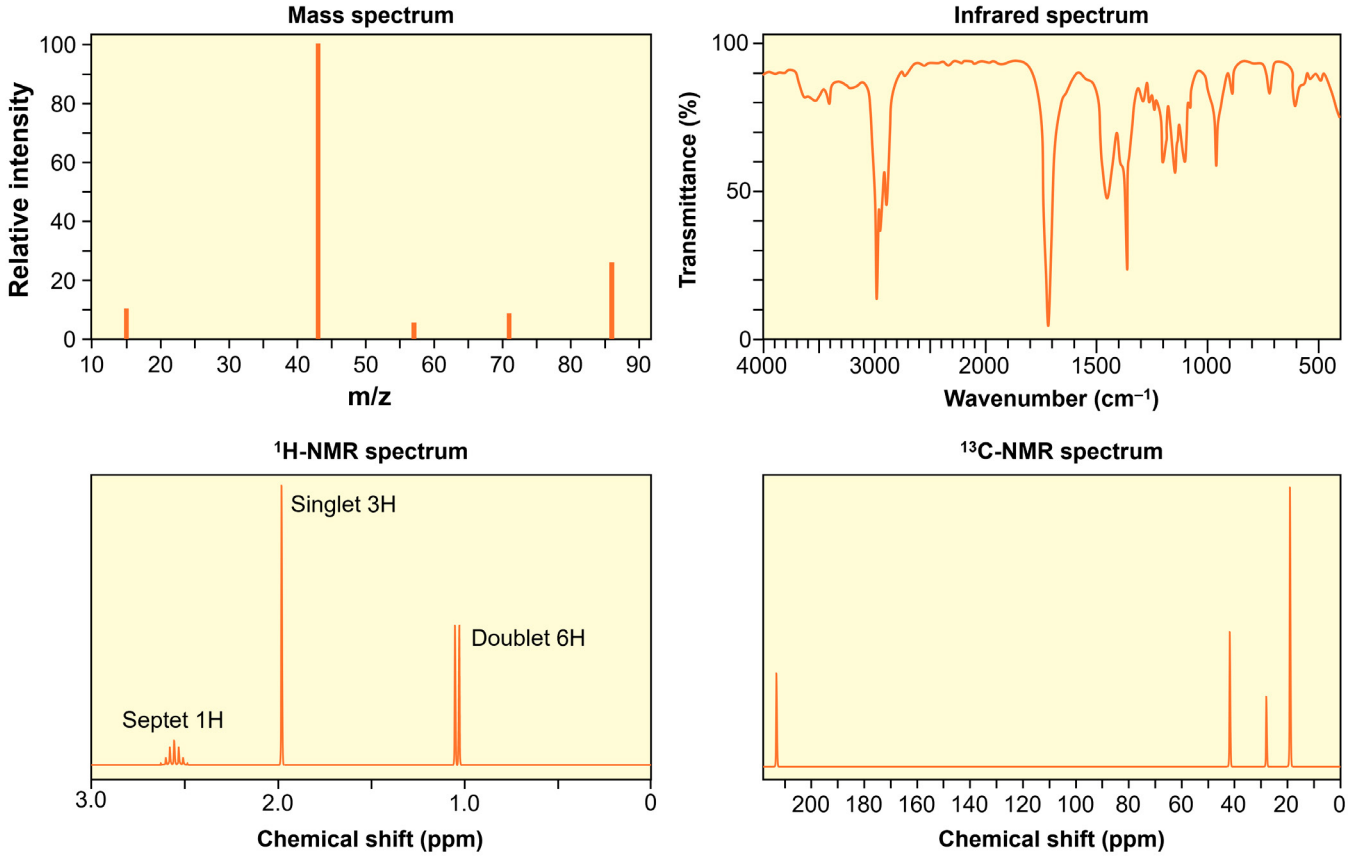
4.4.5 Spectrometry

Deduce the structures of simple organic compounds using a combination of mass spectrometry (MS), infrared spectroscopy (IR), proton nuclear magnetic resonance (^{1}H-NMR) and carbon-13 nuclear magnetic resonance (^{13}C-NMR) (limited to data analysis).

