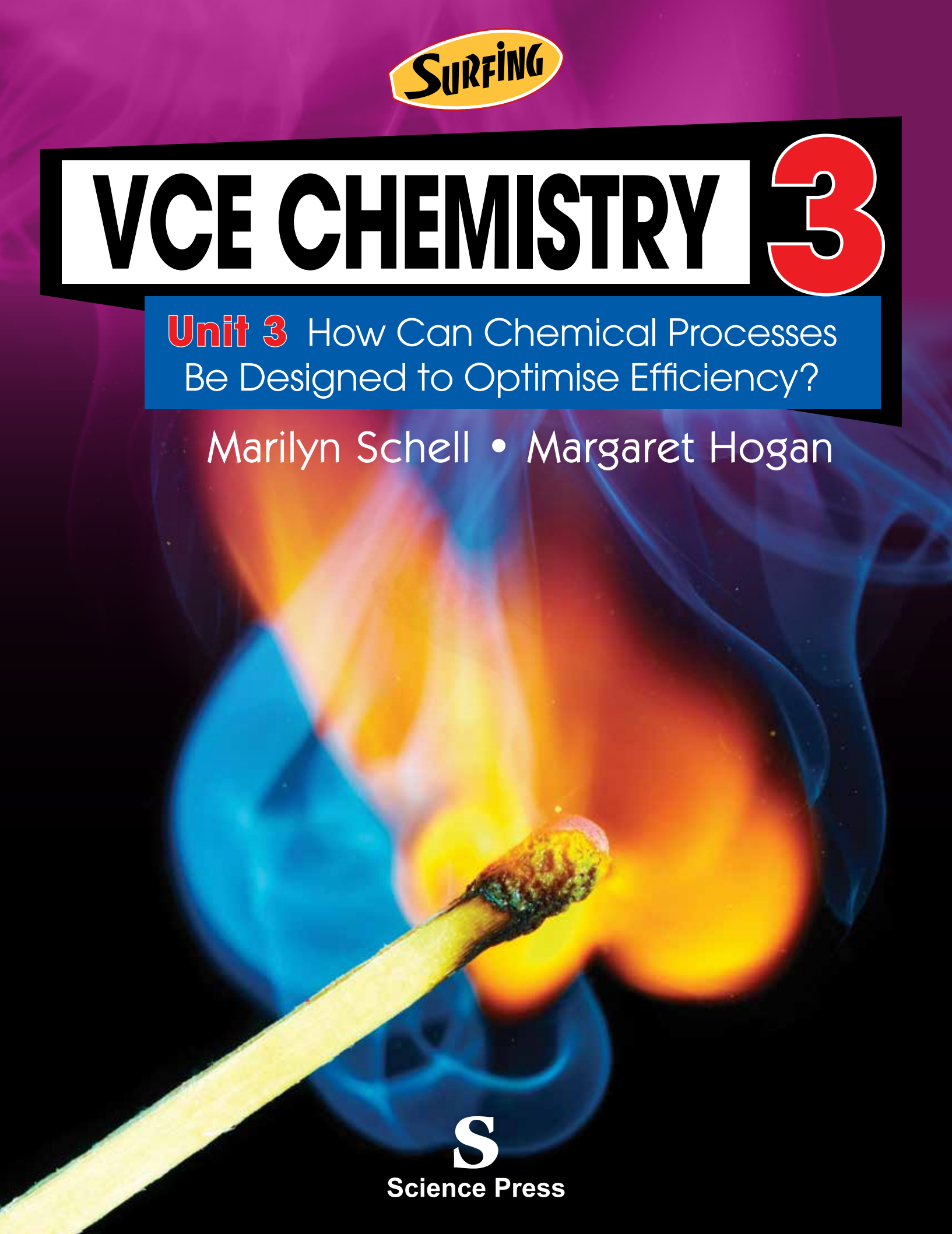
SURFING

VCE CHEMISTRY 3

Unit 3 How Can Chemical Processes Be Designed to Optimise Efficiency?

Marilyn Schell • Margaret Hogan



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Contents

Production of Chemicals by Electrolysis

Rechargeable Batteries

Introduction

This book covers the Chemistry content specified in the Victorian Certificate of Education Chemistry Study Design. Sample data has been included for suggested experiments to give you practice to reinforce practical work in class.

Each book in the *Surfing* series contains a summary, with occasional more detailed sections, of all the mandatory parts of the syllabus, along with questions and answers.

All types of questions – multiple choice, short response, structured response and free response – are provided. Questions are written in exam style so that you will become familiar with the concepts of the topic and answering questions in the required way.

Answers to all questions are included.

A topic test at the end of the book contains an extensive set of summary questions. These cover every aspect of the topic, and are useful for revision and exam practice.

Words To Watch

account, account for State reasons for, report on, give an account of, narrate a series of events or transactions.

analyse Interpret data to reach conclusions.

annotate Add brief notes to a diagram or graph.

apply Put to use in a particular situation.

assess Make a judgement about the value of something.

calculate Find a numerical answer.

clarify Make clear or plain.

classify Arrange into classes, groups or categories.

comment Give a judgement based on a given statement or result of a calculation.

compare Estimate, measure or note how things are similar or different.

construct Represent or develop in graphical form.

contrast Show how things are different or opposite.

create Originate or bring into existence.

deduce Reach a conclusion from given information.

define Give the precise meaning of a word, phrase or physical quantity.

demonstrate Show by example.

derive Manipulate a mathematical relationship(s) to give a new equation or relationship.

describe Give a detailed account.

design Produce a plan, simulation or model.

determine Find the only possible answer.

discuss Talk or write about a topic, taking into account different issues or ideas.

distinguish Give differences between two or more different items.

draw Represent by means of pencil lines.

estimate Find an approximate value for an unknown quantity.

evaluate Assess the implications and limitations.

examine Inquire into.

explain Make something clear or easy to understand.

extract Choose relevant and/or appropriate details.

extrapolate Infer from what is known.

hypothesise Suggest an explanation for a group of facts or phenomena.

identify Recognise and name.

interpret Draw meaning from.

investigate Plan, inquire into and draw conclusions about.

justify Support an argument or conclusion.

label Add labels to a diagram.

list Give a sequence of names or other brief answers.

measure Find a value for a quantity.

outline Give a brief account or summary.

plan Use strategies to develop a series of steps or processes.

predict Give an expected result.

propose Put forward a plan or suggestion for consideration or action.

recall Present remembered ideas, facts or experiences.

relate Tell or report about happenings, events or circumstances.

represent Use words, images or symbols to convey meaning.

select Choose in preference to another or others.

sequence Arrange in order.

show Give the steps in a calculation or derivation.

sketch Make a quick, rough drawing of something.

solve Work out the answer to a problem.

state Give a specific name, value or other brief answer.

suggest Put forward an idea for consideration.

summarise Give a brief statement of the main points.

synthesise Combine various elements to make a whole.

