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GCSE (9–1)

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**Introduction**

When a solid (a solute) dissolves in a liquid (a solvent), the bonds holding the solute molecules together break. This allows the molecules of both the solute and solvent to mix, producing a solution.

Solubility is how much solute will dissolve in a solvent. For most solid substances, the solubility increases as the temperature increases. In this core practical, you will measure the solubility of a solute. The solute is normally sodium chloride but your teacher may provide a different one.

**Method**

- 1 Weigh the mass of an empty evaporating basin. Record this and any other values in the results section on the next page.
- 2 Add approximately 50 cm<sup>3</sup> of distilled water into a beaker. Record the temperature of this water.
- 3 Using a spatula, add the solid you have been given to the water in the beaker. Stir with a glass rod and continue to add solid until no more will dissolve. The solid will be in excess (i.e. more than can be dissolved). You will know this because some of the solid will be undissolved at the bottom of the beaker.
- 4 Set up a conical flask with a filter funnel and filter paper. Filter the solution.
- 5 Put the filtrate collected into the evaporating basin and weigh both together. Record this value in the results section on the next page. Make sure that you do not pour any of the undissolved solid into the evaporating basin.
- 6 Heat the evaporating basin and contents gently until the water has evaporated. Your teacher may guide you to use a beaker as a water bath to ensure gentle heating.
- 7 When it looks like all of the water has evaporated, weigh the evaporating basin and its contents. Record this value.
- 8 To make sure all of the water has actually evaporated, heat the evaporating basin and the contents again and reweigh.
- 9 Repeat steps 1 to 8 at least three more times but with different temperatures of water.

**Learning tips**

- The solution you will produce each time is known as a saturated solution.
- There is no need to use a measuring cylinder to measure the volume of water, it is sufficient for this experiment to use the markings on the beaker.
- You can use the following equation to calculate solubility:

$$\text{solubility (g per 100 g of solvent)} = \frac{\text{mass of solid (g)}}{\text{mass of water removed (g)}} \times 100$$

**Objectives**

- To investigate how the temperature of water affects the solubility of a solute.

**Equipment**

- eye protection
- a powdered solid (e.g. sodium chloride)
- distilled water
- 100 cm<sup>3</sup> beaker
- 250 cm<sup>3</sup> beaker
- Bunsen burner
- heat-resistant mat
- gauze and tripod
- conical flask
- digital top pan balance
- evaporating basin
- filter funnel
- filter paper
- glass stirring rod
- spatula
- thermometer
- water bath

**! Safety notes**

- Wear eye protection.
- Gently heat the evaporating basin to prevent spitting of hot solid.
- The solute suggested (sodium chloride) is not harmful. However, you may be given a different solute to use. If so, your teacher will provide additional guidance about using it safely.

**Practical skills**

- Measuring mass
- Safe handling of solids
- Using a Bunsen burner
- Using filtration apparatus

Results

- 1 Complete the results table below. Your teacher may instruct you to carry out your experiment at a particular temperature. You would then be asked to share your results.

Temp. of water (°C)	Mass of basin (g)	Mass of basin and filtrate (g)	Mass of basin and solid (g)	Mass of solid left (g)	Mass of water removed (g)	Solubility (g per 100 g of solvent)

Analysis

- 1 Plot a graph of temperature (on the x-axis) against solubility (on the y-axis). Join up the points with a smooth curve. This is known as a solubility curve.



2 What can you conclude from your solubility curve?

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**Evaluation**

1 Compare your solubility curve with those from other groups. Comment on the similarities and differences between the graphs.

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2 What changes could be made to the method to improve the accuracy of the results?

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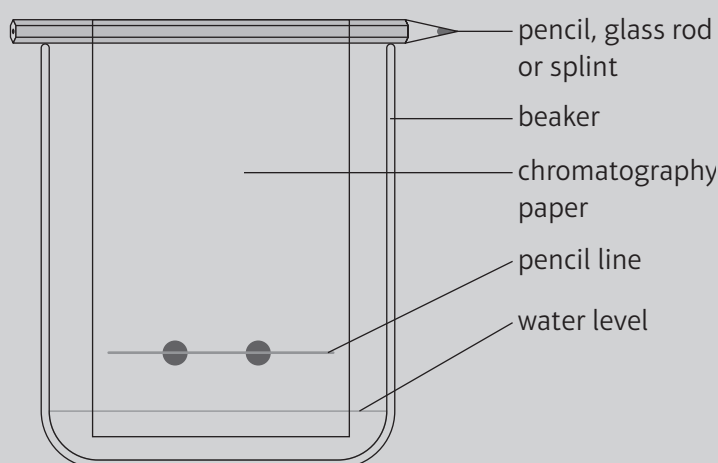
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**Introduction**

Many inks contain a mixture of dyes. Chromatography can be used to identify inks. Some examples would be inks from crime scenes or from documents that may have been forged (illegally copied to cheat somebody).

**Method**

- 1 Check that your chromatography paper hangs close to the bottom of the empty beaker but without touching it, as shown in the diagram.



- 2 Take the paper out of the beaker and draw a pencil line on the paper, about 2 cm from the bottom. This is the baseline.
- 3 Put a small spot of ink from two different pens on the baseline.
- 4 Write the name of each pen or ink colour below each spot. Use a pencil.
- 5 Pour some water into the beaker to a depth of about 1 cm.
- 6 Lower the chromatography paper into the beaker so that the bottom of the paper is in the water, but the water level is below the spots, as shown in the diagram.
- 7 Leave the paper in the beaker until the water soaks up the paper and reaches near the top of the paper. The water is the solvent for the different coloured dye compounds in each ink. The solvent is called the mobile phase in chromatography because it is the part that is moving.
- 8 Take the paper out and immediately use a pencil to mark the location of the solvent front (the level the water has reached) before the water evaporates. Leave the paper to dry.

**Learning tip**

- You could use your own pens, but they need to be water soluble if you use water as a solvent. Permanent pens and colouring pencils will not dissolve in water, so they will not work.

**Results**

- 1 Complete the results table on the next page. There is space in the table for two inks. You will be instructed how to complete the  $R_f$  value row in the **Analysis** section.
  - The 'Distance between baseline and solvent front (cm)' is the value from the lower pencil line to the solvent front.
  - The 'Ink colour' row is the starting colour spotted on the pencil line.

**Objectives**

- To test some inks to see how many dyes they contain and calculate their  $R_f$  values.

**Equipment**

- solvent (water)
- 100 cm<sup>3</sup> beaker
- chromatography paper attached to a pencil, rod or splint
- pencil and ruler
- two marker pens or felt-tip pens

**! Safety notes**

- If the solvent chosen is water, then this is a low-risk experiment. However, if a different solvent is used, your teacher will provide any extra safety guidance.

**Practical skills**

- Measuring distance travelled by solvent
- Measuring height of dye above baseline (estimate to centre of spot)
- Ability to manipulate apparatus for chromatography, recording observations (e.g. number of dyes in each ink, distance travelled by solvent, height of each dye above baseline)
- Ability to carry out investigations safely

- The ‘Colour of dye spots’ row is the various colours the ink has separated into. *There are four spaces here but you may not need all of them.*
- The ‘Distance of spot from baseline (cm)’ row is the distance from the pencil line to the dye spot. Again, you only need to fill this in where you have filled in the box above. You should measure from the middle of the dye spot to make sure that your measuring is consistent (happening in the same way each time) throughout your measuring.

Distance between baseline and solvent front (cm)

Ink colour 1				
Colour of dye spots				
Distance of spot from baseline (cm)				
R <sub>f</sub> value				
Ink colour 2				
Colour of dye spots				
Distance of spot from pencil line (cm)				
R <sub>f</sub> value				

When your paper is dry, attach your chromatogram to this page. This will be a useful reminder of what you need to do to calculate R<sub>f</sub> values when you are revising.

Analysis

- 1 Calculate the R<sub>f</sub> value using the equation below for each separate colour in the inks. Add each R<sub>f</sub> value to your table above.

$$R_f = \frac{\text{distance moved by the coloured spot}}{\text{distance moved by the solvent}}$$

- 2 Were any of the inks a pure colour? Explain your answer.



- 3 Did the same coloured dyes appear in more than one ink? If so, do you think they were the same chemical compound? Explain your answer.

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### Evaluation

- 1 Why was the baseline drawn in pencil?

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- 2 Why did you have to label the spots?

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- 3 Why is the chromatography paper hung with the bottom only just in the water?

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- 4 Why must the water level in the beaker be below the spots?

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- 5 How easy was it to identify the level to which each coloured dye had travelled? How would this affect the accuracy of the  $R_f$  values that you calculated?

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### Introduction

The empirical formula gives you the smallest whole number ratio of atoms in a compound. The empirical formula can be different from the molecular formula. This is because the molecular formula provides the actual number of each atom in a single molecule. For example, the molecular formula of ethanoic acid is  $\text{C}_2\text{H}_4\text{O}_2$  (although it would be more commonly written as  $\text{CH}_3\text{COOH}$ ) but its empirical formula is  $\text{CH}_2\text{O}$ .

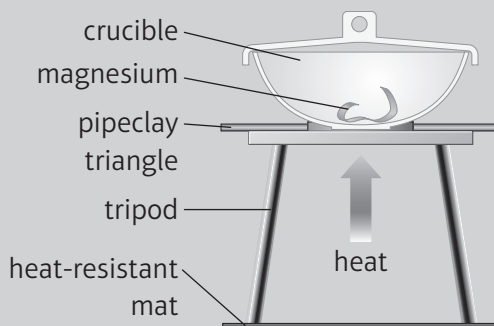
In this core practical, you will determine the empirical formula of a metal oxide using two different techniques. In Part 1, you will carry out a combustion experiment with magnesium. In Part 2, you will carry out or watch a reduction experiment with copper oxide.

### Part 1: Combustion of magnesium

You are going to find the mass of oxygen that combines with a given mass of magnesium to form magnesium oxide. You will then use these masses to determine the empirical formula for magnesium oxide.

### Method

- 1 Find the mass of the crucible and lid. Record this in your results table.
- 2 Clean the magnesium ribbon with emery paper.
- 3 Place the clean magnesium ribbon into the crucible. It might need to be coiled around a pencil to fit it in.
- 4 Find the mass of the crucible, lid and magnesium ribbon. The mass of magnesium will be this reading minus the reading from step 1.
- 5 Set up the apparatus as shown in the diagram.



### 6 Always use tongs to move the crucible and lid – they will be hot.

Heat the crucible strongly with the lid on top for a few minutes. Approximately every minute, lift the lid with a pair of tongs. Be careful not to raise the lid too high otherwise some of the solid will escape. The magnesium should glow red and burn with a bright white flame. It should turn into a white solid. **Do not look directly at the burning magnesium.**

- 7 Allow the crucible to cool enough so that it is safe to move with tongs.
- 8 Find the mass of the crucible, lid and magnesium oxide.
- 9 Reheat the crucible for a further 2 minutes. Record the new mass. If the mass has increased, reheat for a further 2 minutes. Continue to do this until the mass readings remain constant (not changing).

### Objectives

- To determine the empirical formula of magnesium oxide by combustion.
- To determine the empirical formula of copper oxide by reduction.

### Equipment

- eye protection
- magnesium ribbon
- Bunsen burner
- crucible with lid
- digital top pan balance
- emery paper
- heat-resistant mat
- pipeclay triangle
- tongs
- tripod



### Safety notes

- Wear eye protection.
- Use tongs to handle the hot crucible.
- Do not look directly at the burning magnesium as it produces a very bright flame.

### Practical skills

- Measuring mass
- Safe handling of solids
- Using a Bunsen burner
- Safe handling of hot apparatus

**Learning tip**

- Make sure that the balance is reading zero each time you use it.

**Results**

- 1 Complete the results table below.