Modern organic chemistry depends on spectroscopic techniques to identify and confirm the structures of compounds. Four of the most widely used methods are mass spectrometry (MS), infrared spectroscopy (IR), proton nuclear magnetic resonance (^{1}H-NMR), and carbon-13 nuclear magnetic resonance (^{13}C-NMR). Each technique provides unique and complementary information: MS determines the molecular mass and characteristic fragmentation patterns, IR spectroscopy identifies functional groups through their distinct bond vibrations, and NMR spectroscopy reveals the number, type, and chemical environments of hydrogen and carbon atoms within the molecule. By interpreting these data collectively, chemists can eliminate incorrect structural possibilities and deduce the correct molecular structure with high confidence. Combined spectroscopic analysis, therefore, represents one of the most powerful and reliable tools in modern organic chemistry.

Example 4.09

The following spectra were obtained for an unknown organic compound.



The mass, infrared, ^{1}H-NMR, and ^{13}C NMR spectra together provide comprehensive evidence for the compound's molecular structure. A systematic approach involves analysing the information obtained from each spectrum and summarising the key evidence and inferences in a table, as shown on the following page.

Mass Spectrum

Evidence	Inference
Molecular ion at m/z 86	Molar mass is approximately 86 g mol^{-1}
Base peak at m/z = 43	The molecule likely contains a CH_3CO^+ fragment, characteristic of a carbonyl group adjacent to a methyl group.
Peaks at m/z 15 and 71	Two pieces of evidence that indicate the presence of a methyl group in the molecule.

Infrared Spectrum

Evidence	Inference
Strong peak at 1715-1725 cm^{-1}	Characteristic of a carbonyl (C=O) group in a ketone.
No broad O–H band between 2500 and 3300 cm^{-1}	Excludes carboxylic acids and alcohols.
No aldehydic C–H bands between 2720 and 2820 cm^{-1}	Excludes aldehydes.

^{1}H-NMR Spectrum

Evidence	Inference
Septet, 1H (~2.4–2.8 ppm)	A single –CH– split by six equivalent neighbours, which is only possible if that CH is attached to two equivalent CH_3 groups. This indicates the molecule has a $–CH(CH_3)_2$ group. The large chemical shift is consistent with being near a carbonyl group.
Doublet, 6H (~1.0–1.2 ppm)	Two equivalent methyl groups each coupled to one neighbouring proton (the –CH–), further evidence of a $–CH(CH_3)_2$ group.
Singlet, 3H (~2.0–2.2 ppm)	A methyl with no hydrogens on the adjacent carbon atom. This indicates the methyl group is likely adjacent to the carbonyl group ($–CO–CH_3$).

^{13}C-NMR Spectrum

Evidence	Inference
~220 ppm	A single strong resonance from a ketone carbonyl carbon.
~35–45 ppm	A –CH– next to a carbonyl group
~28 ppm	A methyl group adjacent to a carbonyl
~18 ppm	A pair of CH_3 groups on the same carbon atom.

All of the spectral evidence supports the identification of the compound as **3-methylbutan-2-one**. The data show a molecule containing a carbonyl group, a CH group bonded to two identical methyl groups, and another methyl group directly attached to the carbonyl carbon. The mass spectrum includes a molecular ion peak at m/z = 86, corresponding to a molar mass of 86 g mol^{-1}, which matches the molecular formula $C_5H_{10}O$.

The ^{1}H-NMR spectrum shows three signals that fit perfectly with the structure of $CH_3–CO–CH(CH_3)_2$. This pattern differs from that of the straight-chain isomer 3-pentanone, which would display triplet–quartet patterns from the $–CH_2–CH_3$ groups rather than the distinctive septet–doublet pattern of an isopropyl group.

