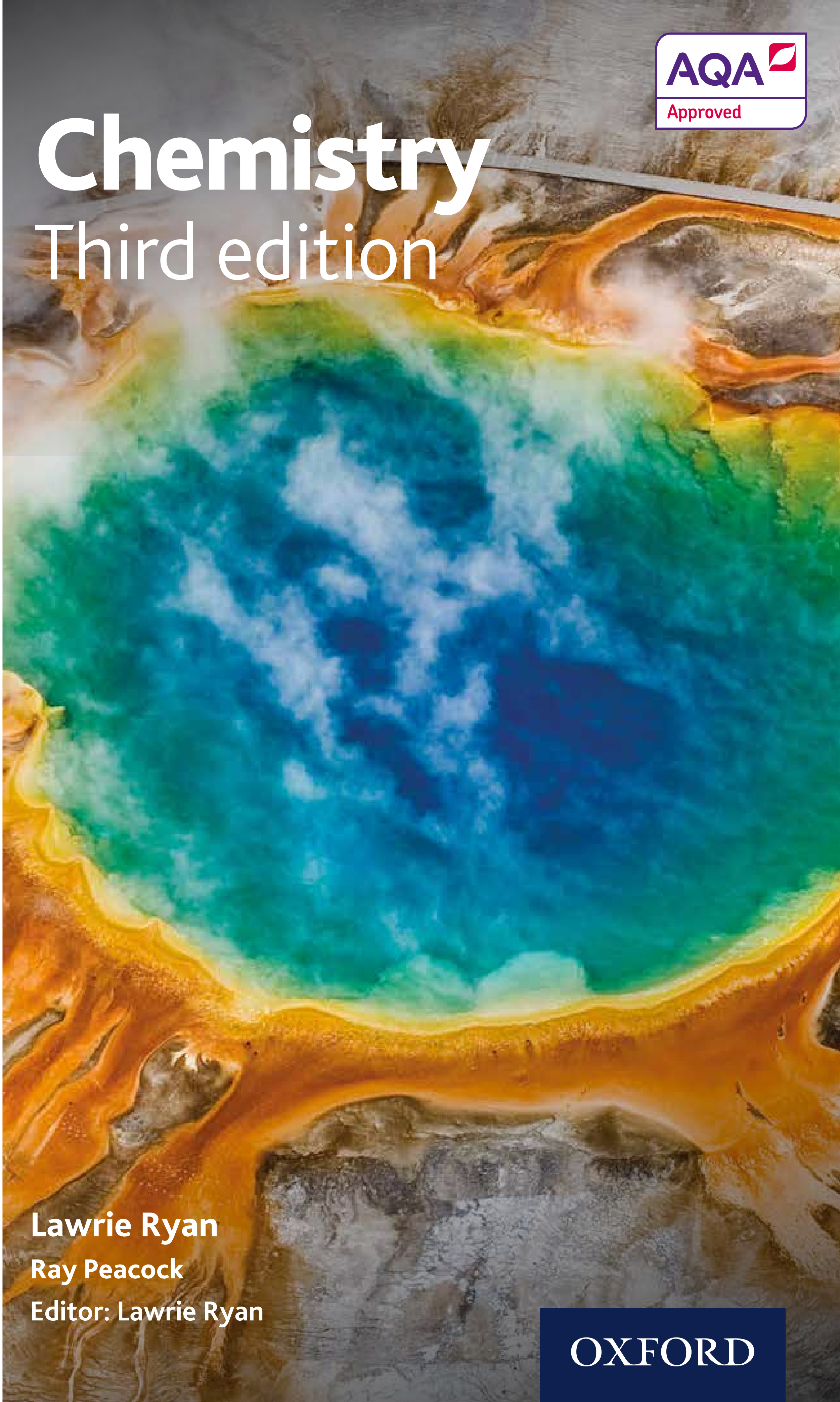
AQA Approved

AQA GCSE Chemistry

Third edition



Lawrie Ryan
Ray Peacock
Editor: Lawrie Ryan

OXFORD

2 Chemical reactions and energy changes

In the early 19th century, people began experimenting with chemical reactions in a systematic way, organising their results logically. Gradually they began to predict exactly what new substances would be formed and used this knowledge to develop a wide range of different materials and processes. They could extract important resources from the Earth, for example, by using electricity to decompose ionic substances. This is how reactive metals such as aluminium and sodium were discovered.