Mass of empty crucible and lid (g)	
Mass of crucible, lid and magnesium (g)	
Mass of crucible, lid and magnesium oxide after several minutes (g)	
Mass of crucible, lid and magnesium oxide after further 2 minutes (g)	
Mass of crucible, lid and magnesium oxide after further 2 minutes (g)	
Highest mass of crucible, lid and magnesium oxide (g)	

**Analysis**

- 1 Mass of *magnesium* used: \_\_\_\_\_ g  
(mass of crucible, lid and magnesium minus mass of empty crucible and lid)
- 2 Mass of *magnesium oxide* formed: \_\_\_\_\_ g  
(highest mass of crucible, lid and magnesium oxide found minus mass of empty crucible and lid)
- 3 Mass of *oxygen* that has reacted: \_\_\_\_\_ g  
(highest mass of crucible, lid and magnesium oxide found minus mass of crucible, lid and magnesium)
- 4 Complete the table below to calculate the empirical formula for magnesium oxide.

Element	magnesium	oxygen
Chemical symbol	Mg	O
Mass reacted in experiment (g)		
Relative atomic mass ( $A_r$ )	24	16
Divide the mass of each element by its $A_r$		
Divide both answers by the smaller of the two to find the simplest ratio		
Find the simplest whole number ratio		

- 5 The empirical formula of magnesium oxide is: \_\_\_\_\_

**Evaluation**

**1** Why was it necessary to lift the lid of the crucible occasionally during heating?

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**2** Describe how you knew when all the magnesium had completely burned.

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**3** All groups should have used the same length of magnesium ribbon. Comment on similarities and differences of the results obtained by different groups, including the empirical formula.

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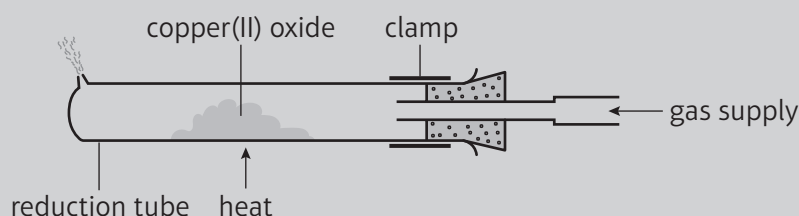
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**Part 2: Reduction of copper oxide**

You can find the formula of copper(II) oxide by removing oxygen from it and looking at how the mass changes.

**Method**

- 1 Weigh a reduction tube with its rubber bung attached.
- 2 Put approximately 3 g of copper oxide into the middle of the reduction tube and weigh the tube again.
- 3 Set up the apparatus as shown in the diagram.



- 4 Pass methane gas over the copper oxide. Allow the gas to run through for 10 seconds before igniting it.
- 5 Use a lit wooden splint to ignite the excess methane, which comes out of the small hole in the reduction tube. You **must** do this before you light the Bunsen burner.  
Do not lean over the apparatus when lighting the excess methane gas.
- 6 Light the Bunsen burner.
- 7 Heat the copper(II) oxide strongly until the reaction is finished. The colour should change from black to pink-brown.
- 8 Turn off the Bunsen burner and leave the reduction tube to cool.
- 9 Turn off the gas supply once the reduction tube has cooled.
- 10 Weigh the reduction tube with the bung and its contents. Use tongs to handle the reduction tube.

**Learning tip**

- It is likely that this experiment will be carried out as a demonstration or shown as a video for safety reasons. The data obtained will still be valid (proved to be true) even if you do not obtain it directly by carrying out an experiment yourself.

**Results**

- 1 Complete the results table below.

Mass of empty reduction tube (g)	
Mass of reduction tube and copper(II) oxide (g)	
Mass of reduction tube and copper (g)	

**Equipment**

- eye protection
- copper(II) oxide powder
- bung and tubing to connect to gas supply
- Bunsen burner
- clamp, stand and boss
- digital top pan balance
- heat-resistant mat
- reduction tube
- wooden splint
- tongs

**! Safety notes**

- Wear eye protection.
- Use a safety screen to shield equipment from those watching.
- Copper(II) oxide is harmful if swallowed or inhaled and causes skin irritation (pain).
- Methane gas is highly flammable, and mixtures of air and methane are explosive. Ensure the reduction tube is flushed with methane gas for 10 seconds before igniting the gas coming out of the end. You **must** light the excess methane gas coming out of the small hole in the reduction tube before lighting the Bunsen burner. Do not lean over the apparatus when lighting the excess methane gas.
- The reduction tube will be hot when removing it from the clamp. Wait for it to cool and use tongs to pick it up to prevent burns.

**Practical skills**

- Measuring mass
- Safe handling of solids
- Using a Bunsen burner
- Safe handling of hot apparatus

**Analysis**

- 1 Mass of *copper oxide* used: \_\_\_\_\_g  
(mass of reduction tube and copper(II) oxide minus mass of empty reduction tube)
- 2 Mass of *copper* formed: \_\_\_\_\_g  
(mass of reduction tube and copper minus mass of empty reduction tube)
- 3 Mass of *oxygen* that has reacted: \_\_\_\_\_g  
(mass of empty reduction tube and mass of copper(II) oxide used minus mass of reduction tube and copper)
- 4 Complete the table to calculate the empirical formula for copper(II) oxide.

Elements	copper	oxygen
Chemical symbol	Cu	O
Mass found in experiment (g)		
Relative atomic mass ( $A_r$ )	63.5	16
Divide the mass of each element by its $A_r$		
Divide both answers by the smaller of the two to find the simplest ratio		
Find the simplest whole number ratio		

- 5 The empirical formula of copper(II) oxide is: \_\_\_\_\_
- 6 What happens to the appearance of the copper oxide during the experiment? Why does this happen?

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**Evaluation**

- 1 What would happen to the mass of the solid after the experiment if it were left for a week and reweighed?

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### Introduction

The electrolysis of molten or dissolved ionic salts is carried out using inert (unreactive) electrodes (usually graphite or platinum). When a molten salt is electrolysed, ions are discharged as atoms or molecules at the electrodes. However, electrolysis of dissolved ionic salts is more complex. This is because water ionises to a very small extent. So, in an aqueous solution of a salt there are some hydrogen ions ( $\text{H}^+$ ) and hydroxide ions ( $\text{OH}^-$ ), as well as the ions of the dissolved solid.

You are going to predict the substances that are formed at the electrodes when different substances are electrolysed. You will then carry out the electrolysis of these substances to confirm whether you were right.

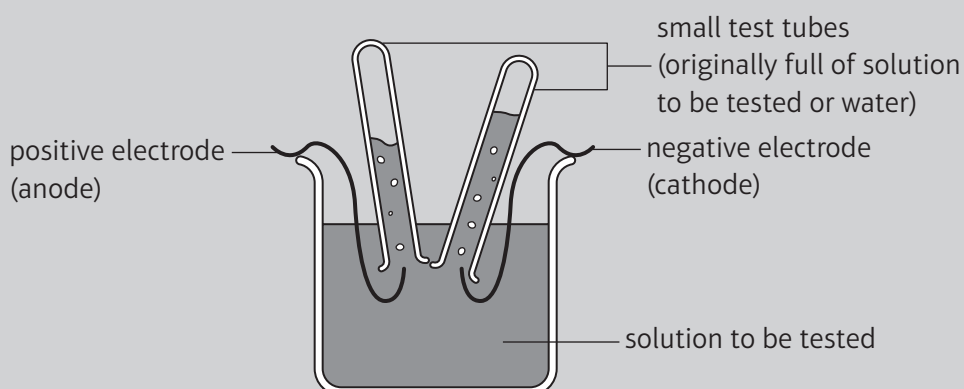
### Prediction

Predict what products will be formed at each electrode. There are two blank rows in case your teacher provides you with additional solutions to investigate.

Solution	Anode (positive) electrode	Cathode (negative) electrode
copper(II) sulfate		
sodium chloride		
sulfuric acid		

### Method

- 1 Place about 50 cm<sup>3</sup> of copper(II) chloride solution into a 100 cm<sup>3</sup> beaker.
- 2 Set up the apparatus as shown in the diagram below. The test tubes over the electrodes can be filled with water or the solution being tested.



### Objectives

- To investigate the substances that are formed at the electrodes when different salt solutions are electrolysed.

### Equipment

- eye protection
- 0.1 mol/dm<sup>3</sup> copper(II) sulfate solution
- 0.1 mol/dm<sup>3</sup> sodium chloride solution
- 0.1 mol/dm<sup>3</sup> sulfuric acid
- 100 cm<sup>3</sup> beaker
- Bunsen burner
- heat-resistant mat
- connecting leads
- damp litmus paper
- electrode holder
- low-voltage power supply
- sandpaper
- S-shaped nichrome electrodes
- wooden splints

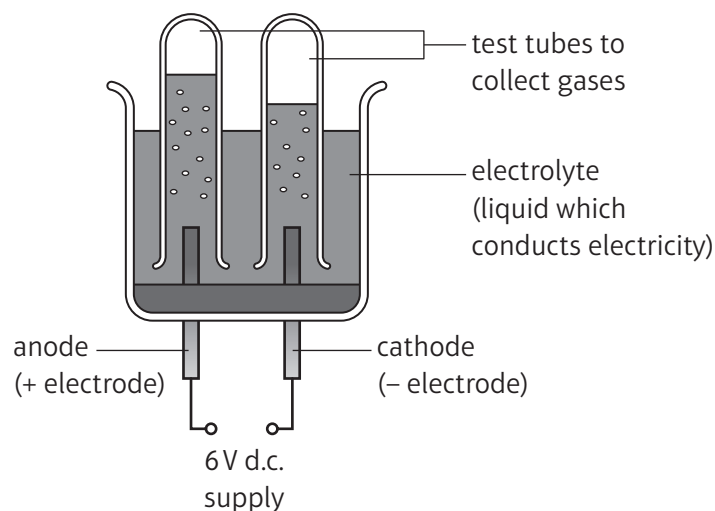
### Safety notes

- Wear eye protection.
- 0.1 mol/dm<sup>3</sup> sulfuric acid and 0.1 mol/dm<sup>3</sup> copper(II) sulfate are irritants.
- Chlorine gas is toxic and is produced in some of the experiments. Only run the experiment to test the gas. Ensure the lab is well ventilated (open to fresh air).
- Hydrogen is extremely flammable.
- Take care with aqueous solutions and electricity. Do not use a voltage higher than 6 V and only switch on the circuit when it is fully set up. If you are unsure, ask your teacher to check your set-up.
- If your teacher provides you with other solutions to test, they will provide you with additional safety guidance if required.

### Practical skills

- Setting up an electrical circuit for electrolysis
- Safe handling of solutions
- Collecting gases and testing for them

Your teacher may provide you with an electrolytic cell to carry out the experiment. This is shown below. If so, they will show you how to set this up.



- 3 Use the leads to connect the electrodes to the low-voltage power supply. Make sure you connect the electrodes to the d.c. terminals. These are normally coloured red and black.
- 4 Set the voltage to 6 V and switch on.
- 5 Look at the electrodes. Record any colour change or fizzing in your results table.
- 6 Test any gas that has been produced. Start by testing for chlorine. Hold a piece of damp litmus paper at the mouth of the test tube. If the litmus paper bleaches, the gas is chlorine. Add this information to your results table.
- 7 If the gas is not chlorine, test for hydrogen (lit splint goes out with a squeaky pop) and then for oxygen (glowing splint relights). Record any information in your results table.
- 8 Switch off the low-voltage supply.
- 9 Remove the liquid from the beaker. Your teacher will tell you where to put the waste liquid. You will need to clean both the electrodes. Rinsing them under running water should be fine. If a solid has formed on the electrodes, use sandpaper to clean them before running them under water.
- 10 Repeat steps 1 to 9 using the other solutions.

### Learning tip

- Graphite electrodes are often used in electrolysis in school as they are cheap. However, they are not really suitable for any electrolysis experiments where oxygen is evolved. This is because the graphite (carbon) will react with the oxygen to produce carbon dioxide.

### Results

- 1 Complete the results table on the next page with the observations taken during the electrolysis and the results of any gas tests you have performed. Your teacher may instruct you use additional solutions. These can be entered in the two rows at the bottom of the table.



Solution	Anode (positive) electrode		Cathode (negative) electrode	
	Observations	Identity of product	Observations	Identity of product
copper(II) sulfate				
sodium chloride				
sulfuric acid				

**Analysis**

- 1 Compare your final results with the predictions that you made at the start. How accurate were your predictions?