SURFING

VCE CHEMISTRY 3

Area of Study 1

What Are the Options for Energy Production?

1 Fuels

In this area of study you will be examining a range of energy resources and technologies, starting with fossil fuels and biofuels. You will be comparing these in terms of energy transformations that take place in chemical reactions, their efficiency, environmental impacts and applications.

A **fuel** refers to any substance that can react to release energy. The unit of energy is the joule (J). In early times, the main source of energy for warmth and cooking was wood from trees. A few thousand years ago coal began to be used and the development of the steam engine in the 1700s led to the industrial use of coal. The Industrial Revolution saw a huge increase in the consumption of coal as a fuel source and later the use of petroleum as a fuel source.

Fossil fuels

Fossil fuels are energy sources that have been **formed by geological processes**, such as heat and pressure, compressing and changing layers consisting of the remains of once living plants and animals. Fossil fuels include coal, liquid petroleum (crude oil) and natural gas, coal seam gas, shale oil and shale gas.

In Australia, fossil fuels provide 95% of the energy consumed – approximately 40% coal, 34% petroleum and 21% natural gas.

Coal is the main source of energy used in the manufacture of electricity, although some power plants use other fossil fuels such as petroleum (oil and natural gas), or nuclear energy or renewable energy sources such as solar and wind power.

Fossil fuels are **non-renewable** – the fossil fuels we are burning for energy have taken millions of years to form and cannot be replaced by any natural process within a relatively short period of time. Fossil fuel resources are finite – they are being used up and we cannot replace them, so they will not last forever.

Recently, scientists have been developing biofuels synthetically to try to supplement the use of fossil fuels so as to preserve these resources.

Biofuels

A biofuel is any fuel **produced from living**, **or recently living**, **materials s**uch as waste plant and animal matter (biomass). Biofuels are **renewable resources**. These include bioethanol, biogas and biodiesel.

Alkanes as fuels

In the home, the most common constituents of the fuels burned are alkanes. Alkanes are present in fossil fuels such as liquid petroleum (e.g. petrol and LPG), natural gas and also biofuels.

Table 1.1 Alkanes as fuels.

Natural gas 	Mostly methane (75% to 90%), with small amounts of ethane (5% to 10%), propane (2% to 7%) and butane (3% to 6%). It can also contain nitrogen, water vapour, carbon dioxide and sometimes hydrogen sulfide.
LPG 	LPG is liquefied petroleum gas and it contains propane and/or butane. LPG is used as a fuel in heating appliances, cooking and in vehicles. It is also used as an aerosol propellant. LPG is heavier than air, so leaks will flow along the floor and settle in low spots.
Petrol, kerosene and diesel 	Used in industry and for transport. These are mixtures of carbon compounds, mostly alkanes. The composition varies, but petrol for cars is made mostly of alkanes from about C4 to C12, kerosene and aviation fuel from C9 to C18, and diesel from about C14 to C20.
Liquefied butane 	Used for cigarette lighters and camp stoves and as a propellant and refrigerant. It is a gas at room temperature but is easily liquefied under pressure.

You will be learning more about fossil fuels and biofuels in the next few chapters.

QUESTIONS

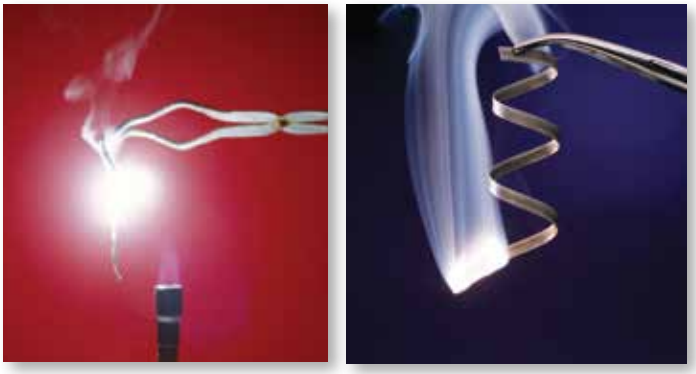
1. A fuel is a substance that can be used to release energy.
   (a) What is the unit of energy?
   (b) Distinguish between fossil fuels and biofuels in terms of how they are formed.
   (c) Distinguish between fossil fuels and biofuels in terms of their renewability.
   (d) Name two fossil fuels and two biofuels.
2. (a) Identify four fuels which are derived from liquid petroleum or natural gas.
   (b) Identify four alkanes which are constituents of fuels.
3. (a) Outline the composition of natural gas.
   (b) Is natural gas a fossil fuel or a biofuel? Justify your answer.

2 Energy and Chemical Reactions

In a chemical reaction, **atoms are never created or destroyed, they are rearranged**. This rearranging of atoms involves **energy changes**.

The law of conservation of energy

The law of conservation of energy states that energy cannot be made or destroyed, but it can change from one type to another. Types of energy include heat, light, sound, electricity, kinetic energy and stored energy. All chemicals contain energy stored in their bonds.



Bunsen starts magnesium burning

Magnesium continues to burn, giving out heat and a brilliant white light

Figure 2.1 Burning magnesium.

In any chemical reaction there is always energy involved, but the law of conservation of energy is obeyed – energy is never created or destroyed, it just changes from one type to another. **Energy is used to break bonds** so that atoms or ions become available to react and be rearranged. **Energy is released when new bonds are formed**. Figure 2.1 shows the energy changes occurring during the combustion of magnesium.

2Mg	+	O=O	⟶	2MgO
↑		↑		↓
Energy breaks metallic bonds		Energy splits covalent bonds		Energy released as ionic bonds form

Exothermic and endothermic reactions

When magnesium burns, energy is needed to break metallic bonds so atoms of magnesium are free to react with oxygen, and energy is also needed to split covalent bonds in oxygen molecules so that oxygen atoms are available to react with magnesium. This energy can come from a Bunsen flame.

When atoms of magnesium and oxygen combine, bonds form between them and energy is released to the environment – heat and light are given out.

This energy can instantly free up more magnesium and oxygen atoms. So once started, the reaction goes on without any more energy input – until you run out of magnesium or oxygen of course.