Energy changes are also an important part of chemical reactions. Transfers of energy take place due to the breaking and formation of bonds. The heating or cooling effects of reactions are used in a range of everyday applications.

Key questions

- How can we extract metals from their ores?
- How can we make and prepare pure, dry samples of salts?
- How can we decompose ionic compounds to get useful products?
- Why do chemical reactions always involve transfers of energy?

Making connections

- Reaction profile diagrams will be used in **C8 Rates and equilibrium** to explain the effect of catalysts on the rate of a chemical reaction.
- The calculations of energy changes using bond energy values relies on you drawing the 2D structures of the molecules involved in the reaction, which was covered in **C3 Structure and bonding**.
- Displacement reactions and the use of electrolysis will be applied in **C14 The Earth's resources**.

I already know...	I will learn...
how to define acids and alkalis in terms of neutralisation reactions.	how to represent neutralisation using an ionic equation.
how to use the pH scale for measuring acidity and alkalinity.	how to calculate the concentration of hydrogen ions in a solution given its whole number pH value.
about displacement reactions and the reactions of acids with metals to produce a salt plus hydrogen.	to interpret displacement reactions and the reaction between an acid and a metal in terms of reduction and oxidation.
the reactions of acids with alkalis to produce a salt plus water.	to calculate the concentration of an unknown acid or alkali from experimental results.
combustion and rusting are examples of oxidation reactions.	to identify and describe oxidation and reduction reactions in terms of electron transfer.
that chemical reactions are exothermic and endothermic.	to use bond energy values to calculate the approximate energy change accompanying a reaction.

Required Practicals

Practical		Topic
1	Making a copper salt	C5.5
	Making a salt from a metal carbonate	C5.6
3	Investigating the electrolysis of a solution	C6.4
4	Investigating temperature changes	C7.1

C 7 Energy changes

7.1 Exothermic and endothermic reactions

Learning objectives

After this topic, you should know:

- energy cannot be created or destroyed in a chemical reaction
- that energy is transferred to or from the surroundings in chemical reactions, and some examples of these exothermic and endothermic reactions
- how to distinguish between exothermic and endothermic reactions on the basis of the temperature change
- how to carry out an investigation into energy changes in chemical reactions.

Synoptic link

Remind yourself about neutralisation reactions in Topic C5.7.

Figure 1 *When a fuel burns in oxygen, energy is transferred to the surroundings. You usually don't need a thermometer to know that there is a temperature change!*

Have you ever warmed your hands near a fire? If so you will have felt the effects of energy being transferred during a chemical reaction. In fact, whenever chemical reactions take place, energy is always transferred, as chemical bonds in the reactants are broken and new ones are made in the products. However, the total energy remains the same before and after a reaction. So energy cannot be created or destroyed in any chemical reaction.

Many reactions transfer energy from the reacting chemicals to their surroundings. These are called **exothermic** reactions. The energy transferred from the reacting chemicals often heats up the surroundings. As a result of this, you can measure a rise in temperature as the reaction happens.

Other reactions transfer energy from the surroundings to the reacting chemicals. These are called **endothermic** reactions. As they take in energy from their surroundings, these reactions cause a fall in temperature as they happen.

Exothermic reactions

The burning of fuels, such as the combustion of methane gas, is an obvious example of exothermic reactions. When methane (the main gas present in natural gas) burns, it gets oxidised and releases energy to its surroundings.

Neutralisation reactions between acids and alkalis are also exothermic. You can easily measure the rise in temperature using simple apparatus (see the practical Investigating temperature changes).

The products of exothermic reactions have a lower energy content than the reactants. The actual differences in energy are usually expressed in kilojoules per mole (kJ/mol).

For example, for the reaction:

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

the energy transferred to the surroundings = 890 kJ/mol.

Figure 2 *All warm-blooded animals rely on exothermic reactions in respiration to keep their body temperatures steady*

Endothermic reactions

Endothermic reactions are much less common than exothermic ones. The reaction between citric acid and sodium hydrogencarbonate is a good example to try in the lab, as it is easy to measure the fall in temperature.