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- 2 The test for chlorine is that it bleaches damp litmus paper. Write down the standard tests for hydrogen gas and for oxygen gas.

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- 3 Explain, in terms of the reactivity series, the pattern in the products formed at the cathode.

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- 4 Write a general rule for predicting the product formed at the anode.

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## CORE PRACTICAL 5: RUSTING

### DETERMINE THE APPROXIMATE PERCENTAGE BY VOLUME OF OXYGEN IN AIR USING A METAL OR A NON-METAL

SPECIFICATION  
REFERENCE

2.14

(2.10)

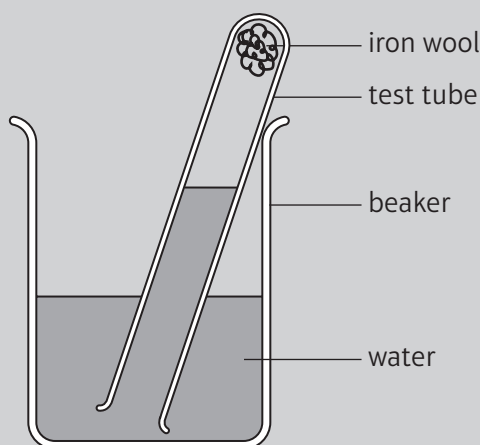
#### Introduction

You probably know that the Earth's atmosphere is composed mostly of nitrogen and oxygen. Nitrogen accounts for about 78% of the atmosphere and oxygen 21%. The remaining 1% is largely argon. Carbon dioxide accounts for only 0.04%.

In this experiment, you are going to use iron to determine the percentage of oxygen in the air.

#### Method

- 1 Soak some iron wool in water. Your teacher or lab technician will have already divided the iron wool into suitable sized pieces.
- 2 Use a stirring rod to push the iron wool to the bottom of a test tube.
- 3 Half fill a beaker with water and invert (turn upside down) the test tube in the beaker as shown in the diagram.



- 4 Make sure the test tube is vertical (*not* at an angle) and then measure the height of air in the test tube in mm. This is measured from the water level *in the tube* to the tip of the test tube
- 5 Leave the test tube at an angle, as shown in the diagram above, for a week.
- 6 After a week, measure the height of the air in the test tube.

#### Learning tip

- Be careful when pushing the iron wool into the test tube. Try not to compress the iron wool too much otherwise there will be parts that cannot react with the oxygen in the test tube.

#### Results

- 1 Complete the results table below.

Height of air in test tube at start of experiment (mm)	
Height of air in test tube at end of experiment (mm)	
Decrease in height of air in test tube (mm)	

#### Objectives

- To determine the percentage by volume of oxygen in the air.

#### Equipment

- eye protection
- iron wool
- water
- 250 cm<sup>3</sup> beaker
- glass stirring rod
- ruler
- test tube



#### Safety notes

- Wear eye protection.

#### Practical skills

- Measuring the length of a column of air

**Analysis**

- 1 Calculate the percentage of oxygen using this equation:

$$\frac{\text{decrease in height of air (mm)}}{\text{starting height of air (mm)}} \times 100 = \text{percentage decrease in volume of air}$$

percentage of oxygen in air: \_\_\_\_\_ %

**Evaluation**

- 1 Compare your results to the known percentage of oxygen in the air (see Introduction). How close were you?

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- 2 Why should the test tube be at an angle and not vertical on its end while the reaction is taking place?

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- 3 A student obtained a result of 18% oxygen in air. What is the most likely reason for the discrepancy (difference) between this value and the actual value?

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# CORE PRACTICAL 6: ACID-METAL REACTIONS

## INVESTIGATE REACTIONS BETWEEN DILUTE HYDROCHLORIC AND SULFURIC ACIDS AND METALS

SPECIFICATION  
REFERENCE

2.21

(2.15, 2.17)

### Introduction

Some metals react with acids to make a salt and hydrogen gas. If you use the same concentration of acid with different metals, it is possible to work out the relative reactivity of the metals. The more vigorous the reaction, the more reactive the metal is. In this experiment you will react various metals with two different acids. By observing the reactions, you will be able to determine the order of reactivity of the metals.

### Method: Testing calcium

- 1 Half fill a 100 cm<sup>3</sup> beaker with 0.4 mol/dm<sup>3</sup> hydrochloric acid.
- 2 Add one or two pieces of calcium to the acid and record your observations. Use a spatula or tweezers when handling calcium as it can cause irritation when it reacts with water on skin.
- 3 When the reaction has finished, dispose of the leftover liquid. Your teacher will explain what to do with it.
- 4 Repeat the experiment with another 100 cm<sup>3</sup> beaker but this time use 0.2 mol/dm<sup>3</sup> sulfuric acid.

### Method: Testing the other metals

- 1 Place five test tubes in a test tube rack and quarter fill each tube with 0.4 mol/dm<sup>3</sup> hydrochloric acid.
- 2 Add one of the other metals to one of the test tubes. Notice whether there is any effervescence (fizzing) and, if there is, test the gas with a lighted wooden splint. Record your results.
- 3 Your teacher will show you how to dispose of the contents of the test tube.
- 4 Repeat step 2 with each of the other metals. Use a new test tube for each different metal.
- 5 Repeat steps 1 to 4 but use 0.2 mol/dm<sup>3</sup> sulfuric acid. Again, test the metals one at a time. Test for hydrogen gas if any effervescence is seen.

### Learning tips

- The exact amount of acid is not important. The acid is in excess (more than is needed) in each experiment.
- Do not pour your mixtures straight into a sink as some of the solids remaining will not dissolve in water and may block the waste pipes.

### Results

- 1 Complete the results table below.

Metal	Reaction with hydrochloric acid	Reaction with sulfuric acid
calcium		
aluminium		
copper		
iron		
magnesium		
zinc		

### Objectives

- To investigate the reactions of some metals with dilute hydrochloric acid and sulfuric acid.

### Equipment

- eye protection
- pieces of aluminium, calcium, copper, iron, magnesium and zinc
- 0.4 mol/dm<sup>3</sup> hydrochloric acid
- 0.2 mol/dm<sup>3</sup> sulfuric acid
- 100 cm<sup>3</sup> beakers
- Bunsen burner
- heat-resistant mat
- spatula or tweezers
- test tube rack
- test tubes
- wooden splints



### Safety notes

- Wear eye protection.
- The reaction with calcium must take place in a beaker; the other reactions can take place in test tubes. Calcium can cause irritation when it reacts with water on skin – use a spatula or tweezers when handling calcium.
- 0.4 mol/dm<sup>3</sup> hydrochloric acid and 0.2 mol/dm<sup>3</sup> sulfuric acid are irritants.
- Magnesium metal is flammable.

### Practical skills

- Safe handling of dilute acids
- Safe handling of metals
- Testing for hydrogen

**Analysis**

1 Arrange the six metals in order of reactivity, with most reactive at the top.

.....

.....

.....

.....

.....

.....

2 What do you notice about the reaction between calcium and sulfuric acid? Can you explain your observation?

.....

.....

3 What do you notice about the reaction between aluminium and either acid? Can you explain your observations?

.....

.....

4 Write balanced chemical equations for the following reactions:

- calcium and hydrochloric acid

.....

- magnesium and hydrochloric acid

.....

- zinc and sulfuric acid

.....

# CORE PRACTICAL 7: PREPARATION OF COPPER SULFATE

## PREPARE A SAMPLE OF PURE, DRY HYDRATED COPPER(II) SULFATE CRYSTALS STARTING FROM COPPER(II) OXIDE

SPECIFICATION  
REFERENCE

2.42

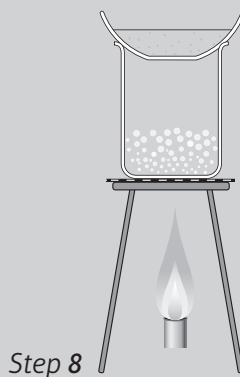
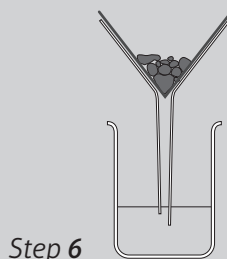
(1.4, 2.39)

### Introduction

Salts, such as copper sulfate, are compounds formed by reacting an acid with a base. Copper(II) oxide reacts with warm dilute sulfuric acid to produce a blue solution of the salt copper(II) sulfate. In this practical, you will use these reactants to prepare pure, dry, hydrated copper(II) sulfate crystals.

### Method

- 1 Pour about 25 cm<sup>3</sup> of 1 mol/dm<sup>3</sup> sulfuric acid into a conical flask.
- 2 Place the conical flask into a water bath at 50 °C and heat for 3–4 minutes to allow the acid to heat up.
- 3 Use the spatula to add a little copper(II) oxide to the acid and stir or swirl the contents of the flask.
- 4 Keep repeating step 3 until the black powder does not disappear after stirring. This makes sure that the copper(II) oxide is in excess.
- 5 Return the mixture to the water bath for a few minutes, to make sure that there is no acid left.
- 6 Filter the mixture into a beaker.
- 7 Pour the filtrate into an evaporating basin.
- 8 Place the evaporating basin on top of a beaker half full of water. Heat the beaker, evaporating basin and contents using a Bunsen burner on a blue flame.
- 9 Heat until about half of the water has evaporated from the evaporating basin. Then allow the evaporating basin to cool.
- 10 When cool, filter the crystals from the remaining liquid and rinse the crystals that are in the filter paper with some distilled water.
- 11 Use tweezers to remove the crystals and press them between filter papers.
- 12 Place the crystals in a warm place to dry.
- 13 Observe the shape and colour of the copper(II) sulfate crystals formed.
- 14 Record the final mass of the copper(II) sulfate crystals formed.



### Objectives

- To prepare a sample of pure, dry hydrated copper(II) sulfate crystals starting from copper(II) oxide.

### Equipment

- eye protection
- copper(II) oxide
- 1 mol/dm<sup>3</sup> sulfuric acid
- 100 cm<sup>3</sup> beaker
- 100 cm<sup>3</sup> conical flask
- 100 cm<sup>3</sup> measuring cylinder
- Bunsen burner
- clamp, stand and boss
- evaporating basin
- filter funnel
- filter paper
- gauze and tripod
- heat-resistant mat
- Petri dish or watch glass
- spatula
- stirring rod
- tongs
- tweezers
- water bath (set at 50 °C)



### Safety notes

- Wear eye protection.
- 1 mol/dm<sup>3</sup> sulfuric acid, copper(II) oxide and copper(II) sulfate are irritants.

### Practical skills

- Measuring the volume of sulfuric acid
- Safe use of a Bunsen burner for warming the sulfuric acid
- Safe use of a water bath or electric heater for evaporating some of the water from the copper sulfate solution
- Safe use of filtration to separate unreacted copper(II) oxide from copper(II) sulfate solution
- Safe use of evaporation to evaporate some of the water from the copper(II) sulfate solution
- Safe use and handling of sulfuric acid, copper(II) oxide and copper(II) sulfate

**Results**

- 1** Mass of dry, hydrated copper(II) sulfate: \_\_\_\_\_ g
- 2** Describe the colour, shape and size of the copper(II) sulfate crystals produced.

.....

.....

- 3** Describe the appearance of:

**a** the sulfuric acid

.....

**b** the copper(II) oxide

.....

**c** the solution at the end of the reaction.

.....

- 4** The maximum theoretical yield of hydrated copper(II) sulfate for this reaction is 6.24 g. Use the equation below to calculate your percentage yield.

$$\frac{\text{actual yield (g)}}{\text{theoretical yield (g)}} \times 100 = \text{percentage yield}$$

percentage yield: \_\_\_\_\_%

**Analysis**

- 1** Write a word equation to show the reaction you have carried out.

.....

- 2** State why you need to be sure that excess copper(II) oxide is added in step 4.

.....

.....

- 3** What would happen in step 5 if there was still some acid left?

.....

- 4** Name the substance left in the filter paper in step 6.

.....

- 5** What is dissolved in the solution that went through the filter paper in step 6?

.....

**6** Explain why this reaction is an example of a neutralisation reaction.

.....

.....

**7** What substance acts as a base in this reaction?

.....