The ^{13}C-NMR spectrum also supports this conclusion. While 3-pentanone would produce only three carbon signals (for the equivalent CH_3, CH_2, and carbonyl carbons), the observed four distinct signals confirm the presence of a branched isopropyl ketone, consistent with 3-methylbutan-2-one.

4.5: Medicinal Chemistry

4.5.1 Extraction and Purification of Natural Compounds

Investigate the extraction and purification of natural plant compounds as possible active ingredients for medicines, using solvent extraction and distillation.

Plants have long been a source of biologically active compounds that form the basis of many traditional and modern medicines. Compounds such as alkaloids, essential oils, and flavonoids can serve as active ingredients with therapeutic properties; however, they are often found in complex mixtures with other plant materials. To study and utilise these compounds effectively, reliable extraction and purification methods are required. Solvent extraction is a widely used technique that exploits the solubility of target molecules in specific solvents to separate them from unwanted components. Following extraction, distillation can be applied to purify volatile substances, such as essential oils, by exploiting differences in boiling points. Together, these methods enable chemists to isolate, concentrate, and evaluate plant-derived compounds for their potential medicinal properties. This section examines the practical applications of solvent extraction and distillation, offering insight into the scientific processes that underpin natural product chemistry and drug discovery.

Extraction and Purification of Natural Compounds

The extraction and purification of natural compounds from plants begins with selecting a suitable solvent. The solvent must dissolve the desired compound efficiently while leaving behind most unwanted materials, such as cellulose or pigments. For example, ethanol and water are commonly used for polar compounds, such as alkaloids, while non-polar solvents, like hexane, are suitable for oils and terpenes. In a typical solvent extraction process, the plant material is crushed or ground to increase its surface area, mixed with the solvent, and then filtered to remove insoluble matter (see Figure 4.76). The solution containing the target compound can be concentrated by gentle heating or by evaporating the solvent under reduced pressure, leaving behind the crude extract.

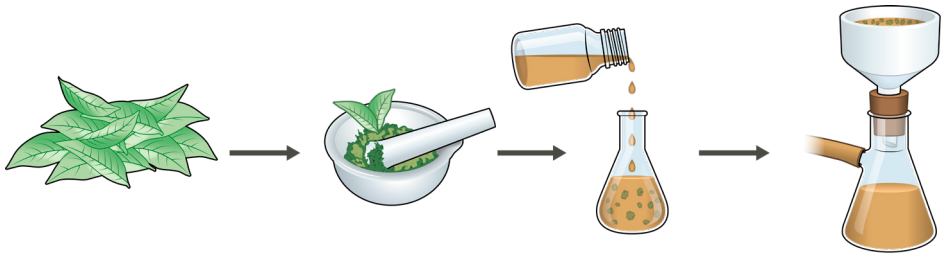


Figure 4.76: Extraction and filtration of a natural compound from plant leaves

Following extraction and filtration, the filtrate is transferred to a **separating funnel** (see Figure 4.77). A solvent of opposite polarity to the desired product is added, forming two layers. When the natural product is non-polar, water is added as a polar solvent that dissolves any polar impurities, separating them from the mixture. In contrast, if the natural product is polar, a non-polar solvent is added, which dissolves any non-polar impurities, separating them from the desired product. Following separation, excess solvent is removed by **rotary evaporation** using a rotary evaporimeter (see Figure 4.78). In a rotary evaporation, the sample is placed in a round-bottom flask that rotates in a temperature-controlled water bath while a vacuum is applied. Rotation spreads the liquid into a thin film, greatly

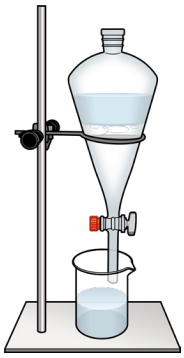


Figure 4.77: Separation

increasing the surface area, and reduced pressure lowers the liquid's boiling point, so the solvent vaporises at modest temperatures. Vapours are condensed in a cooled coil and collected in a receiving flask, allowing easy separation of the solvent from the natural product. This makes rotary evaporation faster, gentler on heat-sensitive natural products, and more energy-efficient than distillation. Rotary evaporation is for removing solvent, not for fine separations; if two or more solutes are present once the solvent has been discarded, they must be separated and the natural product isolated using distillation.