Thermal decomposition reactions are also endothermic. An example is the decomposition of calcium carbonate. When heated, it forms calcium oxide and carbon dioxide. This reaction only takes place if you keep heating the calcium carbonate strongly. The calcium carbonate needs to absorb energy from the surroundings (such as the energy provided by a roaring Bunsen flame) to be broken down.

In endothermic reactions the products have a higher energy content than the reactants, so energy is transferred from the surroundings.

For example, for the reaction:

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

the energy transferred from the surroundings = 178 kJ/mol.

Figure 3 *When you eat sherbet you can feel an endothermic reaction. Sherbet dissolving in the water in your mouth takes in energy. It provides a slight cooling effect*

Investigating temperature changes

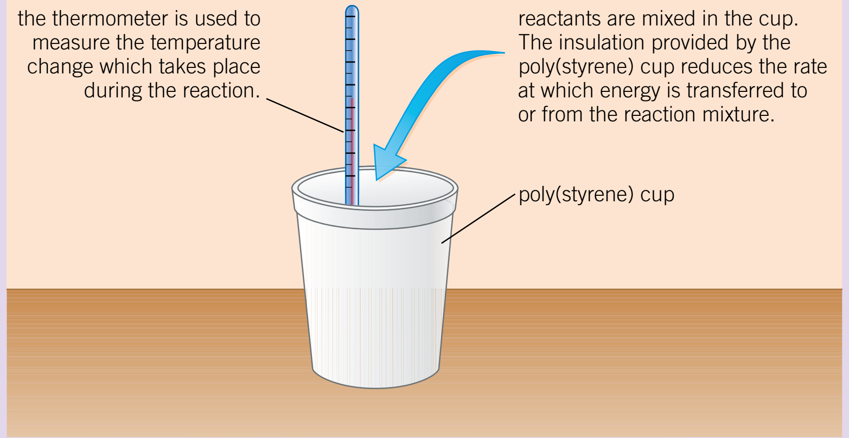
You can use very simple apparatus to investigate the energy changes in reactions involving at least one solution. Often you do not need to use anything more complicated than a poly(styrene) cup and a thermometer.

For example, you can investigate the temperature changes when you mix different combinations of reactants.

- Record the initial temperatures of any solutions and the maximum or minimum temperature reached in the course of the reaction.

To extend yourself, you could consider how the quantities of reactants used might affect the temperature change.

- Make a quantitative prediction and explain your reasoning. This hypothesis could include a sketch graph.
- Test your hypothesis, recording and displaying your data using a suitable table and graph.
- Draw your conclusions.



- Evaluate your investigation, including at least two ways in which you could make the data you collect more accurate.

Safety: Wear eye protection.

Key points

- Energy is conserved in chemical reactions. It is neither created nor destroyed.
- A reaction in which energy is transferred from the reacting substances to their surroundings is called an exothermic reaction.
- A reaction in which energy is transferred to the reacting substances from their surroundings is called an endothermic reaction.

1 **a** What do you call a reaction that transfers energy to its surroundings? [1 mark]
  b What do you call a reaction that takes in energy transferred from its surroundings? [1 mark]
  c Give *two* examples of:
    i an exothermic reaction [2 marks]
    ii an endothermic reaction. [2 marks]

2 Potassium nitrate dissolving in water is an endothermic process. Explain what you would feel if you held a beaker of water in your hand as you stirred in potassium nitrate. [2 marks]

3 Two solutions are added together and the temperature changes from 19 °C to 27 °C. Explain what you can deduce about the energy transferred between the reaction mixture and its surroundings. [2 marks]

4 The energy required for the thermal decomposition of 16.8 g of magnesium carbonate is 23.4 kJ.
  a Write a balanced symbol equation, including state symbols, for the reaction. [2 marks]
  H **b** Calculate the number of moles of magnesium carbonate broken down. [2 marks]

C7.2 Using energy transfers from reactions

Learning objectives

After this topic, you should know:

- how you can make use of the energy from exothermic reactions
- how you can use the cooling effect of endothermic reactions
- how to evaluate uses and applications of exothermic and endothermic reactions given appropriate information.