**8** Write a symbol equation to show the reaction you have carried out. Include the state symbols. Use your answer to question 1 to help you.

.....



# CORE PRACTICAL 8: PREPARATION OF LEAD SULFATE

## PREPARE A SAMPLE OF PURE, DRY LEAD(II) SULFATE

SPECIFICATION  
REFERENCE

2.43C

(2.41C)

### Introduction

Salts can be soluble or insoluble. The salt you want to prepare determines the method that you need to use. Is the salt soluble or insoluble? In Core Practical 7, you made a soluble salt and had to allow the water to evaporate in order to obtain the salt crystals. Insoluble salts are made with precipitation reactions and can be easily obtained by filtering.

### Method

- 1 Measure 25 cm<sup>3</sup> of lead(II) nitrate solution using a measuring cylinder. Pour this into a 100 cm<sup>3</sup> beaker.
- 2 Measure 30 cm<sup>3</sup> of sodium sulfate using a measuring cylinder. Add this to the lead(II) nitrate and stir well. Lead(II) sulfate should precipitate out.
- 3 Set up equipment for filtering.
- 4 Filter the contents of the beaker.
- 5 Rinse out the beaker with distilled water and add to the filter paper.
- 6 Rinse the residue in the filter paper with distilled water.
- 7 Remove the filter paper from the filter funnel and allow to dry in an oven or desiccator.

### Results

- 1 Describe the appearance of the lead(II) sulfate produced.
- .....

### Analysis

- 1 Write a word equation to show the reaction you have carried out.
- .....

- 2 Write a balanced symbol equation for this reaction.

- lead nitrate is Pb(NO<sub>3</sub>)<sub>2</sub>
  - sodium sulfate is Na<sub>2</sub>SO<sub>4</sub>
- .....

- 3 Write an ionic equation for this reaction.
- .....

- 4 What are the spectator ions in this reaction?
- .....

### Evaluation

- 1 Why do you need to make sure that the solution does not go above the top of the filter paper when you are filtering?
- .....

### Objectives

- To prepare a sample of pure, dry, lead(II) sulfate crystals.

### Equipment

- eye protection
- 0.1 mol/dm<sup>3</sup> lead(II) nitrate
- 0.1 mol/dm<sup>3</sup> sodium sulfate
- distilled water
- 25 cm<sup>3</sup> and 50 cm<sup>3</sup> measuring cylinders
- 100 cm<sup>3</sup> beaker
- conical flask
- filter funnel
- filter paper
- glass stirring rod
- oven or desiccator

### Safety notes

- Wear eye protection.
- Lead salts are very dangerous – a serious hazard to health. Avoid allowing dust to rise. Take particular care to avoid inhalation or skin contact. Wash hands thoroughly after use. Clear up spills and wipe over surfaces after using lead salts.

### Practical skills

- Safe use of filtration to separate residue of lead(II) sulfate from the reaction mixture
- Safe use and handling of solutions of lead salts and solid lead(II) sulfate

# CORE PRACTICAL 9: ENDOTHERMIC AND EXOTHERMIC REACTIONS

## INVESTIGATE TEMPERATURE CHANGES ACCOMPANYING SOME OF THE FOLLOWING TYPES OF CHANGE: NEUTRALISATION REACTIONS, SALTS DISSOLVING IN WATER, DISPLACEMENT REACTIONS AND COMBUSTION REACTIONS

SPECIFICATION  
REFERENCE

3.8

(3.1, 3.2, 3.3, 3.4)

### Introduction

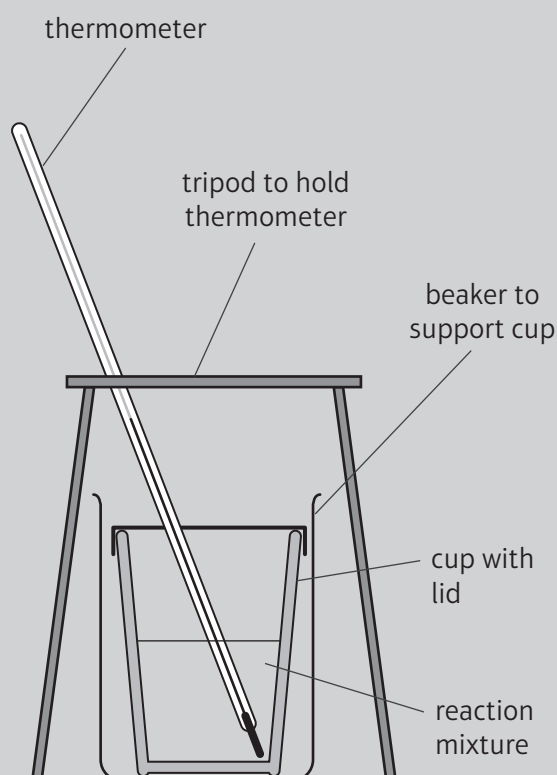
During a chemical reaction, energy is transferred between the reacting substances and their surroundings. This energy transfer usually occurs by heating, particularly if a reaction takes place in solution. The stored thermal (heat) energy in the solution increases during an exothermic reaction, and it decreases during an endothermic reaction. This means that you can determine whether a reaction in solution is exothermic or endothermic. The change in energy is called the enthalpy change and we use calorimetry experiments to measure this energy change.

This core practical is divided into four experiments, each one with a different focus: neutralisation, dissolving, displacement or combustion.

### Part 1: Neutralisation

#### Method

- 1 Place one polystyrene cup into a glass beaker to make it more stable. Enclose (surround) the thermometer with a tripod to stop the cup and thermometer falling over.
- 2 Use a measuring cylinder to put  $30\text{ cm}^3$  of  $1.0\text{ mol/dm}^3$  hydrochloric acid into the polystyrene cup.
- 3 Wash out the measuring cylinder with water.
- 4 Take the temperature of the acid and write this in your results table. This should go in the column for experiment 1.
- 5 Use the measuring cylinder to put  $40\text{ cm}^3$  of  $1.0\text{ mol/dm}^3$  sodium hydroxide solution into a second polystyrene cup.
- 6 Remove  $5\text{ cm}^3$  of sodium hydroxide solution from the  $40\text{ cm}^3$  you measured out in step 5.
- 7 Add this  $5\text{ cm}^3$  to the polystyrene cup containing the acid. Put the lid on the cup containing the mixture and mix gently using the stirring rod.



### Objectives

- To investigate the temperature changes accompanying the reaction between hydrochloric acid and sodium hydroxide solution in a neutralisation reaction.

### Equipment

- eye protection
- $1.0\text{ mol/dm}^3$  hydrochloric acid
- $1.0\text{ mol/dm}^3$  sodium hydroxide
- $250\text{ cm}^3$  beaker
- lid for one cup with holes for stirring rod and thermometer
- measuring cylinders
- polystyrene cups (2)
- thermometer
- stirring rod
- tripod

### Safety notes

- Wear eye protection.
- $1.0\text{ mol/dm}^3$  sodium hydroxide is corrosive and very damaging to eyes.
- Care is needed with solutions of acid. Wash off splashes immediately.

### Practical skills

- Measure volume of solution
- Measure temperature rise
- Measure mass

- 8 Keep stirring until the temperature reaches a maximum and starts to fall. Record the highest temperature in your table.
- 9 Repeat steps 6 to 8 until all 40 cm<sup>3</sup> of sodium hydroxide has been added.
- 10 When you have finished, rinse out your cups and measuring cylinders and repeat the experiment at least one time. If time is short, you could use results from other groups because everyone will have used the same amounts and the same strengths of solutions.

### Learning tips

- You can use the following equation to calculate heat energy transferred:

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

where  $Q$  is heat energy transferred (J),  $m$  is mass of liquid being heated (g),  $c$  is specific heat capacity (J/g/°C) and  $\Delta T$  is the change in temperature of the liquid (°C). This experiment uses water; the specific heat capacity of water = 4.2 J/g/°C.

- You can use the following equation to calculate number of moles,  $n$ :

$$n = \text{concentration} \times \text{volume}$$

- You can use the following equation to calculate enthalpy change:

$$\Delta H = \frac{-Q}{n}$$

where  $\Delta H$  is enthalpy change (kJ/mol),  $Q$  is heat energy transferred (J) and  $n$  is number of moles.  $Q$  is in joules (J) so needs to be converted to kilojoules (kJ); divide  $Q$  by 1000 to use in this equation.

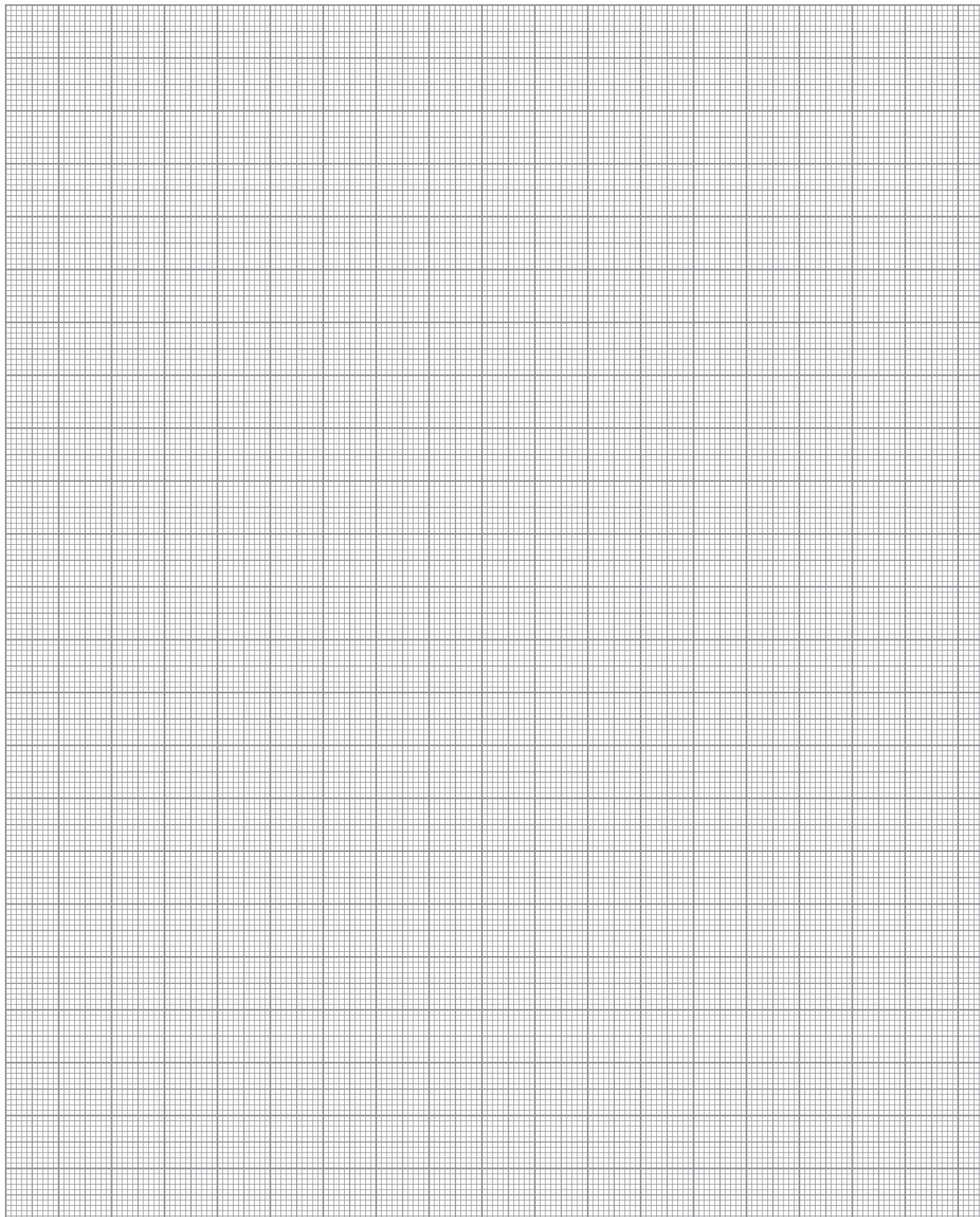
### Results

- 1 Put your results into the table below. There are some extra columns for either repeated results from your experiments or results from other groups. To calculate the mean, add up the results in a row and divide by the number of results that you used.

