Cambridge IGCSE®

Chemistry

Laboratory Practical Book

Bryan Earl
Doug Wilford

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Contents

Experimental skills and abilities

Experiments

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The aim of this book is to help you develop the skills and abilities needed to perform practical laboratory work in chemistry. Before you start any practical work, you need to make sure you are aware of the paramount importance of working **safely** and so this is covered first. Then we introduce the apparatus and measuring techniques that you will use most often.

This is followed by a section on how to make and record measurements accurately. Methods for handling the observations and data you have collected will then be described.

Finally, we discuss how to plan, carry out and evaluate an investigation. You should then be ready to work successfully through the experiments and laboratory activities that follow.

Safety

In all your practical exercises and investigations, materials will be used which, although familiar in many cases, are of a potentially hazardous nature, and appropriate care and precautions should be taken. If in doubt, ask your teacher to make the final decision depending on the circumstances at the time. Also in certain circumstances disposable gloves and fume cupboards will be required. Eye protection should be worn at all times.

Here are a few simple precautions to help ensure your safety when carrying out experiments in the laboratory.

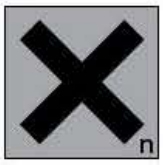
- **Always wear shoes** – to protect your feet if a heavy weight should fall on them.
- **Hot liquids and solids** – set in a safe position where they will not be accidentally knocked over; handle with caution to avoid burns.
- **Toxic materials** – materials such as mercury are toxic; take care not to allow a mercury thermometer to roll onto the floor and break. Other toxic substances such as bromine, chlorine and lead may be present and appropriate information will be given to you on the worksheet and/or by your teacher.
- **Tie back long hair** – to prevent it being caught in a flame.
- **Personal belongings** – leave in a sensible place so that no one will trip over them!
- **Protect eyes and skin from contact with corrosive and harmful chemicals** – any reagent used for any of the experiments in this book must be treated with caution. **Ask for your teacher's advice before handling them**. Sodium hydroxide, hydrochloric acid, sulfuric acid, iodine, and many other chemicals suggested in this book must be handled with care. Some are also flammable, such as alcohol/ethanol. Alcohol, hexane and other solvents, as well as a variety of gases produced in some of the experiments are flammable.

- **Bunsen flames and flammable liquids** – use the safety flame, or turn the Bunsen burner off when not in use. Make sure the Bunsen flame is out before handling flammable liquids, such as alcohol/ethanol hexane, cyclohexane and some other solvents. An alternative may be to heat water using a kettle.

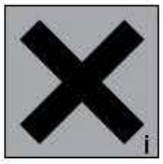
International hazard warning symbols

You will need to be familiar with these symbols when undertaking practical experiments in the laboratory. Make sure you fully understand the hazards indicated (Figure 1a).

Corrosive
These substances attack or destroy living tissues, including eyes and skin.



Harmful/noxious
These substances are similar to toxic substances but less dangerous.



Irritant
These substances are not corrosive but can cause reddening or blistering of the skin.

Explosive
These substance, If treated incorrectly, they may explode. Explosive substances must be handled very carefully.

Oxidising
These substances provide oxygen, which allows other materials to burn more fiercely.

Toxic
These substances can cause death.

Highly flammable
These substances can easily catch fire.

Figure 1a International hazard warning symbols

NOTE: a new system for labelling hazards is being introduced between 2010 and 2015 and, in due course, you will need to become familiar with these new symbols in Figure 1b:

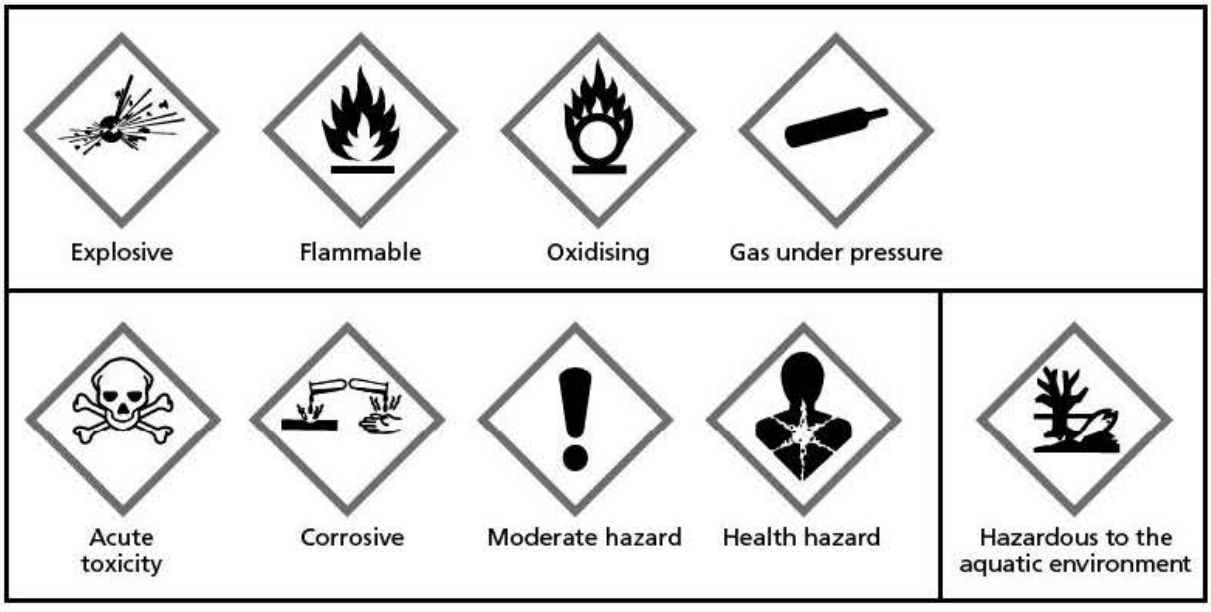


Figure 1b New hazard labelling system

The bottles of chemicals you will be using will have one or more of these symbols on them like the one shown in Figure 2. Remember, laboratories are safe places if you work carefully, tidily and safely!

Figure 2 Sulfuric acid is corrosive.

Special note to teachers

In the suggested practical exercises, materials are used which, although familiar in many cases, are of a potentially hazardous nature, and appropriate care and precautions should be taken. We believe that the experiments can be carried out safely in school laboratories. However, it is the responsibility of the teacher to make the final decision depending on the circumstances at the time. Eye protection should be worn at all times. In certain cases, disposable gloves and fume cupboards will be required. Teachers must ensure that they follow the safety guidelines set down by their employers, and a risk assessment must be completed for *any* experiment that is carried out. Teachers should draw students' attention to the hazards involved in the particular exercise to be performed. The hazards are shown within the 'Safety' section of the individual practicals.

It is recognised that, in some cases, there may not be sufficient apparatus to carry out a class practical. If there is insufficient apparatus, perhaps the experiment can be carried out as a demonstration (consider using some pupil assistance).

In some cases, certain pieces of apparatus may not be available. If possible, use alternatives, **as long as the safety precautions are not overlooked**. For example, burettes, pipettes and gas syringes could be replaced by measuring cylinders of suitable sizes. If you substitute equipment, attention should be drawn to the accuracy of using the alternative – usually much lower than using that suggested.

It is recommended that certain experiments are carried out in a fume cupboard. If this is not available, an alternative maybe to use a well-ventilated room for the experiment or take this experiment outside. However, where relevant, it must be pointed out that the fumes produced are noxious and may cause an asthmatic attack. All the same safety precautions should still be followed. Plasticine and cocktail sticks can be used as an alternative to molecular models. Marshmallows can also be used.

Using and organising techniques, apparatus and materials

In an experiment, you will first have to decide on the measurements to be made and then collect together the apparatus and materials required. The quantities you will need to measure most often in laboratory work are **mass**, **length** and **time**.

- What apparatus should you use to measure each of these?
- Which measuring device is most suitable for the task in hand?
- How do you use the device correctly?

Balances

A **balance** is used to measure the mass of an object. There are several types available.

- In a beam balance the unknown mass is placed in one pan and balanced against known masses in the other pan.
- In a lever balance a system of levers acts against the mass when it is placed in the pan.
- A digital top-pan balance, which gives a direct reading of the mass placed on the pan, is shown in Figure 3.
- The unit of mass is the kilogram (kg).
- The gram (g) is one-thousandth of a kilogram: $1\,g = 1/1000\,kg = 10^{-3}\,kg = 0.001\,kg$

Figure 3 A digital top-pan balance

How accurately do your scales measure?

- A beam balance is accurate to the size of the smallest mass that tilts the balanced beam.
- A digital top-pan balance is accurate to the size of the smallest mass that can be measured on the scale setting you are using, probably 0.01 g.
- When using an electronic balance you should wait until the reading is steady before taking it.

Ruler and vernier scales

- The unit of length is the metre (m).
- Multiples are:
  - 1 decimetre (dm) = 10^{-1} m
  - 1 centimetre (cm) = 10^{-2} m
  - 1 millimetre (mm) = 10^{-3} m
  - 1 micrometre (μm) = 10^{-6} m
  - 1 kilometre (km) = 1000 m
- A **ruler** is often used to measure lengths in the centimetre range.
- The correct way to measure with a ruler is shown in Figure 4, with the ruler placed as close to the object as possible.

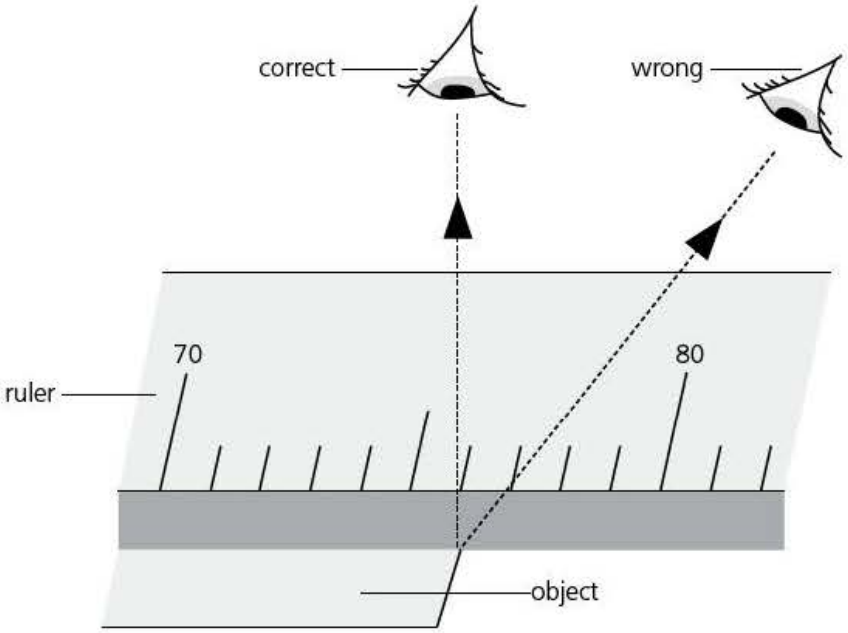


Figure 4 Using a ruler: The reading is 76 mm or 7.6 cm. Your eye must be directly above the mark on the scale or the thickness of the ruler causes parallax errors. The accuracy of the measurement will be about 1 mm.

- Some instruments, such as barometers and microscopes, have a **vernier scale** (Figure 5) to enable small lengths to be measured, usually to 0.1 mm.
- One end of the length to be measured is made to coincide with the zero of the millimetre scale and the other end with the zero of the vernier scale.

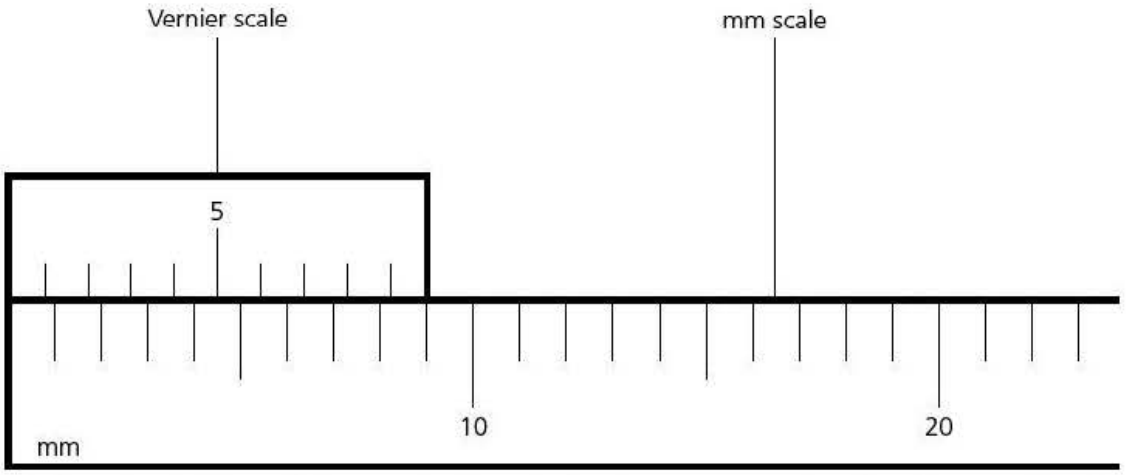


Figure 5a A vernier scale is a small sliding scale that is 9mm long but divided into 10 equal divisions: 1 vernier division = 9/10mm = 0.9mm = 0.09cm

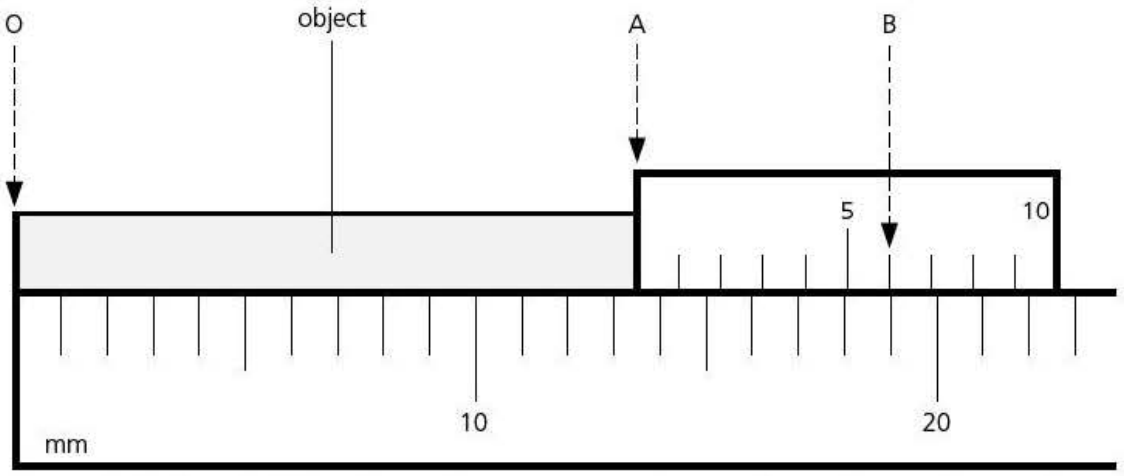


Figure 5b The length of the object is between 1.3 and 1.4cm. The reading to the next decimal place is determined by finding the vernier mark that exactly lines up with a mark on the millimetre scale. Here it is the 6th mark and so the length of the object is 1.36cm since:

OA = OB – AB
= 1.90cm – (6 vernier divisions)
= 1.90 – (6 × 0.09)cm
= 1.90 – 0.54cm
= 1.36cm

Clocks and timers

- Clocks, watches and timers can be used to measure time intervals. In an experiment it is important to choose the correct timing device for the required measurement.
- The unit of time is seconds (s).
- A stopwatch will be sufficient if a time in minutes or seconds is to be measured, but if times of less than a second are to be determined then a digital timer is necessary (Figure 6).

How accurate are your timings?

Figure 6 This stopwatch can be used to measure the time passed in a chemical reaction.

- When using a stopwatch, human reaction times may influence the reading, and an accuracy of about 0.5 s is the best that is likely to be achieved.
- For time intervals of the order of seconds, a more accurate result will be obtained by measuring longer time intervals – for example, time reaction rate over 60 seconds rather than over 15 seconds.
- To measure very short time intervals, a digital timer that can be triggered to start and stop by an electronic signal from a microphone, photogate or mechanical switch is useful.

Changing measurements

- Take readings more frequently if values are changing rapidly.
- It will often be helpful to work with a partner who watches the timer and calls out when to take a reading.
- Pressing the lap-timer facility on the stopwatch at the moment you take a reading freezes the time display for a few seconds and will enable you to record a more accurate time measurement.
- For rapidly changing measurements, it may be necessary to use a datalogger and computer.

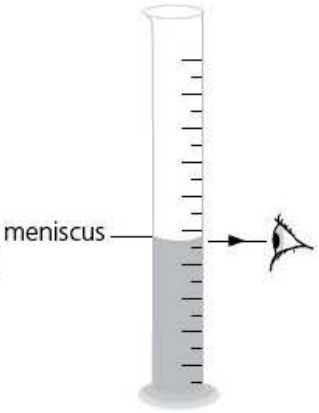


Figure 7 When making a reading the measuring cylinder should be vertical and your eye should be level with the bottom of the curved liquid surface – the meniscus.

Measuring cylinders

- The volume of a liquid can be obtained by pouring it into a measuring cylinder (Figure 7).
- Measuring cylinders are often marked in millilitres (ml) where 1 millilitre = 1 cm^3.
- The accuracy of the reading will be 1 cm^3.
- Note that 1 litre = 1000 cm^3 = 1 dm^3.

Observing, measuring and recording

Having collected together and familiarised yourself with the equipment and materials needed for an experiment, you are now ready to start making some observations and measurements.

- It will be helpful at this stage to draw a clearly labelled diagram of the experimental set-up.
- You should also record any difficulties encountered in carrying out the experiment and any precautions taken to achieve accuracy in your measurements.
- Do not dismantle the equipment until you have completed the analysis of your results and are sure you will not have to repeat any measurements!
- What degree of accuracy will your measurements have?

- How many significant figures will your data have? What are the colour changes and other observations?
- How will you record your results?

Degree of accuracy

Make a list of the apparatus you use in an experiment and record the smallest division of the scale of each measuring device. This will be the accuracy of your measurements.

- For example, the smallest division on a metre rule is 1 mm, so the accuracy of any length measured with the rule will be 1 mm.
- The degree of accuracy will be greater, the longer the length measured:
  - For a measured length of 1 m = 1000 mm, the degree of accuracy will be 1 part in 1000.
  - For a measured length of 1 cm = 10 mm, the degree of accuracy will be 1 part in 10.
  - Similarly, if the gradations on a thermometer are at 1 °C intervals, the accuracy of a temperature reading will be 1°C.

Significant figures

Every measurement of a quantity is an attempt to find its true value and is subject to errors arising from the limitations of the apparatus and the experimental procedure.

- The number of figures given for a measurement, called **significant figures**, indicates how accurate we think it is. More figures should not be given than are justified.
- For example, a measurement of 6.7 has two significant figures. The measurement 0.235 has three significant figures, the 2 being most significant and the 5, which we are least sure about (since it could be 4 or 6), being the least significant.
- When doing calculations, your answer should have the same number of significant figures as the measurements used in the calculation. For example, if your calculator gives an answer of 1.23578, this would be 1.2 if your measurements have two significant figures and 1.24 if your measurements have three significant figures.
- Note that when rounding the least significant figure you look at the following figure. If that is less than 5, you round down (1.23 becomes 1.2) but if it is 5 or above, you round up (1.235 becomes 1.24).
- If a number is expressed in standard notation, the number of significant figures is the number of digits before the power of 10. For example, 6.24×10^2 has three significant figures.
- If values with different numbers of significant figures are used to calculate a quantity, quote your answer to the smallest number of significant figures.

Systematic errors

Figure 8 shows part of a ruler used to measure the height of a point P above the bench.

- The ruler has a space of length x before the zero of the scale.
- The height of the point P = scale reading + x = 5.9 + x.
- By itself, the scale reading is not equal to the height of P; it is too small by the amount x.

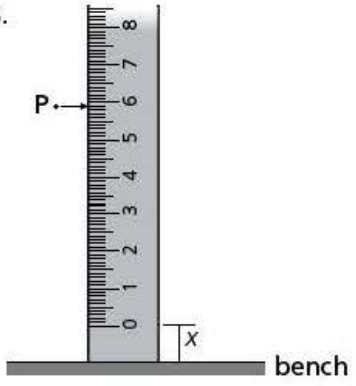
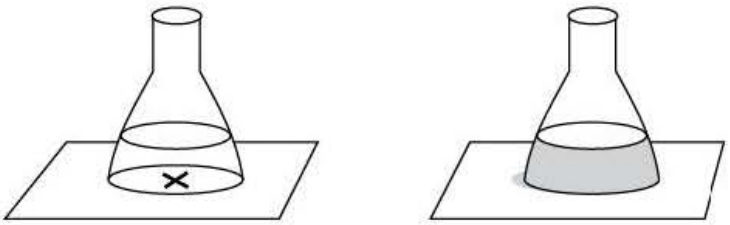


Figure 8

- An error of this type is called a **systematic error** because it is introduced by the system.
- A half-metre ruler does not have a systematic error because its zero is at the end of the rule.
- When using a rule to measure a height, the rule must be held so that it is vertical. If it is at an angle to the vertical, then a systematic error will be introduced.
- Check for any zero error when using a measuring device. If it cannot be eliminated, correct your readings by adding or subtracting the zero error. When carrying out a titration, remember to consider errors associated with both the burette and pipette/measuring cylinder.

Tables

If several measurements of a quantity are being made, draw up a **table** in which to record your results.



At the start the solutions are clear.

Eventually, you can't see through the sulfur in the solution.

Figure 9

- Use the column headings, or start of rows, to **name** the measurement and state its unit.

  For example, the reaction of sodium thiosulfate and hydrochloric acid was observed when varying the temperature by recording the time it took for the cross to disappear (Figure 9). The results were represented in a table (Table 1).
- Repeat the measurement of each observation if possible and record the values in your table. If repeat measurements for the same quantity are significantly different, take a third reading. Calculate an average value from your readings.
- Numerical values should be given to the number of significant figures appropriate to the measuring device.

Table 1 A data table for the study of the variation of rate of reaction between sodium thiosulfate and hydrochloric acid with temperature

Temperature (°C)	Time taken for cross to disappear (seconds)			
	1	2	3	Average
5	290	291	289	290
10	196	194	196	195
15	130	128	128	128
20	95	93	94	94
25	67	66	67	67
30	53	53	53	53
35	39	38	37	38
40	28	30	28	29
45	21	20	19	20
50	15	14	13	14
55	11	11	11	11
60	7	7	6	7

Handling experimental observations and data

Now that you have collected your measurements, you will need to process them. Perhaps there are calculations to be made or you may decide to draw a graph of your results.

Then you can summarise what you have learnt from the experiment, discuss sources of experimental error and draw some conclusions from the investigation.

- What is the best way to process your results?
- Are there some inconsistent measurements to be dealt with?
- What experimental errors are there?
- What conclusions, generalisations or patterns can you draw?

Calculations

You may have to produce an average value to process your results.

The **average** is found by taking all the measurements you have made, adding them together and dividing by the number of measurements taken. When calculating an average, only use concordant values, i.e. values within 0.2.

For example, if you timed a reaction to a particular point on eight occasions and the times were 42.1 s, 44.3 s, 42.2 s, 41.7 s, 42.0 s, 45.3 s, 42.1 s, 44.0 s then:

$$\text{Mean} = \frac{42.1 + 42.2 + 42.0 + 42.1}{4} = 42.1\,\text{s}$$

- The value has been given to three significant figures because that was the accuracy of the individual measurements.

Graphs

Graphs can be useful in finding the relationship between two quantities. Graphs are a pictorial way of looking at data from a table. Certainly, trends can be observed and data that is out of sequence with the rest can be seen. When drawing graphs, ignore anomalous data. **Anomalous data** are readings which fall outside the normal, or expected, range of measurements. In the study of the loss of mass of calcium carbonate lumps when heated with time the anomalous results are clearly seen (Figure 10).

- You will need about six data points taken over as large a range as possible to plot a graph.
- Choose scales that make it easy to plot the points and use as much of the graph paper as possible.
- Make sure you label each axis of the graph with the name and unit of the quantity being plotted.
- Mark the data points clearly with a dot in a circle or a cross using a sharp pencil.
- Join up your points with a smooth line or curve.

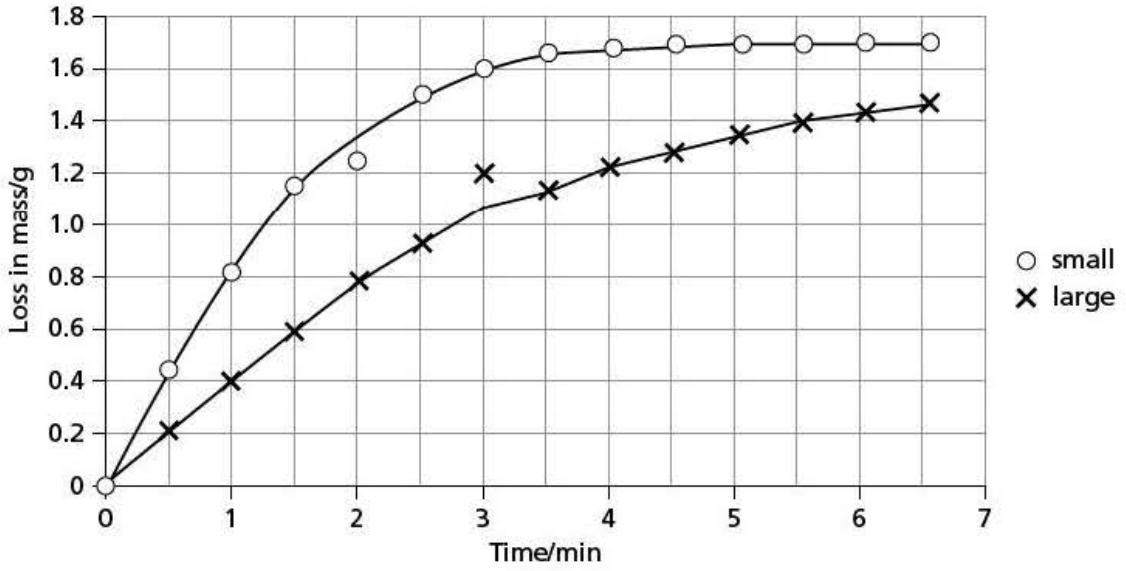


Figure 10

The anomalous points are clearly seen when smooth curves are drawn.

- Always draw straight lines (Figure 11) and curves of best fit when producing graphs.
- Note that the line must go through the origin for the quantities to be proportional.
- If a straight-line graph does not go through the origin, one can only say that there is a linear dependence between *y* and *x*.

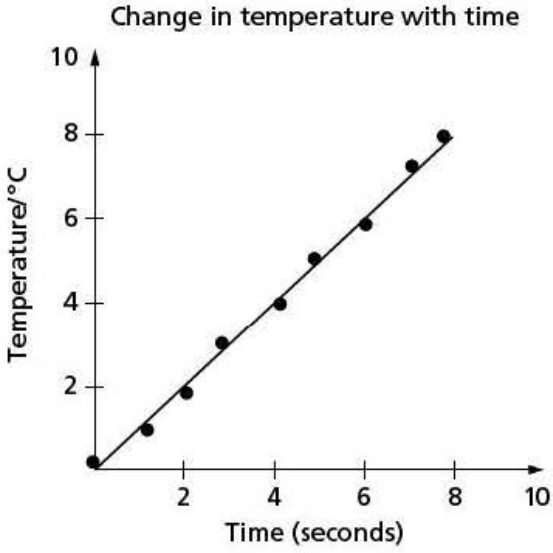


Figure 11 Line of best fit has been plotted to show an accurate trend for the data plotted.

Errors

- In practice, points plotted on a graph from actual measurements may not lie exactly on a straight line or curve of a graph due to experimental errors.
- The 'best straight line' is then drawn through the points, as mentioned earlier.
- If possible, repeat any anomalous measurements to check that they have been recorded properly or try to identify the reason for the anomaly.

Conclusions

Once you have analysed your experimental results, summarise your conclusions clearly and relate them to the aim of the experiment.

- State whether a hypothesis has been verified. If your results do not, or only partially, support a hypothesis, suggest reasons why.
- If a numerical value has been obtained, state it to the correct number of significant figures. Compare your results with known values, if available, and suggest reasons for any differences.
- State any relationships discovered or confirmed between the variables you have investigated.
- Mention any patterns or trends in the data.

- Identify and comment on sources of error in the experiment. For example, it may be very difficult to eliminate all heat losses to the environment in a heat experiment such as the heat of combustion of a reaction; if that is the case, say so. Mention any sources of systematic error in the experiment.

Planning, carrying out and evaluating investigations

- Before you start an experiment, it is important to define an aim and produce a logical and safe plan for the investigation.
- You should identify the variables in the investigation and decide which ones to manipulate and which ones you should try to keep constant. The variable that is manipulated or changed is known as the **independent variable**. The variable that responds and is measured is known as the **dependent variable**. To discover the relationship between variables, you should change only one variable at a time.
- Once you know what you need to measure, you can decide on the apparatus and materials to be used. You should ensure that your measuring devices have sufficient accuracy for the job required.
- Before you start the experiment, familiarise yourself with how to use the apparatus and develop a plan of work. It will be helpful to decide how to record your results; draw up tables in which to record your measurements if appropriate.
- Describe how you carried out the experiment under 'Method' in your laboratory notes. It is useful to include a sketch of the experimental set-up here for future reference.
- When you have obtained your results, manipulate data, draw graphs and carry out the calculations needed to fulfil the aims of the experiment.
- Then, analyse your results and clearly state your conclusions from the investigation.
- Finally, evaluate the experiment and discuss how it could be improved. Could some things have been done better? If so, suggest changes or modifications that could be made to the procedure or the equipment used in the investigation. It helps if you draw a diagram of your apparatus (Figure 12, for example).