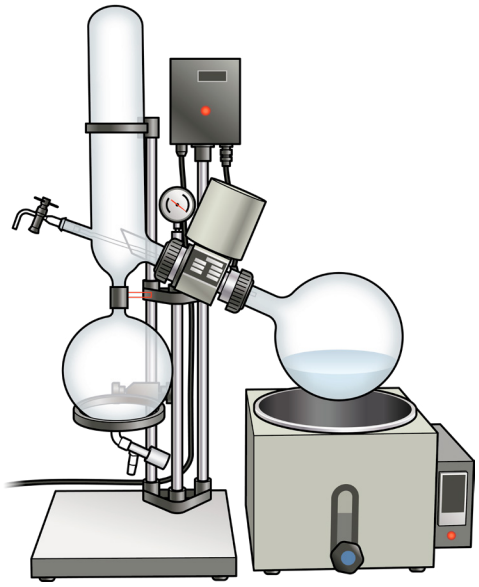


Figure 4.78: Evaporation

Distillation techniques are used to purify the substance. Simple distillation can separate a volatile active ingredient from less volatile residues, while fractional distillation provides finer separation when impurities have boiling points similar to those of the natural product being isolated.

Once extracted and purified, the compound's identity is confirmed by spectroscopy (see **Figure 4.79**). Mass spectrometry (MS) provides the exact mass and isotope pattern of its molecules; infrared (IR) spectroscopy verifies key functional groups; and nuclear magnetic resonance (NMR) spectroscopy establishes atomic connectivity. Taken together, these data verify that the target compound has been isolated. Purity is then assessed by chromatography, typically analytical HPLC, which quantifies the main peak and profiles any impurities.

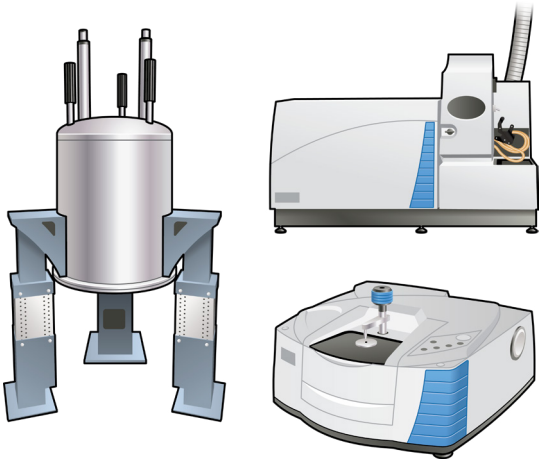


Figure 4.79: Chromatography and spectroscopy instruments

Question 455

Limonene ($C_{10}H_{16}$) is a colourless citrus oil found in orange peel and widely used as a fragrance.

Which laboratory process is not used to isolate limonene from orange peel?

- ○ A Filtration
- ○ B Distillation
- ○ C Separation
- ○ D Chromatography

(1 mark)

Question 456

Caffeine ($C_8H_{10}N_4O_2$) is a stimulant isolated from tea leaves.

Which technique removes aqueous impurities after the tea leaves have been filtered?

- ○ A Evaporation
- ○ B Separation
- ○ C Distillation
- ○ D Chromatography

(1 mark)

Chemists identify structures and functional groups using several approaches. Spectroscopic techniques such as infrared (IR) spectroscopy reveal characteristic absorption bands for functional groups, while nuclear magnetic resonance (NMR) and mass spectrometry provide detailed information about connectivity and molecular mass. These tools confirm the identity of medicinal compounds and ensure their purity during the drug manufacturing process.

By identifying functional groups, chemists can predict how medicines will dissolve, how easily they will cross cell membranes, and how they may be metabolised. This knowledge underpins drug design, enabling chemists to modify structures to enhance activity, reduce side effects, and develop safer, more effective treatments.