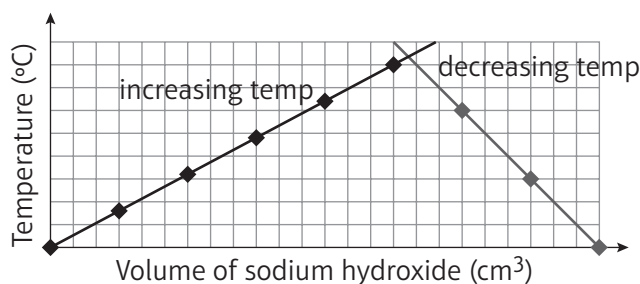
Volume of sodium hydroxide solution added (cm <sup>3</sup> )	Maximum temperature recorded at each stage (°C)				
	Experiment 1	Experiment 2			Mean
0					
5					
10					
15					
20					
25					
30					
35					
40					

**Analysis**

- 1 Draw a graph of your results. On the x-axis, place the volume of sodium hydroxide solution added ( $\text{cm}^3$ ). On the y-axis, place the mean maximum temperature ( $^{\circ}\text{C}$ ).



- 2 Draw two straight lines of best fit. One will be through all the points that are increasing in temperature. The other will be through all the points that are decreasing in temperature. You need to make sure that the lines cross, so you might need to extend the lines. The diagram below should give you an idea what it should look like.



- 3 What volumes of hydrochloric acid and sodium hydroxide would produce the largest temperature rise?

.....

- 4 Why does the temperature start to fall towards the end of the experiment?

.....

.....

- 5 With your data, calculate the energy change ( $Q$ ) and the enthalpy change ( $\Delta H$ ) for this reaction. Determine if this reaction is exothermic or endothermic. Remember, temperature change is measured from the starting temperature of the experiment.

This reaction is .....

- 6 Compare your answer with the data book value of  $-57.2$  kJ/mol for this experiment. How did you get on?

.....

.....

.....

**Evaluation**

**1** How could you make the data you collect from this investigation more accurate?

.....

.....

**2** It is difficult to find the exact volume of sodium hydroxide solution that would give the maximum temperature rise. What further work would need to be carried out to find the exact volume of sodium hydroxide?

.....

.....

.....

.....

# CORE PRACTICAL 9: ENDOTHERMIC AND EXOTHERMIC REACTIONS

## INVESTIGATE TEMPERATURE CHANGES ACCOMPANYING SOME OF THE FOLLOWING TYPES OF CHANGE: NEUTRALISATION REACTIONS, SALTS DISSOLVING IN WATER, DISPLACEMENT REACTIONS AND COMBUSTION REACTIONS

SPECIFICATION  
REFERENCE

3.8

(3.1, 3.2, 3.3, 3.4)

### Part 2: Dissolving salts

#### Method

This investigation uses the same equipment as Part 1. Refer to that experiment for a diagram of the set-up.

- 1 Place the polystyrene cup into a glass beaker to make it more stable. Enclose (surround) the thermometer with a tripod to stop the cup and thermometer falling over.
- 2 Use a measuring cylinder to put 90 cm<sup>3</sup> of distilled water into the polystyrene cup.
- 3 Take the temperature of the water and write this in your results table.
- 4 Weigh out 2.67 g of ammonium chloride and add this to the water in the polystyrene cup. Put the lid on the cup and mix gently using the stirring rod.
- 5 Keep stirring until the temperature stops dropping and starts to rise. Record the lowest temperature in your table.
- 6 When you have finished, rinse out your cup and measuring cylinder and repeat the experiment at least one time. If time is short, you could use results from other groups because everyone will have used the same amounts and the same strengths of solutions.

#### Learning tips

- You can use the following equation to calculate heat energy transferred:

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

where  $Q$  is heat energy transferred (J),  $m$  is mass of liquid being heated (g),  $c$  is specific heat capacity (J/g/°C) and  $\Delta T$  is the change in temperature of the liquid (°C). This experiment uses water; the specific heat capacity of water = 4.2 J/g/°C.

You can use the following equation to calculate number of moles,  $n$ :

$$n = \text{concentration} \times \text{volume}$$

- You can use the following equation to calculate enthalpy change:

$$\Delta H = \frac{-Q}{n}$$

where  $\Delta H$  is enthalpy change (kJ/mol),  $Q$  is heat energy transferred (J) and  $n$  is number of moles.  $Q$  is in joules (J) so needs to be converted to kilojoules (kJ); divide  $Q$  by 1000 to use in this equation.

#### Results

- 1 Put your results into the table below.

	Experiment 1	Experiment 2	Mean
Starting temperature (°C)			
Lowest final temperature (°C)			

#### Objectives

- To investigate the temperature changes accompanying the reaction between ammonium chloride with water in a dissolving salt reaction.

#### Equipment

- eye protection
- ammonium chloride
- distilled water
- 250 cm<sup>3</sup> beaker
- lid for cup
- measuring cylinder
- polystyrene cup
- thermometer
- stirring rod
- tripod
- top pan balance
- weighing boat

#### ! Safety notes

- Wear eye protection.
- Ammonium chloride is harmful. Avoid contact with your skin.

#### Practical skills

- Measure mass
- Measure temperature change

**Analysis**

- 1 With your data, calculate the energy change ( $Q$ ) and the enthalpy change ( $\Delta H$ ) for this reaction. Determine if this reaction is exothermic or endothermic.

This reaction is .....

- 2 Compare your answer with the data book value of +16.4 kJ/mol for this experiment. How did you get on?



# CORE PRACTICAL 9: ENDOTHERMIC AND EXOTHERMIC REACTIONS

## INVESTIGATE TEMPERATURE CHANGES ACCOMPANYING SOME OF THE FOLLOWING TYPES OF CHANGE: NEUTRALISATION REACTIONS, SALTS DISSOLVING IN WATER, DISPLACEMENT REACTIONS AND COMBUSTION REACTIONS

SPECIFICATION  
REFERENCE

3.8

(3.1, 3.2, 3.3, 3.4)

### Part 3: Displacement

#### Method

This investigation uses the same equipment as Part 1. Refer to that experiment for a diagram of the set-up.

- 1 Place the polystyrene cup into a glass beaker to make it more stable. Enclose the thermometer with a tripod to stop the cup and thermometer falling over.
- 2 Use a measuring cylinder to put 25 cm<sup>3</sup> of 1 mol/dm<sup>3</sup> copper(II) sulfate into the polystyrene cup.
- 3 Weigh out about 6 g of zinc powder. It is in excess so does not need to be weighed accurately. Do not add this yet.
- 4 Measure the temperature of the copper(II) sulfate solution every 30 seconds for 2.5 minutes.
- 5 Add the zinc to the copper(II) sulfate in the polystyrene cup after exactly 3 minutes. Put the lid on the cup and mix gently using the stirring rod.
- 6 Measure the temperature after 3.5 minutes and every 30 seconds afterwards for a total of 6 minutes.
- 7 When you have finished, rinse out your cup and measuring cylinder and repeat the experiment at least one time. If time is short, you could use results from other groups because everyone will have used the same amounts and the same strengths of solutions.

#### Learning tips

- You can use the following equation to calculate heat energy transferred:

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

where  $Q$  is heat energy transferred (J),  $m$  is mass of liquid being heated (g),  $c$  is specific heat capacity (J/g/°C) and  $\Delta T$  is the change in temperature of the liquid (°C). This experiment uses water; the specific heat capacity of water = 4.2 J/g/°C.

- You can use the following equation to calculate number of moles,  $n$ :

$$n = \frac{\text{mass}}{M_r}$$

- You can use the following equation to calculate enthalpy change:

$$\Delta H = \frac{-Q}{n}$$

where  $\Delta H$  is enthalpy change (kJ/mol),  $Q$  is heat energy transferred (J) and  $n$  is number of moles.  $Q$  is in joules (J) so needs to be converted to kilojoules (kJ); divide  $Q$  by 1000 to use in this equation.

#### Objectives

- To investigate the temperature changes accompanying the reaction between zinc and copper(II) sulfate solution in a displacement reaction.

#### Equipment

- eye protection
- 1 mol/dm<sup>3</sup> copper(II) sulfate
- zinc powder
- 250 cm<sup>3</sup> beaker
- lid for cup
- measuring cylinder
- thermometer
- tripod
- top pan balance
- weighing boat
- stirring rod
- polystyrene cup



#### Safety notes

- Wear eye protection.
- Copper(II) sulfate is harmful and an irritant.
- Zinc powder is flammable and hazardous to the environment.

#### Practical skills

- Measure mass
- Measure volume of solution
- Measure temperature rise

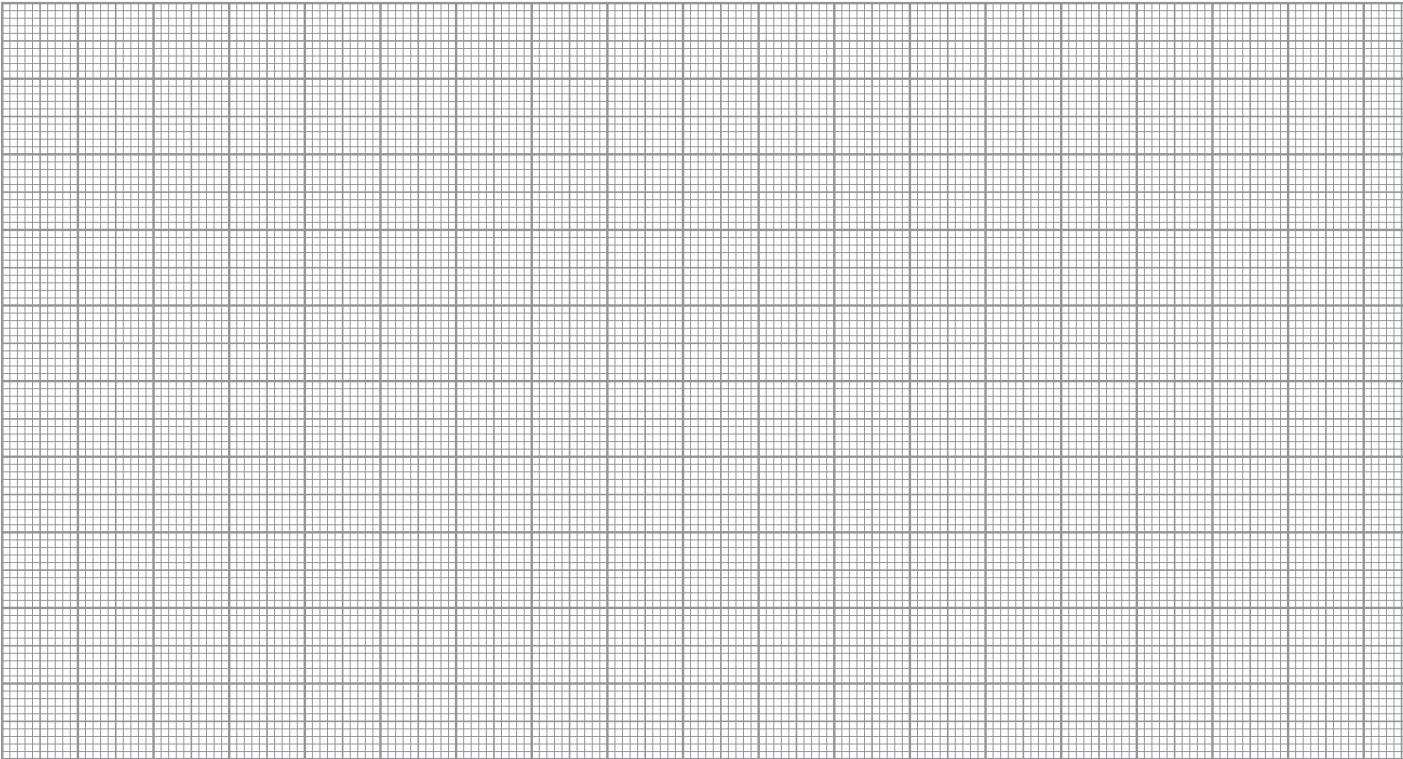
Results

- 1 Put your results into the table below. There are some extra columns either for repeated results from your experiments or for results from other groups. To calculate the mean, add up the results in a row and divide by the number of results that you used.