This is an example of an **exothermic reaction**. The energy in the bonds between magnesium and oxygen is less than the energy in the Mg–Mg bonds and the O=O bonds combined, so energy is released to the environment as heat and light. Combustion (burning) reactions are always exothermic. Other exothermic reactions include the action of dilute acids on active metals such as magnesium, and also neutralisation reactions between an acid and a base such as the reaction of dilute hydrochloric acid on sodium hydroxide.

In an **endothermic reaction**, energy is still put in to break existing bonds, and energy is still released as new bonds form – just like in exothermic reactions. However, in an endothermic reaction, the amount of energy put in is greater than the energy released. If you want the reaction to keep going, you must keep putting in more energy – keep heating the reactants. Photosynthesis is an endothermic reaction, it only proceeds while energy (in the form of sunlight) is available. Electrolysis reactions are also endothermic, for example passing an electric current through a substance such as water or molten sodium chloride to break it down. When the electric current stops, the reaction stops.

Bond energy

Each bond has its own characteristic energy value: the amount of energy needed to break it – and this is the same as the amount of energy released when it forms. This is called its bond energy. Here is an example.

Hydrogen burns to form water.

$2H_2 + O_2 \rightarrow 2H_2O$

We can write this reaction to show the bonds that need to be broken and those that need to be made.

H–H + H–H + O=O → H–O–H + H–O–H

This equation shows that one O=O bond and two H–H bonds must be broken (using energy), and four O–H bonds must be formed (releasing energy). Table 2.1 shows the numerical value of these bonds.

Table 2.1 Bond energies.

Bond type	Energy value per bond (kJ mol^{-1})	Number of bonds	Total value (kJ)
O=O	498	1 broken	498 used
H–H	432	2 broken	864 used
O–H	467	4 formed	1868 released

The final result will be 1868 – (498 + 864) = 506 kJ of energy released. We can write a **thermochemical equation** to summarise the information in Table 2.1.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)\ \Delta H = -506$ kJ

Enthalpy

The term **enthalpy** (H) is used to refer to the total energy involved in a chemical reaction. It is not possible to measure the enthalpy of reactants or products, but we can measure the change in enthalpy during a chemical reaction. The symbol for **change in enthalpy is** ΔH. Change in enthalpy (ΔH) is negative for exothermic reactions and positive for endothermic reactions. Energy is measured in joules (J) or kilojoules (kJ).

Activation energy

Activation energy is the energy needed to start a chemical reaction. This energy breaks the bonds in the chemical reactants. Once the bonds are broken, the atoms can rearrange to form new substances.

The size of the activation energy for a reaction varies. A mixture of hydrogen and oxygen has a low activation energy; it does not need much energy to break the O=O bonds and the H–H bonds. The energy from a match is enough to provide the activation energy for this reaction. However, the reaction of magnesium and oxygen has a higher activation energy. A Bunsen will be needed to produce enough heat to reach the activation energy to get this reaction started.

QUESTIONS

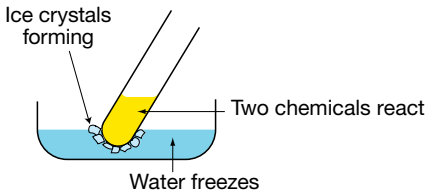
1. Explain the meaning of:
   (a) Endothermic. (b) Exothermic.
   (c) Activation energy. (d) Enthalpy change.
2. Decide if each of the following statements is true or false then justify your answers.
   (a) Energy is used to break chemical bonds.
   (b) Energy is used to make chemical bonds.
   (c) Exothermic reactions release energy.
   (d) In a chemical reaction new atoms are formed.
3. Ethane burns in air to form carbon dioxide and water.
   (a) Write a balanced equation for this reaction.
   (b) Write this equation with structural formulas so all bonds can be clearly seen.
   (c) The following table shows the types of bonds that must be broken in this reaction. Add up how many of each type of bond must be broken and complete the table to show the energy needed to break all these bonds.

Bond type	Energy value per bond (kJ mol^{-1})	Number of bonds broken	Total value (kJ)
O=O	498		
C–H	414		
C–C	346		

   (d) The following table shows the types of bonds that must be formed in this reaction. Add up how many of each type of bond must be formed and complete the table to show the energy released when all these bonds form.

Bond type	Energy value per bond (kJ mol^{-1})	Number of bonds formed	Total value
O=C	804		
O–H	463		

   (e) Calculate the total energy used to break bonds.
   (f) Calculate the total energy released when bonds form.
   (g) Write a thermochemical equation for this reaction.
   (h) Is this an exothermic or endothermic reaction? Justify your answer.
4. Two chemicals were mixed in a test tube. The test tube was standing in a small volume of water as shown below. The water was initially at room temperature, but, as the chemicals inside the test tube reacted, the water outside the test tube froze.



   Is the reaction inside the test tube endothermic or exothermic? Explain.
5. Using your knowledge of chemical reactions, state whether each of the following reactions is exothermic or endothermic.
   (a) Hydrochloric acid and zinc.
   (b) Burning wood.
   (c) Sodium and water.
   (d) Combustion of aluminium.
   (e) Sulfuric acid on magnesium.
   (f) Combustion of hydrogen to form water.
6. Check your knowledge with the following quiz.
   (a) When bonds are broken, energy is .....
   (b) When bonds are formed, energy is .....
   (c) A reaction that releases energy to the environment is said to be (exothermic/endothermic).
   (d) All combustion reactions are (endothermic/exothermic).
   (e) Identify the term for the heat energy needed to start a chemical reaction.
   (f) The total energy change in a chemical reaction is called its ..... change.
   (g) An equation which shows energy changes during a reaction as well as reactants and products is called a ..... equation.

3 Combustion

Combustion (burning) is an exothermic chemical reaction involving oxygen. During combustion:

- Oxygen is used.
- One or more oxides are formed.
- Energy is released.

Combustion of fuels

Fuels are chemicals that undergo combustion to produce usable energy. Combustion of fuels releases the energy stored in their bonds. Most of our fuels are carbon based (organic substances).

Both fossil fuels and biofuels contain **chemical energy** which came originally **from the Sun**. Photosynthesis converts solar energy into chemical energy which is stored within molecules of glucose. The conversion, by living things (mainly plants), of the inorganic carbon in carbon dioxide into complex organic compounds such as glucose is called carbon fixation.

Carbon dioxide + water $\xrightarrow{\text{solar energy}}$ glucose + oxygen

$$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{solar energy}} C_6H_{12}O_6(aq) + 6O_2(g)$$

The product(s) formed during combustion of fuels depend on the fuel burned and also on the amount of oxygen available for its combustion. The amount of oxygen available determines whether the combustion will be complete or incomplete.