Case Study: Aspirin and Paracetamol

Two of the most widely used medicines, **aspirin** and **paracetamol**, illustrate how functional groups define chemical behaviour and therapeutic action.

Aspirin (acetylsalicylic acid) contains both a carboxyl group (–COOH) and an ester group (–COOR) attached to a benzene ring (see **Figure 4.80**). The carboxyl group makes aspirin weakly acidic, enabling it to interact with enzyme active sites through hydrogen bonding. This property underpins its action as a non-steroidal anti-inflammatory drug (NSAID), in which it inhibits cyclooxygenase enzymes that produce prostaglandins, compounds that cause pain and inflammation. The ester group influences aspirin's stability and metabolism, as it can be hydrolysed in the body to release salicylic acid, the true active agent.

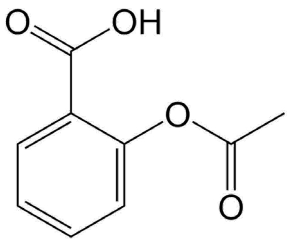


Figure 4.80: Aspirin

Paracetamol (acetaminophen) contains both a hydroxyl group (–OH) and a primary amide group (–$CONH_2$) (see **Figure 4.81**). The hydroxyl group increases solubility in aqueous environments, aiding absorption in the body, while the amide group reduces paracetamol's acidity relative to aspirin, making it gentler on the stomach. Its functional groups enable it to bind to brain enzyme systems that regulate temperature and pain, explaining its analgesic and antipyretic effects.

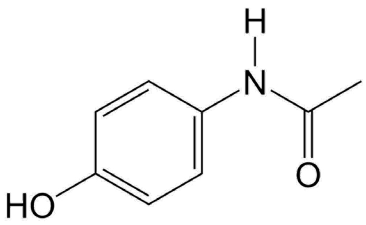
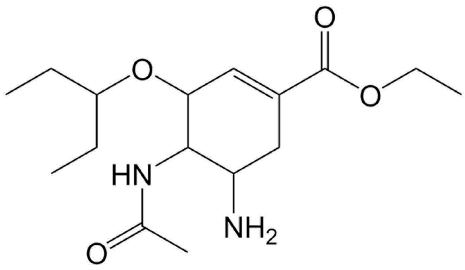


Figure 4.81: Paracetamol

Question 457

Oseltamivir is used to treat influenza virus infections.



Which functional groups are present in oseltamivir?

- ◯ A amine, amide, hydroxyl
- ◯ B alkene, amine, carbonyl
- ◯ C amine, ester, amide
- ◯ D amide, alkene, carbonyl

(1 mark)

SOLUTIONS

SOLUTIONS TO UNIT 3 QUESTIONS

Question	Part		Author's response	Marks
1			D	1
2			C	1
3			D	1
4			B	1
5	(a)		A fuel like petrol stores energy that is released when it undergoes combustion.	1
	(b)	i	Remains of organisms are buried under sediments in anaerobic conditions. Over millions of years, heat and pressure transform this organic matter into mixtures of hydrocarbons.	1 1
		ii	Petroleum forms over geological timescales, far slower than we consume it, so it is non-renewable on a human timescale.	1
	(c)	i	Ethanol is produced from recent biomass (plant carbohydrates) through photosynthesis.	1
		ii	**Any one**	1
			Reduced dependence on imported crude oil supports energy security and reduces vulnerability to fluctuations in oil prices. Provides a market for crops such as sugarcane, corn, or sorghum, benefiting rural economies and farmers. Stimulates employment in farming, biofuel processing, and distribution sectors. Blending can help buffer fuel supply against global oil market volatility, stabilising fuel prices. Alternative domestic fuel sources help safeguard transport fuel availability during supply disruptions.	
		iii	**Any one**	1
			Lower net CO_2 emissions, as carbon released during combustion is partly offset by CO_2 absorbed during crop growth. A higher oxygen content in the fuel leads to more complete combustion and cleaner exhaust. Reduces reliance on finite fossil fuel resources, lessening environmental extraction impacts.	
6	(a)	i	**Any one**	1
			Highly reliable, as coal and natural gas power plants can operate continuously, providing consistent output. Widespread mining, transport, and power plant facilities already exist for coal and natural gas. Coal and natural gas have high energy densities, allowing large-scale generation from relatively small fuel volumes. Both coal and natural gas are inexpensive relative to the cost of the energy they supply.	
		ii	**Any one**	1
			Both release CO_2 when burned, contributing to climate change; natural gas can also leak as methane, a potent greenhouse gas. Both release air pollutants that cause respiratory illnesses, cardiovascular disease, and other health problems in surrounding communities. Both fuels are finite and non-renewable, so they will eventually run out. Both are extracted through mining, which causes habitat destruction, soil erosion, water contamination, and ecosystem disruption. Fuel costs can fluctuate based on global market conditions, which in turn affects electricity prices.	