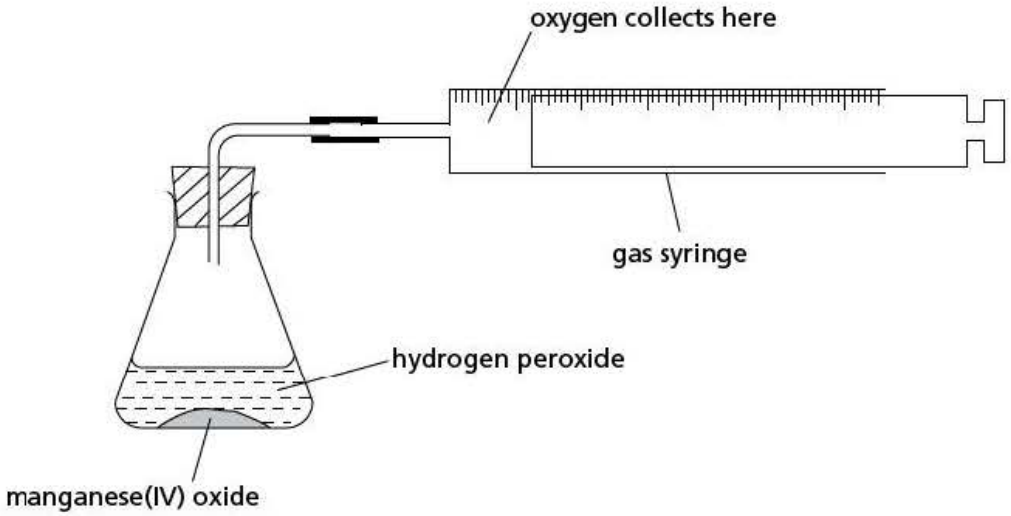


Figure 12 This apparatus was used to obtain the volume of oxygen produced by varying the concentration of hydrogen peroxide using manganese(IV) oxide as a catalyst.

Questions

1 What measuring device would you use to obtain values for the following?

(a) The volume of liquid in a coffee mug

.....

(b) The mass of an apple

.....

(c) The length of the pendulum of a grandfather clock

.....

(d) The temperature of a cup of tea

.....

(e) The time taken to run up 20 stairs

.....

(f) The time taken by an apple to fall one metre

.....

(g) The dimensions of a textbook

.....

2 How would you obtain a value for the following?

(a) The average time for one oscillation of the pendulum of a clock

.....

.....

(b) The average mass of a pin

.....

.....

→

3 Complete the table below by stating the typical accuracy of each of the measuring devices listed.

Table 2

Device	Accuracy
metre ruler	
vernier scale	
stopwatch	
digital timer	
digital balance	
liquid in glass thermometer	
100 cm^3 measuring cylinder	

4 Write the number 9.753864 to:

(a) 3 significant figures .....

(b) 2 significant figures .....

(c) 1 significant figure. .....

5 The measurements in the table below were obtained for the rate of reaction at varying temperatures.

Table 3

Temperature/°C	Rate of reaction
0.0	0.0
5.0	2.0
10.0	4.0
15.0	6.0
20.0	8.0
25.0	10.0

(a) (i) State the variables being measured.

.....

(ii) Name a variable that must be kept fixed.

.....

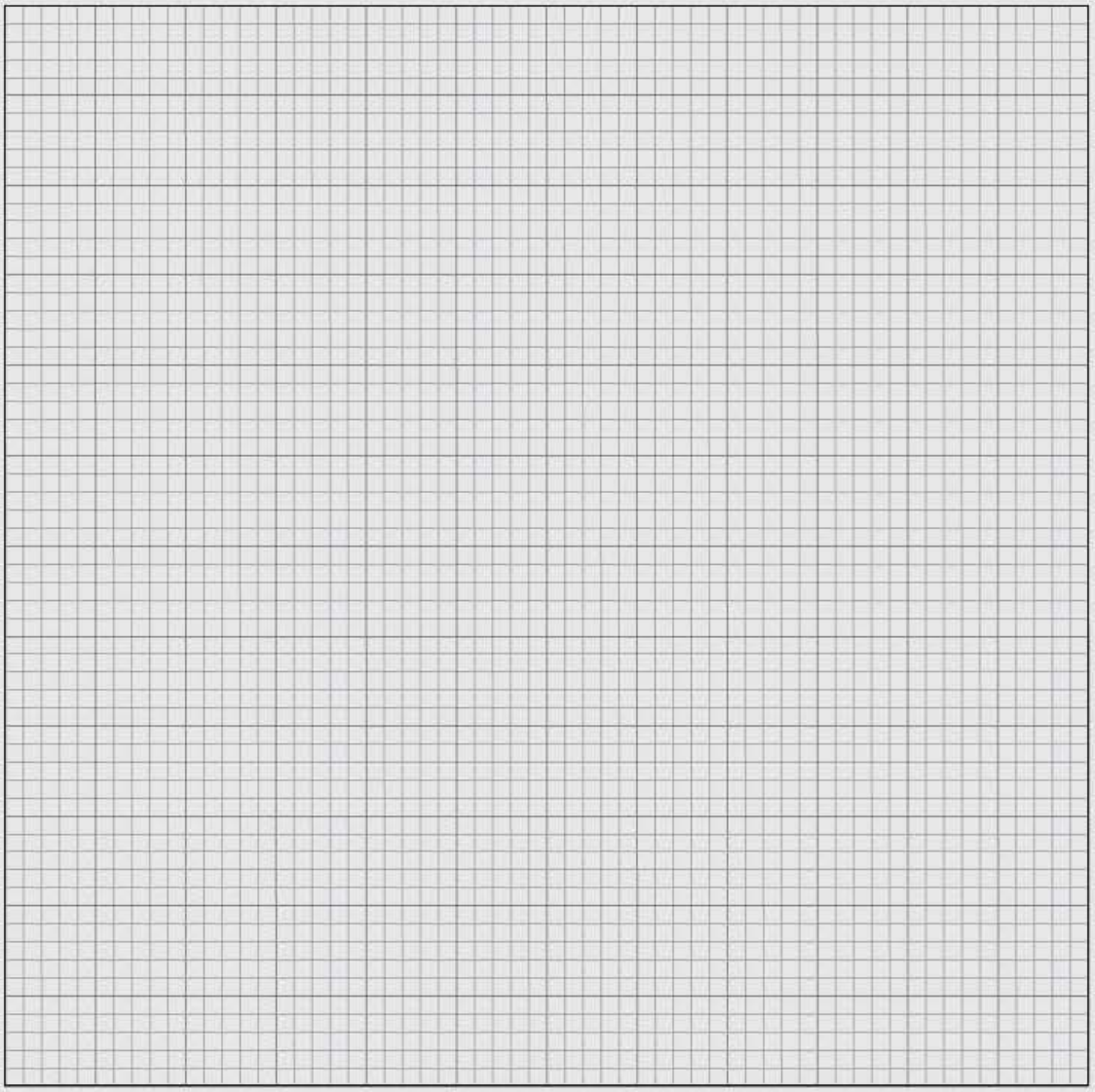
(iii) Complete the following table.

Table 4

Manipulated variable	Fixed variable	Responding variable

→

(b) Using data in Table 3, plot a graph of temperature on the horizontal or *x*-axis and rate of reaction on the vertical or *y*-axis.



(c) What can you conclude about the relationship between temperature and rate of reaction over the temperature range given?

.....

.....

.....

1 The particulate nature of matter

1.1 Rate of diffusion of ammonia and hydrogen chloride

Aim

To determine how the mass of gas particles affects the rate of diffusion of gases.

Theory

In this experiment hydrogen chloride gas, HCl(g), will react with ammonia gas, $NH_3(g)$, to form the white solid ammonium chloride. Hydrogen chloride gas is released from concentrated hydrochloric acid.

ammonia + hydrogen chloride → ammonium chloride

$NH_3(g)$ + HCl(g) → $NH_4Cl(s)$

This is a teacher demonstration and it is recommended that it is performed in a fume cupboard. It is also recommended that a sheet of black paper be placed behind the demonstration so that the formation of the white solid ammonium chloride can be observed clearly.

Apparatus and chemicals

- ☐ eye protection
- ☐ access to a fume cupboard
- ☐ protective gloves
- ☐ stopwatch
- ☐ a length of glass tubing of recommended length 50–100 cm with internal diameter of 2–3 cm
- ☐ 2 × retort stands complete with bosses and clamps
- ☐ 2 × cotton wool buds
- ☐ 2 × bungs to fit the ends of the glass tube
- ☐ concentrated hydrochloric acid
- ☐ concentrated ammonia solution

Safety!

It is very important that the glass tubing is clean. The concentrated hydrochloric acid and concentrated ammonia solution should be supplied in small bottles. Care should be taken when opening the bottle of the ammonia solution because pressure can build up inside the bottle. Do not keep the bottle of ammonia open for a long period of time.

Concentrated ammonia solution — corrosive, dangerous for the environment

Ammonia gas — toxic, dangerous for the environment, use in a well-ventilated laboratory or fume cupboard

Concentrated hydrochloric acid, a few drops — corrosive

Ammonium chloride — harmful

Disposal: Leave the tube in the fume cupboard for an hour. Then wash with water.

Procedure 1

Throughout, the teacher/demonstrator must wear eye protection and safety gloves.

1 While working in a fume cupboard, clamp the glass tubing at both ends. Ensure that the glass tubing lies horizontally.

2 Open the bottle of the concentrated ammonia solution and leave it in the fume cupboard. Point the bottle away from yourself and the students. Then open the bottle of concentrated hydrochloric acid, again in the fume cupboard, and hold the stopper near the mouth of the bottle of ammonia solution. Note the white clouds of ammonium chloride that are produced.

Procedure 2

Throughout, the teacher/demonstrator must wear eye protection and safety gloves.

1 Put the end of one of the cotton buds into the concentrated ammonia solution. Quickly repeat this procedure with the second cotton bud and the concentrated hydrochloric acid. Push the buds into opposite ends of the tube and start a stopwatch. Replace the tops on the bottles of the ammonia and hydrochloric acid solutions.

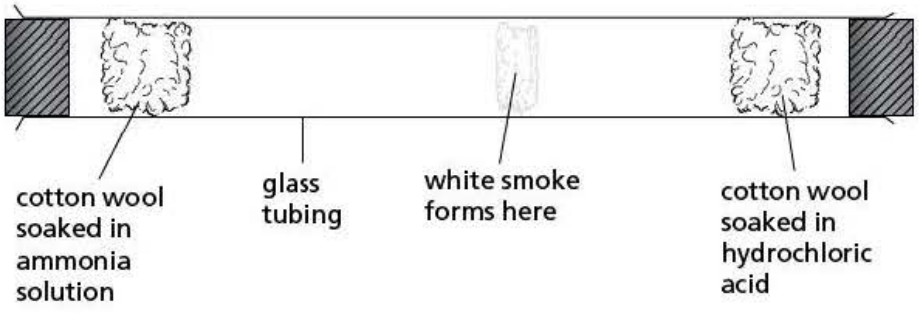


Figure 1 Diagram of the apparatus

2 Watch the ring of white solid forming in the tube, and stop the stopwatch when it is initially observed.

3 Measure the distance each gas has travelled from its end of the tube to where the white solid formed.

Method

1 Why is it important that the glass tube is clamped horizontally before the experiment begins? [1]

.....

2 Why are the solutions which produce the gases placed at the end of a glass tube, rather than just placing them at opposite sides of a closed fume cupboard? [2]

.....

3 Why is it important that the experiment takes place in a fume cupboard? [2]

.....

Results and calculations

Time taken for the white solid to form: ..... seconds

Distance travelled by ammonia gas: ..... cm

Distance travelled by hydrogen chloride gas: ..... cm

Rate of diffusion of the ammonia gas: ..... cm/s

Rate of diffusion of the hydrogen chloride gas: ..... cm/s

Conclusions

1 How will the length of the tube affect the time of formation of the white solid? [2]

.....

.....

.....

2 How would the diameter of the tube affect the time? [1]

.....

3 What other variables would affect the time of formation of the white solid? [2]

.....

.....

.....

.....

4 Does the ring form closer to the ammonia or hydrogen chloride end of the tube? [1]

.....

5 Why do you think that the white solid forms closer to this end of the tube? [2]

.....

.....

.....

Evaluation

Discuss how the experiment could be improved to give more reliable results. [1]

.....

.....

.....

Extension

Do you think that the rate of diffusion of the ammonia and hydrogen chloride gases you have calculated is the same as the speed of the particles (molecules) of the gases? Explain your answer. [3]

.....

.....

.....

.....

.....

1.2 Reaction of potassium iodide with lead(II) nitrate

Aim

To discover how the speed at which a reaction occurs between solids compared to the same reaction using solutions.

Theory

When potassium iodide and lead(II) nitrate react together, one of the products formed is lead(II) iodide, which is a yellow coloured compound.

potassium iodide + lead(II) nitrate → potassium nitrate + lead(II) iodide

$2KI + Pb(NO_3)_2 \rightarrow 2KNO_3 + PbI_2$

In this experiment you will perform this reaction with both solid reactants and then again with solutions.

Apparatus and chemicals

- ☐ eye protection
- ☐ boiling tube
- ☐ 2 × small spatulas
- ☐ 2 × dropping pipettes
- ☐ rubber bung for boiling tube
- ☐ test tube
- ☐ boiling-tube rack
- ☐ 2 × 1 cm^3 dropping pipettes
- ☐ lead(II) nitrate solid and 0.1 mol/dm^3 solution
- ☐ potassium iodide solid and 0.1 mol/dm^3 solution

Safety!

Potassium iodide (solid and 0.1 mol/dm^3 solution) – low hazard

Lead(II) nitrate (solid and 0.1 mol/dm^3 solution) – toxic, dangerous for the environment

Lead(II) iodide (solid) – toxic, dangerous for the environment

Potassium nitrate (solution) – oxidising and extremely combustible with reactive metals

Disposal: solids in a bottle – labelled 'Lead waste'
solutions – filter; then dry residue and place in lead waste bottle

Procedure 1

Throughout the practical the student should wear eye protection.

1 Put a spatula full of solid potassium iodide into the boiling tube.
2 Using the other spatula, place a spatula full of solid lead(II) nitrate into the same boiling tube.
3 Push the rubber bung into the boiling tube.
4 Taking care that you are not near any solid surface, shake the boiling tube for a few minutes.
5 After shaking, place the boiling tube in the boiling-tube rack.
6 Write down your observations in the Results section below.

Procedure 2

Throughout the practical the student should wear eye protection.

1 Using one of the dropping pipettes, place 1 cm^3 of potassium iodide solution into the test tube.
2 Using the other dropping pipette, add 1 cm^3 of lead(II) nitrate to the same test tube.
3 Place the test tube into the rack.
4 After five minutes write down your observations in the Results section below.

Method

1 In Procedure 1 why was the boiling tube shaken and not simply placed into the rack after the two solids had been added to it? [1]

.....

.....

.....

2 In Procedure 1 why was it important to check that you were not near any solid surfaces before you started shaking the tube? [1]

.....

.....

3 In both procedures why were different spatulas and different pipettes used to measure out each substance? [1]

.....

.....

Results

Observation when solid lead(II) nitrate is shaken with solid potassium iodide. [2]

.....

.....

Observation when solutions of lead(II) nitrate and potassium iodide are mixed together. [2]

.....

.....

.....

.....

Conclusions

1 What was observed in Procedure 1 which tells you that a chemical reaction has occurred? [1]

.....

2 What was observed in Procedure 2 which tells you that a chemical reaction has occurred? [1]

.....

3 Which of the two reactions occurred faster? [1]

.....

4 Use ideas about particles in solids and solutions to explain why the reaction you have stated in 3 occurred faster. [3]

.....

.....

.....

5 What colour is potassium nitrate solution? [1]

.....

Evaluation

What could have been done to ensure that, in both procedures, exactly the same amount of each chemical was mixed together? [2]

.....

.....

Extension

How could the reaction between solid reactants used in Procedure 1 be speeded up? Explain why the change you have suggested would speed up the reaction. [3]

.....

.....

.....

1.3 Sublimation of iodine

Aim

To discover what happens when a substance sublimes.

Theory

Iodine is one of a few chemical elements which sublimes when it is heated. Sublimation is the change of state which occurs when a solid turns directly to a gas/vapour when heated, or when a gas is cooled and it turns directly into a solid.

iodine solid $\rightleftharpoons$ iodine vapour

$I_2(s) \rightleftharpoons I_2(g)$

Apparatus and chemicals

- ☐ eye protection
- ☐ access to fume cupboard
- ☐ Bunsen burner
- ☐ test tube containing two small crystals of iodine
- ☐ test-tube holder
- ☐ test-tube rack
- ☐ mineral wool

Safety!

Iodine – harmful, dangerous for the environment

Disposal: add sodium thiosulphate solution to dissolve the iodine. Pour down the foul water drain with more water.

Procedure

Throughout the practical the student should wear eye protection.

You will be given a test tube with two small crystals of iodine in it.

1. Place a small plug of glass wool in the neck of the test tube.
2. Place the test tube in the test-tube holder.
3. Heat the iodine crystals gently in the fume cupboard using a colourless flame. You may find that it is better to hold the tube a little higher above the flame. If iodine vapour is escaping from the tube, you are heating it too strongly.
4. After heating for two minutes, place the test tube in the test-tube rack in the fume cupboard.

Method

1 Why did you carry out this experiment in the fume cupboard? [1]

.....

.....

2 Why is a plug of glass wool placed in the test tube? [1]

.....

.....

3 Could a rubber bung have been used instead of the glass wool? Explain your answer. [2]

.....

.....

Results

Write down what you observed when you heated the iodine crystals and then when the tube cooled down in the test-tube rack. [3]

.....

.....

.....

Conclusions

1 What colour is solid iodine? [1]

.....

2 What colour is iodine vapour? [1]

.....

3 Where did all the iodine collect at the end of the experiment? Explain why. [2]

.....

.....

Evaluation

How could the experiment have been improved to see more clearly the sublimation of the iodine? [1]

.....

Extension

Use your textbook or the Internet to find out which other elements are known to sublime when heated. [2]

.....

.....

.....

2 Elements, compounds and experimental techniques

2.1 Rock salt: an important raw material

Aim

To purify rock salt to obtain sodium chloride crystals.

Theory

Rock salt contains sodium chloride, which is a very important raw material. Because it is a soluble salt, it can be separated from the contaminating earthy materials, which are insoluble, by straightforward and simple techniques. Sodium chloride crystals may then be produced.

Apparatus and chemicals

- ☐ eye protection
- ☐ retort stand, clamp and boss
- ☐ pestle and mortar
- ☐ tripod
- ☐ gauze
- ☐ glass rod
- ☐ evaporating basin
- ☐ 100 m^3 beaker
- ☐ 2 × 400 cm^3 beakers
- ☐ Bunsen burner
- ☐ heat-resistant mat
- ☐ rock salt
- ☐ filter funnels
- ☐ filter papers

Safety!
Rock salt – low hazard

Procedure

1 Put on your eye protection.

2 If you are provided with large, rock-like samples then you will have to crush these before you start, using a pestle and mortar (Figure 1).

Figure 1 Using a pestle and mortar

3 Place some of the rock salt powder you have produced into a 100 cm^3 beaker and half fill with water.

4 Stand the beaker on a tripod and gauze and warm it gently with a Bunsen burner, stirring all the time with a glass rod.

5 After approximately 15 minutes, turn off the Bunsen burner and leave the beaker to stand until it has cooled down, and the insoluble sandy/earthy material has had a chance to settle.

6 Set up the filter funnel and filter paper as shown in Figure 2 and filter the cooled solution from step 5 into a 400 cm^3 beaker.

7 Pour some of the filtrate into an evaporating basin. Half fill a 400 cm^3 beaker with water and stand it on a tripod and gauze.

8 Place the evaporating basin on top of this beaker as shown in Figure 3, light the Bunsen burner and evaporate the filtrate slowly until there is only a small amount of solution left.

9 Set aside and leave to crystallise.

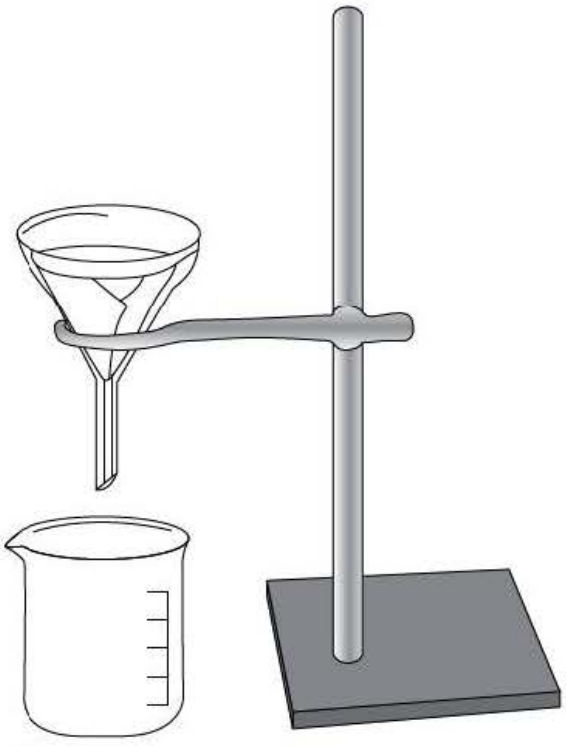


Figure 2

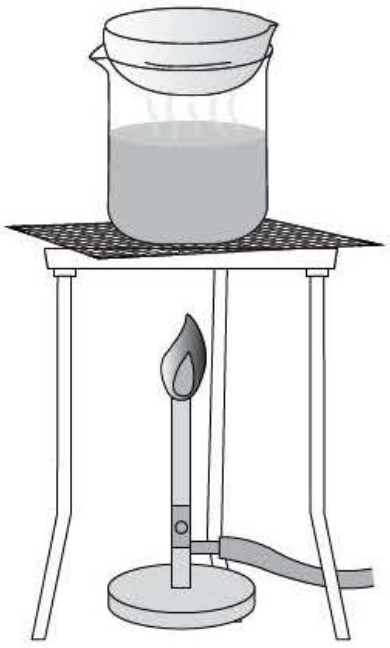


Figure 3

Method

1 Why was water added to the ground-up mixture? [1]

.....

2 Why did you allow the sample to stand? [2]

.....

.....

3 Why was the evaporated solution put to one side, rather than being evaporated to dryness? [2]

.....

.....

Results and calculations

Make a note of your observations below. [2]

.....

.....

Conclusions

Can you explain what you have observed? [2]

.....

.....

Evaluation

Outline how this experiment could be improved. [2]

.....

.....

.....

.....

Extension

1 Repeat this experiment using a sugar/sand mixture instead of rock salt. Explain the reasons for differences, if any, there may be in the procedure. [2]

.....

.....

.....

.....

.....

.....

2 The experiment you have carried out would be suitable for obtaining very small quantities of salt. However, companies need to extract very large quantities of salt. How could you do this on the large scale? [5]

.....

.....

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.....

.....

2.2 Ascending chromatography

Aim

To investigate the use of ascending chromatography in separating substances in a mixture.

Theory

You may have two or more solids, which are soluble, to separate. This situation arises, for example, when you have mixtures of coloured materials such as dyes. There are also mixtures of substances, such as proteins, which are colourless. The technique that is widely used to separate these materials so they can be identified is chromatography. There are several types of chromatography; however, they all follow the same basic principles. The simplest kind is paper chromatography, either flat-bed or ascending chromatography.

In the case of ascending chromatography, as the solvent moves up the paper, the dyes, for example, in a mixture are carried with it and begin to separate. They separate because the substances have different solubilities in the solvent and are absorbed to different degrees by the chromatography paper. As a result, they are separated gradually as the solvent moves up the paper. Sometimes a mixture of solvents is needed to improve the separation of the substances in the mixture.

The **chromatogram** in Figure 1 shows how a mixture of coloured substances in black felt-tip ink contains three dyes, P, Q and R.

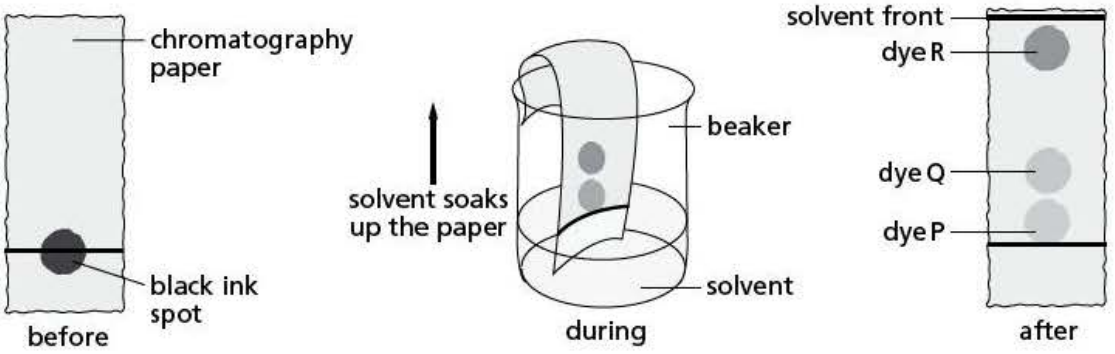

Figure 1

Numerical measurements (retardation factors) known as R_f values can be obtained from chromatograms. An R_f value is defined as the ratio of the distance travelled by the solute (for example P, Q or R) to the distance travelled by the solvent. These R_f values are found in data books for a variety of substances. This is useful for identification purposes.

Apparatus and chemicals

- ☐ eye protection
- ☐ piece chromatography paper (14 × 4 cm)
- ☐ 250 cm^3 beaker
- ☐ glass rod
- ☐ 2 × pieces of sticky tape
- ☐ ethanol (or colourless methylated spirits (IMS or IDA))
- ☐ water
- ☐ plastic bag
- ☐ black felt-tip pen

Safety!

Propanone – highly flammable, irritant

Ethanol – highly flammable

Please note that there should be no flames in the room. The room is well ventilated. Inhalation of solvents should be avoided.

Procedure

1. Put on your eye protection.
2. Rule a pencil line 1 cm from the bottom of a 14 × 4 cm strip of chromatography paper. This line is called the origin. Place one spot of black felt-tip ink in the middle of this line and leave it to dry.
3. Put 0.5 cm depth of an equal mixture of water and ethanol in the bottom of a 250 cm^3 beaker and hang the chromatography paper as shown in Figure 2. Cover the apparatus with an inverted plastic bag.
4. When the solvent has risen almost to the glass rod, take out the chromatogram and put it to one side to dry. Remember to mark the solvent front with a pencil as you are going to work out the R_f values later.

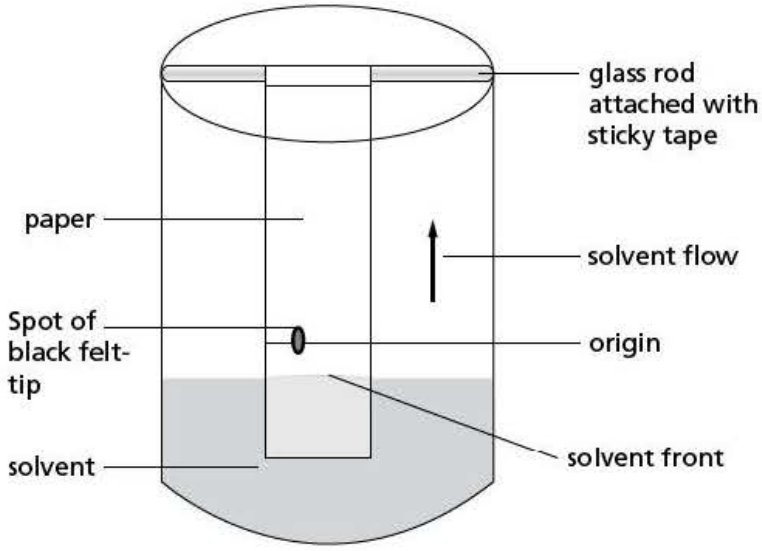

Figure 2

Method

1 Why did you use a pencil line for the origin instead of a ball-point pen? [1]

.....

2 Why did you allow the sample spot to dry? [1]

.....

3 Why was ethanol added to the water? [1]

.....

.....

4 Why was a plastic bag placed over the top of the experimental set-up? [1]

.....

5 Why did you allow the chromatogram to dry? [1]

.....

Results and calculations

Make a note of your observations on Figure 3 below and in Table 1 opposite.

To calculate R_f values you must make measurements of:

- distance solvent front moved
- distance dyes in mixture travelled – remember to measure to the point where there is the densest colour.

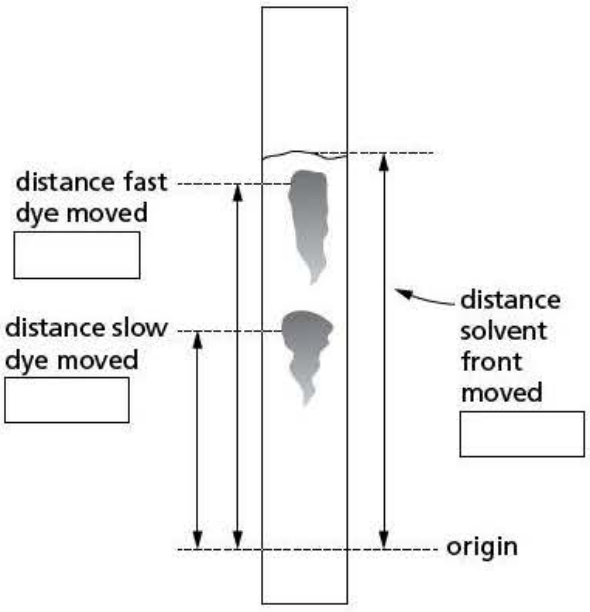

Figure 3

Table 1

Dye	Distance moved by solvent/mm	Distance moved by dye/mm	R_f value = distance travelled by dye/distance travelled by solvent

[12]

Conclusions

1 Can you explain what you have observed? [2]

.....

.....

.....

.....

2 Which dye was the least soluble and which was the most soluble in the solvent mixture? [2]

.....

.....