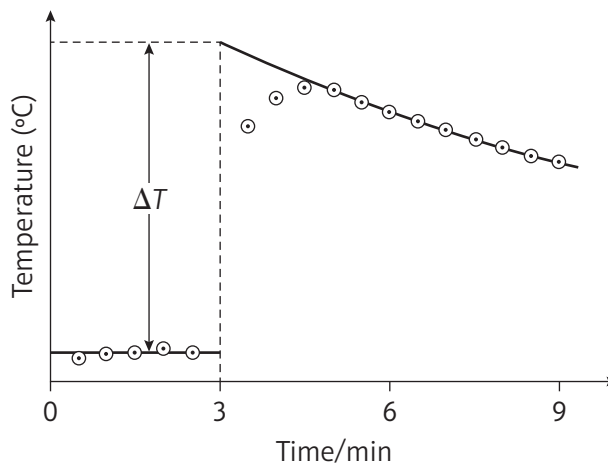
Time (seconds)	Maximum temperature recorded at each stage (°C)				
	Experiment 1	Experiment 2			Mean
0 (start)					
30					
60					
90					
120					
150					
180 (zinc added)					
210					
240					
270					
300					
360					
420					

Analysis

- 1 Draw a graph of your results. Time is on the x-axis and the mean maximum temperature (°C) is on the y-axis. Use a larger piece of graph paper if it makes it easier for you.



- 2 On your graph, find the temperature change ( $\Delta T$ ) by extrapolating the curve back to 3 minutes. The graph below shows you what your graph might look like.



- 3 Using your data, calculate the energy change ( $Q$ ) and the enthalpy change ( $\Delta H$ ) for this reaction. Determine if this reaction is exothermic or endothermic.

This reaction is .....

- 4 Compare your answer with the data book value of  $-217 \text{ kJ/mol}$  for this experiment. How did you get on?

.....

.....

.....

# CORE PRACTICAL 9: ENDOTHERMIC AND EXOTHERMIC REACTIONS

## INVESTIGATE TEMPERATURE CHANGES ACCOMPANYING SOME OF THE FOLLOWING TYPES OF CHANGE: NEUTRALISATION REACTIONS, SALTS DISSOLVING IN WATER, DISPLACEMENT REACTIONS AND COMBUSTION REACTIONS

SPECIFICATION  
REFERENCE

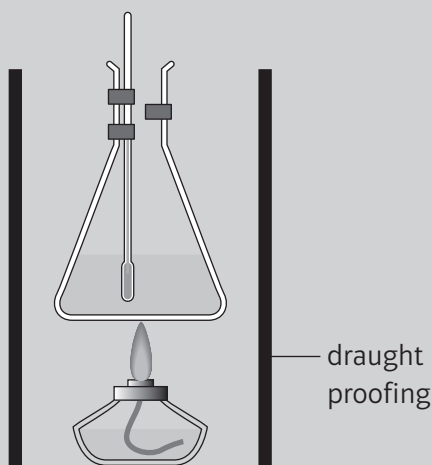
3.8

(3.1, 3.2, 3.3, 3.4)

### Part 4: Combustion

#### Method

- 1 Measure the mass of a spirit burner and cap. Record the mass in your results table.
- 2 Place the spirit burner in the centre of a heat-resistant mat.
- 3 Use a measuring cylinder to add 100 cm<sup>3</sup> of cold water to a conical flask.
- 4 Measure and record the initial temperature of the water and carefully clamp the flask above the spirit burner.
- 5 Surround the apparatus with a draught screen/insulation.



- 6 Remove the cap from the spirit burner, then light the wick and allow the water to heat up by around 40 °C.
- 7 Replace the cap on the burner and measure and record the final temperature of the water.
- 8 Measure the mass of the spirit burner and cap again and record this value.
- 9 Wash out the flask with cold water and repeat steps 1 to 8 using fresh cold water and a different alcohol.

#### Learning tip

- You can use the following equation to calculate energy released from 1 g of fuel:

$$\frac{\text{mass of water (g)} \times \text{temp. rise (}^{\circ}\text{C)} \times \text{shc of water}}{\text{mass of fuel burned (g)}} = \text{energy released from 1 g of fuel (J)}$$

- Mass of water in this experiment is 100 g.
- Specific heat capacity (shc) of water is 4.2 J/g/°C.

#### Objectives

- To investigate the temperature rise produced in a known mass of water by the combustion of the alcohols ethanol, propanol, butanol and pentanol.

#### Equipment

- eye protection
- 100 cm<sup>3</sup> beaker
- 250 cm<sup>3</sup> conical flask
- heat-resistant mat
- insulation/draught shield
- spirit burners with caps containing ethanol, propanol, butanol and pentanol
- clamp, stand and boss
- thermometer
- top pan balance

#### ! Safety notes

- Wear eye protection.
- All alcohols are flammable: handle with care and keep the tops on burners when not in use.

#### Practical skills

- Measure mass
- Measure volume of solution
- Measure temperature rise
- Measure time
- Monitor these reactions by measuring the temperature rise of water
- Make and record measurements for these reactions
- Safe use and handling of alcohols

**Results**

1 Record your results in the table below.

Alcohol	Mass of burner + lid at start (g)	Mass of burner + lid at end (g)	Temperature of water at start (°C)	Temperature of water at end (°C)
ethanol CH <sub>3</sub> CH <sub>2</sub> OH				
propan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> OH				
butan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH				
pentan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH				

2 For each alcohol:

- Calculate the change in mass of the burner and lid. This is the mass of alcohol burned. Fill in your answers in the table below.
- Calculate the change in temperature of the water. Fill in your answers in the table below.

Alcohol	Mass of alcohol burned (g)	Temperature rise in water (°C)
ethanol CH <sub>3</sub> CH <sub>2</sub> OH		
propan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> OH		
butan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH		
pentan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH		

**Analysis**

1 Calculate the energy released from 1 g of each alcohol.

Alcohol	Energy released from 1 g of fuel
ethanol CH <sub>3</sub> CH <sub>2</sub> OH	
propan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> OH	
butan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH	
pentan-1-ol CH <sub>3</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH	

**2** How does the energy released relate to the number of carbon atoms in the alcohol molecules?

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### Evaluation

**1** What are the main errors in this investigation?

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## CORE PRACTICAL 10: RATES OF REACTION

### INVESTIGATE THE EFFECT OF CHANGING THE SURFACE AREA OF MARBLE CHIPS AND OF CHANGING THE CONCENTRATION OF HYDROCHLORIC ACID ON THE RATE OF REACTION BETWEEN MARBLE CHIPS AND DILUTE HYDROCHLORIC ACID

SPECIFICATION  
REFERENCE

3.15

(3.9, 3.10, 3.11)

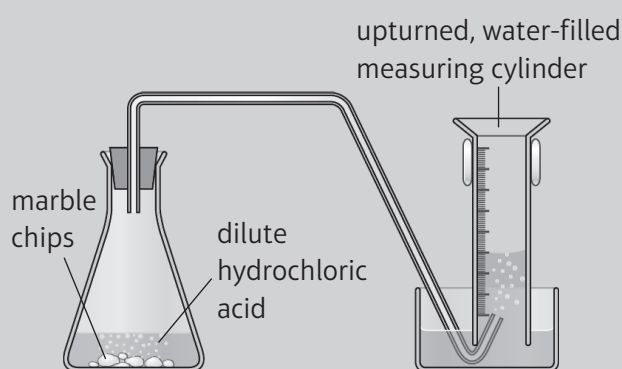
#### Introduction

The rate of a chemical reaction is how fast the reactants are changed into products. You can find the rate of a chemical reaction by looking at the product forming, or the reactant being used up. There are various factors that affect the rate of reaction. In this core practical, you are going to investigate two of them: concentration and surface area.

#### Part 1: Surface area

#### Method

- 1 Set up the apparatus as shown in the diagram.



- 2 Measure out  $40\text{ cm}^3$  of  $1\text{ mol/dm}^3$  hydrochloric acid into a conical flask.
- 3 Add 5 g of small marble chips to the flask.
- 4 Immediately stopper the flask and start the stop clock.
- 5 Note the total volume of gas produced after every 30 seconds for 5 minutes or until the reaction has finished.
- 6 Repeat steps 1 to 5 using 5 g of larger marble chips.

#### Objectives

- To investigate the effect on the rate of reaction of changing the surface area of marble chips.

#### Equipment

- eye protection
- $1\text{ mol/dm}^3$  hydrochloric acid
- marble chips (large)
- marble chips (small)
- $100\text{ cm}^3$  measuring cylinders
- conical flask
- delivery tube and bung
- stop clock
- top pan balance
- water trough or beehive shelf
- clamp
- stand

#### Safety notes

- Wear eye protection.
- Care is needed with solutions of acid. Wash off splashes immediately.

#### Practical skills

- Use appropriate apparatus to make and record measurements of mass, volume of solutions, time, temperature and volume of gas
- Use appropriate apparatus and techniques for monitoring chemical reactions, for example, a gas syringe or collecting gas over water in an upturned measuring cylinder
- Make and record observations and measurements of rate of reaction when a gas is produced
- Safe use and handling of hydrochloric acid and marble chips

**Results**

- 1 Record your results in the table below.

Time (min)	Volume of gas (cm <sup>3</sup> ) for small chips	Volume of gas (cm <sup>3</sup> ) for large chips
0		
0.5		
1.0		
1.5		
2.0		
2.5		
3.0		
3.5		
4.0		
4.5		
5.0		

**Analysis**

- 1 Use your results to draw a graph of volume of gas against time for small chips and a similar graph for large chips on the same axes. Put time on the horizontal axis and volume of gas on the vertical axis. Join the points with a smooth line. Use different coloured lines and labels for each graph. Use the graph paper on the next page.
- 2 Explain how you can tell from the graphs when the reactions were finished.

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- 3 Describe how the size of the marble chips is related to the surface area for a fixed mass of chips.

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4 Describe how increasing the surface area affects the rate of reaction.

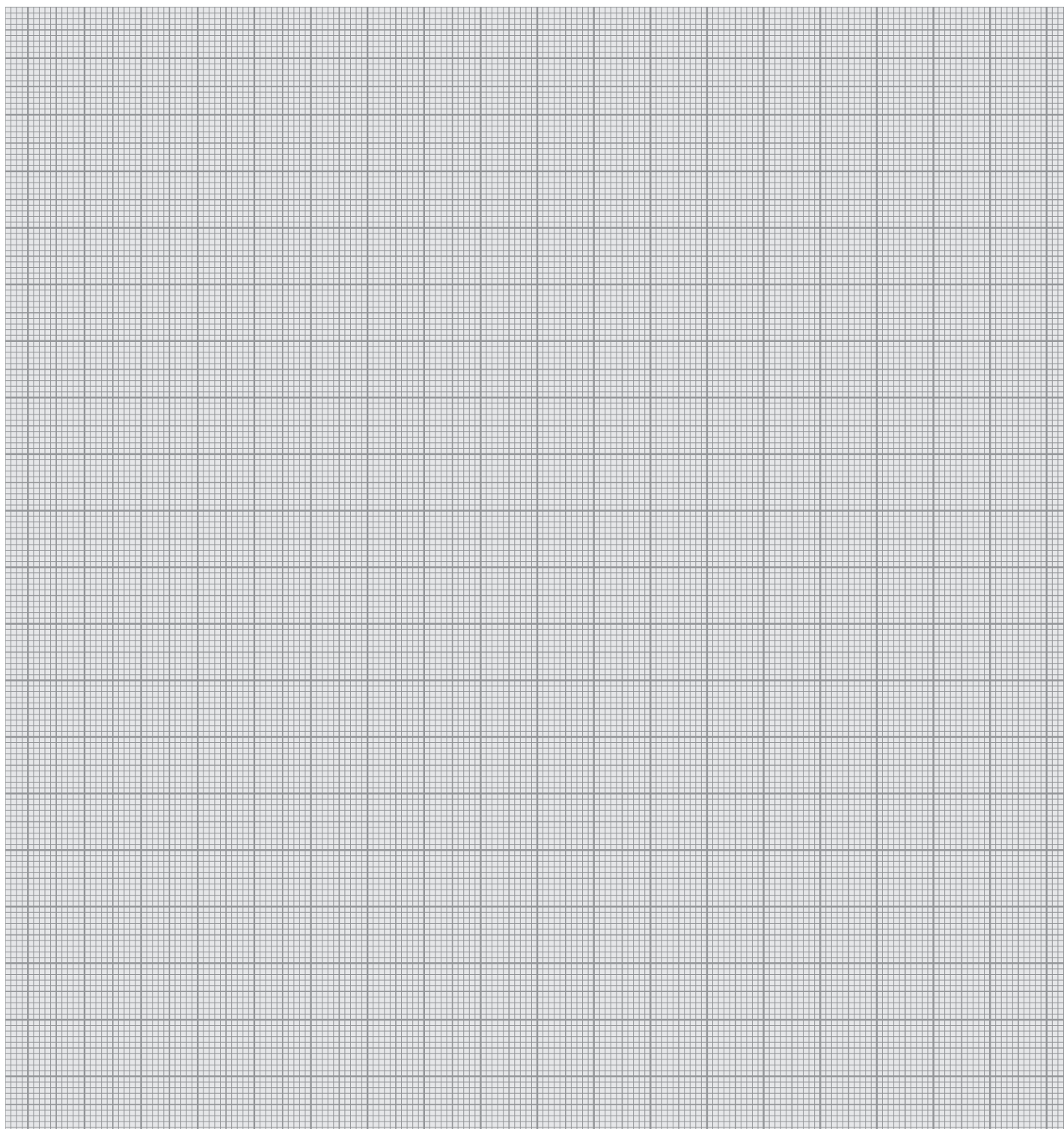
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5 Explain how your results and graphs fit with your conclusion in question 4.