Warming up

Chemical hand and body warmers can be very useful. These products use exothermic reactions to warm you up. People can take hand warmers to places they know will get very cold. For example, spectators at outdoor sporting events in winter can warm their hands. People usually use the chemical body warmers to help ease aches and pains.

Some hand warmers can only be used once. An example of this type makes use of the energy transferred to the surroundings in the oxidation of iron. Iron turns into hydrated iron(III) oxide in an exothermic reaction. The reaction is similar to rusting. Sodium chloride (common salt) is used as a catalyst. This type of hand warmer is disposable. It can be used only once but it lasts for hours.

Other hand warmers can be reused many times. These are based on the formation of crystals from solutions of a salt. The salt used is often sodium ethanoate, $CH_3COO^-Na^+$. A supersaturated solution is prepared by dissolving as much of the salt as possible in hot water. The solution is then allowed to cool.

A small metal disc in the plastic pack is used to start the exothermic change (see Figure 1). When you press this a few times, small particles of metal are scraped off. These 'seed' (or start off) the crystallisation. The crystals spread throughout the solution, transferring energy to the surroundings in an exothermic change. They work for about 30 minutes.

Figure 1 *During the recrystallisation of sodium ethanoate in a hand warmer, energy is transferred from the reaction system to the surroundings, increasing the temperature of the surroundings – your hand*

To reuse the warmer, you simply put the solid pack into boiling water to re-dissolve the crystals. Once it has cooled down, the pack is ready to activate again.

Exothermic reactions are also used in self-heating cans (Figure 2) that make drinks like hot coffee without any external heating device (e.g., a kettle). The reaction used to transfer energy to the food or drink is usually:

calcium oxide + water → calcium hydroxide

You press a button in the base of the can. This breaks a seal and lets the water and calcium oxide mix. Then the exothermic reaction can begin.

Development of the self-heating coffee can took years and cost millions of pounds. Even then, over a third of the can was taken up with the reactants needed to transfer enough energy to the coffee. Also, in some early versions, the temperature of the coffee did not rise high enough in cold conditions.

Cooling down

Endothermic processes can be used to cool things down. For example, chemical cold packs usually contain ammonium nitrate and water. When ammonium nitrate dissolves, it absorbs energy from its surroundings, making them colder.

These cold packs are used as emergency treatment for sports injuries. The decrease in temperature reduces swelling and numbs pain.

The ammonium nitrate and water (sometimes present in a gel) are kept separate in the pack. When squeezed or struck, the bag inside the water pack breaks, releasing ammonium nitrate. The instant cold packs work for about 20 minutes.

Instant cold packs can only be used once, but are ideal where there is no ice available to treat a knock or strain. This type of cold pack is often included in the first aid kit at venues used for amateur sports or outdoor pursuits.

The same endothermic change can also be used to chill cans of drinks.

1 **a** Give *two* uses of endothermic changes. [2 marks]
b Which endothermic change is often used in cold packs? [1 mark]
c The solid used in cold packs is often ammonium nitrate.
i Give the formula of ammonium nitrate. [1 mark]
ii State another use of ammonium nitrate. [1 mark]

2 **a** Which solid is usually used in the base of self-heating coffee cans? [1 mark]
b Write a balanced symbol equation, including state symbols, for the reaction of water with the solid in part **a**. [3 marks]
c Why is it essential that the coffee stays out of contact with the solid in part **a**? [1 mark]

3 **a** Describe the chemical reaction which takes place in a disposable hand warmer. [4 marks]
b Describe how a reusable hand warmer works. [4 marks]
c Give an advantage and a disadvantage of each type of hand warmer. [2 marks]
d Name *one* use of an exothermic reaction in the food industry. [1 mark]