Evaluation

Outline how this experiment could be improved. [2]

.....

.....

Extension

1 Repeat this experiment using a different solvent mixture. Explain the reasons for differences, if any, there may be in the procedure as well as results obtained. [2]

.....

.....

.....

.....

→

2 Outline how chromatography techniques can be applied to colourless substances by exposing chromatograms to substances called locating agents. [2]

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.....

2.3 Elements, mixtures and compounds

Aim

To identify the differences between elements, their mixtures and compounds.

Theory

In this experiment the metal element iron and the non-metal element sulfur will be mixed and reacted together to form the compound iron sulfide.

iron + sulfur → iron sulfide

Fe(s) + S(s) → FeS(s)

The elements iron and sulfur, the mixture of these two elements and the iron sulfide produced will then be reacted with water and dilute hydrochloric acid. Also, the effect of a magnet will be demonstrated. Using the results of these experiments, the properties shown by the elements, mixtures and compound will be identified.

Part 1 is a teacher demonstration and it is recommended that it is performed in a fume cupboard, or at least in a well-ventilated laboratory with a safety screen.

Part 2 is a student experiment and runs alongside part 1.

Apparatus and chemicals

- ☐ eye protection
- ☐ protective gloves
- ☐ spatula
- ☐ 9 × heat-resistant test tubes
- ☐ heat-resistant mat
- ☐ Bunsen burner
- ☐ test-tube holder
- ☐ test-tube rack
- ☐ magnet
- ☐ access to fume cupboard
- ☐ powdered sulfur
- ☐ fine iron filings
- ☐ water
- ☐ 2 mol/dm^3 hydrochloric acid
- ☐ filter paper soaked in lead(II) ethanoate
- ☐ paper towels
- ☐ mineral wool

Safety!

Noxious fumes are produced during several of the experiments and so a fume cupboard is recommended, but if that is not available then a well-ventilated laboratory is needed and the use of a safety screen would then be essential.

Throughout, the teacher/demonstrator must wear eye protection. Put on plastic gloves when testing the gases produced from iron(II) sulfide reacting with dilute hydrochloric acid.

Iron filings – highly flammable

Sulfur – low hazard

Hydrochloric acid (2 mol/dm^3) – irritant

Lead(II) nitrate – toxic, dangerous for the environment

Iron(II) sulfide – low hazard

Hydrogen sulfide – a toxic foul smelling gas is produced. Use a fume cupboard.

Procedure 1

Throughout, the teacher/demonstrator must wear eye protection.

1 Initially, test the iron filings and sulfur powder with a magnet wrapped around by a paper towel, by putting it very close to the iron filings and the sulfur powder.

2 One spatula measure of iron filings with one spatula measure of sulfur powder on a paper towel are then mixed and tested with a magnet as in step 1.

3 Mix 1 g of iron filings with 1 g of sulfur and heat strongly in a heat-resistant test tube which has a mineral wool plug in it. This should be done until the red glow disappears.

4 Put the test tube aside, on a heat-resistant mat, to cool.

5 When it has cooled, tip the contents onto a paper towel and re-test with a magnet.

6 Iron sulfide should be tested with water and dilute hydrochloric acid. Test the gas given off in the reaction with dilute acid with moist lead ethanoate paper.

Procedure 2

1 Put on your eye protection.

2 Note down in Table 1:

(a) the appearance of the iron (element) filings, sulfur (element) powder and the mixture produced by your teacher.

(b) the effect of a magnet on these substances.

3 Your teacher will carry out a reaction between the iron filings and the sulfur powder.

4 Note down in the table:
   (a) the appearance of the product (iron sulfide) produced by your teacher.
   (b) the effect of a magnet on this substance.

5 Your teacher will give you six test tubes and a test-tube rack.

6 Into two test tubes add a small amount of iron filings. In another two test tubes add a small amount of sulfur powder and in a third pair a small amount of the iron/sulfur mixture.

7 To a sample of the iron, sulfur and the mixture add a little water. Note down in the table what happens.

8 In a fume cupboard repeat step 7 with the samples left, but this time use dilute hydrochloric acid. Note down in the table what happens.

Method

1 Do the elements keep their own properties when they form a mixture? [1]

.....

2 Do the elements keep their own properties when they are joined together as a compound? [1]

.....

3 There is a red glow created when the mixture is heated. What does this tell you is happening? [2]

.....

.....

Results and calculations

Table 1

	Sulfur	Iron	Mixture	Compound (iron sulfide)
appearance				
effect of magnet				
effect of water				
effect of dilute hydrochloric acid				

[16]

Conclusions

Mixtures behave as the individual ..... they are made up from. Compounds differ from the elements or ..... in that they are new ..... and so behave ..... . [2]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

.....

.....

Extension

1 Calcium oxide is a compound.

(a) Identify the elements present in this compound. [2]

.....

(b) Write a word and balanced chemical equation for the formation of this compound from the elements you have identified in 1a. [3]

.....

.....

2 Explain the difference between an element, a mixture and a compound. [3]

.....

.....

.....

3.1 Structure of substances

Aim

To identify small versus macromolecular substances.

Theory

Compounds containing covalent bonds have molecules whose structures can be classified as either **simple molecular** or **giant molecular** (macromolecular). Simple molecular structures are formed from only a few atoms. They have strong covalent bonds between the atoms within a molecule (**intramolecular bonds**) but have weak bonds between the molecules (**intermolecular bonds**).

Giant molecular or macromolecular structures contain many hundreds of thousands of atoms joined by strong covalent bonds.

You are going to investigate some of these substances which are spread around the laboratory and try to identify those which are small molecules versus those which macromolecular.

Apparatus and chemicals

- ☐ eye protection
- ☐ stoppered labelled bottles containing:
  - ☐ water
  - ☐ polyethene
  - ☐ pencil lead
  - ☐ bromine (only used by the teacher)
  - ☐ buckminster fullerene (soot)
  - ☐ nylon
  - ☐ oxygen
  - ☐ sand

Safety!

Bromine – toxic, corrosive. Should be used in a fume cupboard.

Soot – irritant. Should be used in a fume cupboard.

Procedure

1 Put on your eye protection.

2 Your teacher will tell you which substance to start with. **Do not handle the samples that have been placed in the fume cupboard**. For each substance note down its appearance.

3 Complete the table for the sample.

4 Suggest whether it is a substance containing small molecules or one which is macromolecular (inference).

5 Repeat this process with the other seven samples you will find around the laboratory.

Method

1 Why did you have to put on your eye protection? [1]

.....

2 Why was bromine placed in the fume cupboard? [1]

.....

Results and calculations

Table 1

Substance	Appearance	Inference
water		
polyethene		
pencil lead		
bromine		
buckminster fullerene(soot)		
Nylon		
oxygen		
sand		

[16]

Conclusion

The substances ..... contain small molecules while the substances .... are macromolecular. [8]

Evaluation

Outline how this experiment could be improved or made more reliable. [2]

.....

.....

Extension

1 Define the terms that are highlighted in bold in the theory section. [8]

.....

.....

.....

.....

2 Choose one of the small molecules and draw a dot and cross diagram of its bonding scheme. [3]

3 Choose one of the macromolecular substances and draw a dot and cross diagram of part of its bonding scheme? [3]

3.2 Properties of ionic and covalent substances

Aim

To identify ionic and covalent types of substances by observing their properties.

Theory

Ionic and covalent substances have different properties.

For ionic compounds:

- They are usually solids at room temperature, with high melting points. They also have high boiling points.
- They are usually crystalline and hard substances.
- They mainly dissolve in water.
- They usually cannot conduct electricity when solid, but they will conduct electricity when in aqueous solution or in the molten state.

For covalent compounds:

- As simple molecular substances, they are usually gases, liquids or solids with low melting and boiling points.
- Generally, they do not conduct electricity when molten or dissolved in water.
- Covalent substances are generally soluble in organic solvents. However, it should be noted that, generally, they do not dissolve in water. Water can interact with and dissolve *some* covalent molecules better than others.

You are going to investigate some of these properties and identify which of the substances that are spread around the laboratory are ionic and which are covalent.

Apparatus and chemicals

- ☐ eye protection
- ☐ stoppered bottles labelled A, B, C, D, E, F, G, H

Safety!

D – highly flammable, harmful

F – toxic, dangerous to the environment

G – highly flammable, harmful

H – harmful, irritant

Procedure

1 Put on your eye protection.

2 Your teacher will tell you which substance to start with. Write down its appearance in Table 1.

3 Complete the table for the sample and check out the melting point and boiling point and solubility in water, which are also given in the table.

4 Suggest whether the sample is an ionic or covalent substance (inference).

5 Repeat this process with the other seven samples you find around the laboratory.

Method

Why did you have to put on your eye protection? [1]

.....

Results and calculations

Table 1

Substance	Appearance	Melting point/°C	Boiling point/°C	Solubility in water/g/dm^3	Inference
A		–218.80	–83.00	0.008	
B		186.00	Decomposes	2115.0	
C		801.00	1413.00	350.0	
D		–45.00	156.00	5.90	
E		771.00	1500.00	342.00	
F		370.00	914.00	9.73	
G		–95.30	68.70	0.0095	
H		614.00	1382.00	830.50	

[16]

Conclusion

Substances ..... are ionic while substances ..... are covalent. [8]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

Extension

What other test could be carried out to check your answer as to whether the substance was ionic or covalent? [1]

.....

.....

3.3 Electrolysis of solutions

Aim

To use electrolysis to identify ionic substances.

Theory

Ionic substances usually conduct electricity when in aqueous solution. The forces of attraction between the ions are weakened and the ions are free to move to the appropriate electrode. This allows an electric current to be passed through the aqueous solution.

Generally, covalent substances do not conduct electricity when dissolved in water (if they dissolve in water!) This is because they do not contain ions. However, some molecules actually react with water to form ions. For example, hydrogen chloride gas produces aqueous hydrogen ions and chloride ions when it dissolves in water:

$$HCl(g) \xrightarrow{\text{water}} H^+(aq) + Cl^-(aq)$$

In this experiment you will identify substances that are ionic or covalent by their ability to conduct or not in aqueous solution.

Apparatus and chemicals

- ☐ eye protection
- ☐ 2 × carbon electrodes
- ☐ electrode clip holder
- ☐ 100 cm^3 beaker
- ☐ 2.5 V bulb in a bulb holder
- ☐ three leads plus crocodile clips
- ☐ 6 V d.c. power supply
- ☐ distilled/deionised water
- ☐ 0.2 M solutions of:
  - ☐ sugar
  - ☐ magnesium sulfate
  - ☐ copper(II) sulfate
  - ☐ 1 mol/dm^3 sulfuric acid

Safety!

Magnesium sulfate – low hazard

Copper(II) sulfate – toxic, dangerous to the environment

Sulfuric acid (1 mol/dm^3) – irritant

Procedure

1 Put on your eye protection.

2 Set up the circuit as shown in the diagram but do not connect or switch on your constant 6 V voltage supply.

3 Half fill the beaker with dilute sulfuric acid and set the voltage supply to 6 V.

4 Before switching on ask your teacher to check your circuit.

5 Switch on the 6 V supply and allow the current to pass through the solution for at least 2 minutes.

6 Record your observations in the results table.

7 Now pour away the sulfuric acid and rinse the electrodes with distilled/deionised water.

8 Repeat steps 2 to 7 with magnesium sulfate, copper(II) sulfate and sugar solution.

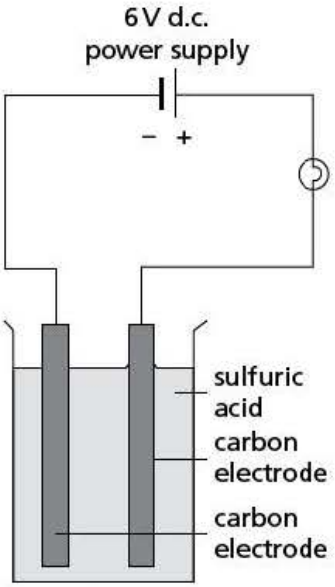

Figure 1

Method

1 When the bulb lights what does this tell you? [2]

.....

2 Why is it necessary to rinse the carbon electrodes after each experiment? [2]

.....

.....

Results and calculations

Table 1

Solution	Does the bulb light?	Observation at cathode (–ve)	Observation at anode (+ve)	Inference
dilute sulfuric acid				
magnesium sulfate				
copper(II) sulfate				
sugar solution				

[16]

Conclusion

When an electric current is passed through a solution and the bulb ..... then the dissolved substance is ..... . If the bulb does not ..... then it is ..... . The substances are broken down by the passage of an electric current – this is ..... . [5]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

Extension

1 What other tests could be carried out to check your answers to whether the substance was ionic or covalent? [1]

.....

2 Define the terms:

anode [1]

.....

cathode [1]

.....

electrode [2]

.....

electrolysis [2]

.....

4 Stoichiometry – chemical calculations

4.1 Determination of the formula of magnesium oxide

Aim

To find, by experiment, the formula of the compound magnesium oxide by burning a known mass of magnesium metal in air.

Theory

Magnesium metal is a reactive metal which will, when heated, react with oxygen gas in air to form solid magnesium oxide. When the reaction occurs, an intense white light is observed.

magnesium + oxygen → magnesium oxide

Even at room temperature, magnesium metal reacts slowly with oxygen to produce the oxide, so before the experiment starts it is necessary to remove the oxide layer on the surface of the magnesium metal using some emery paper.

The formula of magnesium oxide can be found from the moles of oxygen gas which will react with one mole of magnesium metal, using an empirical formula calculation.

Apparatus and chemicals

- ☐ eye protection
- ☐ crucible and lid
- ☐ Bunsen burner
- ☐ heat-resistant mat
- ☐ tripod
- ☐ pipe-clay triangle
- ☐ emery paper
- ☐ accurate balance
- ☐ tongs which meet
- ☐ 10 cm of magnesium ribbon

Safety!

Make sure that you do not look directly at the intense white light produced when the magnesium burns because it can cause an after image.

Allow the crucible to cool before you record its mass on the balance.

Magnesium ribbon – low hazard

Magnesium oxide – low hazard

Procedure

Throughout the practical the student should wear eye protection.

1 Clean a 15 cm piece of magnesium ribbon by pulling the ribbon through a folded up piece of emery paper a few times – it should become shinier as you remove the magnesium oxide and other impurities from its surface.

2 Weigh the crucible and lid. Make sure that the lid you are using overlaps the edges of the crucible. Record the mass in the results section.

3 Coil up your magnesium ribbon so that is fits onto the bottom of the crucible.

4 Weigh the crucible, lid and magnesium, again recording the mass in the results section.

5 Place the crucible containing the magnesium, and the lid onto a pipe-clay triangle on a tripod above a Bunsen burner.

6 Using a colourless Bunsen flame, heat the crucible strongly for about 15 minutes. Every 5 minutes lift off the crucible lid, with the tongs, for 2–3 seconds. During this time the magnesium metal will have reacted with the oxygen in the air. If it does not, keep heating for another five minutes.

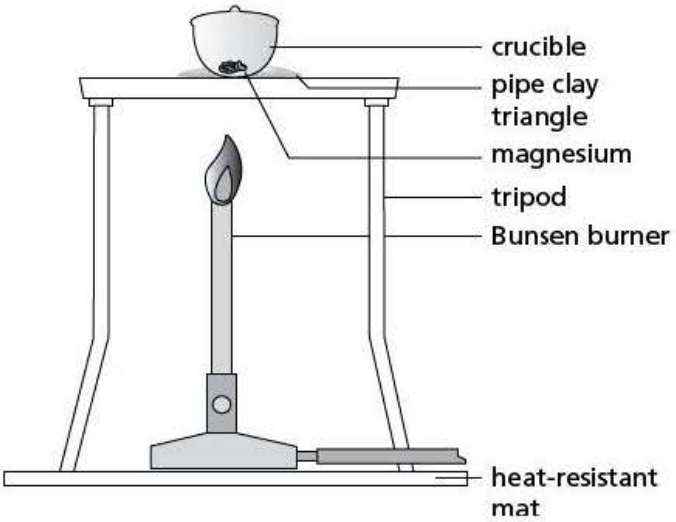

Figure 1

7 After the reaction has occurred, allow the crucible, magnesium oxide and lid to cool.

8 Reweigh the crucible, magnesium oxide and lid on the balance and record its mass in the results section.

9 Re-heat the crucible for another 5 minutes, raising the lid once for a few seconds. Allow the crucible, magnesium oxide and lid to cool before re-weighing them. Record the mass in the results section.

10 If the mass recorded in step 8 is the same as that in step 9, then the experiment is finished. If the mass has changed, repeat step 9 until it is constant. Record the final constant mass in the results section.

Method

1 Why is the magnesium ribbon coiled up at the start of the experiment? [1]

.....

.....

2 Why is it important to use a colourless Bunsen flame to heat the crucible? [2]

.....

.....

3 Why is a pipe-clay triangle used and not a gauze? [1]

.....

.....

4 Why is it important to raise the lid during the experiment? [1]

.....

.....

5 Why is it important only to raise the lid for a few seconds? [1]

.....

.....

6 Why does the heating have to be repeated until there is no change in the final mass of the apparatus and its contents? [2]

.....

.....

Results and calculations

Mass of crucible and lid = ... g

Mass of crucible, lid and magnesium ribbon = ... g

Mass of crucible, magnesium oxide and lid after 15 minutes = ... g

Final mass of crucible, magnesium oxide and lid after a further heating =

..... g

1 From your results, find the mass of magnesium used in the experiment. [1]

.....

2 From your results, find the mass of oxygen gas from the air which has reacted with the magnesium in your experiment. [1]

.....

3 Use the masses from 1 and 2 to complete the following calculation to obtain the formula of magnesium oxide from your experiment. [4]

Table 1

	Mg	O
mass/g		
moles	$\frac{\quad}{24}$ =	$\frac{\quad}{16}$ =
simplest ratio of moles		

Conclusions

What is the empirical formula of magnesium oxide from your data? [1]

.....

Evaluation

1 There are other gases in air which magnesium can react with to produce other products. How could the procedure be changed to remove this problem? [2]

.....

.....

2 How could the procedure be improved to increase the accuracy of the data obtained? [1]

.....

.....

Extension

Try to find out the names of at least two other gases found naturally in air which the magnesium metal could have reacted with. For each, write a balanced chemical equation for the reactions which occur. [4]

.....

.....

.....

.....

4.2 Determination of the volume occupied by one mole of a gas

Aim

To determine the volume occupied by one mole of hydrogen gas at the temperature and pressure of your chemistry laboratory.

Theory

It is known that at 25 °C (298 K) and 1 atmosphere of pressure, one mole of any gas occupies a volume of 24 dm^3. This is known as the molar volume.

In this experiment we will use the reaction between a known mass of magnesium metal and excess hydrochloric acid to give hydrogen gas, which will be collected and its volume recorded.

magnesium + hydrochloric acid → magnesium chloride + hydrogen

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

Apparatus and chemicals

- ☐ eye protection
- ☐ 100 cm^3 glass gas syringe
- ☐ clamp stand, boss and clamp
- ☐ boiling tube
- ☐ ignition tube
- ☐ bung with single hole fitted with glass tube
- ☐ rubber tubing to connect boiling tube to gas syringe
- ☐ 10 cm^3 measuring cylinder
- ☐ boiling-tube rack
- ☐ accurate balance
- ☐ scissors
- ☐ emery paper
- ☐ 10 cm of magnesium ribbon
- ☐ 2 mol/dm^3 hydrochloric acid
- ☐ thermometer

Safety!

Magnesium ribbon – low hazard

Hydrochloric acid (2 mol/dm^3) – irritant

Magnesium chloride solution – low hazard

Hydrogen gas – extremely flammable

Please note there should be no flames in the room.

Procedure

Throughout the practical the student should wear eye protection.

1 Using the measuring cylinder, place 5 cm^3 of the 2 mol/dm^3 hydrochloric acid into the boiling tube and set it into the boiling-tube rack.

2 Clean the magnesium ribbon with the emery paper.

3 Weigh the magnesium ribbon using the balance. Use scissors to get the mass of the ribbon you will use as close to 0.05 g as possible. Record the actual mass of the ribbon you use in the results section.

4 Place the magnesium ribbon you have weighed into an ignition tube.

5 Carefully lower the tube into the hydrochloric acid in the boiling tube. The level of the acid should not be above the open end of the tube.

6 Put the bung onto the boiling tube and connect the glass tube to the gas syringe using the rubber tubing. Ensure that the gas syringe reading is 0.0 cm^3 before connecting the rubber tubing.

7 Lift the boiling tube from the rack and gently tilt it to allow the hydrochloric acid to react with the magnesium ribbon.

8 Collect the gas in a syringe, as shown in Figure 1.

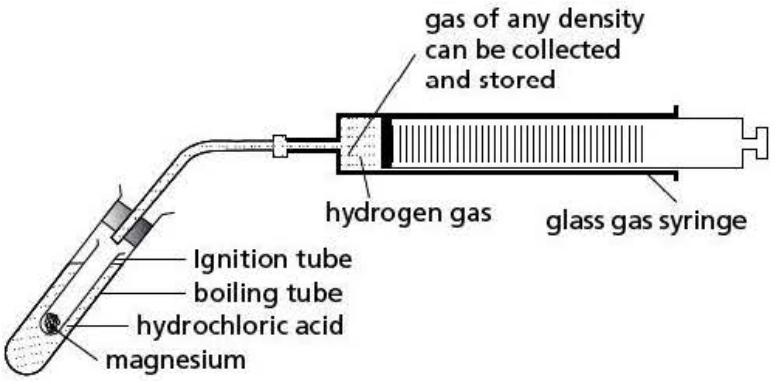

Figure 1

9 After all of the magnesium ribbon has reacted, record the final volume of hydrogen gas collected in the gas syringe.

10 Using the thermometer, record the temperature of the laboratory.

Method

1 Why is the magnesium ribbon cleaned with emery paper? [2]

.....

.....

2 Why is an excess of hydrochloric acid used in the experiment? [1]

.....

Results and calculations

Mass of magnesium ribbon used = ... g

Volume of gas collected = ... cm^3

1 From your results, calculate the number of moles of magnesium you used in your experiment. [2]

.....

.....

2 Using the balanced chemical equation, work out the number of moles of hydrogen gas that could be produced from the number of moles of magnesium you have calculated. [2]

.....

.....

.....

3 The number of moles you have just calculated is equivalent to the volume of gas collected in the syringe. Determine what the volume would have been if you had produced one mole of hydrogen gas. [2]

.....

.....

4 Convert the volume into dm^3. [1]

.....

Conclusions

1 From your results, what is the molar volume of hydrogen gas under the conditions at which the experiment was carried out? [1]

.....

2 Comment on your answer in relation to the known molar volume of a gas. [3]

.....

.....

.....

Evaluation

1 Discuss how the experiment could be improved to give more reliable results. [2]

.....

.....

2 As you may have seen during the experiment, sometimes the gas syringe gets stuck and the barrel needs to be rotated to help the collection of the gas. Can you think of a different way of collecting the hydrogen gas produced from this reaction? [2]

.....

.....

Extension

The volume of a given mass of a gas changes at different temperatures and pressures. Use your research skills to find the names of the two 'laws' which explain how the volume of a gas varies with temperature and pressure. [2]

.....

4.3 Determination of the percentage yield of a chemical reaction

Aim

To carry out a precipitation reaction and then to determine the mass of the precipitate formed. The percentage yield for the reaction will then be determined.

Theory

The percentage yield of a chemical reaction is determined using the relationship:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

The actual yield is the mass of the desired product obtained by carrying out the experiment. The theoretical yield is the maximum amount which could be produced if the reaction was carried out perfectly.

In this experiment, we will react sodium carbonate with excess calcium nitrate to produce calcium carbonate, which is insoluble in water, a precipitate.

sodium carbonate + calcium nitrate → calcium carbonate + sodium nitrate

$Na_2CO_3(aq) + Ca(NO_3)_2(aq) \rightarrow CaCO_3(s) + 2NaNO_3(aq)$

Apparatus and chemicals

- ☐ eye protection
- ☐ 25 cm^3 pipette and filler
- ☐ 50 cm^3 measuring cylinder
- ☐ 150 cm^3 beaker
- ☐ 250 cm^3 conical flask
- ☐ glass rod
- ☐ filter paper
- ☐ filter funnel
- ☐ accurate balance
- ☐ 1 mol/dm^3 sodium carbonate solution
- ☐ 1 mol/dm^3 calcium nitrate solution
- ☐ distilled water

Safety!

When you are putting the pipette filler onto the pipette, make sure that you hold the pipette close to the end the filler will be attached to. This should stop the pipette from breaking as you push on the filler.

Sodium carbonate solution (1 mol/dm^3) – low hazard

Calcium nitrate solution (1 mol/dm^3) – irritant

Calcium carbonate solid – irritant

Sodium nitrate solution – low hazard

Procedure

Throughout the practical the student should wear eye protection.

1 Using the pipette, with a filler, place 25 cm^3 of 1 mol/dm^3 sodium carbonate solution into the beaker.

2 Using the measuring cylinder, pour 50 cm^3 of 1 mol/dm^3 calcium nitrate into the beaker.

3 Using the glass rod, stir the contents of the beaker gently for two minutes.

4 Use the distilled water to ensure that all the solid is removed from the glass rod into the beaker.

5 Use the balance to weigh a piece of the filter paper, and record the mass in the results section.

6 Put the filter paper into the funnel, place the funnel in the top of the conical flask and filter the contents of the beaker. Use the distilled water to ensure that all of the solid has been removed from the beaker and is tipped onto the filter paper.

7 The calcium carbonate and filter paper now need to dry. This can be done by placing the filter paper over a radiator, using a drying oven or simply by leaving it for a few days.

8 When dry, find the mass of the filter paper and the calcium carbonate, and record the mass in the results section.

Method

1 Why was an excess of calcium nitrate solution used? [1]

.....

2 Why was a pipette used to measure out the sodium carbonate solution, but a measuring cylinder used for the calcium nitrate solution? [2]

.....

.....

.....

3 Why was it necessary to dry the filter paper and calcium carbonate before weighing them? [2]

.....

.....

.....

4 Why is the calcium carbonate not simply removed from the filter paper to find the mass produced? [1]

.....

.....

Results and calculations

Mass of filter paper = ... g

Mass of dry filter paper and calcium carbonate = ... g

1 From your results, calculate the mass of calcium carbonate produced in the experiment. This is the actual yield. [1]

.....

We now need to determine the theoretical yield using the amount of sodium carbonate used and the balanced chemical equation.

2 Calculate the number of moles of sodium carbonate in 25 cm^3 of 1 mol/dm^3 solution. [2]

.....

.....

3 Using the equation, determine the number of moles of calcium carbonate that should be produced. [2]

.....

.....

4 Using the number of moles of calcium carbonate which should have been produced, calculate the mass that should have been obtained. This is the theoretical yield. (A_r: Ca = 40, C = 12, O = 16) [2]

.....

.....

5 Now work out the percentage yield for the experiment. [2]

.....

.....

Conclusions

1 Was the percentage yield good or bad? [1]

.....

.....

2 In which part of the experiment do you think you lost the most calcium carbonate? Why did this happen? [2]

.....

.....

Evaluation

How could the experiment have been improved to increase the percentage yield? [2]

.....

.....

.....

Extension

Write a balanced ionic equation for the reaction you have just carried out. [2]

.....

.....

5.1 Electrolysis of lead(II) bromide

Aim

To investigate the electrolysis of molten lead(II) bromide and to identify, from given data, the substances produced as a result of passing an electric current through them.

Theory

During **electrolysis** the electric current enters and leaves the **electrolyte** through **electrodes**, which are usually made of unreactive metals such as platinum or of the non-metal carbon (graphite). These are said to be **inert** electrodes. The names given to the two electrodes are **cathode** (which attracts **cations**) and **anode** (which attracts **anions**).

Electrolysis is very important in industry. An experiment to get an understanding of this process is the electrolysis of a compound such as lead(II) bromide.

Apparatus and chemicals

- ☐ eye protection
- ☐ access to fume cupboard
- ☐ 2 × carbon electrodes
- ☐ electrode clip holder
- ☐ pipe-clay triangle
- ☐ tripod, gauze and heat-resistant mat
- ☐ Bunsen burner
- ☐ small crucible
- ☐ 2.5 V bulb in a bulb holder
- ☐ 3 × leads plus crocodile clips
- ☐ 6 V d.c. power supply
- ☐ lead(II) bromide

Safety!

Lead(II) bromide – toxic, dangerous for the environment

Bromine – toxic, corrosive

Lead metal – toxic

Procedure

The teacher or demonstrator should wear eye protection throughout the demonstration experiment. It should take place in a fume cupboard.

The results should be filled in during the demonstration.

1 Set up the circuit as shown in Figure 1 with the crucible set in a pipe-clay triangle which is set on a tripod on a heat-resistant mat.

2 Half fill the crucible with lead(II) bromide and set the voltage supply to 6 V.

3 Switch on the 6 V supply and allow the current to pass through the solid.

4 Now heat the crucible strongly and melt the lead(II) bromide.

5 After observations have been made, allow the apparatus to cool. Remove the cathode and examine it and the place in the crucible where it was.

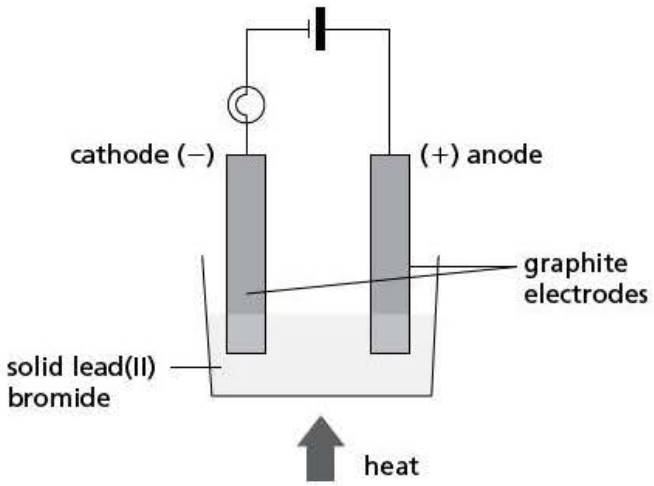

Figure 1

Method

1 Why is the experiment carried out in a fume cupboard? [2]

.....