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## CORE PRACTICAL 10: RATES OF REACTION

### INVESTIGATE THE EFFECT OF CHANGING THE SURFACE AREA OF MARBLE CHIPS AND OF CHANGING THE CONCENTRATION OF HYDROCHLORIC ACID ON THE RATE OF REACTION BETWEEN MARBLE CHIPS AND DILUTE HYDROCHLORIC ACID

SPECIFICATION  
REFERENCE

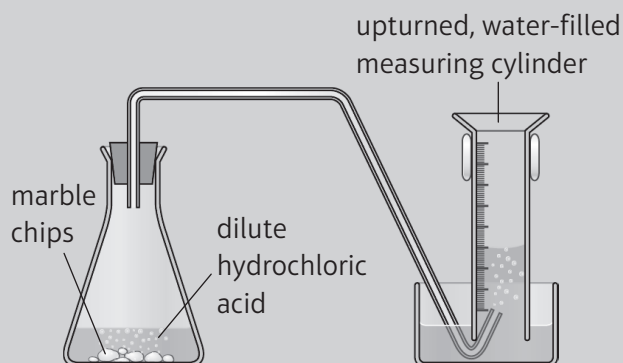
3.15

(3.9, 3.10, 3.11)

#### Part 2: Concentration

##### Method

- 1 Set up the apparatus as shown in the diagram.



- 2 Measure out  $40\text{ cm}^3$  of  $1\text{ mol/dm}^3$  hydrochloric acid into a conical flask.
- 3 Add 5 g of small marble chips to the flask.
- 4 Immediately stopper the flask and start the stop clock.
- 5 Note the total volume of gas produced after 1 minute.
- 6 Repeat steps 1 to 5 using  $0.8\text{ mol/dm}^3$ ,  $0.6\text{ mol/dm}^3$ ,  $0.4\text{ mol/dm}^3$  and  $0.2\text{ mol/dm}^3$  hydrochloric acid.

##### Results

- 1 Record your results in the table below.

Concentration of acid ( $\text{mol/dm}^3$ )	1.0	0.8	0.6	0.4	0.2
Volume of gas produced in 1 minute ( $\text{cm}^3$ )					

##### Objectives

- To investigate the effect on the rate of reaction of changing the concentration of hydrochloric acid.

##### Equipment

- eye protection
- $1\text{ mol/dm}^3$  hydrochloric acid
- $0.8\text{ mol/dm}^3$  hydrochloric acid
- $0.6\text{ mol/dm}^3$  hydrochloric acid
- $0.4\text{ mol/dm}^3$  hydrochloric acid
- $0.2\text{ mol/dm}^3$  hydrochloric acid
- marble chips (small)
- $100\text{ cm}^3$  measuring cylinders
- conical flask
- delivery tube and bung
- stop clock
- top pan balance
- water trough or beehive shelf

##### ! Safety notes

- Wear eye protection.
- Care is needed with solutions of acid. Wash off splashes immediately.

##### Practical skills

- Use appropriate apparatus to make and record measurements of mass, volume of solutions, time, temperature and volume of gas
- Use appropriate apparatus and techniques for monitoring chemical reactions, for example, a gas syringe or collecting gas over water in an upturned measuring cylinder
- Make and record observations and measurements of rate of reaction when a gas is produced
- Safe use and handling of hydrochloric acid and marble chips

**Analysis**

- 1 Use your results to draw a graph of concentration of acid against volume of gas produced in one minute. Put volume of gas on the horizontal axis and concentration of acid on the vertical axis. Use a larger piece of graph paper if it makes it easier for you.



- 2 Describe how the rate of the reaction is related to the concentration of the acid.

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- 3 Explain how your results and graph fit with your conclusion in question 2.

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**Evaluation for Parts 1 and 2**

- 1 Suggest possible sources of error in these investigations.

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- 2 Suggest possible changes to the methods of both tasks that could improve the reliability of the results.

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# CORE PRACTICAL 11: CATALYTIC DECOMPOSITION

## INVESTIGATE THE EFFECT OF DIFFERENT SOLIDS ON THE CATALYTIC DECOMPOSITION OF HYDROGEN PEROXIDE SOLUTION

SPECIFICATION  
REFERENCE

3.16

(3.9, 3.10, 3.12, 3.13)

### Introduction

Catalysts increase the rate of chemical reactions but are not used up themselves. At the end of a chemical reaction, catalysts can be recovered and reused in another reaction.

Catalysts provide an alternative pathway for a reaction at a lower activation energy.

Catalysts are specific. This means that a particular catalyst increases the rate of a specific reaction.

In this core practical, you will investigate the decomposition (the gradual process of being destroyed) of hydrogen peroxide with different catalysts.

### Method

- 1 Using a measuring cylinder, measure out 10 cm<sup>3</sup> of hydrogen peroxide. Place this into a boiling tube. Repeat, so that you have four boiling tubes containing hydrogen peroxide.
- 2 Weight out 1 g copper(II) oxide on a top pan balance. Add this to one of your boiling tubes. Record what happens in the results table.
- 3 Repeat step 2 with the three other solids (copper turnings, manganese(IV) oxide and potassium iodide).
- 4 When you have completed all four experiments, weigh a piece of filter paper. Record this value in your results table in the row for copper(II) oxide. Set up apparatus for filtering. Use the filter paper that you have just weighed. Filter the contents of the copper(II) oxide boiling tube. Use distilled water to wash out the boiling tube and add these washings to the already filtering copper(II) oxide.
- 5 Remove the filter paper from the funnel, place it somewhere warm to allow it to dry out. This may take a few days. Label your filter paper in pencil with your name and the solid that was filtered.
- 6 Repeat steps 4 and 5 with the other boiling tubes. Make sure to write down the mass of the filter paper each time.
- 7 Reweigh each piece of filter paper and add the values to the correct places in the results table.

### Learning tips

- Your teacher might provide you with other catalysts. If so, they will provide you with additional safety guidance.
- The decomposition of hydrogen peroxide to oxygen and water is exothermic and is slow at room temperature. The rate can be altered easily in a science laboratory with the addition of a catalyst.

### Objectives

- To investigate the effect of different solids on the catalytic decomposition of hydrogen peroxide solution.

### Equipment

- eye protection
- copper(II) oxide
- copper turnings
- dilute hydrogen peroxide solution
- manganese(IV) oxide
- potassium iodide
- distilled water
- 10 cm<sup>3</sup> measuring cylinder
- boiling tubes
- conical flasks
- filter funnel
- filter paper
- top pan balance

### ! Safety notes

- Wear eye protection.
- Hydrogen peroxide can cause serious eye irritation.
- Copper(II) oxide and manganese(IV) oxide are harmful.
- Potassium iodide may stain your skin.

### Practical skills

- Safe handling of chemicals, such as hydrogen peroxide, which is a strong oxidising agent
- Planning and then carrying out an investigation

**Results**

- 1 Complete the results table below. There are some blank rows in case your teacher wants you to carry out extra experiments using other solids. Subtract the mass of the filter paper from the mass of the filter paper plus dry residue to obtain the mass of the dry residue.

Solid	Observation when added to hydrogen peroxide	Mass of filter paper (g)	Mass of filter paper and dry residue (g)	Mass of dry residue (g)
copper(II) oxide				
copper turnings				
manganese(IV) oxide				
potassium iodide				

**Analysis**

- 1 With the information above, complete the table below. If the solid is a catalyst, it should increase the rate of the reaction and remain unchanged at the end.

Solid	Did the reaction rate increase?	Did the mass change?	Is it a catalyst?
copper(II) oxide			
copper turnings			
manganese(IV) oxide			
potassium iodide			

- 2 Which was the best catalyst? Looking at the vigour of reaction, determine the order from most effective to least effective. Explain how you determined this order.

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**Evaluation**

- 1** Why was it important to use the same volume of hydrogen peroxide and the same mass of catalyst in each reaction?

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- 2** Comment on the accuracy of your measurements.

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- 3** Suggest a better way of measuring any gas produced in each reaction.

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## CORE PRACTICAL 12: PREPARATION OF ETHYL ETHANOATE

### PREPARE A SAMPLE OF AN ESTER SUCH AS ETHYL ETHANOATE

SPECIFICATION  
REFERENCE

4.43

(4.39C)

#### Introduction

Ethyl ethanoate is an ester that is made by a condensation reaction between ethanol (an alcohol) and ethanoic acid (a carboxylic acid). Esters can be used as solvents, in perfumes and as flavourings. In this core practical, you will make ethyl ethanoate.

#### Method

- 1 Your teacher will place one drop of concentrated sulfuric acid into a test tube and pass the tube to you.
- 2 Add 10 drops of ethanol into the test tube containing the acid.
- 3 Your teacher will dispense (give out) 10 drops of concentrated ethanoic acid into your test tube.
- 4 Place cotton wool in the top of the test tube to prevent any vapours from escaping.
- 5 Pour approximately 100 cm<sup>3</sup> of hot water from the kettle into a large beaker. Place the test tube into the beaker. Alternatively, you may have access to a water bath. If you do use a water bath, make sure your tube is marked in some way to identify it from others.
- 6 Pour 25 cm<sup>3</sup> sodium carbonate solution into a small beaker.
- 7 After 10 minutes, remove the test tube from the hot water and pour the contents of the test tube into the sodium carbonate solution in the small beaker.
- 8 Smell the ester that has been produced by carefully wafting the vapours towards you. Your teacher will show you how to smell substances in a science laboratory.
- 9 Complete the results table for this ester. Your teacher may provide you with additional substances to make different esters.

#### Learning tips

- There are many carboxylic acids and alcohols that can be combined to make different esters. You may be able to make some more esters. Your teacher will provide you with additional safety guidance if it is needed.
- Do not use a rubber bung in the top of the test tube as the bung could fly off if the pressure increases.

#### Objectives

- To prepare a sample of an ester such as ethyl ethanoate.