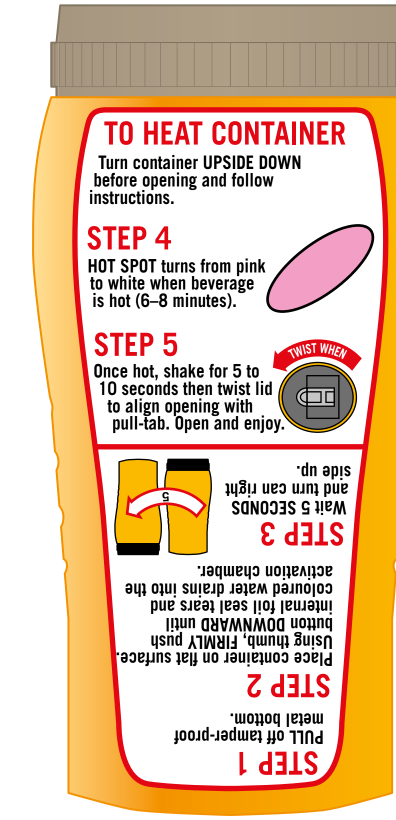


Figure 2 *Development of the self-heating can in the USA took about 10 years*

Figure 3 *Instant cold packs can be applied as soon as an injury occurs to minimise damage*

Key points

- Exothermic changes can be used in hand warmers and self-heating cans. Crystallisation of a supersaturated solution is used in reusable warmers. However, disposable, one-off warmers heat up the surroundings for longer.
- Endothermic changes can be used in instant cold packs for sports injuries.

C7 Energy changes

Summary questions

1 Two solutions are mixed and react in an endothermic reaction. When the reaction has finished, the reaction mixture is allowed to stand until it has returned to its starting temperature.

- **a** Sketch a graph of temperature (*y*-axis) against time (*x*-axis) to show how the temperature of the reaction mixture changes. [1 mark]
- **b** Label the graph clearly and explain what is happening wherever you have shown that the temperature is changing. [6 marks]

2

- **a** Draw a reaction profile to show the exothermic reaction between nitric acid and sodium hydroxide, including its activation energy. [3 marks]
- **b** Draw a reaction profile to show the endothermic change when ammonium nitrate dissolves in water, including its activation energy. [3 marks]

3 When you eat sugar, you break it down to eventually produce water and carbon dioxide.

- **a** Complete the balanced symbol equation:
  $C_{12}H_{22}O_{11} + O_2 \rightarrow$ [2 marks]
- **b** Why must your body *supply* energy in order to break down a sugar molecule? [1 mark]
- **c** When you break down sugar in your body, energy is released. Explain where this energy comes from in terms of the bonds in molecules. [3 marks]
- **d** You can get about 1700 kJ of energy by breaking down 100 g of sugar. If a heaped teaspoon contains 5 g of sugar, how much energy does this release when broken down by your body? [1 mark]

H 4 Hydrogen peroxide has the structure H—O—O—H. It decomposes slowly to form water and oxygen:

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$

The table shows the bond energies for different types of bond.

Bond	Bond energy in kJ/mol
H—O	464
H—H	436
O—O	144
O═O	498

- **a** Use the bond energies to calculate the energy change for the decomposition of hydrogen peroxide, as shown in the equation above. [4 marks]
- **b** Explain exothermic and endothermic reactions in terms of bond breaking and bond making. [4 marks]

5 A student carried out an experiment to find out the voltage produced when different combinations of three metals (A, B, and C – not their chemical symbols) were connected in electrical cells.
Two of their results are shown below:

Donator of electrons (attached to negative terminal of the voltmeter)	Acceptor of electrons (attached to positive terminal of the voltmeter)	Voltage in volts
B	A	1.1
C	B	0.5

- **a** The students also tested metals A and C.
  - **i** Add in the missing row of the table, with your prediction of the voltage produced. [3 marks]
  - **ii** Draw a labelled diagram of the apparatus the student would use to test metals A and C. [3 marks]
- **b** Put the three metals in order of their reactivity, with the most reactive metal first. [1 mark]
- **c** Which metal is the least powerful reducing agent? [1 mark]