2 Why do you have to melt the lead(II) bromide? [2]

.....

.....

Results and calculations

1 Does the bulb light when the lead(II) bromide is solid? ..... [1]

Explain your answer. [2]

.....

.....

2 Does the bulb light when the lead(II) bromide becomes molten? ..... [1]

Explain your answer. [2]

.....

.....

3 When the lead(II) bromide is molten, what do you see at the anode? ..... [1]

Explain your answer [2]

.....

.....

4 When the cathode is allowed to cool what do you see at this electrode?

..... [1]

Explain your answer. [2]

.....

.....

Conclusion

Solid lead(II) bromide does ..... conduct electricity. However, when it is ..... it does conduct and ..... takes place. ..... is produced at the cathode and ..... is produced at the anode. [5]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Define the terms highlighted in bold in the theory section of the demonstration. [16]

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

2 Write word and chemical equations for the reactions taking place at the anode and cathode. [6]

.....

.....

→

.....

.....

3 The following data was obtained from a series of electrolysis experiments.

Table 1

Substance	Physical state	Conductivity	Products	
			Anode	Cathode
A	liquid	yes	–	–
B	liquid	yes	pink metal	oxygen
C	solid	yes	–	–
D	liquid	yes	silvery metal	chlorine

(a) What type of substance do you think C is? ..... [1]

(b) Which of these substances could be mercury? ..... [1]

(c) Identify substance B. ..... [1]

(d) Identify substance D. ..... [1]

5.2 Electrolysis of water

Aim

To electrolyse water and confirm the formula for water is H_2O.

Theory

Several industrial processes involve the electrolysis of aqueous solutions. Pure water is a very poor conductor of electricity because there are so few ions in it.

Water $\rightleftharpoons$ hydrogen ions + hydroxide ions

$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

However, water can be made to decompose if an electric current is passed through following the addition of some sulfuric acid. The sulfuric acid provides ions to the solution.

The two gases produced are, of course, hydrogen and oxygen!

The standard apparatus used to carry out this experiment is the Hofmann voltameter (Figure 1).

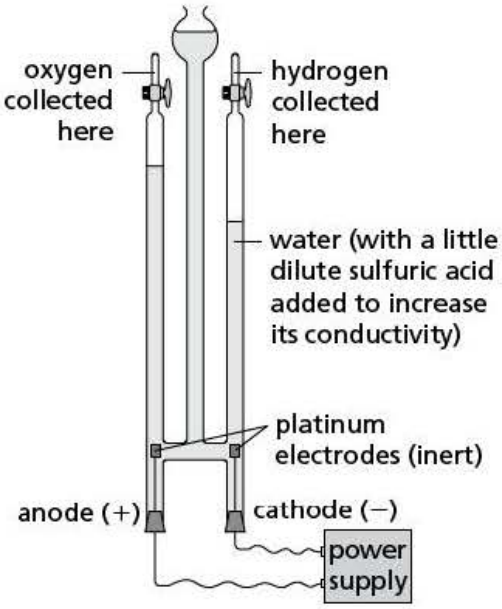

Figure 1

Apparatus and chemicals

- ☐ eye protection
- ☐ 400 cm³ beaker
- ☐ Hofmann voltameter
- ☐ 2 × test tubes
- ☐ 2 × wooden spills
- ☐ power supply
- ☐ deionised water
- ☐ 1 mol/dm³ sulfuric acid

Safety!
Sulfuric acid (1 mol/dm³) – irritant

Procedure

The teacher or demonstrator should wear eye protection throughout the demonstration experiment.

1 Some deionised water ($\frac{2}{3}$ by volume) is mixed with dilute sulfuric acid ($\frac{1}{3}$ by volume) in a 400 cm³ beaker.

2 This solution is poured into the apparatus through the reservoir (a thistle-headed centre tube) until both of the outside tubes are full.

3 The power supply is connected up and switched on. The gases produced are collected in the anode tube and cathode tube for approximately 10 minutes.

4 Note the approximate ratio of the volumes in the cathode tube : the anode tube.

5 Test tubes are then placed over these tubes and the gases collected in the test tubes.

6 The gas from the anode tube is tested with a glowing splint and the gas collected from the cathode tube is tested with a lighted splint.

Method

1 Why is deionised water used in this experiment? [2]

.....

.....

2 Why is the power supply connected up *after* the apparatus is filled by the solution? [2]

.....

.....

Results and calculations

Approximate ratio of gases produced: cathode : anode .....

Effect of glowing splint with anode tube gas: .....

The gas is .....

Effect of lighted splint with cathode tube gas: .....

The gas is .....

Do the results fit in with the known formula for water? .....

Explain your answer.

.....

.....

[8]

Conclusion

Pure water is a very bad ..... of electricity. This is because it has very few ..... present in the liquid. If dilute ..... acid is added to water, ... are added and this makes the solution a ..... one. An electric current passes through this solution and ..... takes place. Hydrogen gas is produced at the ..... and oxygen gas is produced at the ..... . The ratio of the volumes of the gases produced is ..... hydrogen to oxygen. This fits in with the formula for water being ..... . [10]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Write word and balanced chemical equations for the processes that take place at the anode and cathode. [6]

.....

.....

.....

2 Use your textbook to identify two very important industrial processes which use electrolysis of solutions. Outline the details of the processes involved. [12]

.....

.....

.....

5.3 Electrolysis of brine

Aim

To electrolyse brine and identify the products.

Theory

The electrolysis of saturated sodium chloride solution (brine) is the basis of a major industry – the chlor-alkali industry. The electrolytic process is a very expensive one, requiring vast amounts of electricity. The process is economic only because all three products, hydrogen, chlorine and sodium hydroxide, have a large number of uses.

The standard apparatus used to carry out this experiment is the Hofmann voltameter (Figure 1).

Figure 1

Apparatus and chemicals

- ☐ eye protection
- ☐ Hofmann voltameter
- ☐ 2 × test tubes
- ☐ 2 × wooden spills
- ☐ power supply
- ☐ saturated sodium chloride solution
- ☐ universal indicator solution

Safety!

Saturated sodium chloride solution (brine) – low hazard

Universal indicator solution – low hazard

Chlorine – toxic. Chlorine even in small quantity can induce asthmatic attack. Therefore it should not be inhaled.

Hydrogen – extremely flammable

Sodium hydroxide – irritant

Procedure

The teacher or demonstrator and students should wear eye protection throughout the demonstration experiment. It should take place in a fume cupboard or a well-ventilated laboratory.

1 The results of any experiments with collected gas and colour changes should be filled in during the demonstration.

2 Add a little universal indicator solution to the saturated sodium chloride solution.

3 Pour this solution into the apparatus through the reservoir (a thistle-headed centre tube) until both of the outside tubes are full.

4 The power supply is connected up and switched on. Any gases produced are collected in the anode tube and cathode tube for approximately 10 minutes.

5 A test tube is then placed over the cathode tube and the gas produced collected in it.

6 The gas collected from the cathode tube is tested with a lighted splint.

Method

1 Why is universal indicator solution used in this experiment? [2]

.....

.....

2 Why is the experiment carried out in a fume cupboard or well-ventilated laboratory? [2]

.....

.....

3 Why is the power supply connected up *after* the apparatus is filled by the solution? [2]

.....

.....

Results and calculations

Colour changes of the solution in the anode tube:

Initially .....

During the experiment .....

Colour changes of the solution in the cathode tube:

Initially .....

During the experiment .....

Effect of lighted splint with cathode tube gas:

.....

The gas is .....

Do the results fit in with the expectation of the electrolysis?

.....

Explain your answer.

.....

[10]

Conclusion

Brine is a very good ..... of electricity. This is because it has a lot of ..... present in the liquid. This makes the solution an ..... one. An electric current passes through this solution and ..... takes place. Hydrogen gas is produced at the ..... and ..... gas is produced at the .... is left in solution. [8]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Write word and balanced chemical equations for the processes that take place at the anode and cathode. [3]

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2 Use your textbook to identify three important uses of the three products of the electrolysis of brine. [9]

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6.1 Electrochemical cells: chemical energy to electrical energy

Aim

To make simple electrochemical cells, and to use the results obtained to put metals in an order of reactivity.

Theory

An electrochemical cell is one in which chemical energy is changed into electrical energy. To produce an electrochemical cell, two different metals are needed, an electrolyte and, in this experiment, a voltmeter to show the voltage produced. Alternatively a bulb could be used, which would light if the voltage produced was large enough. The voltage produced is large if the reactivity of the two metals is very different.

In an electrochemical cell, the more reactive metal loses electrons to form metal ions, while the less reactive metal gains electrons to convert its ions, from the solution, to form the metal.

To construct an electrochemical cell two half-cells are set up. These are connected together by a salt bridge through which ions can move.

Apparatus and chemicals

- ☐ eye protection
- ☐ voltmeter
- ☐ 2 × wire leads with plugs to connect to the voltmeter
- ☐ 2 × crocodile clips for the other end of the leads
- ☐ 2 × 150 cm^3 beakers
- ☐ filter paper
- ☐ saturated potassium chloride solution
- ☐ cleaned pieces (1 cm × 3 cm) of metals: copper, lead, magnesium, iron, zinc, tin
- ☐ 0.1 mol/dm^3 solutions of the nitrates (or other suitable solutions) of copper, lead, magnesium, iron, zinc and tin

Safety!

Be very careful of the sharp edges of the pieces of metal; they can give deep cuts.

Sulfuric acid (0.5 mol/dm³) – irritant

Procedure

Throughout the practical the student should wear eye protection.

1 Take the beaker and fold the piece of copper metal over the top. Clip a crocodile clip to the copper metal, holding it in place in the beaker. Connect the other end of the lead to the positive terminal of the voltmeter.

2 Fill the beaker with the 1 mol/dm³ copper nitrate solution so that the copper metal dips into it. This is known as a copper half-cell.

3 Take your piece of lead, fold it over the top of the second beaker. Fill the beaker with the 1 mol/dm³ lead(II) nitrate solution so that the lead metal dips into it. Connect the lead metal to the other terminal on the voltmeter. This is a lead half-cell.

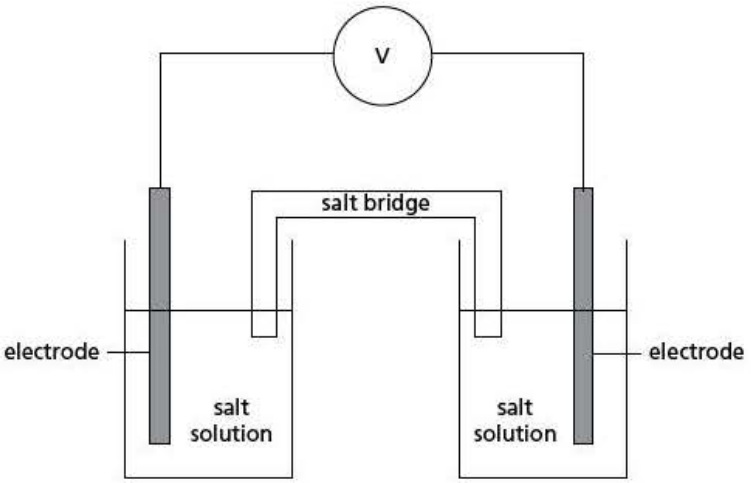

Figure 1

4 Finally, connect the two beakers together with a salt bridge which is made from a folded piece of filter paper soaked in potassium chloride solution.

5 You should now see a voltage being recorded on the voltmeter. You have just made an electrochemical cell – a battery. Record the voltage in the results table.

6 Repeat steps 3–5 for the other four metals, each time replacing the beaker and metal connected to the copper half-cell (the lead half-cell in the first case) with half-cells of the other metals.

Method

1 Which metal is being used as the standard in this experiment? [1]

.....

2 From the apparatus list can you see a piece of information about the solutions which is used to ensure a fair test? [1]

.....

3 Why is it important that the pieces of the metals are cleaned? [1]

.....

Results

1 Complete the table as you carry out the experiment. [2]

Table 1

Positive terminal	Negative terminal	Voltage/V
copper	lead	
copper	magnesium	
copper	iron	
copper	zinc	
copper	tin	

Conclusions

1 Which combination of metals produces the greatest voltage? [1]

.....

2 Which combination of metals produces the smallest voltage? [1]

.....

3 What do you think the voltage would have been if you had used a copper half-cell connected to both terminals of the voltmeter? [1]

.....

4 Which of the metals is the most reactive, and which is the least reactive (not including copper)? [2]

.....

5 Put the metals in order of reactivity with the least reactive first. [2]

.....

6 For the most reactive metal, write an ionic half-equation to show what happens when it is connected to the copper half-cell. [3]

.....

Evaluation

How could the procedure have been improved to increase the reliability of the results? [2]

.....

.....

Extension

1 What do you think the voltage would have been if you had connected a lead half-cell to a magnesium half-cell to make an electrochemical cell? [2]

.....

2 The experiment using the copper half-cell and the zinc half-cell was one of the first made, and was called the Daniel cell. Use your research skills to find out about the Daniel cell and the voltage it produced. Was your voltage similar? If not, what could have affected your value? [6]

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6.2 Calculating the energy of combustion of methanol and ethanol

Aim

To determine the energy of combustion of ethanol.

Theory

The energy of combustion of any fuel is the energy given off when one mole of the fuel is completely burned in oxygen at 25 °C (298 K) and 1 atmosphere of pressure.

In this experiment, a known mass of water will be heated by a fuel in a spirit burner. The temperature rise of the water and the mass of fuel burned can then be used to find the energy transferred from the fuel to the water using:

energy transferred (J) = mass of water used (g) × 4.2 (J/g °C) × rise in temperature (°C)

where the specific heat capacity of water is 4.2 J/g °C.

The density of water is 1 g/cm^3.

Apparatus and chemicals

- ☐ eye protection
- ☐ copper calorimeter
- ☐ glass rod
- ☐ spirit burner containing methanol and lid
- ☐ spirit burner containing ethanol and lid
- ☐ clamp stand, boss and clamp
- ☐ 100 cm^3 measuring cylinder
- ☐ thermometer
- ☐ accurate balance

Safety!

Take care when using spirit burners. Depending on the length of wick sticking out of the top of the burner, the flame size can get very large.

The copper calorimeter will be very hot at the end of the first experiment.

Methanol – highly flammable, toxic

Ethanol – highly flammable

Procedure

Throughout the practical the student should wear eye protection.

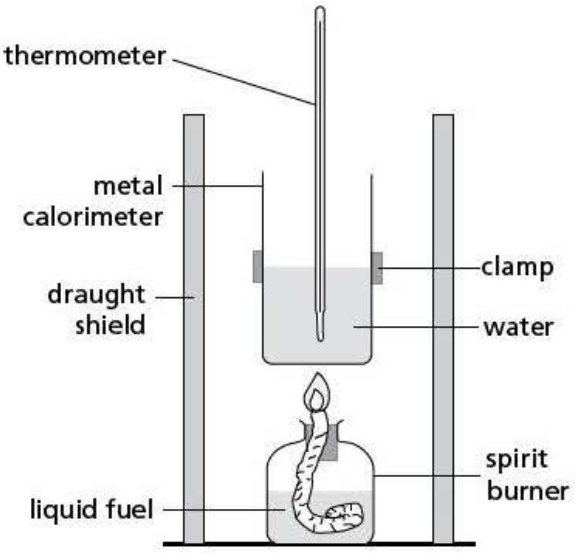

Figure 1

1 Using a measuring cylinder, put 150 cm^3 (or a known volume) of cold water into the copper calorimeter. Take the temperature of the water using a thermometer, and record it in the results section.

2 Support the calorimeter over a spirit burner containing methanol. Ensure that the height of the calorimeter is set so that the flame from the burner will meet the bottom of the calorimeter.

3 Weigh the spirit burner with the lid on. Record your results in the results section.

4 Replace the burner under the calorimeter and light the wick. Use the thermometer to stir the water all the time it is being heated. Continue heating until the temperature has risen by between 25 and 30 °C.

5 Extinguish the burner and replace the lid.

6 Keep stirring the water and record the highest temperature reached in the results section.

7 Weigh the spirit burner and lid again.

8 Repeat steps 1–7 using a spirit burner containing ethanol. Put new cold water into the calorimeter and ensure that the flame size is the same as in the first experiment.

Method

1 Why is a copper calorimeter used instead of a glass beaker? [2]

.....

.....

2 What is the purpose of the lid on the spirit burner? [2]

.....

.....

.....

3 What are the major problems with the procedure used? [2]

.....

.....

4 At the end of the experiment what can you see has happened to the bottom of the copper calorimeter? What is this caused by? [2]

.....

.....

Results

Methanol: Volume of water used = ... cm^3

Initial mass of spirit burner, methanol and lid = ... g

Final mass of spirit burner, methanol and lid = ... g

Initial temperature of water = ... °C

Maximum temperature of water = ... °C

Ethanol: Volume of water used = ... cm^3

Initial mass of spirit burner, ethanol and lid = ... g

Final mass of spirit burner, ethanol and lid = ... g

Initial temperature of water = ... °C

Maximum temperature of water = ... °C

	Methanol	Ethanol	
1 Find the temperature rise in the experiment (°C)			[2]
2 Energy transferred (J) (mass of water × 4.2 × temp. rise)			[2]
3 Mass of fuel burned (g)			[2]
4 Moles of fuel burned (moles) (mass/M_r)	___ / 32 =	___ / 46 =	[2]
5 Energy transferred by 1 mole (J/mol) (energy transferred/moles burned)			[2]
6 Energy of combustion (kJ/mol) (energy transferred by 1 mole/1000) =	___ / 1000 =	___ / 1000 =	[1]

Conclusions

The data-book value for the energy of combustion of methanol is –726 kJ/mol and for ethanol it is –1371 kJ/mol.

1 What does the '–' indicate in front of these numbers? [2]

.....

.....

Your values should have a '–' sign inserted at the front.

2 Why are the values you have obtained for the energy of combustion so small compared to the actual values? [3]

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Evaluation

As you have seen from your results, in comparison with the actual values, the experimental values are not perfect. Think of some ways in which the experiment could be improved to increase the accuracy of the results obtained. [3]

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Extension

Use your research skills to find out how the data-book values are obtained. [2]

.....

.....

6.3 Determination of the energy change of a displacement reaction

Aim

To determine the energy given off when the displacement reaction between iron metal and copper(II) sulfate occurs.

Theory

Iron is a more reactive metal than copper. So when iron metal is added to copper(II) sulfate solution, a displacement reaction occurs.

iron + copper(II) sulfate → iron(II) sulfate + copper

$Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$

By measuring the temperature rise which occurs when the iron and copper(II) sulfate are reacted together, it is possible to work out the energy change for the reaction.

Apparatus and chemicals

- ☐ eye protection
- ☐ polystyrene cup
- ☐ 25 cm^3 pipette and filler
- ☐ glass rod
- ☐ thermometer
- ☐ stopwatch
- ☐ accurate balance
- ☐ 2 g of iron filings
- ☐ 1 mol/dm^3 copper(II) sulfate

Safety!

Be careful – when the thermometer is placed in the polystyrene cup to record the temperature it may fall over unless it is held up.

Copper(II) sulfate (1 mol/dm^3) – harmful

Iron(II) sulfate solution – harmful

Copper metal – low hazard

Procedure

Throughout the practical the student should wear eye protection.

1. Using the pipette and filler, place 25 cm^3 of the 1 mol/dm^3 copper(II) sulfate solution into the polystyrene cup.
2. Using the balance, weigh out 2 g of iron filings. This is an excess.
3. Place the thermometer in the copper(II) sulfate solution and record the temperature in the results table. Repeat every 30 seconds, recording the results in the table.
4. After 2 minutes, add all of the iron filings to the copper(II) sulfate solution. Continue to record the temperature of the reaction mixture, every 30 seconds until 12 minutes have passed, in the results table. Stir the mixture with the thermometer.

Method

1 Why is a polystyrene cup used instead of a glass beaker? [2]

.....

.....

2 Why was a pipette used to measure the copper(II) sulfate and not a measuring cylinder? [2]

.....

.....

.....

.....

3 Why was an excess of iron filings used? [1]

.....

Results and calculations

Table 1

Time/minutes	Temperature/°C
0	
0.5	
1.0	
1.5	
2.0	
2.5	
3.0	
3.5	
4.0	
4.5	
5.0	
5.5	
6.0	

Time/minutes	Temperature/°C
6.5	
7.0	
7.5	
8.0	
8.5	
9.0	
9.5	
10.0	
10.5	
11.0	
11.5	
12.0	

Plot the results you have obtained on the graph paper below.

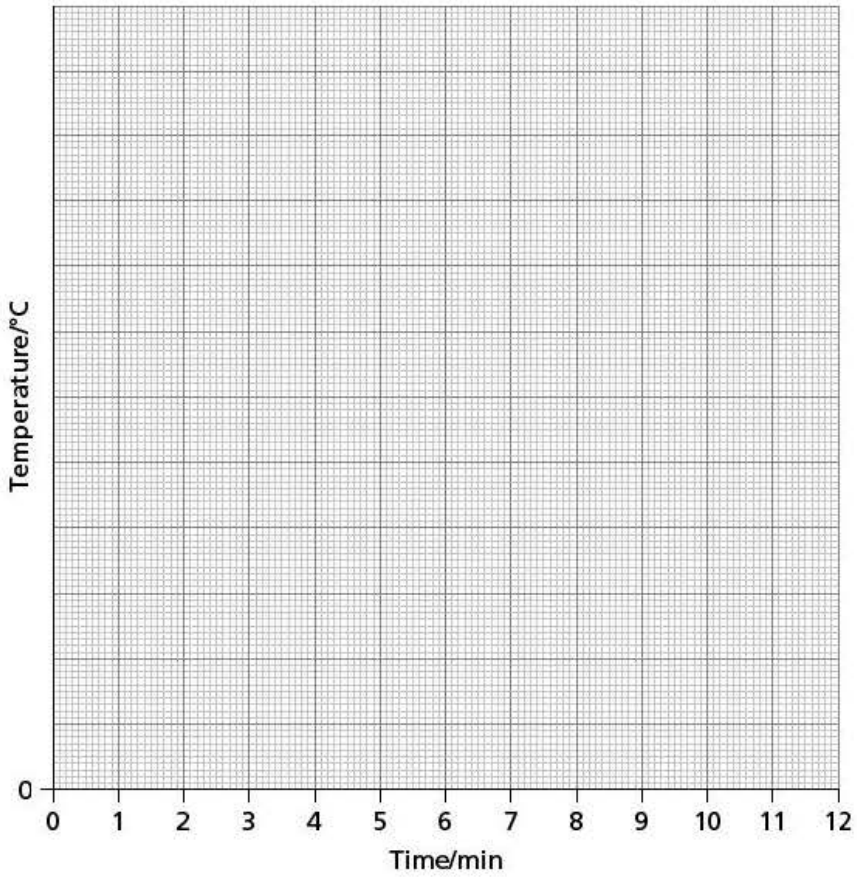

Figure 1

You will notice, from the shape of your graph, that the maximum temperature rise does not occur exactly when the iron filings have been added to the copper(II) sulfate. To work out the maximum temperature rise follow Figure 2.

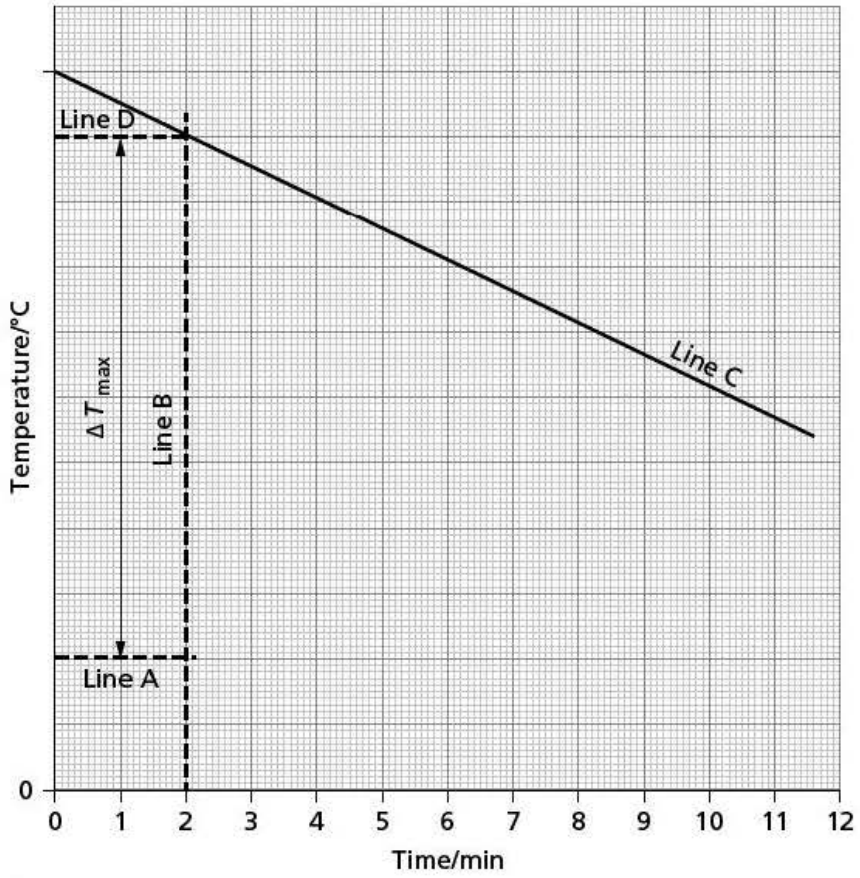

Figure 2

1 Using a ruler draw lines A, B, C and D as shown on Figure 2. [4]

2 The theoretical maximum temperature change which could have occurred, had all the copper(II) sulfate reacted with the iron when it was added, is shown as ΔT_{max} on the diagram. Determine what ΔT_{max} would have been for your results. [2]

.....

.....

3 Find the energy produced in the reaction using: [2]

Energy produced (J) = $25 \times 4.2 \times \Delta T_{max}$

.....

4 This is the amount of energy produced when 0.025 moles of copper(II) sulfate is used. What would have been the energy produced if you had started with 1 mole of copper(II) sulfate? [2]

.....

.....

Conclusions

1 The accepted value for the energy of displacement you have just carried out is –152 kJ/mol. Comment on the energy change you determined from your experiment. [2]

.....

.....

2 The energy change of displacement for the reaction between copper(II) sulfate solution and zinc metal is –219 kJ/mol. Can you explain why this reaction is more exothermic? [2]

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3 What can you say about the size of the energy change for the displacement reaction between copper(II) sulfate and tin metal, in comparison to the energy changes for iron and zinc? [3]

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Evaluation

Write down at least two ways in which this experiment could have been improved and state why they would have improved the experiment. [4]

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Extension

To determine the ΔT_{max} value from your graph you have used an extrapolation technique. What is this? [2]

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7.1 How does changing surface area affect the rate of a reaction?

Aim

To find out how changing the surface area of a specific mass of solid reactant affects the rate of a chemical reaction.

Theory

To show the effect of changing the surface area of a solid reactant on the reaction rate, we will use the reaction between limestone (calcium carbonate, $CaCO_3$) and hydrochloric acid.

calcium carbonate + hydrochloric acid → calcium chloride + water + carbon dioxide

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

When this reaction occurs, carbon dioxide gas is produced. The rate of the reaction can be followed in a number of ways: one way is to look at the loss in mass of the apparatus against time.

Apparatus and chemicals

- ☐ eye protection
- ☐ accurate balance
- ☐ 2 × 250 cm^3 conical flasks
- ☐ 100 cm^3 measuring cylinder
- ☐ cotton wool
- ☐ stopwatch
- ☐ 4 mol/dm^3 hydrochloric acid solution
- ☐ 10 g of large limestone chips
- ☐ 10 g of smaller limestone chips

Safety!
Hydrochloric acid (2 mol/dm^3) – irritant

Procedure

Throughout the practical the student should wear eye protection.

1 Place the 10 g of large limestone chips carefully into a conical flask.

2 Take a piece of cotton wool and check that it fits into the neck of the flask without falling in, and then remove it.

3 Using the measuring cylinder, measure out 100 cm^3 of the 4 mol/dm^3 hydrochloric acid solution.

4 Place the conical flask on the balance, carefully pour in the hydrochloric acid from the measuring cylinder, place the cotton wool in the flask and start the stopwatch. Record the initial mass of the apparatus in the results table.

5 Using the table, record the mass at the time intervals shown.

6 After 15 minutes, remove the flask from the balance.

7 Repeat steps 1–6 for the 10 g of smaller limestone chips.

Method

What is the purpose of the cotton wool in the neck of the conical flask? [2]

.....

.....