Complete combustion

Complete combustion occurs when there is plenty of oxygen available. The reactant combines with as much oxygen as possible and produces carbon dioxide.

Carbon + oxygen → carbon dioxide

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Ethane + oxygen → carbon dioxide + water

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$$

Ethanol + oxygen → carbon dioxide + water

$$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$$

Incomplete combustion

Incomplete combustion occurs when less oxygen is available. The products of incomplete combustion of carbon and its compounds can include carbon (soot) and/or carbon monoxide.

Carbon + oxygen → carbon monoxide

$$2C(s) + O_2(g) \rightarrow 2CO$$

Ethane + oxygen → carbon monoxide + water

$$2C_2H_6(g) + 5O_2(g) \rightarrow 4CO(g) + 6H_2O(l)$$

Ethane + oxygen → carbon + water

$$2C_2H_6(g) + 3O_2(g) \rightarrow 4C(g) + 6H_2O(l)$$

Ethanol + oxygen → carbon monoxide + water

$$C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l)$$

Complete combustion releases more energy than incomplete combustion.

Burning vapour

Notice that it is not the solid or liquid fuel that burns, but its vapour (gas). You can see this if you look carefully at a burning candle.

To burn a solid fuel, such as the large chain hydrocarbons that make up candle wax (e.g. $C_{30}H_{62}$), it must be both melted and vaporised (two changes of state). The wax melts, the molten wax moves up the wick and then this molten wax vaporises. The wax vapour (gas) around the wick is what actually burns.

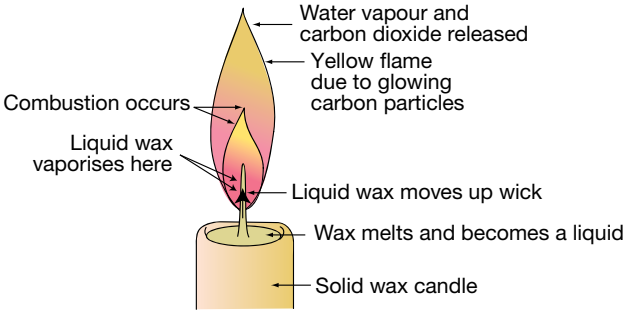


Figure 3.1 Burning a candle.

The flame colour from burning hydrocarbons depends on the completeness of combustion and the temperature reached. Incomplete combustion produces tiny soot particles which glow with a yellow flame. Cooler flames are red, hotter flames are white. Blue flames are produced by complete combustion.

Combustion terminology

The following terms are frequently associated with the use of fuels.

- **Volatile** – a volatile fuel is one that has a low boiling point and thus turns to a vapour easily. Volatile fuels may have low flash points.
- **Flash point** – is the lowest temperature at which a fuel's vapour, when mixed with air or oxygen, can be ignited by a spark or a flame. Diesel fuel has a higher flash point than petrol, so diesel does not burn as readily as petrol.

- **Ignition temperature** – is the temperature to which a fuel must be heated in air or oxygen in order for it to **ignite** in the **absence of a spark or flame**. Spontaneous combustion occurs at this temperature. Diesel engines do not have spark plugs, so diesel fuel is heated to its ignition temperature by compression of the air/fuel mixture.
- The **heat of combustion** refers to the energy released when a substance undergoes complete combustion with oxygen. This is always written as a positive value.
- **Enthalpy of combustion** (ΔH) refers to the change in energy of a system during combustion. Combustion is always exothermic (heat is released), so the enthalpy of combustion is always written as a negative value.
- The **efficiency of any energy transformation** relies on all of the energy used being converted to useful energy. If you burn a fuel to heat water, then you want all of the energy released to actually heat the water. If you burn a fuel in a car, you want all of the energy from the fuel to be converted to mechanical energy to bring about movement of the car. When energy is converted to other forms such as heat and sound, the efficiency is reduced.

QUESTIONS

1. Distinguish between complete and incomplete combustion.
2. Combustion always involves the addition of oxygen. Write balanced chemical equations for the following combustion reactions.
   (a) Complete combustion of butane.
   (b) Incomplete combustion of methane forming carbon monoxide.
   (c) Incomplete combustion of methane to produce carbon.
   (d) Complete combustion of methanol.
   (e) Incomplete combustion of methanol to produce carbon monoxide.
3. Petrol used in cars contains octane (C_8H_{18}). Write equations to show the combustion of octane under the following conditions.
   (a) Excess oxygen available.
   (b) Limited oxygen available.
4. (a) When you use a Bunsen burner to heat chemicals, do you have the hole in the barrel open or closed? Explain.
   (b) If you heat a beaker of water using the safety flame of the Bunsen burner, the beaker becomes black. Explain.

5. Pentane is burned in a limited supply of oxygen and forms water vapour, carbon monoxide and particles of carbon. Write an equation for this reaction.
6. (a) Identify the changes of state involved in the combustion of a burning candle.
   (b) Research the reason why a candle flame has an orange glow.

7. Identify whether each of the following statements is true or false and justify your answer.
   (a) When a candle burns, it is the wick that burns to make light, not the candle wax.
   (b) More oxygen must be available for complete combustion than for incomplete combustion.
   (c) Spontaneous combustion occurs at the flash point.
   (d) Respiration is an exothermic reaction.
   (e) Carbon burns in plenty of air to form a basic oxide.
8. Methane (CH_4) is present in natural gas and is a common fuel. Water is one product of the combustion of methane. Identify three other possible products of the combustion of methane as the supply of oxygen decreases. Use equations to justify your answer.
9. (a) What is meant by the term fossil fuel?
   (b) Name three fossil fuels.
   (c) For one of these fossil fuels, explain the photosynthetic origin of its energy.
   (d) Define carbon fixation.
10. (a) What is meant by a biofuel?
    (b) Name an example of a biofuel.
11. Check your knowledge with this quick quiz.
    (a) Identify the element always used during every combustion reaction.
    (b) Combustion in a plentiful supply of oxygen is called ..... combustion.
    (c) What do we call any substance that can be burnt to produce usable energy?
    (d) Name three fossil fuels.
    (e) Identify the original source of the energy in fossil fuels.

4 Thermochemical Equations

A thermochemical equation includes a balanced equation, showing the stoichiometric proportions in which reactants react and products are formed, and it also states any change in energy, or enthalpy involved in the reaction. You are used to writing balanced (stoichiometric) equations. For a thermochemical equation, you just need to add the ΔH value, obtained experimentally or from a data table.