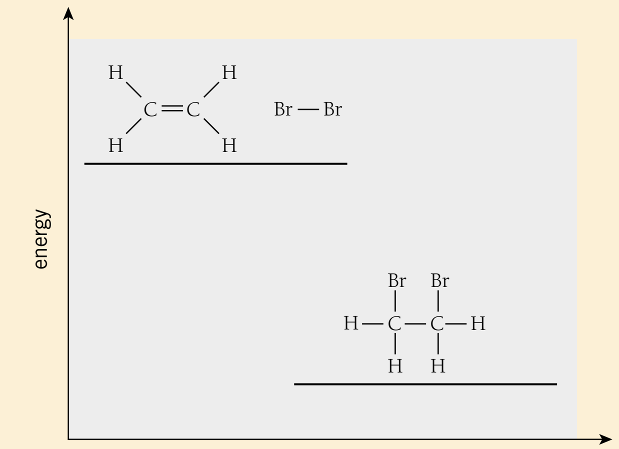
6

- **a** In an alkaline hydrogen fuel cell, name and give the formula for a suitable electrolyte. [2 marks]
- **H b**
  - **i** Write a half equation showing the chemical change involving hydrogen gas that provides the source of the electrons to the external circuit in a hydrogen fuel cell. [1 mark]
  - **ii** What type of reaction does the hydrogen undergo at this electrode in the fuel cell? [1 mark]
- **c**
  - **i** Name the other gas required to operate a hydrogen fuel cell. [1 mark]
  - **ii** Write the balanced symbol equation, including state symbols, for the overall reaction that takes place in a hydrogen fuel cell. [2 marks]
  - **iii** Another type of fuel cell uses methane gas, CH_4, (the main gas in 'natural gas') instead of hydrogen in a fuel cell. Write the balanced symbol equation, including state symbols, for the overall reaction that takes place in a methane fuel cell. [1 mark]
  - **iv** State two advantages a hydrogen fuel cell has compared with a methane fuel cell. [3 marks]

Practice questions

H 01 **Figure 1** shows the energy level diagram for the reaction of reaction of ethene and bromine

Figure 1



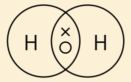
01.1 Copy and complete the diagram and use arrows to label:

- activation energy E_a
- energy released. [3 marks]

01.2 Explain why, in terms of the energy involved in bond breaking and bond making, this reaction is exothermic. [3 marks]

H 02 A molecule of hydrogen is shown in **Figure 2**.

Figure 2



02.1 Describe the attractions in a covalent bond and explain why bond breaking is endothermic. [5 marks]

02.2 Hydrogen reacts with chlorine as shown in the equation.

$$H—H + Cl—Cl \rightarrow 2H—Cl$$

The bond enthalpies are shown in **Table 1**.

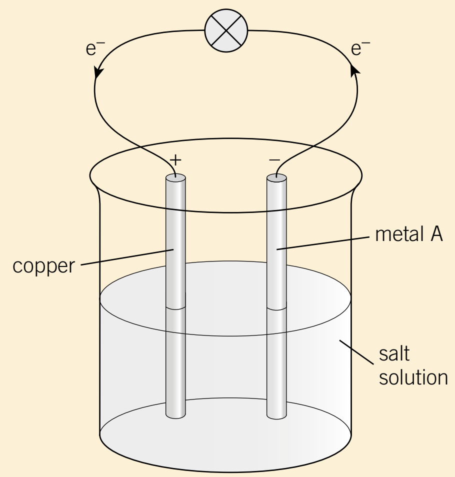
Table 1

Bond	Bond energy in kJ/mol
H—H	436
Cl—Cl	243
H—Cl	432

Calculate the energy change for the reaction in kJ/mol. [3 marks]

03 An electrical cell is made when two different metals are connected together in contact with a salt solution. **Figure 3** shows an electrical cell created by copper and metal **A**.

Figure 3



The bigger the difference in the reactivity of the two metals, the larger the voltage produced.
The voltages produced with four different metals are shown in **Table 2**.

Table 2

Metal A	Voltage in volts
zinc	1.10
iron	0.79
lead	0.47
silver	–0.46

03.1 Why is potassium unsuitable for metal **A**? [2 marks]

03.2 Give **two** variables that should be controlled in this experiment. [2 marks]

03.3 Chromium is less reactive than zinc but more reactive than iron.
Predict the voltage produced when metal **A** is chromium. [1 mark]

03.4 Suggest why the voltage produced when metal **A** is silver is negative. [2 marks]

H 03.5 When metal **A** is zinc, Cu^{2+} ions become copper metal and zinc metal becomes Zn^{2+} ions.
Write half equations for the two processes and explain why Cu^{2+} ions are reduced. [4 marks]