Results and calculations

1 Complete this table as you obtain the results. [2]

2 You will need to work out the total loss in mass, based on the initial mass, after each result is recorded. [2]

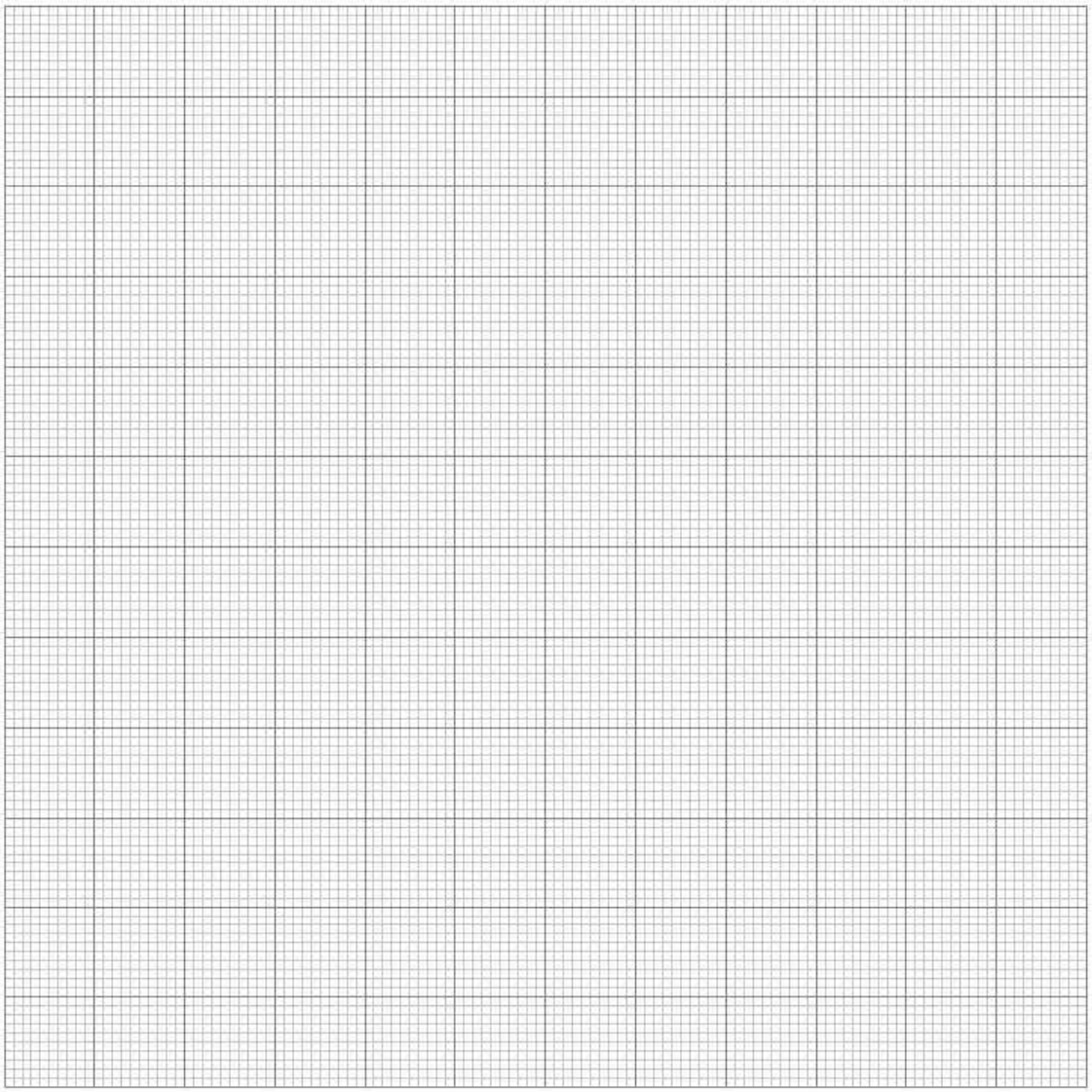
Table 1

Time (min)	10 g of large marble chips		10 g of smaller marble chips	
	Mass (g)	Loss in mass (g)	Mass (g)	Loss in mass (g)
0		–		–
0.5				
1.0				
1.5				
2.0				
2.5				
3.0				
3.5				
4.0				
4.5				

Table 1 continued

Time (min)	10 g of large marble chips		10 g of smaller marble chips	
	Mass (g)	Loss in mass (g)	Mass (g)	Loss in mass (g)
5.0				
6.0				
7.0				
8.0				
9.0				
10.0				
11.0				
12.0				
13.0				
14.0				
15.0				

3 Plot a graph of loss in mass (*y*-axis) against time (*x*-axis) for both experiments. Use the same axes. [4]



Conclusions

1 Which of the experiments is the fastest? [1]

.....

2 How can you tell this from the graph? [1]

.....

3 Use ideas about particles to explain your answer. [4]

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4 What other factors must be kept constant during this experiment to show the effect of surface area on the rate of reaction? [3]

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.....

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Evaluation

Write down at least two ways in which this experiment could have been improved and state why they would have improved the experiment. [2]

.....

.....

Extension

Nanotubes are to be used as catalysts in a variety of chemical reactions. Use your research skills to find out about nanotubes and why they are likely to be very good catalysts. [2]

.....

.....

7.2 What is the effect of changing the temperature on the rate of a reaction?

Aim

To find out how changing the temperature alters the rate of a chemical reaction.

Theory

To show the effect of a change in temperature, we will use the reaction between hydrochloric acid and sodium thiosulfate.

When hydrochloric acid reacts with sodium thiosulfate, one of the products of the reaction is sulfur. The sulfur that is formed makes the reaction mixture cloudy, as it is a yellow solid.

sodium thiosulfate + hydrochloric acid → sodium chloride + water + sulfur dioxide + sulfur

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

In this experiment the concentration of the sodium thiosulfate and hydrochloric acid remains constant but the temperature will be changed. Any changes in the rate of the reaction will be due to the temperature at which the reaction is being carried out.

The rate of the reaction can be found using

rate (/s) = 1/time (s)

because the same amount of sulfur is being produced each time to obscure the pencil cross.

Apparatus and chemicals

- ☐ eye protection
- ☐ white paper
- ☐ 250 cm^3 beaker
- ☐ Bunsen burner
- ☐ gauze
- ☐ tripod
- ☐ thermometer
- ☐ 2 × boiling tubes
- ☐ 25 cm^3 measuring cylinder

- ☐ 10 cm^3 measuring cylinder
- ☐ 250 cm^3 conical flask
- ☐ stopwatch
- ☐ 0.2 mol/dm^3 sodium thiosulfate solution
- ☐ 1 mol/dm^3 hydrochloric acid
- ☐ bucket with sodium carbonate solution

Safety!

The apparatus will be hot, be careful when you are touching it.

Hydrochloric acid (1 mol/dm^3) – low hazard

Sodium thiosulfate (0.2 mol/dm^3) – low hazard

Sulfur dioxide gas – toxic so there should be no inhalation of the contents

Sulfur – low hazard

Procedure

Throughout the practical the student should wear eye protection.

1. On a piece of white paper place a pencil cross. Make sure that it is not too dark.
2. Half fill a beaker with water and place it on a tripod and gauze. Heat the water in the beaker until the temperature is around 30 °C.
3. Into a boiling tube, place 15 cm^3 of the sodium thiosulfate solution using a measuring cylinder. Put this boiling tube into the beaker containing the water.
4. Using a 10 cm^3 measuring cylinder, measure out 10 cm^3 of the hydrochloric acid solution into a boiling tube and put this into the same beaker.
5. Allow the contents of the tubes to reach around 30 °C. Be careful to not contaminate the solutions by putting the same thermometer in both solutions.
6. When you are ready to start the experiment, pour the contents of the two boiling tubes into a conical flask sitting on the pencil cross, swirl the flask once and start the stopwatch. Take the temperature of the reaction mixture and record it in Table 1.
7. Look down through the flask at the cross and stop the watch when the cross is no longer visible. Record the time taken in seconds in the table. Pour contents into the bucket of sodium carbonate solution - this stops the reaction and neutralises the sulfur dioxide.
8. Repeat steps 1–7 for the other temperatures: around 40 °C, 50 °C, 60 °C and 70 °C.
9. Finally, carry out the experiment at room temperature.

Method

1 Why is it important not to make the pencil cross not too dark? [1]

.....

2 Why is it important to use the same pencil cross throughout the experiment? [2]

.....

.....

.....

3 Why will the amount of swirling of the flask affect the time? [1]

.....

4 Why should you look down through the flask towards the cross each time? [1]

.....

Results and calculations

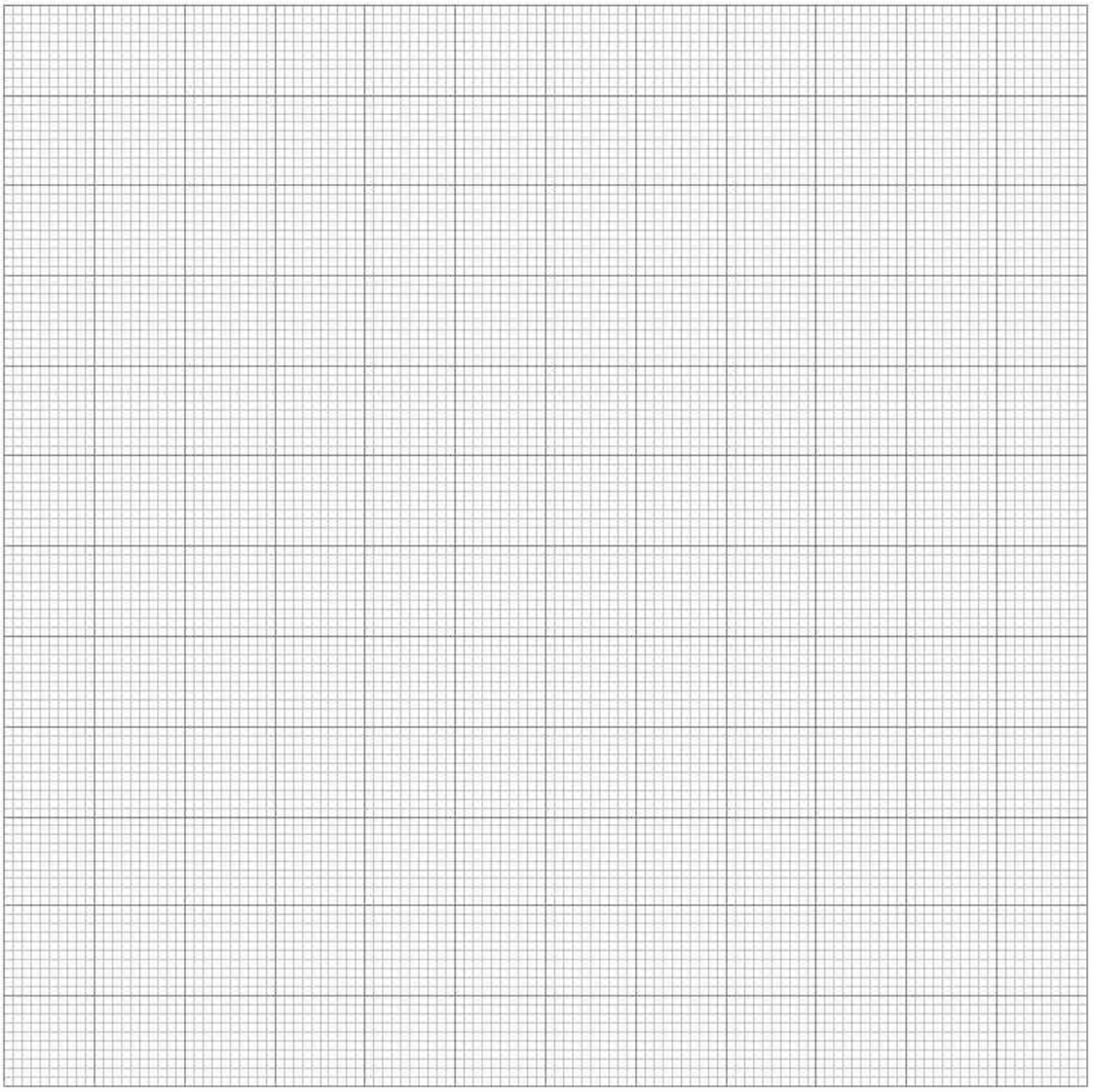
1 Record your times in the table below as you carry out the experiment. Record the times as accurately as possible from your stopwatch. [2]

Table 1

Expt no.	Volume of hydrochloric acid/cm^3	Volume of sodium thiosulfate/cm^3	Temperature/°C	Time/s	Rate/s^{-1}
1	10	15			
2	10	15			
3	10	15			
4	10	15			
5	10	15			
6	10	15			

2 When you have done all the experiments, complete the table by working out the rate of reaction for each experiment. [1]

3 Plot a graph of the rate of reaction (/s) on the *y*-axis against temperature (°C) on the a-axis using the graph paper below. [4]



Conclusions

1 What is the relationship between the time you have recorded and the rate of the reaction? [2]

.....

.....

2 What does your graph tell you about how temperature affects the rate of a reaction? [2]

.....

.....

3 What was kept constant in this experiment to ensure a 'fair test'? [2]

.....

4 Explain your findings in terms of particles. [4]

.....

.....

.....

.....

Evaluation

How could the procedure have been improved to give more accurate and reliable data? [4]

.....

.....

.....

Extension

Use your research skill to find out about the relationship between the speed of the reactant particles and the temperature of the reaction. [2]

.....

.....

7.3 What is the effect of changing the concentration on the rate of a reaction?

Aim

To find out how changing the concentration of a reactant affects the reaction rate.

Theory

To show the effect of a change in concentration of a reactant, the reaction between hydrochloric acid and sodium thiosulfate will be used.

When hydrochloric acid reacts with sodium thiosulfate one of the products of the reaction is sulfur. The sulfur that is formed makes the reaction mixture cloudy, as it is a yellow solid.

sodium thiosulfate + hydrochloric acid → sodium chloride + water + sulfur dioxide + sulfur

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

In this experiment the concentration of the sodium thiosulfate will be changed while the concentration of the hydrochloric acid remains constant. Any changes in the rate of the reaction will, therefore, be due to the change in the concentration of the sodium thiosulfate.

The rate of the reaction can be found using

Rate (/s) = 1/time (s)

because the same amount of sulfur is being produced each time to obscure the pencil cross.

Apparatus and chemicals

- ☐ eye protection
- ☐ white paper
- ☐ 25 cm^3 measuring cylinder
- ☐ 10 cm^3 measuring cylinder
- ☐ 250 cm^3 conical flask
- ☐ stopwatch
- ☐ 0.2 mol/dm^3 sodium thiosulfate solution
- ☐ 2 mol/dm^3 hydrochloric acid
- ☐ bucket containing sodium carbonate solution

Safety!

Hydrochloric acid (1 mol/dm³) – low hazard

Sodium thiosulfate (0.2 mol/dm³) – low hazard

Sulfur dioxide gas – toxic so no inhalation of the contents

Sulfur – low hazard

Procedure

Throughout the practical the student should wear eye protection.

1. On a piece of white paper place a pencil cross. Make sure that it is not too dark.
2. Into a conical flask place 15 cm^3 of the sodium thiosulfate solution using a measuring cylinder.
3. Place the conical flask on the paper, over the pencil cross.
4. Using a 10 cm^3 measuring cylinder, measure out 5 cm^3 of the hydrochloric acid solution.
5. When you are ready to start the experiment, pour the hydrochloric acid into the conical flask, swirl the flask once and start the stopwatch.
6. Look down through the flask at the cross and stop the watch when the cross is no longer visible. Record the time taken in seconds in Table 1. Pour liquid into the bucket of sodium carbonate solution.
7. Repeat steps 1–6 for the other experiments shown in the table, remembering to add water as indicated to ensure that the volume remains constant.

Method

1 Why is it important not to make the pencil cross not too dark? [1]

.....

2 Why is it important to use the same pencil cross throughout the experiment? [2]

.....

.....

3 Why will the amount of swirling of the flask affect the time? [1]

.....

4 Why should you look down through the flask towards the cross each time? [1]

.....

Results and calculations

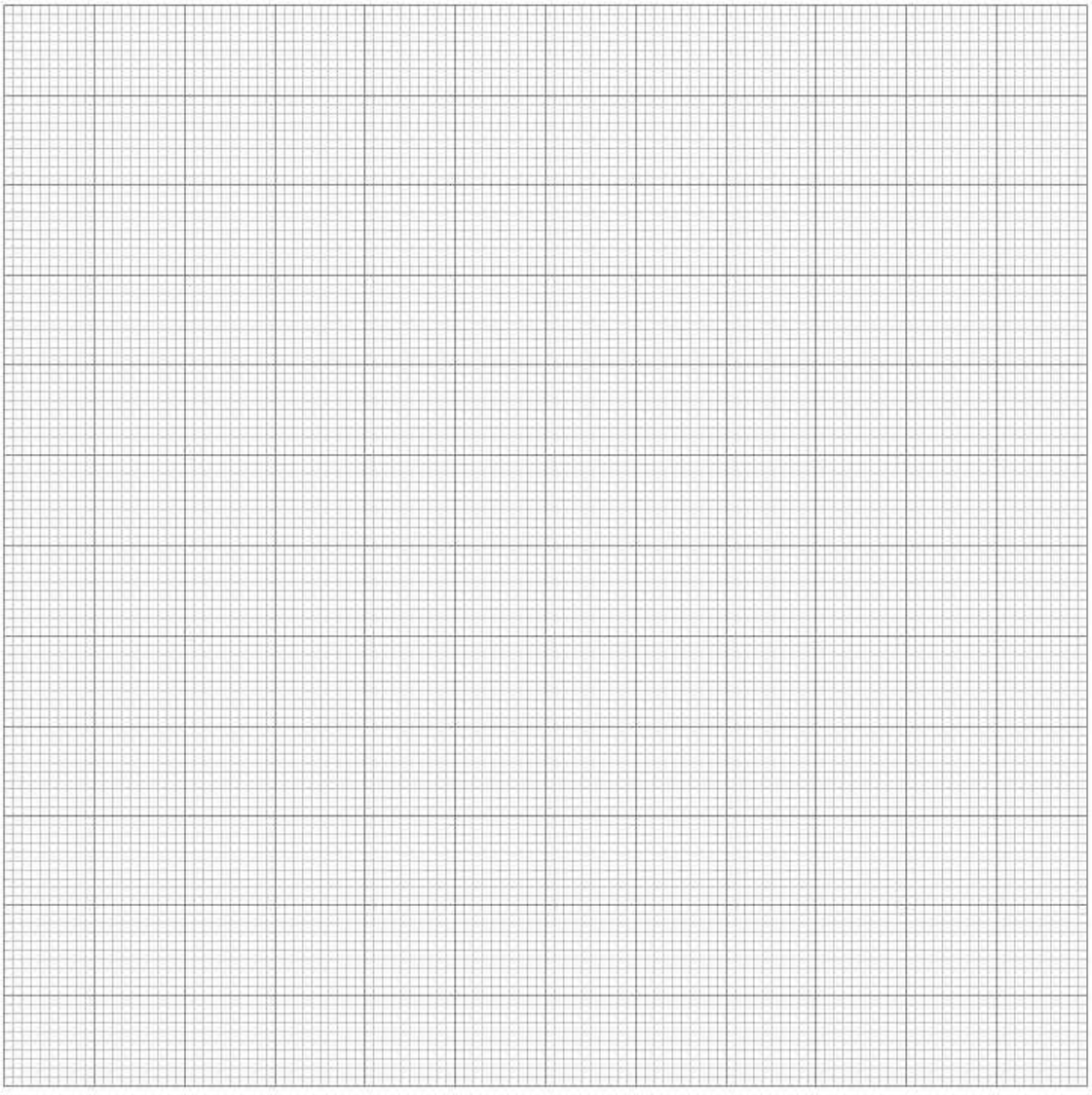
1 Record your times in the table below as you carry out the experiment. Record the times as accurately as possible from your stopwatch. [1]

Table 1

Expt no.	Volume of hydrochloric acid/cm^3	Volume of sodium thiosulfate/cm^3	Volume of water/cm^3	Time/s	Rate/s^{-1}
1	5	15	0		
2	5	12	3		
3	5	10	5		
4	5	8	7		
5	5	6	9		
6	5	4	11		
7	5	2	13		

2 When you have done all the experiments, complete the table by working out the rate of reaction for each experiment. [1]

3 Plot a graph showing the rate (*y*-axis) against the volume of sodium thiosulfate (*x*-axis). [4]



Conclusions

1 What is happening to the concentration of the sodium thiosulfate used from Experiment 1 to Experiment 7? [1]

.....

2 What does your graph tell you about how increasing the concentration of sodium thiosulfate affects the rate of a reaction? [2]

.....

.....

3 What was kept constant in this experiment to ensure a 'fair test'? [2]

.....

.....

4 Explain your findings in terms of particles. [4]

.....

.....

.....

.....

Evaluation

How could the procedure have been improved to give more accurate and reliable data? [2]

.....

.....

.....

Extension

At Advanced level Chemistry it would be stated that the reaction above is first order with respect to the concentration of sodium thiosulfate. Use your research skills to find out what 'first order' means. [2]

.....

8.1 Hydrated salts: how much water do they contain?

Aim

To determine the amount of water of crystallisation in a hydrated salt to obtain its correct formula.

Theory

Many salts produce crystals which contain some water locked inside their structure. These water molecules are referred to as water of crystallisation. Salts which contain water of crystallisation are called hydrated salts. This water can be removed from the hydrated salt by heating it. Once removed, the salt is known as an anhydrous salt – one which is without water.

hydrated magnesium sulfate → anhydrous magnesium sulfate + water

$MgSO_4.xH_2O(s) \rightarrow MgSO_4(s) + xH_2O(g)$

When the water is removed from the hydrated salt, the mass will decrease as the anhydrous salt is formed. By recording the change in mass as this process occurs and knowing the initial mass of the hydrated salt, the formula of the hydrated salt can be found.

We will be using hydrated magnesium sulfate, $MgSO_4.xH_2O$, in the experiment which is more commonly known as Epsom salts, a commonly used pharmaceutical.

Apparatus and chemicals

- ☐ eye protection
- ☐ crucible and lid
- ☐ tongs
- ☐ pipe-clay triangle
- ☐ tripod
- ☐ Bunsen burner
- ☐ heatproof mat
- ☐ balance
- ☐ hydrated magnesium sulfate (Epsom salts)

Safety!

While heating, the water being removed from the hydrated salt might spit out of the crucible.

The apparatus will be hot.

Hydrated magnesium sulfate – low hazard

Anhydrous magnesium sulfate – low hazard

Procedure

Throughout the practical the student should wear eye protection.

1. Find the mass of the crucible using the balance and record its mass in the results section.
2. Place three spatulas of the hydrated magnesium sulfate into your crucible and weigh it. Record the new mass in the results section.
3. Place the crucible on the pipe-clay triangle on the tripod.
4. Heat the crucible with a low, colourless Bunsen flame for about 5 minutes. If the contents of the crucible start spitting as the water is driven off, use the tongs and crucible lid to cover the crucible partially until it ceases.
5. Allow the crucible to cool and re-weigh it. Record the mass in the table.
6. Place the crucible back onto the pipe-clay triangle and heat it again for another 2 minutes, allow it to cool and record the new mass.
7. Repeat step 6 until the mass of the crucible and its contents does not change.

Method

1 Why was the crucible heated on a pipe-clay triangle rather than on a gauze? [2]

.....

.....

2 Why was a colourless flame used to heat the crucible? [2]

.....

.....

3 Why couldn't a lid be left on the crucible to prevent the spitting? [2]

.....

.....

4 Why was it important to keep heating the crucible until a constant mass was obtained? [2]

.....

.....

Results and calculations

1 Complete this section as you obtain the masses during the experiment. [2]

(a) Mass of crucible = ... g

(b) Mass of crucible and hydrated magnesium sulfate = ... g

(c) Mass after heating for 5 minutes = ... g

(d) Mass after second heating = ... g

(e) Mass after final heating = ... g

2 Calculate the mass of hydrated magnesium sulfate used. [2]

.....

.....

3 Calculate the mass of water removed by heating the hydrated magnesium sulfate. [2]

.....

.....

4 Calculate the mass of anhydrous magnesium sulfate. [2]

.....

.....

5 Determine the number of moles of anhydrous magnesium sulfate you had at the end of the experiment. (A_r: Mg = 24, S = 32, O = 16) [2]

.....

.....

6 Determine the number of moles of water driven off during the experiment. (A_r: H = 1, O = 16) [2]

.....

.....

7 Using the information from parts 5 and 6 work out the value of '*x*' in the formula $MgSO_4.xH_2O$. [2]

.....

.....

8 Write down the correct formula of hydrated magnesium sulfate. [1]

.....

Conclusions

1 The correct formula of hydrated magnesium sulfate is $MgSO_4.7H_2O$. Comment on your answer compared to this correct answer. [1]

.....

2 How could the procedure used in this experiment have caused your answer to be different? [1]

.....

Evaluation

How could the procedure have been improved to give more accurate and reliable data? [2]

.....

.....

.....

.....

Extension

Use your research skills to find a use for anhydrous magnesium sulfate. [2]

.....

.....

8.2 Determination of the concentration of a solution of hydrochloric acid

Aim

To find the accurate concentration of a solution of hydrochloric acid using the titration method.

Theory

The titration technique is used to react an acid with an alkali to produce a neutral solution. In this experiment we will use a solution of hydrochloric acid of unknown concentration and react it with a known volume of a solution of sodium hydroxide whose concentration is also known.

hydrochloric acid + sodium hydroxide → sodium chloride + water

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

From the balanced chemical equation, it can be seen that one mole of hydrochloric acid reacts with one mole of sodium hydroxide. Because we know the volume and concentration of the sodium hydroxide used and the volume of acid needed to neutralise the alkali, we will be able to work out the accurate concentration of the acid.

Apparatus and chemicals

- ☐ eye protection
- ☐ burette
- ☐ 25 cm^3 pipette and filler
- ☐ 250 cm^3 conical flask
- ☐ white tile
- ☐ hydrochloric acid solution with a concentration of between 0.075 and 0.125 mol/dm^3
- ☐ 0.1 mol/dm^3 sodium hydroxide solution
- ☐ phenolphthalein indicator

Safety!

When you are putting the pipette filler onto the pipette, make sure that you hold the pipette close to the end the filler will be attached to. This should stop the pipette from breaking as you push on the filler.

Hydrochloric acid solution (0.75–1.25 mol/dm^3) – low hazard

Sodium hydroxide solution (1 mol/dm^3) – corrosive

Phenolphthalein – low hazard

Procedure

Throughout the practical the student should wear eye protection.

1 Clamp a burette vertically and fill it with some of the hydrochloric acid solution, ensuring that some of the acid solution runs through the tap. Record the initial reading of the burette into Table 1 to two decimal places.

2 Using a pipette and filler, put 25 cm^3 of the 1 mol/dm^3 sodium hydroxide solution into a conical flask, on a white tile.

3 Add four drops of the indicator phenolphthalein to the sodium hydroxide solution in the conical flask. Make sure that you use the same number of drops each time.

4 Slowly add the acid from the burette, swirling the flask as you do, until the indicator *just* changes colour from pink to colourless. Record the final volume reading from the burette into the results table to two decimal places.

5 Repeat steps 2–4 until you have obtained three readings within 0.10 cm^3 of one another.

Method

1 Why is it important that the burette is clamped vertically? [1]

.....

2 What is the purpose of the white tile? [2]

.....

3 Why should the same number of drops of phenolphthalein be used each time? [1]

.....

.....

.....

4 Why is it important to add the acid slowly to the alkali? [2]

.....

.....

.....

Results and calculations

1 Complete the results table as you carry out your experiment. [2]

Table 1

	Rough	1	2	3	4
final burette reading/cm^3					
initial burette reading/cm^3					
volume of acid used/cm^3					

2 Work out an average volume (average titre) of acid used from the three results within 0.10 cm^3 of one another. [2]

Average volume (titre) = ... cm^3

3 Calculate the number of moles of sodium hydroxide used in each experiment. [2]

.....

.....

4 Find the number of moles of hydrochloric acid that would react with this number of moles of sodium hydroxide. [2]

.....

.....

5 Using the average volume of hydrochloric acid used and the number of moles of it used in the reaction, calculate the concentration of the acid solution. [3]

.....

.....

Conclusions

Does your answer fit in with the information given to you about the strength of the acid used? [2]

.....

.....

Evaluation

How could the procedure have been improved to give more accurate and reliable data? [1]

.....

.....

Extension

Use your research skills to find out how titrations are used in research. [2]

.....

.....

.....

.....

.....

.....

8.3 Preparation of hydrated magnesium sulfate

Aim

To prepare hydrated magnesium sulfate crystals from the reaction between sulfuric acid and magnesium carbonate.

Theory

Magnesium sulfate can be obtained from the reaction between magnesium carbonate and sulfuric acid:

magnesium carbonate + sulfuric acid → magnesium sulfate + water + carbon dioxide

$MgCO_3(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$

Once the magnesium sulfate solution is formed, the hydrated salt is obtained as a solid by evaporation of some of the water, allowing some to remain and be taken into the crystals as water of crystallisation. The hydrated crystals, once formed, can then be dried further.

Apparatus and chemicals

- ☐ eye protection
- ☐ 150 cm^3 beaker
- ☐ 25 cm^3 measuring cylinder
- ☐ glass rod
- ☐ spatula
- ☐ filter funnel
- ☐ 2 × filter paper
- ☐ evaporating basin
- ☐ tripod
- ☐ gauze
- ☐ Bunsen burner
- ☐ heatproof mat
- ☐ magnesium carbonate solid
- ☐ 1 mol/dm^3 sulfuric acid solution

Safety!

Be aware that much of the apparatus will be too hot to touch at the end of the experiment – allow it to cool.

Magnesium carbonate solid – irritant

Sulfuric acid (1 mol/dm^3) – irritant

Procedure

Throughout the practical the student should wear eye protection.

1 Using a measuring cylinder, measure out 25 cm^3 of sulfuric acid into the 150 cm^3 beaker.

2 Add the magnesium carbonate crystals, using the spatula, to the acid, stirring gently with a glass rod.

3 When the effervescence has stopped and you can see some magnesium carbonate crystals at the bottom of the beaker, stop adding the magnesium carbonate.

4 Filter the contents of the beaker, using a funnel and filter paper, allowing the filtrate to run into an evaporating basin.

5 Place the evaporating basin on a gauze on a tripod. Heat the contents of the evaporating basin, with a colourless Bunsen flame, until half of the volume of the basin has evaporated.

6 Place another piece of filter paper over the evaporating basin.

7 Leave the remainder of the solution to cool and allow crystallisation to occur over the next few days.

Method

1 Why did you stir the reaction mixture with the glass rod? [2]

.....

.....

2 What caused the effervescence? [1]

.....

.....

3 Why did you stop adding the magnesium carbonate when the effervescence stopped? [2]

.....

.....