The value of ΔH (change in enthalpy) can have a **positive or a negative sign**. A positive sign indicates that the system is using energy and is described as endothermic. A negative sign means that overall the system is releasing energy to the surroundings and it is described as exothermic.

Endothermic: $A + B + \text{energy} \rightarrow C \qquad \Delta H > 0$

Exothermic: $A + B \rightarrow C + \text{energy} \qquad \Delta H < 0$

For thermochemical equations, certain rules apply.

- **The reverse reaction** has the opposite sign for its ΔH value, for example:

  $H_2O(s) \rightarrow H_2O(l) \qquad \Delta H = 6\text{ kJ}$

  $H_2O(l) \rightarrow H_2O(s) \qquad \Delta H = -6\text{ kJ}$

- The value of ΔH refers to the **equation as it is written**, for example:

  $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)\ \Delta H = -286\text{ kJ}$

  This can mean that 286 kJ of energy is produced by the combustion of 2 mole of $H_2(g)$, or by using 1 mol $O_2(g)$ or by producing 2 mol water.

- If you use or produce more of a **chemical**, then the **enthalpy change** (ΔH) differs by the same amount. For example, the combustion of 2 mol H_2 gas produces 286 kJ energy, so the combustion of 1 mol hydrogen would produce $\frac{286}{2} = 143$ kJ of energy. And the combustion of 10 mol H_2 would produce $143 \times 10 = 1430$ kJ of energy

Table 4.1 shows **enthalpies of combustion** for some common fuels. This is the energy released by the **complete combustion** of **1 mole** of each of the substances listed, with the final products being **carbon dioxide and water**. Notice that the enthalpy change can be expressed in different units such as kJ /mol, kJ/g and MJ/tonne. In this table the values are expressed as kJ per mole. These values represent the difference between the energy absorbed in breaking bonds and the energy released when new bonds form. They are all negative values as combustion is an exothermic reaction.

Table 4.1 Enthalpy of combustion (measured at a constant pressure of 100 kPa and at a temperature of 25°C).

Fuel	Formula	State	ΔH (kJ mol^{-1})
Hydrogen	H_2	Gas	–286
Carbon	C	Solid	–394
Methane	CH_4	Gas	–890
Ethane	C_2H_6	Gas	–1560
Propane	C_3H_8	Gas	–2220
Butane	C_4H_{10}	Gas	–2874
Octane	C_8H_{18}	Liquid	–5460
Methanol	CH_3OH	Liquid	–726
Ethanol	C_2H_5OH	Liquid	–1368
Propan-1-ol	C_3H_7OH	Liquid	–2021
Butan-1-ol	C_4H_9OH	Liquid	–2671
Glucose	$C_6H_{12}OH$	Solid	–2816

Note: Tables of enthalpy may vary in values. In an examination always use the values in the data sheet provided.

Energy per mole and per gram

It is interesting to compare the energy released by a mole of various fuels with the energy released by a gram of the same fuels. For example, let's compare hydrogen and octane.

Hydrogen

1 mole of hydrogen gas (H_2) has a mass of 2 grams. So, on combustion,

1 mol H_2 (2 grams) releases 286 kJ

1 gram H_2 releases 143 kJ of energy.

Octane

Mass of 1 mole octane $C_8H_{18} = 8 \times 12 + 18 \times 1 = 114$ g

So 114 g octane releases 5460 kJ of energy.

And 1 gram octane will release $\frac{5460}{114} = 47.89$ kJ.

Although the molar heat of combustion of octane is numerically much greater than that of hydrogen, 1 gram of hydrogen will produce much more energy (143 kJ) than 1 gram of octane (48 kJ). When the mass of the fuel is critical, such as in rocket ships, it makes sense to use a fuel such as hydrogen with its large heat release from a very small mass of fuel.

Figure 4.1 Launch of a rocket ship using hydrogen fuel.

QUESTIONS

1. Before you do any mole calculations involving equations, you should practise calculating the formula masses of chemicals. You will recall that the formula mass is the sum of atomic masses of the atoms in the formula. Atomic masses are found in the periodic table. Copy and complete the following table.

Name of chemical	Formula	Working	Formula mass
Hydrogen gas			
Carbon dioxide			
Methane			
Ethane			
Propane			
Butane			
Methanol			
Ethanol			

2. Write thermodynamic equations for the combustion of the following chemicals.
   (a) Solid glucose.
   (b) Gaseous methane.
   (c) Ethane gas.
   (d) Butane gas.
   (e) Liquid ethanol.
3. (a) Of the chemicals listed in Table 4.1, which two would release the most heat energy per mole when they undergo complete combustion?
   (b) Carbon undergoes complete combustion to form carbon dioxide.
   $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H = -394$ kJ
   In a limited supply of oxygen, carbon undergoes incomplete combustion, forming carbon monoxide.
   $2C(s) + O_2(g) \rightarrow 2CO(g) \quad \Delta H = -444$ kJ
   Which of these reactions would release more heat energy per mole of carbon?
   (c) Given the following equation:
   $C(s) + O_2(g) \rightarrow CO_2(g)) \quad \Delta H = -394$ kJ
   State the ΔH value for the reaction:
   $CO_2(g) \rightarrow C(s) + O_2(g)$
   (d) Does a negative ΔH value indicate a reaction which is endothermic or exothermic?
4. Methanol is toxic and can cause blindness. It is present in methylated spirits, making it unsuitable to drink.
   (a) Write the thermochemical equation for the complete combustion of liquid methanol.
   (b) How much energy would be released by the complete combustion of 1 mol of methanol?
   (c) Deduce the energy released by the complete combustion of 2 mol methanol.
   (d) Calculate the molar mass of methanol.
   (e) Calculate the energy that would be released by the complete combustion of 500 g of methanol.
5. Using the values in Table 4.1, calculate the heat released per gram during combustion of the following fuels.
   (a) Methane.
   (b) Ethane.
   (c) Propane.
   (d) Butane.
   (e) Ethanol.
6. Using the information in Table 4.1, and your answers to Question 5, list the fuels methane, ethane, propane and ethanol in decreasing order of:
   (a) Molar heats of combustion.
   (b) Heat of combustion per gram.
7. The SI unit of energy is the joule (J) and enthalpy change values are usually given in kJ/mol, kJ/g or MJ/tonne.
   (a) What is the relationship between a joule and a kilojoule?
   (b) What is the relationship between a joule and a megajoule?
   (c) How many joules in 3.5 MJ?
   (d) A tonne is a megagram. How many grams in a megagram?
   (e) If a fuel releases 500 MJ per tonne of fuel burned, how much energy, expressed in kJ/gram, would be released by 500 grams of this fuel?
   (f) Using values in Table 4.1, how much heat energy would be released by the combustion of 1 tonne of ethanol? Express this value in MJ/tonne.