	(b)	i	Both are produced from biomass, which can be regrown or replenished on a short timescale.	1
		ii	The CO_2 released during the combustion of bioethanol and biodiesel is offset by the CO_2 absorbed by plants during photosynthesis as they grow, making the process more carbon neutral. Coal and natural gas release carbon that was stored underground for millions of years, adding new CO_2 to the atmosphere.	1 1
		iii	**Any two** There is limited availability of sustainable biomass, as large-scale cultivation of feedstocks often competes with food production, leading to land-use changes, deforestation, and biodiversity loss. These crops require high inputs of water, fertilisers, and pesticides, which contribute to soil degradation and runoff pollution. Clearing land for crops may release more CO_2 than burning fossil fuels. Biofuels have lower energy density than coal or natural gas, so more fuel is required to generate the same amount of electricity. Biomass supply is seasonal and insufficient to provide a continuous baseload of power at the national level. Production costs are higher than those of fossil fuels, making the industry dependent on subsidies and vulnerable to market fluctuations. Biomass must be collected, transported, and pre-processed, raising both costs and emissions.	1+1
7	(a)		Methane obtained from biomass is more sustainable because it is part of the short-term carbon cycle; the CO_2 released on combustion is roughly balanced by the CO_2 absorbed when the plants grow.	1
	(b)		Methane from natural gas deposits is less sustainable because it comes from finite fossil reserves and releases carbon that has been stored underground for millions of years, adding new greenhouse gases to the atmosphere.	1
8			D	1
9			B	1
10			A	1
11			D	1
12			B	1
13			A	1
14			C	1
15			D	1
16			C	1
17			C	1
18			B	1
19			C	1
20			D	1
21			A	1
22	(a)		$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$	1
	(b)		$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$	1
	(c)		The cellulose in *M. giganteus* is first hydrolysed to glucose. The glucose is then fermented by yeast under anaerobic conditions to produce ethanol, which is separated by distillation.	1 1

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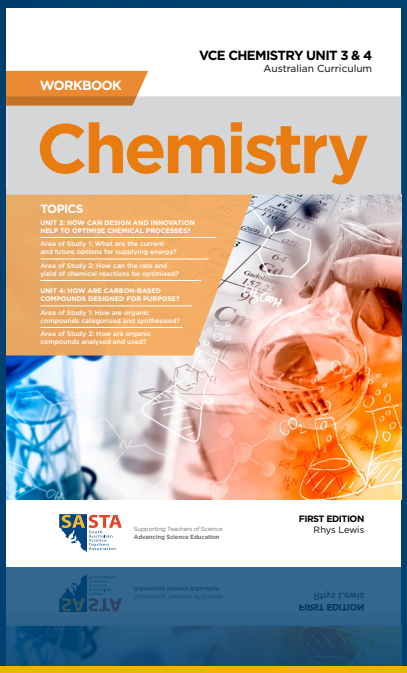

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