4 Why is it important to be able to see some magnesium carbonate at the bottom of the beaker in step 3? [2]

.....

.....

5 Why was only half of the volume of the solution evaporated? [3]

.....

.....

6 What was the purpose of the filter paper being placed over the evaporating basin in step 6? [1]

.....

.....

Results

What colour are the hydrated magnesium sulfate crystals you have produced? [1]

.....

Conclusions

1 Is hydrated magnesium sulfate a soluble or insoluble salt? ..... [1]

2 What would have been formed if all of the water had been evaporated from the evaporating basin? What would it look like? [2]

.....

.....

3 Why is water needed to allow the crystals to form? [2]

.....

.....

Evaluation

Think of some ways in which this procedure could have been improved. State why your suggestion would improve the procedure. [2]

.....

.....

.....

Extension

1 Think of three other ways, using other salt preparation techniques from your course, in which you could have made crystals of hydrated magnesium sulfate. [3]

.....

.....

2 For each of the three ways, write a balanced chemical equation to show the formation of magnesium sulfate. [6]

.....

.....

.....

9.1 Reactions of the Group 1 metals

Aim

To demonstrate the reactivity trend of the Group 1 metals lithium, potassium and sodium by looking at their reaction with water. To be able to use the trend obtained to predict the reactions of the other Group 1 metals.

Theory

The metals in Group 1 of the Periodic Table are known as the alkali metals. This is because when they react with water they all form the metal hydroxide, producing alkaline solutions. For example, when lithium reacts with water, lithium hydroxide is formed.

lithium	+	water	→	lithium hydroxide	+	hydrogen
2Li(s)	+	$2H_2O(l)$	→	2LiOH(aq)	+	$H_2(g)$

The three alkali metals available for use in school are lithium, sodium and potassium. They are all very reactive metals and are stored under oil to prevent them from reacting with water and oxygen in the air.

All of the Group 1 metals are soft metals and are easily cut with a knife. When they are cut, they show a bright silvery colour which rapidly tarnishes on exposure to the air.

Apparatus and chemicals

- ☐ eye protection
- ☐ safety screen
- ☐ 3 × water troughs
- ☐ 3 × scalpels
- ☐ 3 × tweezers
- ☐ 3 × boiling tubes
- ☐ 3 × dropping pipettes
- ☐ white tile
- ☐ universal indicator solution and chart
- ☐ boiling-tube rack
- ☐ filter paper
- ☐ sodium metal
- ☐ lithium metal
- ☐ potassium metal

Safety!

All of these metals are very reactive with water hence place a safety screen close to the trough with students 3 m away. Only small pieces of the metals should be used. The use of large pieces of the metals may result in explosions.

Lithium metal – highly flammable, corrosive

Sodium metal – highly flammable, corrosive

Potassium metal – highly flammable, corrosive

Solutions of the hydroxides of sodium, lithium and potassium – likely to be an irritant at the concentrations produced in the demonstration

Procedure

Throughout the practical the teacher should wear eye protection.

1 Half fill all three of the water troughs and place them side by side on the bench, behind a safety screen.

2 Using tweezers, take out a piece of lithium metal from its container and use a scalpel to cut off a small piece, not more than 0.5 cm^3. Replace the remainder of the lithium metal back into its container, into oil.

3 Show the students the newly cut surface of the metal. They should record their observations in the results table.

4 Use a piece of filter paper to remove some of the oil from the piece of metal you have cut off.

5 Use the tweezers to drop the lithium into the middle of the water in one of the water troughs. Students should record their observations in the results table.

6 When the reaction is finished, remove some of the solution formed using a dropping pipette into a boiling tube.

7 Add a few drops of universal indicator solution to the solution in the tube and ask students to record colour and the pH of the solution formed in the results table.

8 Repeat steps 1–7 for potassium and sodium, using different water troughs each time.

Method

1 What is the purpose of the safety screen? [1]

.....

.....

2 Why should the metal which is *not* going to be used in the experiment be returned to its container? [2]

.....

.....

3 What happened to the newly cut surfaces of the metals you were shown? Explain why this happens. [3]

.....

.....

.....

4 Why was only a small piece of each metal used? [1]

.....

.....

5 Why was the filter paper used to remove most of the oil from the surface of each metal? [2]

.....

.....

.....

6 Why was each metal placed into the water at the centre of the trough? [2]

.....

.....

.....

.....

Results

Complete this table as your teacher carries out the demonstration. [3]

Table 1

Metal	Colour of metal	Observations when added to water	Colour of solution formed with universal indicator solution	pH of the solution formed
lithium				
potassium				
sodium				

Conclusions

1 From your observations, what can be said about the density of the three metals? [2]

.....

.....

2 Which of the metals was, based on your observations, the most reactive? [1]

.....

3 Which of the metals was, based on your observations, the least reactive? [1]

.....

4 Place the three metals in order of reactivity with the least reactive first. [1]

.....

5 Explain this trend. [3]

.....

.....

.....

.....

6 Write a balanced chemical equation for the reaction of potassium metal with water. [3]

.....

7 Why do all three of the solutions produced have a pH above 7? [2]

.....

.....

8 Why were the other three Group 1 metals, francium, caesium and rubidium, not used in the experiment? [1]

.....

.....

9 Which of francium, caesium and rubidium would be the most reactive? Explain how you know this. [2]

.....

.....

Evaluation

1 In order to describe the trend in reactivity of the three metals used in the experiment you used your observations. Did the procedure used give clear evidence of the trend you observed? [2]

.....

.....

2 Think of another way in which numeric data could be produced to show the trend. [3]

.....

.....

Extension

1 Use your research skills to find videos showing the effect of putting pieces of caesium, rubidium and francium into water. Give the web addresses of the sites you used. [3]

.....

.....

.....

2 Find a use for each of the first three Group 1 metals, or their compounds. [3]

.....

.....

.....

9.2 Halogen displacement reactions

Aim

Displacement reactions can be used to determine the reactivity of the Group 7 elements, the halogens. In this experiment you will use these reactions to determine an order of reactivity for iodine, chlorine and bromine. You will then use this to predict the reactivity of the other two halogens, fluorine and astatine.

Theory

When a more reactive metal reacts with a compound of a less reactive metal, a displacement reaction will occur with the more reactive metal displacing the less reactive metal from the salt. The same idea works for the Group 7 elements. If a more reactive halogen reacts with a compound of a less reactive halogen, the less reactive halogen will be displaced and it will form the halogen molecule, while the more reactive halogen becomes a halide ion.

To observe the presence of the different halogens in solution, it is best to use cyclohexane as the halogens produce more vivid and unique colours in this solvent compared to water.

In the first procedure, you will find out the colours of the chlorine, bromine and iodine in both water and cyclohexane.

In the second procedure, you will use these colours to determine whether a reaction has occurred when a displacement reaction is carried out.

Apparatus and chemicals

- ☐ eye protection
- ☐ 12 × test tubes
- ☐ 12 × test-tube bungs
- ☐ 24 × dropping pipettes
- ☐ test-tube rack
- ☐ chlorine water
- ☐ bromine water
- ☐ iodine water
- ☐ 1 mol/dm^3 sodium chloride solution
- ☐ 1 mol/dm^3 sodium bromide solution
- ☐ 1 mol/dm^3 sodium iodide solution
- ☐ cyclohexane

Safety!

Chlorine water – low hazard

Bromine water – toxic so vapours should not be inhaled

Iodine water – harmful

Sodium chloride, sodium bromide and sodium iodide solutions (1 mol/dm^3) – low hazard

Cyclohexane – highly flammable

Procedure 1

Throughout the practical the student should wear eye protection.

1 Use a dropping pipette to put a 1 cm depth of chlorine water into a test tube. Record the colour of the solution in Table 1.

2 Using a different pipette, add the same depth of cyclohexane to the tube. Put a bung on the tube and shake the contents gently.

3 After shaking, you will see two layers forming. The lower layer is the water (aqueous layer) and the upper layer is the cyclohexane layer. Record the colour of the halogen in the cyclohexane layer in the results table.

4 Repeat steps 1–3 for the other two halogens.

Procedure 2

Throughout the practical the student should wear eye protection.

You are now ready to use the colour of the halogens in the cyclohexane layer to determine the presence of the halogens after a displacement reaction has occurred.

1 Using a new test tube and dropping pipette, place 1 cm depth of chlorine water into the tube.

2 Using a different pipette, add into the same tube a 1 cm depth of sodium iodide solution.

3 Now add a 1 cm depth of cyclohexane. Bung the tube and gently shake it.

4 Record the colour of the upper and lower layers in Table 2.

5 Repeat steps 1–4 using the solution combinations shown below.
   (a) bromine water and sodium iodide
   (b) chlorine water and sodium bromide
   (c) iodine water and sodium bromide
   (d) bromine water and sodium chloride
   (e) iodine water and sodium chloride

6 Record the colours of the upper and lower layers in each tube in the table.

The contents of all your test tubes should be emptied into a beaker of sodium thiosulfate solution in a fume cupboard, *not* poured down the sink.

Method

1 Why is it important to use a bung on the tube, rather than putting your thumb over the open end before shaking? [2]

.....

.....

2 Why were you given solutions of the halogens to use, rather than the halogens themselves? [3]

.....

.....

.....

3 Why was fluorine not used in this experiment? [1]

.....

4 Why was astatine not used in this experiment? [2]

.....

5 What do we call solutions which do not mix together, instead forming two layers, such as water and cyclohexane? [1]

.....

Results

Complete the results table as you carry out Procedure 1. [3]

Table 1

Halogen	Colour in water	Colour in cyclohexane
chlorine		
bromine		
iodine		

Complete this results table as you carry out Procedure 2. [6]

Table 2

		Chlorine water	Bromine water	Iodine water
colour after shaking with sodium iodide solution				
colour of each layer with cyclohexane	Upper			
	lower			
colour after shaking with sodium bromide solution				
colour of each layer with cyclohexane	upper			
	lower			
colour after shaking with sodium chloride solution				
colour of each layer with cyclohexane	upper			
	lower			

Conclusions

1 In how many of the six reactions you carried out has a displacement reaction occurred? [1]

.....

2 For each of these displacement reactions, write a balanced chemical equation. [3]

.....

.....

.....

3 What type of chemical change has happened to the halogen molecules in the displacement reactions? [2]

.....

4 State the order of reactivity of the three halogens used in the experiment. [1]

.....

5 What would be the colour of fluorine and astatine in water? [2]

.....

.....

6 Which of all the five halogen elements is the most reactive? Explain your answer. [2]

.....

.....

Evaluation

1 Did the experiment clearly show you the order of reactivity of the three halogens used? [1]

.....

.....

2 Could the procedures used have been improved? [1]

.....

.....

Extension

1 Use your research skills to find out a use for each of the five halogens or their compounds. [5]

.....

.....

.....

2 Can you think of another chemical which can be used to determine the presence of iodine in solution? Explain what you would see. [2]

.....

.....

9.3 Using transition metal ions as catalysts

Aim

To find d-block metal compounds which will act as catalysts for the iodine clock reaction.

Theory

The iodine clock reaction is a reaction between potassium iodide and potassium peroxodisulfate.

potassium iodide + potassium peroxodisulfate → potassium sulfate + iodine

$2KI(aq) + K_2S_2O_8(aq) \rightarrow 2K_2SO_4(aq) + I_2(aq)$

When the reaction occurs, iodine is formed. In the presence of iodine indicator (or starch), the production of the iodine results in the formation of a dark blue/black colour.

The reaction is, however, so fast that another chemical, sodium thiosulfate, is added to slow down the formation of the blue/black colour.

You will first carry out the iodine clock reaction without a catalyst and then add the five potential catalyst solutions, one at a time, to the reaction mixture, to see whether they speed up the formation of the blue/black colour and increase the rate of the reaction.

The rate of the reaction can be found using:

Rate (/s) = 1/time (s)

because the same amount of iodine is being produced each time to bring about the formation of the blue/black colour.

Apparatus and chemicals

- ☐ eye protection
- ☐ 12 × boiling tubes
- ☐ 18 × 10 cm^3 measuring cylinders
- ☐ boiling-tube rack
- ☐ spatula
- ☐ stopwatch
- ☐ 5 × dropping pipettes
- ☐ 0.5 mol/dm^3 potassium iodide solution
- ☐ 0.04 mol/dm^3 potassium peroxodisulfate solution
- ☐ 0.01 mol/dm^3 sodium thiosulfate solution
- ☐ iodine indicator (or freshly made starch solution)
- ☐ 0.1 mol/dm^3 iron(III) sulfate solution
- ☐ 0.2 mol/dm^3 ammonium molybdate solution
- ☐ 0.2 mol/dm^3 chromium(III) chloride solution
- ☐ 0.2 mol/dm^3 iron(II) sufate solution
- ☐ 0.2 mol/dm^3 copper(II) sulfate solution

Safety!

Potassium iodide solution (1 mol/dm³) – low hazard

Potassium peroxodisulfate solution (0.04 mol/dm³) – harmful

Sodium thiosulfate solution (0.01 mol/dm³) – low hazard

Iodine indicator (or freshly made starch solution) – low hazard

Iron(III) sulfate solution (0.1 mol/dm³) – low hazard

Ammonium molybdate solution (0.2 mol/dm³) – harmful

Chromium(III) chloride solution (0.2 mol/dm³) – low hazard

Iron(II) sulfate solution (0.2 mol/dm³) – low hazard

Copper(II) sulfate solution (0.2 mol/dm³) – harmful

Procedure

Throughout the practical the student should wear eye protection.

1. Using a measuring cylinder, measure out 5 cm^3 of the potassium iodide solution and pour it into a boiling tube.
2. Into the same boiling tube, but using a different measuring cylinder, pour in 2 cm^3 of the sodium thiosulfate solution.
3. Into the same boiling tube, add a spatula full of iodine indicator powder (or 1 cm^3 of starch solution).
4. Into a different boiling tube, using another measuring cylinder, place 2 cm^3 of potassium peroxodisulfate solution.
5. When you are ready to start the experiment, pour the tube containing the potassium peroxodisulfate solution into the other tube, swirl once, place it back into the rack and start the stopwatch.
6. Stop the stopwatch when you observe the formation of the blue/black colour. Record the time in the results table, as accurately as your stopwatch allows. This is the time for the uncatalysed reaction.
7. You will now carry out exactly the same experiment, but this time you will add 10 drops of a potential catalyst solution to the tube containing the potassium peroxodisulfate, using a dropping pipette. Each time, record the time taken for the blue/black colour to form in the results table.

Method

1 Why is it important to use different measuring cylinders for each of the solutions? [2]

.....

.....

2 Why is the concentration of the iron(III) sulfate half that of the other catalysts? [2]

.....

.....

.....

.....

3 Why is the potassium peroxodisulfate kept separate from the other chemicals in a different boiling tube? [2]

.....

.....

4 Why is it important to swirl the tube once only for each experiment? [2]

.....

.....

Results and calculations

1 Complete the results table as you carry out the experiment. [2]

Table 1

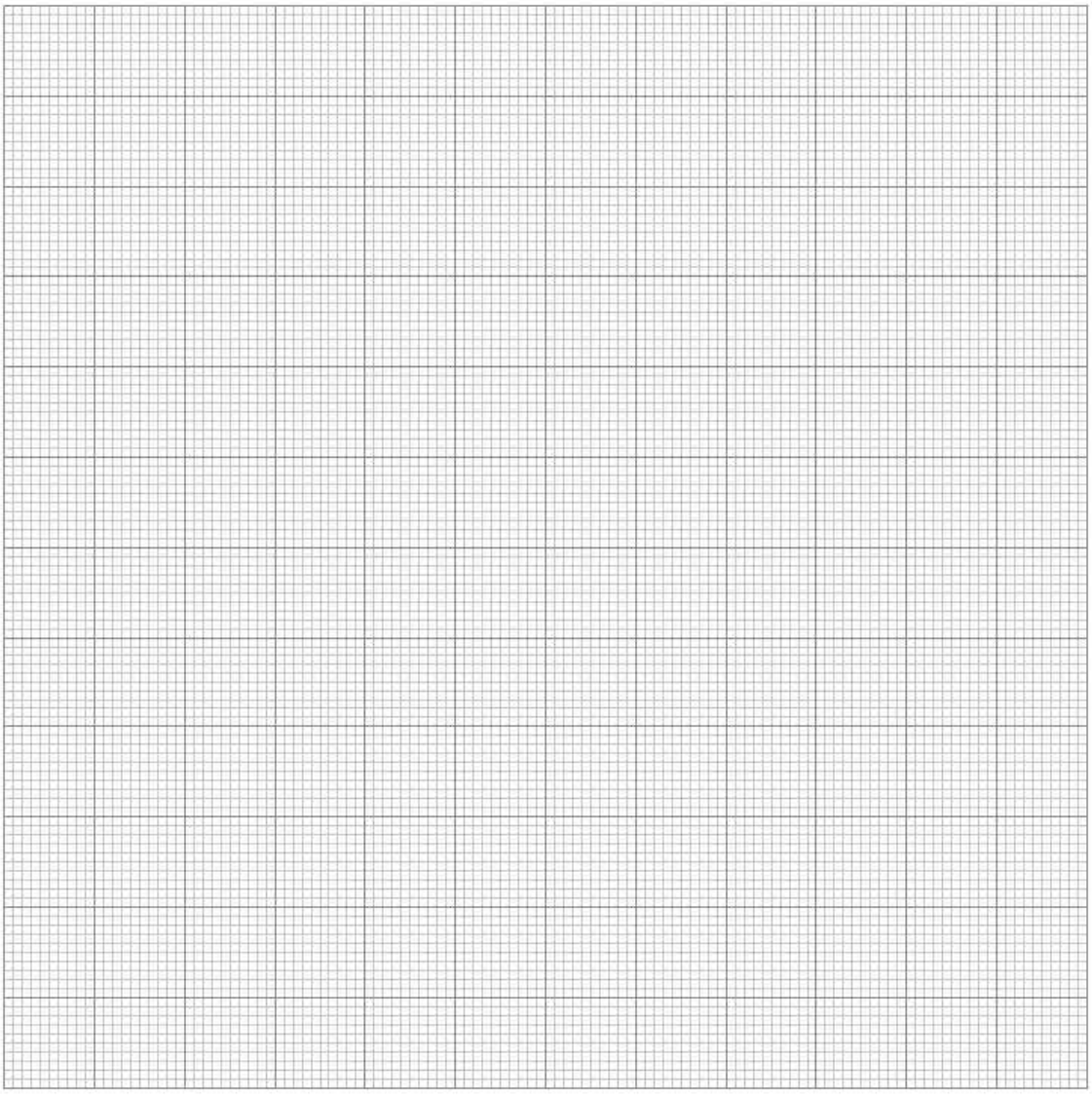
Catalyst	Time (s)	Rate (/s)
uncatalysed reaction		
iron(III) sulfate		
ammonium molybdate		
chromium(III) chloride		
iron(II) sulfate		
copper(II) sulfate		

2 Work out the rate for each of the six experiments. [2]

.....

.....

3 Plot a bar chart to show the rate of reaction for each of the six reactions you have carried out. [4]



Conclusions

1 Which of the d-block metal compounds acted as a catalyst for the iodine clock reaction? [1]

.....

2 Which of the compounds was the best catalyst? [1]

.....

3 Did any of the compounds slow the rate of the reaction down? [1]

.....

4 Apart from speeding up a chemical reaction, what is the other important factor when defining a compound as a catalyst? [1]

.....

5 How does a catalyst increase the rate of a chemical reaction? [4]

.....

.....

.....

.....

.....

Evaluation

How could the procedure have been improved to get more accurate results? For each of your suggestions, state why they would improve the procedure. [4]

.....

.....

Extension

Use your research skills to find out which d-block metals or d-block metal compounds are used in the manufacture of ammonia gas, sulfuric acid, margarine and nitric acid. [4]

.....

.....

.....

.....

10.1 Metal displacement reactions

Aim

To use displacement reactions to place six metals, magnesium, copper, iron, tin, lead and iron, into an order of reactivity.

Theory

A more reactive metal will displace a less reactive metal from its salt. For example, copper is a more reactive metal than silver and so would displace silver from its salt.

copper + silver nitrate → copper(II) nitrate + silver

$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$

A reaction is known to have occurred because a colour change is observed. As the reaction proceeds, the solution turns blue as copper(II) nitrate is formed and the pink colour of copper metal is replaced by a more silvery solid product, silver. The change in colour is a way of indicating whether a reaction has occurred or not.

Apparatus and chemicals

- ☐ eye protection
- ☐ dimple tile (or test tubes)
- ☐ 6 × dropping pipettes
- ☐ cleaned small pieces (0.5 cm × 0.5 cm) of magnesium, copper, iron, tin, lead and zinc
- ☐ solutions (1 mol/dm^3) of the nitrates (or other suitable solutions) of magnesium, copper, iron, tin, lead and zinc

Safety!

Lead – toxic

Lead(II) nitrate – toxic

Copper(II) nitrate – harmful

Procedure

Throughout the practical the student should wear eye protection.

1 Using a dropping pipette, fill up six of the dimples in the tile (or 1 cm depth in six test tubes) with magnesium nitrate solution.

2 Add a small piece of each metal to separate dimples (or test tubes). Ensure that the metal is pushed under the solution.

3 Leave the reaction mixtures for 5 minutes.

4 Look at each of the reaction mixtures. Where there has been a change of colour on the surface of the metal or in the colour of the solution, you can assume that a reaction has occurred. In this case, place a tick (✓) in the results table. Where no colour change has occurred, no reaction will have happened so place a cross (✗) in the results table.

5 Repeat the process for the other five nitrate solutions.

Method

1 Why is it important that the metals have been cleaned? [2]

.....

.....

2 Why have the metals potassium and lithium not been used in the experiment? [1]

.....

3 Why is the metal pushed under the surface of the solution? [2]

.....

.....

4 Why is 5 minutes allowed before the results are recorded? [1]

.....

.....

Results

Complete this table as you carry out your experiments. [6]

Table 1

	Metal nitrate solutions					
Metals	**Magnesium**	**Copper**	**Iron**	**Tin**	**Lead**	**Zinc**
Magnesium						
Copper						
Iron						
Tin						
Lead						
Zinc						

Conclusions

1 Of the metals used, which reacted most often? [1]

.....

2 Of the metals used, which reacted least often? [1]

.....

3 Using your results, place the metals in a list with the most reactive metal first. [1]

.....

4 Write a balanced chemical equation for the reaction between magnesium and copper(II) sulfate. [2]

.....

5 In all the displacement reactions which have occurred, what chemical process is happening to each of the more reactive metals? State why you have reached this conclusion. [2]

.....

.....

6 Write an ionic half-equation to show this change for copper in its reaction with silver nitrate solution. [3]

.....

7 What is happening to each of the less reactive metals in the displacement reactions? Explain your answer. [2]

.....

.....

8 Write an ionic half-equation to show this change for the silver present in silver nitrate in its reaction with copper metal. [3]

.....

9 Why is one metal more reactive than another in this type of reaction? [2]

.....

.....

Evaluation

1 Did the results of your experiment match the order of reactivity of metals shown in your chemistry textbook? [1]

.....

2 If one or two of the reactions which gave a 'wrong' result can you give a reason? [1]

.....

3 How could the procedure have been improved to give results which follow the accepted order of reactivity of the metals? [2]

.....

.....

Extension

When displacement reactions occur between more reactive halogen molecules and less reactive halogens in salts, does the same chemical change happen to the most reactive halogen as occurred to the most reactive metal which you have mentioned above? [3]

.....

.....

.....

.....

10.2 Rusting of iron

Aim

To plan an experiment which can be used to show the conditions needed for rusting to occur.

Theory

After a period of time, objects made of iron or steel become coated with rust. The rusting of iron is a serious problem and wastes enormous amounts of money each year. Estimates are difficult to make, but it is thought that upwards of £1 billion a year is spent worldwide on replacing iron and steel structures.

Rust is an orange–red powder consisting mainly of hydrated iron(III) oxide ($Fe_2O_3.xH_2O$). Both water and oxygen are essential for iron to rust and, if one of these two substances is not present, then rusting will not take place. Unfortunately, both of these substances are present in air and in water. It is possible to remove oxygen from water by boiling water for a few minutes.

Rusting is actually speeded up in the presence of salt solution as this helps the movement of electrons in the rusting process, which is a redox reaction.

Rusting of iron can be prevented by putting the iron in contact with a more reactive metal such as zinc.

Apparatus and chemicals

- ☐ eye protection
- ☐ 5 × boiling tubes
- ☐ bungs for the boiling tubes
- ☐ tripod
- ☐ gauze
- ☐ Bunsen burner
- ☐ 5 × iron nails
- ☐ emery paper
- ☐ glass wool
- ☐ cooking oil
- ☐ anhydrous calcium sulfate (drying agent)
- ☐ distilled water
- ☐ solid sodium chloride
- ☐ thin strips of zinc metal

Safety!

Boiling water – hot

Anhydrous calcium sulfate – harmful

Cooking oil – flammable and should be kept away from the Bunsen flame

Sodium chloride – low hazard

Procedure 1

1 Using the information given in the theory section and your knowledge from your chemistry course about rusting, plan a procedure to show that:
   (a) both water and oxygen are needed for rusting to occur
   (b) without oxygen rusting does not occur
   (c) without water rusting does not occur
   (d) rusting speeds up in the presence of salt solution
   (e) a more reactive metal attached to the iron nail will slow down rusting.

.....

.....

.....

.....

.....

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.....

.....

.....

2 When you have written down your procedure, check it with your teacher before you carry it out.

Procedure 2

Throughout the practical the student should wear eye protection.

1 Using your procedure (which has been checked by your teacher) set up the five boiling tubes to show the five points above.

2 You will need to leave the finished tubes a few days for the rust to start to form before they can be checked and your observations entered into the results table.

Method

1 How have you ensured that in one tube no oxygen gas will be present? How does your method remove the oxygen gas? [2]

.....

.....

.....

2 How have you ensured that one of the tubes has no water present? How does your method remove the water from the tube? [2]

.....

.....

.....

Results

Write your observations in the table a week after you set up the tubes. [5]

Table 1

Tube		Observation
1	water and oxygen	
2	without oxygen	
3	without water	
4	with salt solution	
5	with zinc	

Conclusions

1 In which tube(s) did rusting occur? [1]

.....

2 Why did rusting occur in these tubes? [2]

.....

.....

3 In which tube(s) did rusting *not* occur? Explain why. [4]

.....

.....

.....

.....

.....

.....

4 In which tube did most rusting occur? [1]

.....

5 It was stated earlier that rusting is a redox process. When iron rusts, the iron changes into iron(III) ions to form iron(III) oxide. What type of chemical change is this? Write an ionic half-equation to show this change. [4]

.....

.....

6 What type of chemical change does the oxygen undergo to form oxide ions, also present in the iron(III) oxide?. Write an ionic half-equation to show this change. [4]

.....

.....

Evaluation

Did you make any changes to your initial procedure as you carried out the practical itself? If so, state what you did differently and why. [2]

.....

.....

Extension

1 Only iron undergoes rusting. What is the reaction called when other metals react with gases in the environment? [1]

.....

2 Name two metals, other than zinc, which could protect the iron nail from rusting simply by being in contact with it. [2]

.....

10.3 Metal reactivity

Aim

To use the reaction of metals with acid to produce an order of reactivity for metals.

Theory

The majority of the elements we know are metals. Metallic elements show some common properties, as well as having characteristic physical properties. They are also very different in other ways. Iron, for example, will rust quickly if left unprotected, while gold remains totally unchanged after hundreds of years. Iron is said to be reactive in comparison to gold, which is unreactive.

In this experiment, you will be given six different metals and some hydrochloric acid solution. You will use the reaction between the metal and acid to produce an order of reactivity for the six metals. Metals react with hydrochloric acid solution to give the metal chloride and hydrogen gas.

metal + hydrochloric acid → metal chloride + hydrogen

You will need to ensure that you use the same mass of each metal in each experiment.

Apparatus and chemicals

- ☐ eye protection
- ☐ 5 × boiling tubes
- ☐ metal cutters
- ☐ 25 cm^3 measuring cylinder
- ☐ stopwatch
- ☐ access to powders of magnesium, copper, lead, iron and zinc
- ☐ 2 mol/dm^3 hydrochloric acid solution

Safety!

Hydrochloric acid (2 mol/dm^3) – irritant

Lead metal – toxic

Procedure

Throughout the practical the student should wear eye protection.

1 Using the measuring cylinder, put 20 cm^3 of the hydrochloric acid into a boiling tube.

2 Starting with any one of the five metals, weigh out 0.2 g of the metal powder.

3 When you are ready to start the experiment, add the metal powder to the acid and start the stopwatch. Swirl the contents of the tube once.

4 Record the time it takes for all of the metal to react. If it takes over 5 minutes record this in your table. Record your observations for each metal in the table.

5 Repeat steps 1–4 for the other four metals.

Method

What did you do in the experiment to ensure that it was a fair test? [3]

.....

.....

.....

Results

Complete the table as you carry out the experiment. [5]

Table 1

Metal	Observation(s)	Time (s)
magnesium		
copper		
lead		
iron		
zinc		

Conclusions

1 Which of the metals did not react at all with the hydrochloric acid? [1]

.....

2 Which of the metals was the most reactive? [1]

.....

3 Using your results, write down the order of reactivity you have obtained, starting with the least reactive metal. [1]

.....

4 Write a balanced chemical equation for the reaction between zinc and hydrochloric acid. [3]

.....

Evaluation

How could the procedure be improved to give clearer results? [4]

.....

.....

.....

Extension

1 Use your research skills to find the names of three metals which would have been more reactive than the most reactive metal you used. [3]

.....

.....

.....

2 Use your research skills to find the names of three metals which would also have *not* reacted with the hydrochloric acid. [3]

.....

.....

.....

11.1 The active part of the air

Aim

To determine the percentage of oxygen in a sample of air.