5 Behaviour of Gases

We need to look at gases now because many fuels are gases, as are the products of their combustion. Of course they are not the only gases, in fact we are surrounded by gases – the atmosphere is a mixture of gases, and gases such as oxygen and carbon dioxide are essential for life on Earth. Let's start by recalling the properties of gases based on your earlier study of the particle theory of matter.

Gases have **no definite shape**. If you introduce a gas into a container it will spread out to fill the whole container, taking the shape of the container, whatever that may be.

Gases **diffuse** – they spread through other gases. If you release some smelly gas at one end of a room, you will soon smell it at the other end of the room and even outside the room. This also works for liquids that readily vaporise such as perfume.

Gases, like liquids, **can be poured** from one container to another. You can see this when you pour a dense, coloured gas such as nitrogen dioxide.

Figure 5.1 Nitrogen dioxide gas being poured.

Gases **can be compressed** – you can push them into a smaller space, e.g. in bicycle tyres.

Figure 5.2 Air being compressed into a bicycle tyre.

Gases exert pressure. The pressure of the atmosphere can hold up a column of mercury 76 cm high at sea level. We call this atmospheric pressure. If you are at the top of a tall mountain the pressure will be a little lower. On top of Mount Kosciuszko the air pressure can be low enough to boil water at about 70°C.

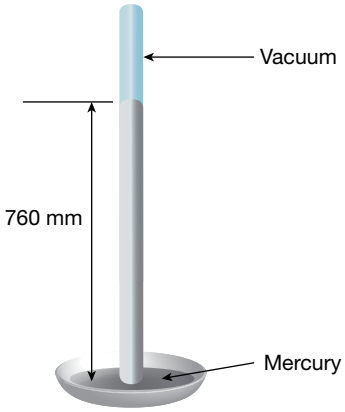


Figure 5.3 The air exerts pressure.

Pressure in a gas acts in all directions – including upwards. It can even hold a piece of paper against an upside down glass of water.

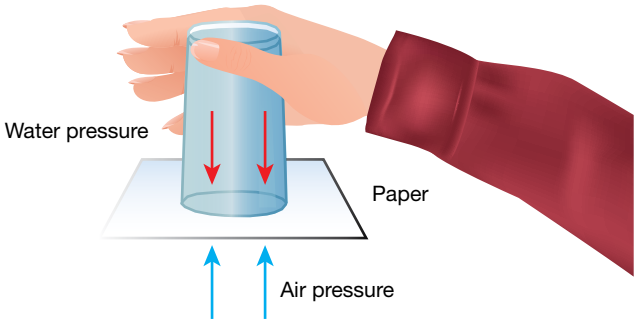
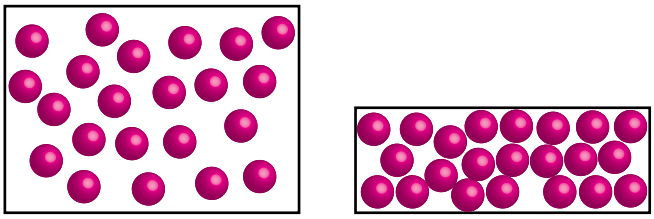


Figure 5.4 Air pressure acts in all directions.

If you push a volume of gas into a smaller space, the particles of gas will be closer together and there will be more collisions with each other and with the walls of the container. The pressure of the gas will be greater. If you increase the volume a gas occupies, the pressure will be less.



In a large volume pressure is less

In a smaller volume pressure is greater

Figure 5.5 Same amount of gas, different volume.

Gases often have **low density** and many are less dense than the mixture of gases that make up air. A gas takes up less space when it is cooled, becoming more dense. It takes up more space when it is heated, and becomes less dense.

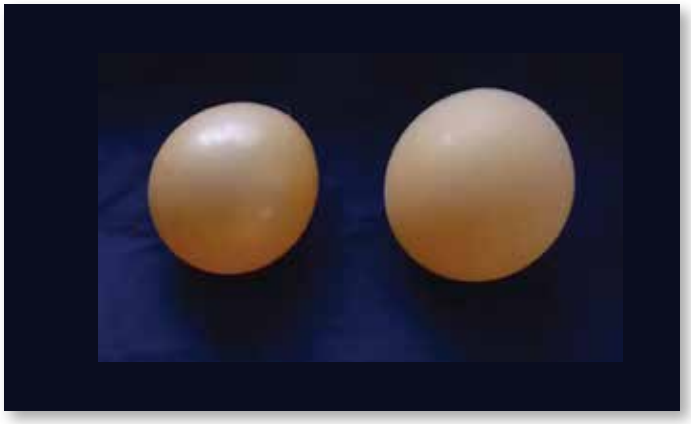
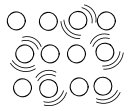
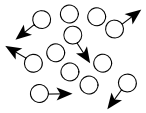
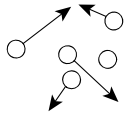


Figure 5.6 These were identical balloons. The one on the left was placed in a freezer for a few minutes.

The behaviour of gases can be explained by using the **particle theory** which states that all matter is made of particles which are constantly moving around at random and interacting with one another. There is a lot of empty space between particles of matter.

If necessary, you should revise the particle theory. You should be able to use this theory to account for the different behaviour of the three states of matter – solids, liquids and gases – in terms of the behaviour of their particles.

Table 5.1 Particle theory of matter.

Solid	Liquid	Gas
Particles are close together and vibrating in fixed positions	Particles are close together and moving more freely	Particles are far apart and moving very freely
		

In the particle theory, gases are seen as being made of tiny particles which are far apart and constantly moving. Relatively small forces hold these particles together so that they are free to move throughout the available space and spread through other gases. As with a liquid, any pressure exerted will be transmitted through a gas in all directions.

QUESTIONS

1. (a) Use a diagram to represent the particles making up a container of a gas and describe the movement of these particles.
   (b) How would the particles change if you heated the container?
   (c) If the container were a balloon, what would you observe on heating it?
2. A hot air balloon such as the one pictured rises when the air inside it is heated.

Account for this observations in terms of the behaviour of particles of gas.