#### Equipment

- eye protection
- concentrated ethanoic acid
- concentrated sulfuric acid
- ethanol
- 0.1 mol/dm<sup>3</sup> sodium carbonate
- 100 cm<sup>3</sup> beaker
- 250 cm<sup>3</sup> beaker
- cotton wool
- kettle or water bath
- pipettes
- thermometer
- water bath

#### ! Safety notes

- Wear eye protection.
- Gloves may be worn in this practical.
- Concentrated sulfuric acid is highly corrosive. Your teacher may add this to your experiment. You might be asked to work in a fume cupboard.
- Concentrated ethanoic acid is corrosive. Ethanoic acid will be dispensed by your teacher in a fume cupboard.
- Ethanol is flammable and harmful. Keep it away from sources of ignition.
- Ethyl ethanoate is flammable and an irritant. Keep it away from sources of ignition.
- When smelling the ester, hold the beaker at arm's length and waft the smell towards you. If you cannot smell it, move the beaker a little closer and repeat.
- If your teacher provides you with other substances to make different esters, they will provide you with additional safety guidance if required.

#### Practical skills

- Safe handling of alcohols
- Safe handling of carboxylic acids
- Safe handling of concentrated sulfuric acid

**Results**

1 Complete the results table below. The extra rows are for any other esters you produce.

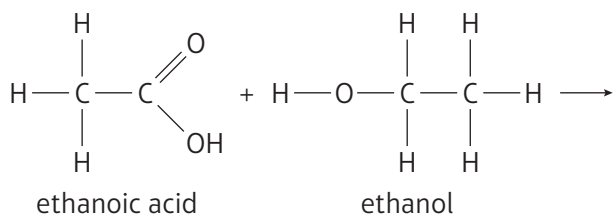
Name of alcohol	Name of carboxylic acid	Ester produced	Smell of ester

**Analysis**

1 Write a word equation for the reaction between ethanol and ethanoic acid.

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2 Complete the displayed formula equation:



3 What is the function of the sodium carbonate solution?

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4 The reaction takes at least 10 minutes. Why is the reaction slow?

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5 Why is concentrated sulfuric acid used?

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6 Why is a water bath used to heat the mixture and not a Bunsen burner?

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7 What could be the disadvantage of using a water bath to heat the mixture instead of using a Bunsen burner?

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## PRACTICAL SKILLS TABLE

Tick each technique when you cover it in your practical work.

Practical skill covered	Core Practical											
	CP1	CP2	CP3	CP4	CP5	CP6	CP7	CP8	CP9	CP10	CP11	CP12
Solve problems set in a practical context												
Apply scientific knowledge and understanding in questions with a practical context												
Devise and plan investigations, using scientific knowledge and understanding when selecting appropriate techniques												
Demonstrate or describe appropriate experimental and investigative methods, including safe and skilful practical techniques												
Make observations and measurements with appropriate precision, record these methodically and present them in appropriate ways												
Identify independent, dependent and control variables												
Use scientific knowledge and understanding to analyse and interpret data to draw conclusions from experimental activities that are consistent with the evidence												
Communicate the findings from experimental activities, using appropriate technical language, relevant calculations and graphs												
Assess the reliability of an experimental activity												
Evaluate data and methods taking into account factors that affect accuracy and validity												

### Core Practical 1

- 1A Recognise and use numbers in decimal form
- 1C Use ratios and fractions
- 2A Use an appropriate number of significant figures

### Core Practical 2

- 1A Recognise and use numbers in decimal form
- 2A Use an appropriate number of significant figures
- 3C Substitute numerical values into an algebraic equation using appropriate units
- 3D Solve a simple algebraic equation

### Core Practical 3

- 1A Recognise and use numbers in decimal form
- 1C Use ratios and fractions
- 2A Use an appropriate number of significant figures

### Core Practical 5

- 1A Recognise and use numbers in decimal form
- 1C Use ratios and fractions
- 2A Use an appropriate number of significant figures

### Core Practical 6

- 1A Recognise and use numbers in decimal form
- 1C Use ratios and fractions
- 2A Use an appropriate number of significant figures

### Core Practical 7

- 1A Recognise and use numbers in decimal form
- 1C Use ratios and fractions
- 2A Use an appropriate number of significant figures

### Core Practical 9

- 1A Recognise and use numbers in decimal form
- 2A Use an appropriate number of significant figures
- 3C Substitute numerical values into algebraic equations using appropriate units for physical quantities
- 3D Solve simple algebraic equations

### Core Practical 10

- 1A Use expressions in decimal form
- 1C Use ratios
- 2A Use an appropriate number of significant figures
- 4A Translate information between graphical and numeric form
- 4C Plot two continuous variables from experimental data
- 4E Draw and use the slope of a tangent to a curve as a measure of rate of reaction

Word	Meaning
<b>acid</b>	A solution that reacts with alkalis, turns litmus red and has a pH of less than 7.
<b>accuracy</b>	A measure of how close values are to the accepted or correct value.
<b>alcohol</b>	An homologous series of compounds that contain the –OH functional group.
<b>anhydrous</b>	Without water.
<b>anion</b>	A negatively charged ion formed by gaining electrons (usually a non-metal ion).
<b>anode</b>	Positive electrode.
<b>anomalous result</b>	A result that does not fit in with the pattern of the others.
<b>atom</b>	The smallest neutral part of an element that can take part in chemical reactions.
<b>burette</b>	A piece of apparatus used to accurately measure the volume of solution that has been added during a titration.
<b>calibrated</b>	Marked with a scale for accurate readings.
<b>calorimetry</b>	Measuring the heat given out or taken in by a chemical reaction.
<b>carboxylic acid</b>	An homologous series of compounds that contain the –COOH functional group.
<b>catalyst</b>	A substance that speeds up the rate of a reaction without itself being used up.
<b>cathode</b>	Negative electrode.
<b>cation</b>	A positively charged ion formed by losing electrons.
<b>chemical analysis</b>	Using chemical reactions or sensitive machines to identify and measure substances in a sample.
<b>chemical cell</b>	A device that produces a voltage.
<b>chromatogram</b>	The piece of paper showing the results of chromatography.
<b>chromatography</b>	A technique for separating the components of a mixture, e.g. different food colouring agents.
<b>combustion</b>	A chemical reaction in which a compound reacts with oxygen.
<b>compound</b>	A substance that can be split into simpler substances because it contains the atoms of two or more elements joined together.
<b>condensation polymerisation</b>	When monomers join together and eliminate a small molecule such as water.
<b>correlation</b>	A relationship between two variables, such that if one variable changes so does the other. This can be positive or negative.
<b>corrosion</b>	The gradual deterioration of a substance when it reacts with oxygen.
<b>crystallisation</b>	Separating the solute from a solution by evaporating the solvent.
<b>crystals</b>	Solids that are made up of a regular, repeated pattern of atoms, molecules or ions which form fixed shapes with flat surfaces and sharp edges.
<b>decomposition reaction</b>	A chemical reaction in which a compound is broken down into its elements or simpler compounds.
<b>dehydration</b>	Removal of water.
<b>dependent variable</b>	What you measure in an experiment.

Word	Meaning
<b>electrode</b>	A rod made of metal or graphite that carries the current into or out of the electrolyte.
<b>electrolysis</b>	The process in which energy transferred by a direct electrical current decomposes electrolytes.
<b>electrolyte</b>	An ionic compound that is molten or dissolved in water.
<b>endothermic</b>	A type of reaction in which energy from the surroundings is transferred to the products, e.g. photosynthesis.
<b>ester link</b>	This link is present in all polyester molecules. It consists of $\text{-COO-}$ .
<b>ethanoic acid</b>	The carboxylic acid which contains two carbon atoms and is the main acid in vinegar.
<b>evaporate</b>	When a liquid turns into a gas.
<b>exothermic</b>	A type of reaction in which energy is transferred to the surroundings from the reactants, e.g. combustion.
<b>excess</b>	Having more than enough of a reactant to react with all of something else.
<b>filtrate</b>	A solution that has passed through a filter.
<b>filtration</b>	Using a filter to separate insoluble substances from a liquid.
<b>functional group</b>	An atom or group of atoms in a molecule that is mainly responsible for the molecule's chemical reactions and properties.
<b>half equation</b>	A chemical equation written to describe an oxidation or reduction half-reaction.
<b>hazard</b>	Something that could cause harm to a user.
<b>ignite</b>	To start burning.
<b>impurity</b>	Unwanted substance found mixed into a useful substance.
<b>independent variable</b>	What you vary in an experiment.
<b>inert</b>	Does not react.
<b>insoluble</b>	A substance that cannot be dissolved in a certain liquid.
<b>mean</b>	An average calculated by adding up the values of a set of measurements and dividing by the number of measurements in the set.
<b>melting point</b>	The temperature at which a substance changes from the solid state to the liquid state when heated or from the liquid state to the solid state when cooled.
<b>mixture</b>	A substance containing two or more different substances that are not joined together.
<b>mobile phase</b>	In paper chromatography, this is the solvent that moves along the paper carrying the dissolved samples with it.
<b>monomer</b>	A small molecule that can join with other molecules like itself to form a polymer.
<b>neutral</b>	A liquid that is neither acidic nor alkaline and has a pH of 7.
<b>neutralisation</b>	A reaction in which an acid reacts with a base to produce a salt and water only.
<b>neutralise</b>	To make a solution that is neither acidic nor alkaline. During neutralisation, a base reacts with an acid, forming a salt and water.

Word	Meaning
<b>organic compound</b>	A compound that has a central framework of carbon atoms onto which hydrogen and other atoms are attached. Methane ( $\text{CH}_4$ ) is organic, but carbon dioxide is not because it contains no hydrogen atoms.
<b>oxidation</b>	A reaction in which oxygen is added to a chemical; loss of electrons by an atom or negative ion.
<b>oxidising agent</b>	A substance that causes another substance to be oxidised in an oxidation reaction.
<b>paper chromatography</b>	Chromatography carried out by spotting drops of the same samples onto paper and then allowing a solvent to move up the paper. Different components in the samples travel up the paper in the solvent at different rates.
<b>pH scale</b>	A scale going up to 14 showing acidity or alkalinity. Numbers below 7 are acids; numbers above 7 are alkalis; pH 7 is neutral.
<b>pipette</b>	A piece of apparatus which can be used in a titration to accurately measure a set volume of a solution.
<b>polyester</b>	A polymer that contains large numbers of ester links.
<b>polymer</b>	A long-chain molecule made by joining many smaller molecules (monomers) together.
<b>polymerisation</b>	A reaction in which a large number of small molecules (monomers) join together to form a long-chain molecule (polymer).
<b>precipitate</b>	An insoluble substance that is formed when two soluble substances react together in solution.
<b>precipitation reaction</b>	A reaction in which an insoluble product is formed from two soluble reactants.
<b>prediction</b>	What you think will happen in an experiment (usually given with a reason of why you think this).
<b>product</b>	A substance formed in a reaction.
<b>pure</b>	A single substance with a fixed composition that does not have anything else mixed with it.
<b>rate</b>	How quickly something happens.
<b>reactant</b>	A substance used up in a chemical reaction.
<b>redox reaction</b>	A reaction in which both oxidation and reduction occur.
<b>reduction</b>	A reaction in which a substance loses oxygen or gains electrons.
<b>resolution</b>	The smallest change that can be measured by an instrument.
<b><math>R_f</math> value</b>	The ratio of the distance travelled by a solute on a chromatogram to the distance travelled by the solvent under the same conditions.
<b>risk</b>	The chance of a hazard causing harm.
<b>risk assessment</b>	Identification of the hazards of doing an experiment and ways of reducing the risk of harm from those hazards.
<b>rusting</b>	The corrosion of iron or steel; water and oxygen must be present for rusting to occur.
<b>salt</b>	A compound formed by neutralisation of an acid by a base.
<b>saturated</b>	A molecule that contains only single bonds between the carbon atoms in its chain.

## GLOSSARY

Word	Meaning
<b>saturated solution</b>	Contains the maximum amount of solute that can dissolve in that amount of solvent at that temperature.
<b>solute</b>	A substance that dissolves in a liquid to make a solution.
<b>solution</b>	Formed when a substance has dissolved in a liquid.
<b>solvent</b>	The liquid in which a substance dissolves to make a solution.
<b>stationary phase</b>	The surface through which the solvent and dissolved substances move in chromatography.
<b>spectator ion</b>	These are ions that do not change in a reaction.
<b>tangent</b>	A straight line that touches a curve but does not cut across it.
<b>toxic</b>	Poisonous.
<b>transition metal</b>	A metal element placed in the block between groups 2 and 3 in the periodic table.
<b>variable</b>	A factor that can change.