Theory

The gas oxygen is known as the most active or reactive part of the air we breathe. It will react with many metals when heated with them. For example, magnesium, when heated in air, produces magnesium oxide as the product.

magnesium + oxygen → magnesium oxide

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

In this experiment, the percentage of oxygen in the air will be found by passing air over heated copper metal.

Apparatus and chemicals

- ☐ eye protection
- ☐ large gas syringes
- ☐ retort stands complete with bosses and clamps
- ☐ 15 cm heat-resistant glass tube
- ☐ 100 cm^3 beaker
- ☐ glass wool
- ☐ Bunsen burner
- ☐ safety screen
- ☐ electronic balance
- ☐ candle
- ☐ copper powder/granules/shavings

Safety!
Copper metal – low hazard

Procedure

1 Put on eye protection.

2 Students should fill in the results section as the demonstration proceeds.

3 Place copper powder/granules/shavings into the heat-resistant glass tube. Place plugs of glass wool in either end.

4 Weigh the glass tube containing the copper powder/granules/shavings.

5 Set up the apparatus as shown in Figure 1 with 100 cm^3 of air in one of the syringes, while the other should read 0 cm^3. Place a safety screen in front of this apparatus.

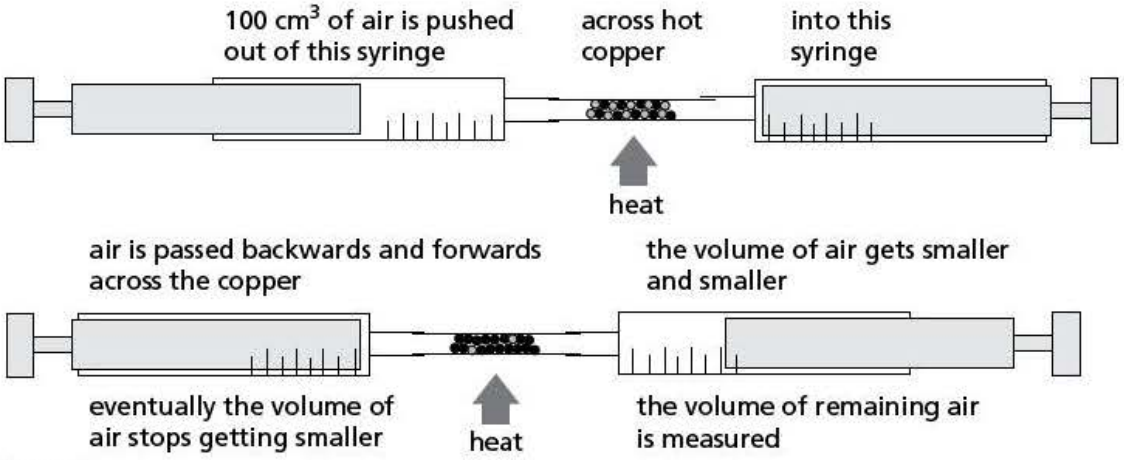


Figure 1

6 Heat the copper strongly for 5 minutes while the air is passed backwards and forwards from syringe to syringe.

7 Leave the apparatus to cool.

8 Reweigh the glass tube.

9 Measure the final volumes on the gas syringes with one at 0 cm^3.

10 Calculate the percentage of oxygen in the sample of air.

11 Set up a small lighted candle in a 100 cm^3 beaker and squeeze the remaining gas carefully onto it.

Method

1 Why is there a volume decrease during the experiment? [1]

.....

2 Why is there a change in colour of the copper? [1]

.....

3 What causes the increase in mass of the copper? [1]

.....

4 Write word and chemical equations for the reaction which has caused the changes in colour and mass of the copper sample in the glass tube. [3]

.....

.....

.....

Results and calculations

Initial colour of the copper is .....

Final colour of the copper after heating in air is .....

Mass of glass tube + copper at start = ... g

Mass of glass tube + copper oxide after heating = ... g

Difference in mass = ... g

Initial volume of air in the syringe = ... cm^3

Final volume of gas left in the syringe after the apparatus has

cooled = ... cm^3

Volume change = ... cm^3

Work out the percentage of oxygen gas in air using the following equation:

$$\text{percentage of oxygen gas in the air} = \frac{\text{volume change}}{100} \times 100$$

What happened to the candle in the beaker when the remaining gas from the cooled syringe was squeezed gently onto it? [10]

.....

.....

.....

.....

Conclusion

Using the apparatus shown in Figure 1 the percentage of oxygen in the air is

.....%. [1]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Dry air is a mixture of gases. Find out the percentages of the components of this mixture. [4]

.....

.....

.....

.....

.....

.....

.....

.....

.....

2 Air, or to be more precise oxygen, is necessary for our existence on this planet. We use oxygen in respiration. Write word and balanced chemical equations for this process. [3]

.....

.....

→

.....

.....

3 How can air be separated into its different gas components on a large scale? [2]

.....

.....

11.2 Making a fertiliser

Aim

To make a sample of a nitrogenous fertiliser.

Theory

The Haber process produces ammonia (NH_3). Some of this ammonia is used to produce ammonia solution (aqueous ammonia, $NH_3(aq)$). If ammonia is then reacted with sulfuric acid, we have the basic neutralisation reaction for the production of many artificial fertilisers. The use of artificial fertilisers has increased over the years. This is because there is an ever-increasing world population that has to be fed and so an increase in crop production is needed.

Crops remove nutrients from the soil as they grow; these include nitrogen, phosphorus and potassium. Artificial fertilisers are added to the soil to replace these nutrients.

Ammonium sulfate is a widely used nitrogenous fertiliser. It is manufactured by the following reaction.

ammonia + sulfuric acid → ammonium sulfate

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

You are going to make a sample of this fertiliser.

Apparatus and chemicals

- ☐ eye protection
- ☐ burette
- ☐ burette stand
- ☐ 2 × 25 cm^3 measuring cylinders
- ☐ 250 cm^3 conical flasks
- ☐ white tile
- ☐ filter funnel
- ☐ glass rod
- ☐ evaporating basin
- ☐ Bunsen burner
- ☐ tripod
- ☐ gauze
- ☐ heat-resistant mat
- ☐ 2 × 400 cm^3 beakers
- ☐ 0.05 M sulfuric acid
- ☐ 1 M aqueous ammonia
- ☐ phenolphthalein

Safety!

0.05 mol/dm^3 sulfuric acid – low hazard

1 M aqueous ammonia solution – low hazard, dangerous for the environment

Phenolphthalein – low hazard

Procedure

1 Put on your eye protection.

2 Pour dilute aqueous ammonia into a 25 cm^3 measuring cylinder and pour this exact amount into a conical flask to which a few drops of phenolphthalein indicator have been added. Phenolphthalein is pink in alkaline conditions but colourless in acid.

3 A 0.05 mol dm^{-3} solution of sulfuric acid is placed in the burette using a filter funnel until it is filled up exactly to the zero mark.

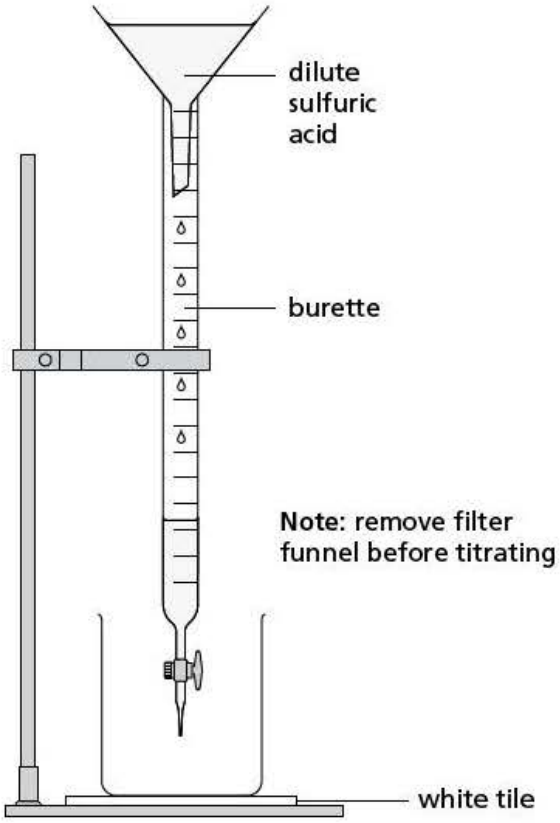


Figure 1

4 The filter funnel is now removed.

5 The sulfuric acid is added to the aqueous ammonia solution in small quantities – usually no more than 0.5 cm^3 at a time (Figure 2). The contents of the flask must be swirled after each addition of acid for thorough mixing.

6 The acid is added until the aqueous ammonia has been neutralised completely. This is shown by the pink colour of the indicator *just* disappearing.

7 Take the final reading on the burette at the end-point (just as neutralisation takes place). This value should be recorded.

8 To obtain crystals of ammonium sulfate without the phenolphthalein present, the following procedure is followed.

(a) To 25 cm^3 of aqueous ammonia add, with stirring, the volume of sulfuric acid you needed in the titration in a 400 cm^3 beaker.

(b) Transfer half of this solution to an evaporating basin.

(c) Half fill a 400 cm^3 beaker with water and stand it on a tripod and gauze.

(d) Place the evaporating basin on top of this beaker as shown in Figure 3, light the Bunsen burner and evaporate the filtrate slowly until there is only a small amount of solution left.

(e) Set aside and leave to crystallise.

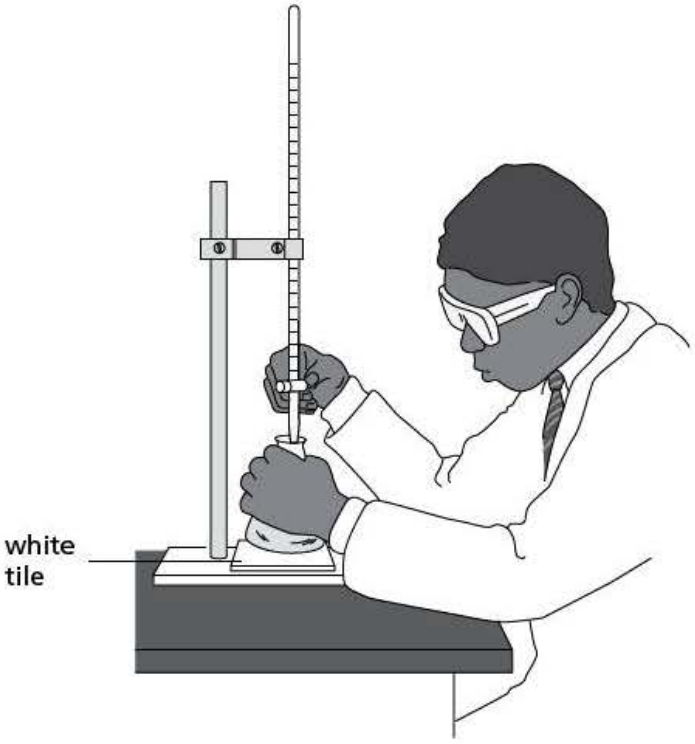


Figure 2

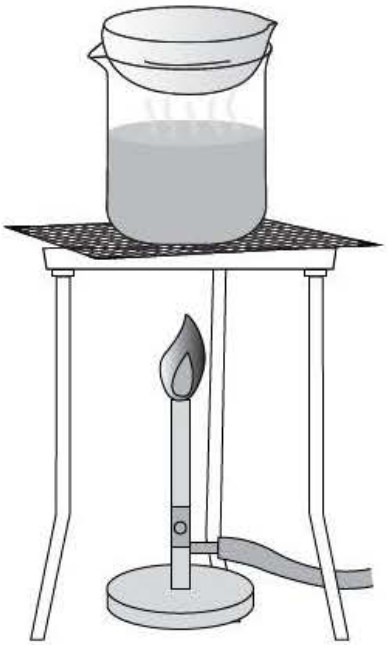


Figure 3

Method

1 Why should the filter funnel be removed from the burette before the titration can be carried out? [2]

.....

.....

2 What is the purpose of the white tile? [1]

.....

3 To make the ammonium sulfate crystals, why was the titration process repeated without the presence of the phenolphthalein? [1]

.....

.....

4 Why was the evaporated solution put to one side? Why was it not evaporated to dryness? [2]

.....

.....

.....

.....

Results and calculations

Volume of dilute sulfuric acid needed for the complete neutralisation of the ammonia solution = ... cm^3 [1]

Conclusion

Crystals of the ..... ammonium sulfate can be made by the ..... of aqueous ammonia by dilute ..... acid. [3]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 How could the experiment be modified to find out the exact concentration of the aqueous ammonia using 0.05 M sulfuric acid? [5]

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

2 Ammonia is a weak alkali. What do understand by the term 'weak alkali'? [4]

.....

.....

.....

.....

.....

.....

3 Calculate the percentage of nitrogen in ammonium sulfate.
(A_r : H = 1, N = 14, O = 16, S = 32) [2]

.....

.....

11.3 The effects of acid rain

Aim

To study the effect of 'acid rain' on various construction materials.

Theory

Air pollution is a major problem in our society. Concentrations of pollutant gases in the atmosphere, such as sulfur dioxide and nitrogen oxides, are increasing with the growing population and industrialisation. As the population rises there is a consequent increase in the need for energy, industries and motor vehicles. Sulfur dioxide and oxides of nitrogen are produced primarily from the combustion of the fossil fuels coal, oil and gas, but they are also produced by the smoking of cigarettes.

These pollutant gases are quite soluble and dissolve in rain water. After dissolving in rain, they produce the acids sulfuric (H_2SO_4) and nitric (HNO_3). The pH of this rain can fall from the natural level of about 5.7 to between 3 and 4.8. Acid rain is responsible for quite a lot of damage to our buildings.

You are going to investigate the effect of 'acid rain' on a range of materials used in the construction industry.

Apparatus and chemicals

- ☐ eye protection
- ☐ 8 × test tubes
- ☐ test-tube rack
- ☐ acid rain
- ☐ pieces of copper, limestone, brick, lead, steel, zinc, aluminium, concrete

Safety!

Acid rain – corrosive

Lead – toxic, dangerous for the environment

Procedure

1. Put on your eye protection.
2. Collect a sample of each of the construction materials.
3. Put a small amount of each sample into separate test tubes in a test-tube rack.
4. To each sample, add approximately one-third of a test tube of 'acid rain'.

5 Observe the samples over a 2 minute period and record your observations in the table below.

6 From your observations decide which of the materials:
(a) is the least affected by the 'acid rain'
(b) is the most affected by the 'acid rain'.

Method

1 Why is it necessary to use quite a large amount of 'acid rain'? [2]

.....

.....

2 Why is it necessary to observe over a period of time? [2]

.....

.....

Results and calculations

Table 1

Material	Observation after 2 minutes	Suitability of the material for construction
copper		
limestone		
brick		
lead		
steel		
zinc		
aluminium		
concrete		

[16]

Conclusion

In your opinion, based on the experiments carried out with 'acid rain', determine the most and least reactive materials.

1 Which materials are least affected by the 'acid rain'? [4]

.....

.....

2 Which materials are most affected by the 'acid rain'? [4]

.....

.....

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Consider your conclusion. Why do you think some of the materials that you have said are the most reactive towards 'acid rain' are still used? [3]

.....

.....

.....

.....

2 How do you think the amount of acid in real acid rain can be reduced? [3]

.....

.....

.....

.....

12.1 Sulfuric acid: a useful quantitative analytical chemical

Aim

To use sulfuric acid to determine the concentration of a solution of sodium hydroxide.

Theory

Sulfuric acid is a cheap and excellent bulk acid. It is used extensively to make other substances such as detergents and fertilisers. It is also used by analytical chemists to determine unknown concentrations of acids. Titration is the important and widely used analytical technique that is used in this situation. It is also used in forensic laboratories and water treatment laboratories, for example.

In titration, the acid is used to *just* neutralise a known quantity of alkali.

hydrogen ions + hydroxide ions → water

In the case of sulfuric acid with sodium hydroxide:

$H^+(aq) + 2OH^-(aq) \rightarrow 2H_2O(l)$

If the volumes of the substances involved are known, as well as the concentration of the acid, then it is possible to calculate the unknown alkali concentration.

Apparatus and chemicals

- ☐ eye protection
- ☐ burette
- ☐ burette stand
- ☐ 2 × 25 cm^3 pipettes
- ☐ 250 cm^3 conical flask
- ☐ 400 cm^3 beaker
- ☐ white tile
- ☐ filter funnel
- ☐ glass rod
- ☐ pipette filler
- ☐ heat-resistant mat
- ☐ 0.05 M sulfuric acid
- ☐ sodium hydroxide solution
- ☐ phenolphthalein

Safety!

0.05 mol/dm³ sulfuric acid – low hazard

Sodium hydroxide solution – irritant

Phenolphthalein – low hazard

Procedure

1. Put on your eye protection.
2. Fill the 25 cm³ pipette with sodium hydroxide. Pour this exact amount into a conical flask to which a few drops of phenolphthalein indicator have been added. Phenolphthalein is pink in alkaline conditions but colourless in acid.
3. A 0.05 mol/dm³ solution of sulfuric acid is placed in the burette using a filter funnel until it is filled up exactly to the zero mark (Figure 1).
4. The filter funnel is now removed.
5. The sulfuric acid is added to the aqueous sodium hydroxide in small quantities – usually no more than 0.5 cm³ at a time (Figure 2). The contents of the flask must be swirled after each addition of acid.
6. The acid is added until the aqueous sodium hydroxide has been neutralised completely. This is shown by the pink colour of the indicator *just* disappearing.
7. Take the final reading on the burette at the end-point (just as neutralisation has taken place). This value should be recorded in Table 1 as the rough value.
8. Further titrations need to be carried out until consistent results are obtained (within 0.1 cm³ of each other).

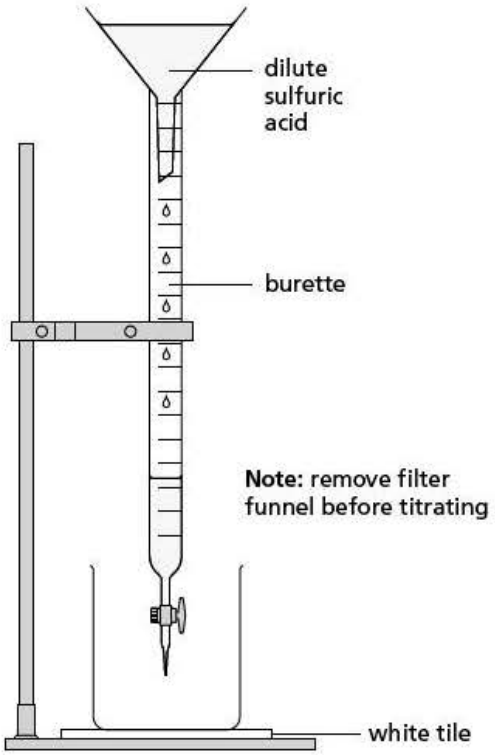


Figure 1

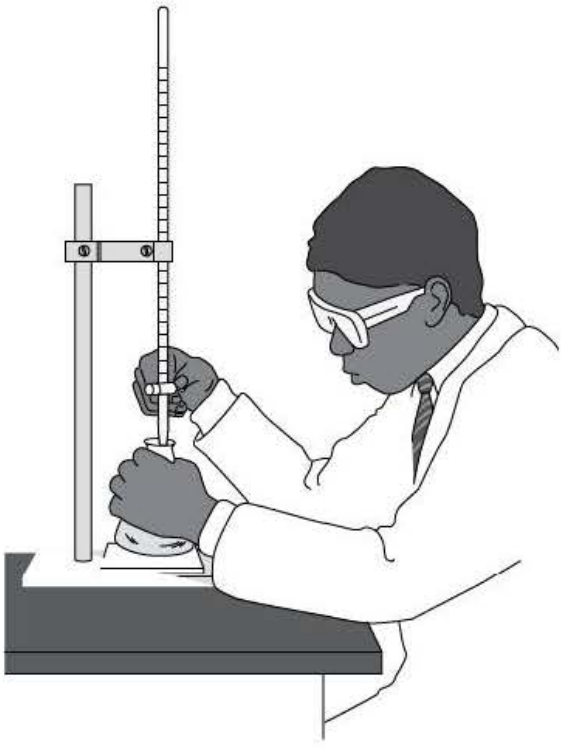


Figure 2

Method

1 Why must the filter funnel be removed from the burette before titration can be carried out? [2]

.....

.....

2 What is the purpose of the white tile? [1]

.....

3 Why do you have to swirl the contents of the conical flask as you add the acid? [1]

.....

4 Why is the first titration figure known as the 'rough' value? [1]

.....

.....

Results and calculations

Table 1

	Rough	1	2	3	4
final burette reading/cm^3					
initial burette reading/cm^3					
volume of acid used/cm^3					

[5]

1 Calculating the concentration of the alkali:

The chemical equation is:

sulfuric acid + sodium hydroxide → sodium sulfate + water

$H_2SO_4(aq)$ + $2NaOH(aq)$ → $Na_2SO_4(aq)$ + $H_2O(l)$

From this balanced equation it can be seen that:

..... [1]

The number of moles of sulfuric acid in 25.00 cm^3:

$$\frac{\text{average titration} \times \text{molarity}}{1000} = \text{.....} \text{ moles} \quad [1]$$

Since 1 mole of sulfuric acid reacts with 2 moles of sodium hydroxide, the number of moles of sodium hydroxide which will react with 2 × ..... moles of sulfuric acid = ... moles which is found in 25 cm^3 of solution. [2]

The concentration of the alkali is:

$$\text{moles of alkali in 25 cm}^3 \times \frac{1000}{25} = \text{.....} \text{ mol/dm}^3 \quad [1]$$

2 Calculating the molarity of the alkali:

The chemical equation is:

sulfuric acid + sodium hydroxide → sodium sulfate + water

$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

The calculation of the unknown concentration of alkali is done by substituting in the following mathematical equation:

$$\frac{M_1V_1}{M_{acid}} = \frac{M_2V_2}{M_{alkali}}$$

where:

M_1 = concentration of the acid used

V_1 = average volume of acid used (cm^3)

M_{acid} = number of moles of acid shown in the chemical equation

M_2 = concentration of the alkali used

V_2 = volume of the alkali used (cm^3)

M_{alkali} = number of moles of alkali shown in the chemical equation. [2]

Conclusion

The concentration of the alkali sodium hydroxide as a molarity = ... M [1]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 How could the experiment be altered to obtain a sample of the salt sodium sulfate? [4]

.....

.....

.....

.....

.....

.....

2 Calculate the amount of sodium sulfate in grams that would be produced in titration where A_r : H=1, Na=23, O=16, S=32. [3]

.....

.....

.....

.....

.....

.....

12.2 Concentrated sulfuric acid

Aim

To show the dehydrating properties of concentrated sulfuric acid.

Theory

Sulfuric acid is a very important bulk chemical made in the contact process. It is so important that the amount of sulfuric acid which a country uses in one year can be seen as a measure of that country's economic development. Beyond its major uses in the manufacture of fertilisers and detergents, the concentrated acid has a number of important properties. One of these will be demonstrated in a fume cupboard – that it is a good dehydrating agent. As a dehydrating agent, it absorbs water or the elements of water from substances such as sugar and hydrated copper(II) sulfate.

Sulfuric acid is so good at absorbing water that it is often used to dry gases.

Apparatus and chemicals

- ☐ eye protection
- ☐ test-tube rack
- ☐ 100 cm^3 beaker
- ☐ boiling tube
- ☐ teat pipette
- ☐ concentrated sulfuric acid
- ☐ sugar
- ☐ hydrated copper(II) sulfate
- ☐ safety gloves
- ☐ access to fume cupboard

Safety!

Concentrated sulfuric acid – corrosive

Hydrated copper(II) sulfate – harmful, dangerous for the environment

Procedure

The teacher or demonstrator should wear eye protection and plastic gloves throughout the demonstration experiment which must take place in a fume cupboard.

1 Half fill a 100 cm^3 beaker with sugar.

2 Add several pipettes full of concentrated sulfuric acid carefully to this sugar.

3 Write your observations of this reaction in the table provided below.

4 Some hydrated copper(II) sulfate is placed in the bottom of a boiling tube.

5 Add half a pipette of concentrated sulfuric acid carefully to this copper(II) sulfate.

6 Write your observations of this reaction in the table provided below.

Method

1 Why must this experiment be carried out in the fume cupboard? [2]

.....

.....

2 Why must plastic gloves be worn? [1]

.....

.....

Results and calculations

Table 1

Substance	Formula	Observation with concentrated sulfuric acid	Inference	Equation for reaction taking place
sugar	$C_6H_{12}O_6$			
copper(II) sulfate hydrate	$CuSO_4.5H_2O$			

[20]

Conclusion

Concentrated sulfuric acid is an excellent dehydrating agent. It will remove water from

salt ..... such as hydrated copper(II) sulfate or the .....

of water from substances such as sugar. When this happens the reactions are very

..... . [3]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Write the word and chemical equation for the dehydration reaction that takes place when concentrated sulfuric acid is added to hydrated cobalt(II) chloride, $CoCl_2.6H_2O$. [3]

.....

.....

2 What might you expect to be produced if concentrated sulfuric acid was reacted with ethanol, C_2H_5OH? [1]

.....

12.3 Properties of dilute sulfuric acid

Aim

To examine the properties of dilute sulfuric acid.

Theory

Dilute sulfuric acid is a typical strong dibasic acid. A dibasic acid is one with two replaceable hydrogen atoms, which allow it to produce two series of salts – **normal** and **acid** salts.

It is made by diluting concentrated sulfuric acid with water. When water is added to the acid, full ionisation takes place as the following process occurs:

$$\text{sulfuric acid} \xrightarrow{\text{water}} \text{hydrogen ions} + \text{sulfate ions}$$

$$H_2SO_4(aq) \xrightarrow{\text{water}} 2H^+(aq) + SO_4^{2-}(aq)$$

It has the typical properties of a strong acid in that it will react with and be **neutralised** by alkalis, metal oxides, **MAZIT** metals and carbonates. You are going to examine these properties in this experiment.

Apparatus and chemicals

- ☐ eye protection
- ☐ test-tube rack
- ☐ 4 × test tubes
- ☐ test-tube holder
- ☐ thermometer(–10–110 °C)
- ☐ Bunsen burner
- ☐ heat-resistant mat
- ☐ delivery tube
- ☐ 1 mol/dm^3 sulfuric acid
- ☐ wooden spills
- ☐ magnesium ribbon
- ☐ copper(II) oxide
- ☐ sodium carbonate
- ☐ 1 mol/dm^3 sodium hydroxide
- ☐ limewater

Safety!

Sulfuric acid (1 mol/dm³) – irritant

Magnesium ribbon – low hazard

Copper(II) oxide – harmful, dangerous for the environment

Sodium hydroxide solution (1 mol/dm³) – corrosive

Sodium carbonate – low hazard

Limewater – irritant

Procedure

1 Put on your eye protection and light your Bunsen burner. As you work through the experiments fill in your observations in the results table.

2 Place a piece of magnesium ribbon into a test tube and add a small amount of sulfuric acid to it. Put your thumb over the end to stop any gas escaping and test the collected gas with a lighted spill.

3 Place a small sample of the sodium hydroxide solution into a test tube in the test-tube rack. Insert the thermometer and read the starting temperature. Add a small amount of sulfuric acid to it. Measure the new temperature as the reaction occurs.

4 Place a little sodium carbonate into a test tube and add a small amount of sulfuric acid to it. Put the delivery tube into the neck and pass any gas through limewater.

5 Place copper(II) oxide into a test tube and add a small amount of sulfuric acid to it. Warm the mixture gently.

Method

1 Why do you put your thumb over the test tube in the reaction between magnesium and sulfuric acid? [1]

.....

2 Why is a thermometer needed in the reaction between sodium hydroxide solution and sulfuric acid? [2]

.....

3 Why is a delivery tube needed in the reaction between sodium carbonate and sulfuric acid? [1]

.....

4 Why is it necessary to heat the reaction between copper(II) oxide and sulfuric acid? [1]

.....

Results and calculations

Table 1

Substance	Reaction with dilute sulfuric acid	Inference	Name of salt produced
magnesium			
sodium hydroxide			
sodium carbonate			
copper(II) oxide			

[12]

Conclusion

Dilute sulfuric acid is a ..... acid. It reacts with MAZIT metals such as magnesium to produce a ..... and release ..... gas. When it reacts with sodium hydroxide the ..... reaction forms the sodium ..... and water only. In the reaction with sodium carbonate it forms sodium ..... and releases the gas ..... . In the final experiment with copper(II) oxide, on warming, copper(II) sulfate a blue ..... salt is produced.

These are general reactions of strong acids. Another example of a strong acid would be ..... acid. [9]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Define the terms shown in bold in the theory section. [8]

.....

.....

.....

.....

.....

.....

.....

.....

2 Write the word and chemical equations for the reactions that take place in the four experiments.

(a) with magnesium [3]

.....

.....

(b) with sodium hydroxide [3]

.....

.....

(c) with sodium carbonate [3]

.....

.....

(d) with copper(II) oxide [3]

.....

.....

→

3 Use your textbook to help you write ionic equations for the above processes:

(a) with magnesium [2]

.....

(b) with sodium hydroxide [2]

.....

(c) with sodium carbonate [2]

.....

(d) with copper(II) oxide. [2]

.....

13.1 Limestone: a useful resource

Aim

To study the limestone cycle.

Theory

Limestone is widely used as a building material, to neutralise acids in soil, to make glass and in the manufacture of iron and steel.

Limestone can be **thermally decomposed** by heating it strongly. When this happens calcium oxide (quicklime), CaO, is produced. Calcium oxide is a base and is still used by some farmers to spread on fields to **neutralise** soil acidity and to improve drainage of water through soils that contain large amounts of clay. It is also used to neutralise industrial waste products, for example in **flue gas desulfurisation**. It also has uses as a drying agent in industry and in the manufacture of mouthwash.