3. For each of the statements decide if it is true or false and justify your decision.
   (a) When a gas is cooled its particles become smaller.
   (b) When a bicycle tyre is pumped up some air is pushed out to make room for the air being pumped in.
4. The gases that make up the Earth's atmosphere above and around you are exerting a pressure equivalent to about 1 kilogram weight pushing on every square centimetre of your body. Why do we not get squashed?
5. Copy and complete the table. For each of the properties of gases listed in the table, write an explanation in terms of the particle theory.

Properties of gases	Explanation
Gases do not have a definite shape, it depends on the container shape.	
Gases can diffuse (spread through other gases).	
Gases can be compressed.	
Gases exert pressure in all directions.	
Gases often have low density.	
Heat increases the temperature and pressure of a gas.	
Heat increases the volume of a gas if it has room to expand.	

Answers

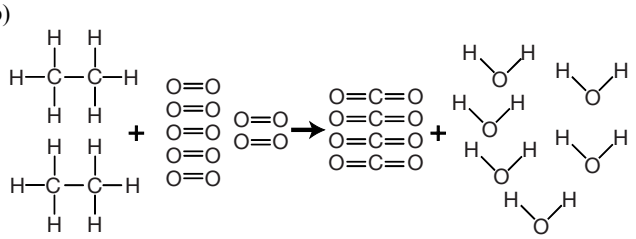
1 Fuels

1. (a) The joule (J).
   (b) Fossil fuels are formed by geological processes such as heat and pressure, compressing and changing layers of the remains of once living plants and animals. Biofuels are produced using living, or recently living materials such as waste plant and animal matter. Biofuels are produced over a short time span but fossil fuels can only be produced over millions of years.
   (c) Fossil fuels cannot be renewed, biofuels can be renewed.
   (d) Various, e.g. Fossil fuels – coal, petroleum; biofuels – bioethanol, biodiesel.
2. (a) Various, e.g. methane, LPG (liquefied petroleum gas), butane, petrol, kerosene or diesel.
   (b) Various, e.g. methane, ethane, propane, butane.
3. (a) Natural gas is mostly (75% to 90%) methane, with small amounts of ethane, propane, and butane. It can also contain nitrogen, water vapour, carbon dioxide and sometimes hydrogen sulfide.
   (b) Natural gas is a fossil fuel. Like liquid petroleum, it is formed by the action of geological processes on the remains of once living plants and animals and it is non-renewable.

2 Energy and Chemical Reactions

1. (a) Endothermic – a reaction in which heat is absorbed from the surroundings.
   (b) Exothermic – a reaction in which heat is released to the surroundings.
   (c) Activation energy is the energy needed to start a chemical reaction.
   (d) Enthalpy change (ΔH) is the energy change during a chemical reaction.
2. (a) True – To break chemical bonds energy is needed and existing bonds need to be broken so that atoms can be free to rearrange and form new substances.
   (b) False – Making bonds releases energy. Bonds are only made if the product contains less energy than the reacting particles, so making bonds always releases energy. No energy is used when bonds are made, instead energy is given out.
   Note: Many students find this idea hard to accept. Probably because when we break things physically we use energy – but we also use energy when we make things. In chemical reactions it is different. You still need energy to split any bonds. But when you make bonds – you are not picking up the atoms and linking them together – that would need energy on your part. Atoms/ions uniting to form a bond when joining to make a new substance will release energy – the product contains less energy than the atoms/ions.
   (c) True – In exothermic reactions more energy is released, from making the new bonds, than is used to break the old bonds. So overall, energy is given out by the reaction.
   (d) False – Atoms are rearranged (not made) in chemical reactions. No new atoms are made in chemical reactions. (Only nuclear reactions can make new atoms, e.g. in a nuclear reactor or atomic bomb.)
3. (a) $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ or

   $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$
   (b)



   (c)

Bond type	Energy value per bond (kJ mol^{-1})	Number of bonds broken	Total value (kJ)
O = O	498	7	3486
C–H	414	12	4968
C–C	346	2	692
			9146

   (d)

Bond type	Energy value per bond (kJ mol^{-1})	Number of bonds formed	Total value (kJ)
O = C	804	8	6432
O–H	463	12	5556
			11 988

   (e) 9146 kJ
   (f) 11 988 kJ
   (g) $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ $\Delta H = -2842$ kJ
   (h) Exothermic reaction. Energy is released by the reaction – the energy released from making bonds (11 988 kJ) is greater than the energy used by breaking bonds (9146 kJ), so ΔH is negative and energy will be released into the environment.
4. Endothermic – heat is taken from its surroundings. For the reaction to proceed, so much heat was taken from the surrounding water that the water froze.
5. (a) Exothermic.
   (b) Exothermic.
   (c) Exothermic.
   (d) Exothermic.
   (e) Exothermic.
   (f) Exothermic.
6. (a) Used.
   (b) Released.
   (c) Exothermic.
   (d) Exothermic.
   (e) Activation energy.
   (f) Enthalpy (ΔH).
   (g) Thermochemical.

3 Combustion

1. Combustion is an exothermic chemical reaction in which oxygen is used and one or more oxides formed. For complete combustion there is a plentiful supply of oxygen. During incomplete combustion less oxygen is used than during complete combustion. The products of combustion also differ, e.g. complete combustion of carbon produces carbon dioxide, incomplete combustion of carbon produces carbon monoxide.
2. (a) $C_4H_{10}(l) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$
   (b) $CH_4(g) + \frac{3}{2}O_2(g) \rightarrow CO(g) + 2H_2O(l)$
   (c) $CH_4(g) + O_2(g) \rightarrow C(g) + 2H_2O(l)$
   (d) $CH_3OH(l) + \frac{3}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
   (e) $CH_3OH(l) + O_2(g) \rightarrow CO(g) + 2H_2O(l)$
3. Various, e.g.
   (a) Octane + oxygen → carbon dioxide + water

   $C_8H_{18}(l) + \frac{25}{2}O_2(g) \rightarrow 8CO_2(g) + 9H_2O(g)$
   (b) Octane + oxygen → carbon monoxide + water

   $C_8H_{18}(l) + \frac{17}{2}O_2(g) \rightarrow 8CO(g) + 9H_2O(g)$

   Note: Other equations are possible here; incomplete combustion of hydrocarbons can produce solid particles of carbon as well as carbon monoxide.

4. (a) Open – this allows in more air (containing oxygen) so complete combustion of the gas can occur and maximum heat can be produced.
Methane + oxygen → carbon dioxide + water
$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
(b) To produce the safety flame, the hole is closed. This means that less air/oxygen enters the Bunsen so incomplete combustion of the gas (mainly methane) occurs. Carbon is produced by incomplete combustion of the methane and black particles of soot (carbon) are deposited on the outside of the beaker.
5. Pentane + oxygen → carbon monoxide + carbon + water
$C_5H_{12}(g) + 4O_2(g) \rightarrow 2CO(g) + 3C(s) + 6H_2O(g)$
6. (a) Candle wax melts (changes from solid to liquid) and then vaporises (changes from liquid to gas).
(b) The orange colour is from the glow of hot carbon particles, indicating that incomplete combustion is occurring.
7. (a) False. The candle wax melts then vaporises, and it is this vapour that burns to produce light.
(b) True. More oxygen is needed for complete combustion.
(c) False. Spontaneous combustion occurs at the ignition temperature. Spontaneous combustion occurs without any flame and ignition temperature is the point at which combustion occurs without a flame.
(d) True. Respiration is the combustion of food in living cells. It is exothermic because energy is released to the surrounding cell.
(e) False. Carbon is a non-metal. It does burn in plenty of air (or oxygen) to form carbon dioxide, but this is an acidic oxide. We know carbon dioxide is acidic because it dissolves in water to form the acid carbonic acid and it will turn blue litmus red.
8. In plentiful oxygen – complete combustion – carbon dioxide forms.
Methane + oxygen → carbon dioxide + water
$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
In less oxygen – incomplete combustion – carbon monoxide forms.
Methane + oxygen → carbon monoxide + water
$CH_4(g) + \frac{3}{2}O_2(g) \rightarrow CO(g) + 2H_2O(g)$
In still less oxygen – incomplete combustion – carbon forms.
Methane + oxygen → carbon + water
$CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(g)$
9. (a) A fossil fuel is a substance which is burned to provide energy and which has been produced from once-living organisms being buried by sediment and compressed for millions of years.
(b) Coal, natural gas, petroleum.
(c) Various. Coal – plants are buried by sediments and compressed for millions of years. Originally, when the plants were alive they obtained energy from the Sun by using solar energy for the process of photosynthesis in (the chloroplasts of) their cells.
Carbon dioxide + water $\xrightarrow{\text{solar energy}}$ glucose + oxygen
$6CO_2(g) + 6H_2O(l) \xrightarrow{\text{solar energy}} C_6H_{12}O_6(aq) + 6O_2(g)$
The solar energy absorbed was converted into the chemical energy in glucose which became part of the structure of the plant. Compression for millions of years converted the buried plant material to peat, then brown coal and black coal.
(d) Carbon fixation is the conversion, by living things such as plants, of inorganic carbon (e.g. in carbon dioxide) into organic compounds, e.g. glucose.
10. (a) A biofuel is a fuel produced from living organisms such as plants and algae.
(b) Various, e.g. bioethanol, biodiesel.
11. (a) Oxygen.
(b) Complete.
(c) A fuel.
(d) Coal, petroleum, natural gas.
(e) The Sun.

4 Thermochemical Equations

1.

Name of chemical	Formula	Working	Formula mass
Hydrogen gas	H_2	2 × 1	2
Carbon dioxide	CO_2	12 + (2 × 16)	44
Methane	CH_4	12 + (1 × 4)	16
Ethane	C_2H_6	(12 × 2) + (1 × 6)	30
Propane	C_3H_8	(12 × 3) + (1 × 8)	44
Butane	C_4H_{10}	(12 × 4) + (1 × 10)	58
Methanol	CH_3OH	12 + 16 + (1 × 4)	32
Ethanol	C_2H_5OH	(12 × 2) + 16 + (1 × 6)	46

2. (a) $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)\ \Delta H = -2816\ \text{kJ mol}^{-1}$
(b) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)\ \Delta H = -890\ \text{kJ mol}^{-1}$
(c) $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)\ \Delta H = -1560\ \text{kJ mol}^{-1}$
Or
$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)\ \Delta H = -1560\ \text{kJ mol}^{-1}$
(d) $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)\ \Delta H = -2874\ \text{kJ mol}^{-1}$
(e) $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)\ \Delta H = -1368\ \text{kJ mol}^{-1}$
3. (a) Octane and butane.
(b) The complete combustion reaction releases more heat energy per mole of carbon than does incomplete combustion, 394 kJ per mole compared to 222 kJ per mol.
(c) +394 kJ
(d) Exothermic.
4. (a) $CH_3OH(l) + \frac{3}{2}O_2(g) \rightarrow CO_2(g) + 2H_2O(g)\ \Delta H = -726\ \text{kJ mol}^{-1}$
Or
$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(g)\ \Delta H = -726\ \text{kJ mol}^{-1}$
(b) 726 kJ
(c) 2 × 726 = 1452 kJ
(d) Molar mass of CH_3OH = 12 + (4 × 1) + 16 = 32
(e) 32 g releases 726 kJ
1 g releases $\frac{726}{32}$ kJ
500 g releases $\frac{726}{32} \times 500 = 11\ 343.4$ kJ
Or to 3 significant figures this is 11 300 kJ.
5. (a) Using the periodic table, molar mass of CH_4 = 16 g
Using Table 4.1, 16 g produces 890 kJ
1 g produces $\frac{890}{16} = 55.6$ kJ
Heat of combustion per gram is 55.6 kJ
(b) Using the periodic table, molar mass of ethane, C_2H_6 = 30 g
Using Table 4.1, 30 g produces 1560 kJ
1 g produces $\frac{1560}{30} = 52$ kJ
Heat of combustion per gram = 52 kJ
(c) Using the periodic table, molar mass of propane, C_3H_8 = 44 g
Using Table 4.1, 44 g produces 2220 kJ
1 g produces $\frac{2220}{44} = 50.5$ kJ
Heat of combustion per gram is 50.5 kJ
(d) Using the periodic table, molar mass of butane, C_4H_{10} = 58 g
Using Table 4.1, 58 g produces 2874 kJ
1 g produces $\frac{2874}{58} = 49.6$ kJ
Heat of combustion per gram = 49.6 kJ
(e) Using the periodic table, molar mass of ethanol, C_2H_5OH = 46 g
Using Table 4.1, 46 g produces 1368 kJ
1 g produces $\frac{1368}{46} = 29.7$ kJ
Heat of combustion per gram = 29.7 kJ
6. (a) Molar heats of combustion: Butane (2874), propane (2220), ethane (1560), ethanol (1368) and methane (890).
(b) Heats of combustion per gram: Methane(55.6), ethane (52), propane (50.5), butane (49.6) and ethanol (29.7).