Large amounts of calcium oxide are also converted into calcium hydroxide, $Ca(OH)_2$, which is called slaked lime. Calcium hydroxide is a cheap industrial **alkali**. It is used in large quantities to make bleaching powder, by some farmers to reduce soil acidity, for neutralising acidic industrial waste products, in the manufacture of whitewash, in glass manufacture and in water purification. Calcium hydroxide, in its white powder form, is produced by adding an equal amount of water to calcium oxide in a carefully controlled reaction. A weak solution of calcium hydroxide in water is called limewater. It is used to test for carbon dioxide gas, as a white solid of calcium carbonate, $CaCO_3$, is formed if carbon dioxide gas is mixed with it.

This brings us round in a circle from limestone (calcium carbonate, $CaCO_3$), through calcium oxide (CaO) and calcium hydroxide, $Ca(OH)_2$, and back to calcium carbonate (limestone). This is known as the 'limestone cycle' (Figure 1).

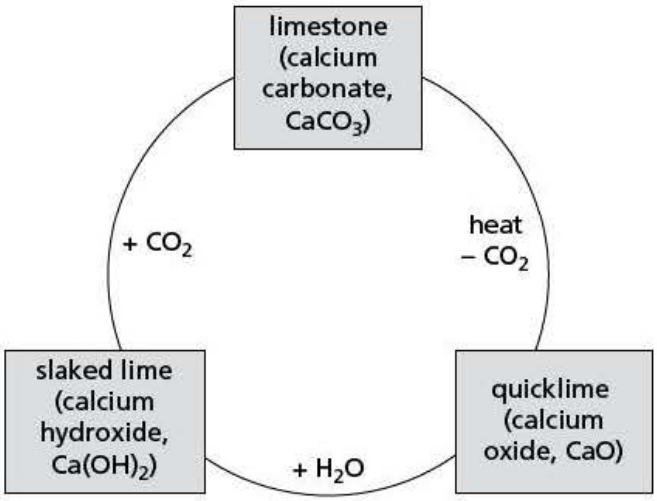


Figure 1

Apparatus and chemicals

- ☐ eye protection
- ☐ tripod
- ☐ gauze
- ☐ heat-resistant mat
- ☐ Bunsen burner
- ☐ tin lid
- ☐ tongs
- ☐ dropping pipette
- ☐ test tube
- ☐ test-tube rack
- ☐ 3 × boiling tubes
- ☐ 400 cm^3 beaker – to hold the boiling tubes
- ☐ filter funnel and filter paper
- ☐ drinking straw
- ☐ universal indicator solution
- ☐ limestone pieces

Safety!

The apparatus will be hot, so be careful.

Universal indicator solution – low hazard

Limestone (calcium carbonate) – low hazard

Calcium oxide – corrosive. The addition of water to calcium oxide is very exothermic.

Calcium hydroxide – irritant

Procedure

1. Put on your eye protection.
2. Place several small pieces of limestone onto a tin lid.
3. Set up the apparatus as shown in Figure 2 and heat the limestone pieces very strongly from above for 10 minutes. Record your observations in Table 1.

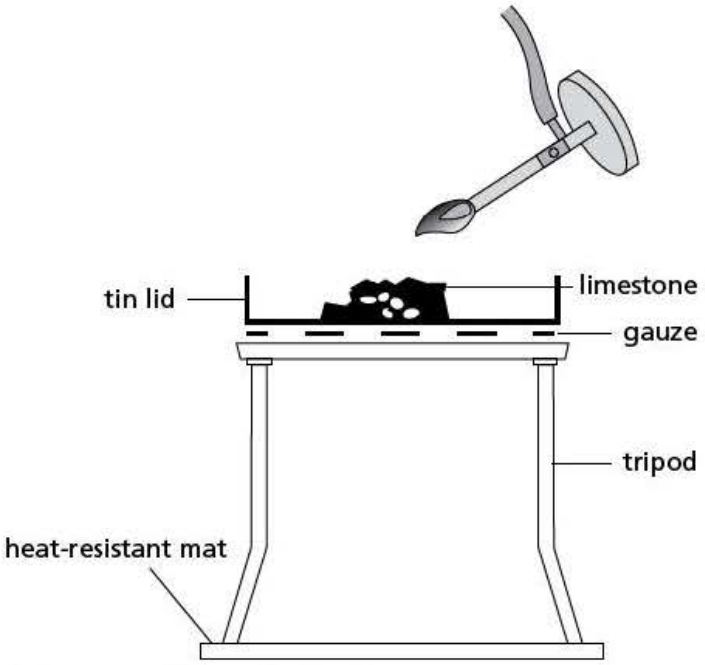


Figure 2 Heating limestone pieces

4 Allow the tin lid and its contents to cool for 5 minutes.

5 Using tongs, transfer the lump(s) of white solid, calcium oxide, ***carefully*** from the tin lid to a dry boiling tube (note that the calcium oxide is *corrosive* and should not be handled).

6 Using a dropping pipette, ***carefully*** add about half a test tube of water dropwise to the calcium oxide in the boiling tube. Record your observations in Table 1.

7 You have now converted calcium oxide (quicklime) into calcium hydroxide (slaked lime).

8 Filter the mixture ***carefully*** into a second boiling tube.

9 Divide the filtrate into two equal portions.

10 Add three drops of universal indicator solution to one half of the filtrate. Record your observations in Table 1.

11 Using a straw, blow gently into the other half of the filtrate. Record your observations in Table 1.

Method

1 In the instructions, why is the word *carefully* used and emphasised? [2]

.....

.....

2 What does the observation from step 10 tell you about calcium hydroxide? [1]

.....

.....

3 What does the observation from step 11 tell you? [2]

.....

.....

Results and calculations

Table 1

Procedure step	Observation	inference
3		
10		
11		

[6]

Conclusion

When limestone is heated strongly it is converted to ..... (quicklime).

When water is added to the ..... a new substance is produced called

..... (slaked lime) which is an ..... . When carbon dioxide is

blown through the solution it forms ..... . [5]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Write word and chemical equations for these reactions:

(a) thermal decomposition of calcium carbonate [3]

.....

.....

(b) production of calcium hydroxide from calcium oxide [3]

.....

.....

(c) the effect of carbon dioxide (in your breath) on the calcium hydroxide solution. [3]

.....

.....

2 When calcium carbonate (limestone) is spread on acid soil, it neutralises the acidity. If the acidity is caused by sulfuric acid, write a word and chemical equation for the neutralisation reaction. [3]

.....

.....

3 Explain the words or phrases shown in bold in the theory section. [8]

.....

.....

.....

.....

.....

.....

13.2 Does the food we eat contain carbon?

Aim

To show that foodstuffs contain the element carbon.

Theory

Carbon occurs in many naturally occurring compounds, for example, natural gas and petroleum (a complex mixture of **hydrocarbons**). Coal is a complex mixture of compounds of carbon, sulfur, hydrogen and oxygen. There is also a large number of carbonates, such as calcium and magnesium carbonate; while in the atmosphere there is carbon dioxide, the carrier gas for the **carbon cycle**.

Another set of naturally occurring substances which contain carbon are sugars, for example, glucose ($C_6H_{12}O_6$). This substance is also involved in the carbon cycle, in the formation of food through **photosynthesis**.

$$\text{carbon dioxide} + \text{water} \xrightarrow[\text{chlorophyll}]{\text{sunlight}} \text{glucose} + \text{oxygen}$$

$$6CO_2(g) + 6H_2O(l) \xrightarrow[\text{chlorophyll}]{\text{sunlight}} C_6H_{12}O_6(aq) + 6O_2(g)$$

Can we show that carbon from substances like glucose is present in the foods we eat?

Apparatus and chemicals

- ☐ eye protection
- ☐ straw
- ☐ 5 × test tubes
- ☐ boiling tube
- ☐ test-tube rack
- ☐ test-tube holder
- ☐ teat pipette
- ☐ Bunsen burner
- ☐ heat-resistant mat
- ☐ bread
- ☐ rice
- ☐ cornflakes
- ☐ macaroni
- ☐ spaghetti
- ☐ limewater

Safety!
The apparatus will be hot, so be careful.

Procedure

1 Put on your eye protection.

2 Place a small piece of bread in a test tube and, holding it in a test-tube holder, heat it very strongly in a Bunsen flame.

3 Squeeze the air out of a teat pipette, place it deep inside the test tube, just above the heated bread, and release the teat. Gas will be sucked up into the pipette.

4 Put some limewater into a boiling tube and expel the gas from the pipette into the limewater.

5 Repeat step 4 several times, each time expelling the gas collected through the limewater.

6 Note down any observations from steps 2–5 in the table below.

7 Repeat steps 2–6 with other foodstuffs, filling in the table after each experiment.

Method

1 Why do you have to heat strongly? [1]

.....

2 Which gas are you testing for in these experiments on food and what is its formula? [2]

.....

3 How is the gas you have named in 1 formed? [1]

.....

Results and calculations

Table 1

Food	Effect of heat	Effect of extracted gas on limewater	Does the food contain carbon?

[15]

Conclusion

From the results of my experiments I can say that the foods I tested contain [1]

..... .

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Write word and chemical equations for the reaction to produce the gas you named earlier. [3]

.....

.....

2 We produce the gas you named earlier by aerobic respiration. What is aerobic respiration? [2]

.....

.....

3 What do you understand by the terms in bold in the theory section. [6]

.....

.....

.....

.....

13.3 Carbon dioxide

Aim

To study the preparation and properties of carbon dioxide.

Theory

The carbon dioxide found in the atmosphere (approximately 0.04% by volume) is produced in a number of ways, such as by the complete combustion of fossil fuels like methane:

methane + oxygen → carbon dioxide + water

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

In the laboratory, the gas is made by the action of an acid on a carbonate. The acid usually used is dilute hydrochloric acid, which is poured onto calcium carbonate as marble chips.

calcium carbonate + hydrochloric acid → calcium chloride + carbon dioxide + water

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

You are going to prepare carbon dioxide and then test its properties.

Apparatus and chemicals

- ☐ eye protection
- ☐ plastic dish (to act as a trough)
- ☐ 6 × boiling tubes
- ☐ 4 × bungs for the boiling tubes
- ☐ conical flask fitted with a bung containing a thistle funnel and a delivery tube
- ☐ tongs
- ☐ small candle
- ☐ Bunsen burner
- ☐ 100 cm^3 beaker
- ☐ wooden spills
- ☐ limewater
- ☐ magnesium ribbon (2 cm length)
- ☐ dilute hydrochloric acid
- ☐ marble chips
- ☐ soda water
- ☐ universal indicator solution

Safety!

Universal indicator solution – low hazard

Hydrochloric acid (2 mol/dm^3) – irritant

Magnesium ribbon – low hazard

Marble chips (calcium carbonate) – low hazard

Procedure

1 Put your eye protection on.

2 Set up the apparatus as shown in Figure 1.

3 Carefully pour a little dilute hydrochloric acid into the conical flask until the level of the acid is above the bottom of the thistle funnel as shown in Figure 1.

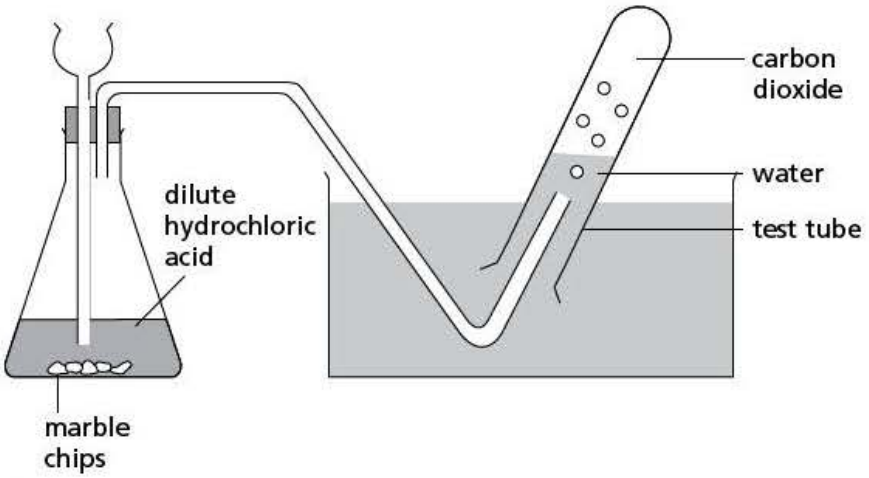


Figure 1

4 Collect the gas in the boiling tubes. Do not collect the gas coming over in the first half minute because it will contain quite a lot of air. When removing these tubes from the trough, either have your thumb over the end or insert a bung to the open end.

5 With the first tube, observe the colour of the gas and smell it. Write your observations in Table 1.

6 With the second tube, plunge a burning splint into the gas. Write your observations of what happens in Table 1.

7 Pour some soda water into a boiling tube and add a few drops of universal indicator solution. Write your observations of what happens in Table 1.

8 With the third tube of the gas, add a little limewater and shake gently. Write your observations in Table 1.

9 Hold a 2 cm piece of magnesium ribbon with tongs and set it alight in a Bunsen flame. Note – do **not** look directly at the bright light. Plunge the burning magnesium into the fourth tube of gas. Write your observations in Table 1.

10 Sit a small piece of candle in a 100 cm^3 beaker and light it. Pour the contents of two tubes of gas onto the candle. Write your observations in Table 1.

11 Invert one boiling tube of carbon dioxide over the top of an empty one and leave them in position for 15 seconds. Pour a little limewater into both tubes and then shake them. Write your observations in Table 1.

12 Repeat step 10, but this time put an empty tube over the full tube of carbon dioxide. Write your observations related to the limewater in Table 1.

Method

1 Can you think of a reason why you would not use sulfuric acid to make carbon dioxide? [2]

.....

.....

.....

2 You collected the carbon dioxide over water. What does this tell you about its solubility? [2]

.....

.....

3 What was the reason for putting a bung in the end of each boiling tube before testing the gas? [1]

.....

.....

Results and calculations

Table 1

Test on the carbon dioxide	Observation	Inference
colour		
smell		
what happens to the burning spill		
pH of the solution		
effect of limewater		
burning magnesium		
burning candle		
density experiments		

[16]

Conclusion

Carbon dioxide gas can be prepared using marble chips and dilute hydrochloric acid. Its properties are given here.

It has ..... colour. [1]

It has ..... smell. [1]

It will not support ..... . [1]

Its solution has a pH of ..... which makes it an ..... gas. [2]

It turns limewater cloudy ..... . This is used as a chemical test for carbon dioxide. [1]

It will support ..... for strongly burning substances such as magnesium. [1]

It is ..... than air. [1]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 When magnesium burns in carbon dioxide it is a redox process.

(a) What do you understand by the term 'redox process'? [2]

.....

.....

(b) What is oxidised and what is reduced in the reaction? [2]

.....

.....

2 When carbon dioxide dissolves in water it produces a *weak acid* called carbonic acid (H_2CO_3). Explain what you understand by a 'weak acid'. [2]

.....

.....

3 Use your research skills to find out four uses of carbon dioxide. [4]

.....

.....

.....

.....

14.1 Is methane a hydrocarbon?

Aim

To show that alkanes, such as methane, are hydrocarbons.

Theory

A fuel is a substance which can be conveniently used as a source of energy. Fossil fuels release energy in the form of heat when they undergo combustion:

fossil fuel + oxygen → carbon dioxide + water + energy

For example, natural gas or methane (a gaseous alkane), burns in a good supply of air (**complete combustion**) forming carbon dioxide and water, as well as plenty of heat energy:

methane + oxygen → carbon dioxide + water + energy

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

How can we use this reaction to show that methane is a **hydrocarbon**? Your teacher will demonstrate how this can be done.

Apparatus and chemicals

- ☐ eye protection
- ☐ water pump
- ☐ *glass* filter funnel
- ☐ safety screen
- ☐ 3 × retort stands with clamps and bosses to hold filter funnel and U-tubes
- ☐ 400 cm^3 beaker to act as ice bath
- ☐ ice
- ☐ Bunsen burner
- ☐ limewater
- ☐ anhydrous cobalt(II) chloride paper

Safety!

This experiment should be carried out behind a safety screen with the teacher/demonstrator wearing eye protection. Care is needed with the glass funnel. It may crack. Plastic gloves should be worn when handling the cobalt(II) chloride paper.

Limewater – irritant

Anhydrous cobalt(II) chloride) – toxic, dangerous for the environment

Procedure

1 The teacher or demonstrator should wear eye protection throughout the demonstration.

2 Set up the apparatus shown in Figure 1 behind a safety screen.

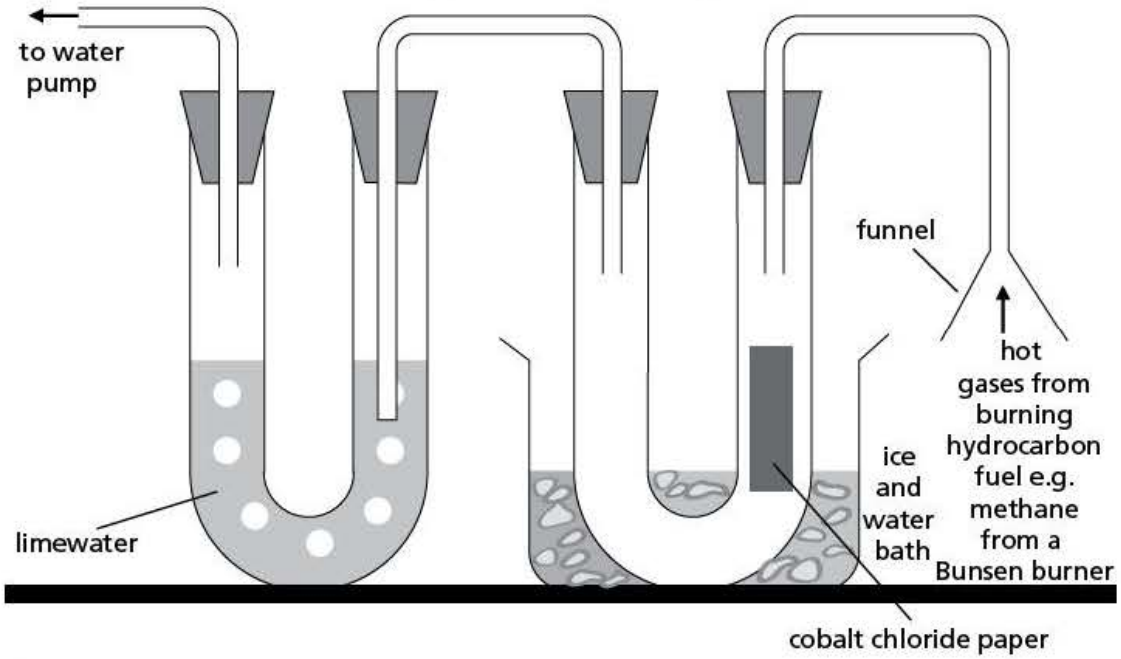


Figure 1

3 Record observations in the results section as the demonstration proceeds.

4 Place a piece of cobalt chloride paper in the first U-tube and some limewater in the second U-tube.

5 Light the Bunsen burner. Set the flame with the air hole half open.

6 Turn on the water pump and then move the Bunsen burner under the filter funnel. Ensure the flame is not too close.

7 Allow the process to proceed for 5 minutes.

Method

1 Why is the cobalt chloride placed in the first U-tube and the limewater put in the second U-tube? [2]

.....

.....

2 Why is a glass filter funnel used instead of a plastic one? [1]

.....

3 Why is the Bunsen flame not too close to the filter funnel? [1]

.....

Results and calculations

1 What happens to the cobalt chloride paper? [1]

.....

Explain your observation. [2]

.....

.....

2 What happens to the limewater? [1]

.....

.....

Explain your observation. [2]

.....

.....

3 What do your observations tell you about the elements present in methane and other alkanes? [2]

.....

Conclusion

When the gases produced by burning methane in air are passed though the apparatus various observation can be made.

The cobalt chloride paper tests for ..... and turns from a

.....

colour to a ..... colour. This shows the presence of the element

..... in methane.

The limewater tests for ..... and goes ..... . This shows

the presence of the element ..... in methane.

Overall this shows that methane and, hence, other alkanes contain the elements

..... and ..... only. They are, therefore,

..... . [9]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Define the phrases shown in bold in the theory section. [4]

.....

.....

→

2 Methane can also undergo incomplete combustion.

(a) What is the difference between complete and incomplete combustion? [2]

.....

.....

(b) Write a word and chemical equation for the *incomplete* combustion of methane. [3]

.....

.....

3 If the organic substance contains sulfur and is flammable, then how could the presence of sulfur be detected? [2]

.....

.....

14.2 Difference between alkanes and alkenes

Aim

To identify the difference between alkanes and alkenes.

Theory

Alkanes and alkenes are **hydrocarbons**. In the case of the alkanes, because these molecules possess only single covalent bonds (Figure 1), they are said to be **saturated**, as no further atoms can be added, e.g. hexane.

This is not the case of the alkenes such as hexene (Figure 2). These molecules possess a double covalent bond and are said to be **unsaturated**, because it is possible to break the double bond to add extra atoms to the molecule. This makes the alkenes more reactive than the alkanes.

It is possible to test to find out whether a hydrocarbon is an alkane or an alkene. When a few drops of a bromine solution in an organic solvent are shaken with the hydrocarbon, if it is an alkene a reaction takes place in which bromine joins to the alkene double bond. The bromine solution loses its red–brown colour. This is due to the addition reaction that has taken place. If an alkane is shaken with a bromine solution, no colour change takes place. This is because there are no double bonds between the carbon atoms of alkanes and so bromine doesn't react.

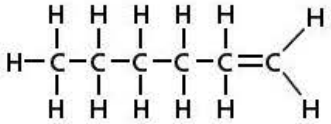


Figure 2

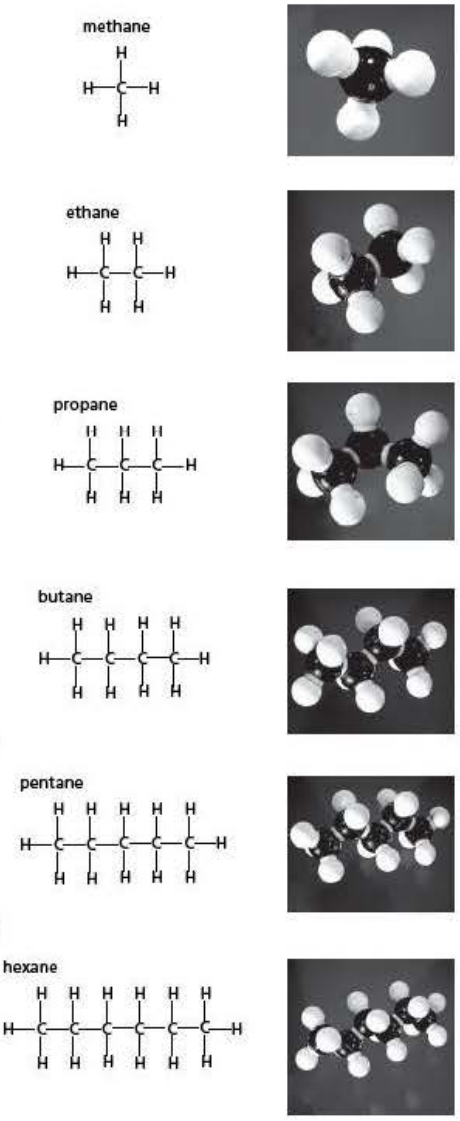


Figure 1

Apparatus and chemicals

- ☐ eye protection
- ☐ set of molecular models
- ☐ 2 × wooden spills
- ☐ 2 × bungs
- ☐ plastic gloves
- ☐ evaporating basin
- ☐ access to fume cupboard
- ☐ 0.02 M bromine water
- ☐ sample of hexane
- ☐ sample of hexene

Safety!

Hexane – highly flammable, harmful

Hexene – highly flammable, harmful

0.02 M bromine water – toxic, corrosive, dangerous for the environment

Disposal: evaporate in a fume cupboard.

Procedure

1. These reactions should be carried out in a fume cupboard.
2. Put on your eye protection and plastic gloves.
3. Pour 5 drops of hexane into a test tube. Waft the fumes towards you and smell them. What does hexane look like? What does hexane smell like? Add your observations to Table 1.
4. Add a little 0.02 M bromine water to hexane. Put a bung in the end and shake. Add the observations to Table 1.
5. Pour a little hexane into an evaporating basin and set fire to it with a burning splint. Add your observations to Table 1.
6. Repeat the steps 1–5 with hexene.
7. Build molecular models of hexane and hexene.

Method

1 Why must you wear plastic gloves? [1]

.....

2 Why have you had to do these reactions in a fume cupboard? [1]

.....

Results and calculations

Table 1

Substance	Appearance	Smell	Shake with bromine in an organic solvent	Effect of a lighted splint
hexane				
hexene				

[8]

Having looked at the theory section and by using your textbook as a resource, build a molecular model of the substance that is formed when hexene reacts with bromine in an organic solvent.

Conclusion

It is possible to identify an alkene using the fact that it will ..... bromine dissolved in an organic solvent. In this reaction the bromine ..... across the ..... bond in the alkene. The reaction is an example of an ..... reaction. [4]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

Alkenes generally undergo addition reactions. Describe the addition reaction between ethene and (a) water and (b) hydrogen. [6]

.....

.....

.....

.....

.....

.....

.....

.....

.....

14.3 Hydrocarbons can form isomers

Aim

To identify and build the isomers of hexane.

Theory

Sometimes it is possible to write more than one structural formula to represent a molecular formula. The structural formula of a compound shows how the atoms are joined together by the covalent bonds. For example, there are two different compounds with the molecular formula C_4H_{10}. The structural formulae of these two substances, along with their names and physical properties, are shown in Figure 1. Compounds such as those in Figure 1 are known as **isomers**.

Isomers are substances which have the same molecular formula but different structural formulae. The different structures of the compounds shown in Figure 1 have different melting and boiling points. Molecule **b** contains a branched chain and has a lower melting point than molecule **a**, which has no branched chain. All the alkane molecules with four or more carbon atoms possess isomers.

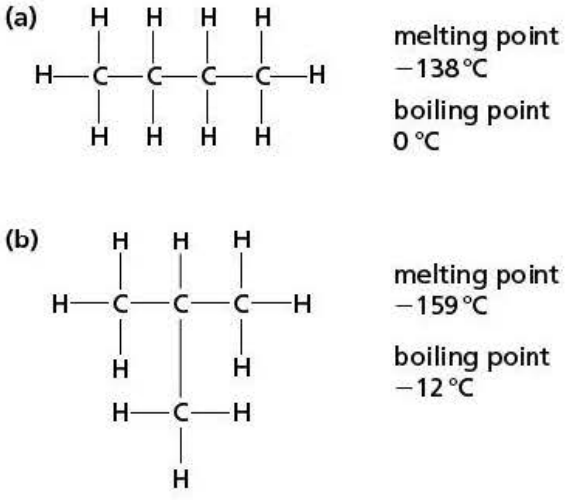


Figure 1

Apparatus and chemicals

☐ molecular modelling kit

Procedure

1 Draw the structural formula to represent hexane (C_6H_{14}). Get this checked by your teacher and then construct this alkane molecule.

2 Draw the structural formulae for the other four isomers of hexane. Check them with your teacher. If they are correct then construct these molecules.

Method

Which of the isomer molecules do you think has the highest melting and boiling point? [1]

.....

Explain your answer. [2]

.....

.....

Results and calculations

1 (a) Drawings (3D and 2D) for hexane, C_6H_{14}: [3]

(b) Drawings (3D and 2D) and names of the other four isomers of C_6H_{14}: [3]

(c) [3]

(d) [3]

(e) [3]

Conclusion

There are ..... isomers of hexane. There is one which has ..... branch chains while the other four ..... branched chains. The molecule with a straight chain has the ..... melting and boiling points. This is the case because the forces of attraction between the molecules are at a ..... when there are ..... branched chains. [6]

Evaluation

Outline how this experiment could be improved. [2]

.....

.....

Extension

1 Explain why there are so many organic molecules in our world. [2]

.....

.....

2 There are isomers that can be created in alkene molecules by changing the position of the C═C. Draw the structural formulae for the isomers created in this way for pentene, C_5H_{10}. [2]

15.1 Organic structures and functional groups

Aim

Study the shape, bond angles and the functional groups of various organic compounds.

Theory

Organic molecules can be represented by stick structures. For example, methane and ethane from the alkanes can be represented as drawings as shown in Figure 1.

Double bonds, as in the case of the alkenes, are shown in the structures as a double line, e.g. ethene (Figure 2).

If you replace one of the hydrogen atoms on an alkane molecule with a group such as –OH, the hydroxyl group, then you get a new **homologous series** called the alcohols. If you replace one of the hydrogen atoms on an alkane molecule with a –COOH group then you get a homologous series called the carboxylic acids or alkanoic acids. Whichever group is attached, it will bring with it a new set of physical and chemical properties. These groups are known as **functional groups**.

(a) methane **(b) ethane**

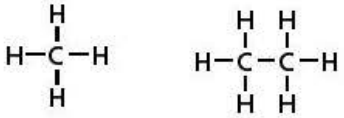


Figure 1

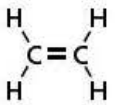


Figure 2 Ethene

Apparatus and chemicals

- ☐ molecular modelling kit
- ☐ protractor

Procedure

1 Build the following molecules using a molecular modelling kit (Figure 3).

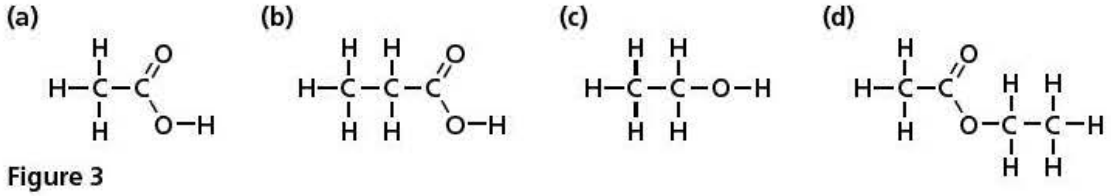


Figure 3

2 For each model, measure the bond angles in the parts of the molecules shown in Figure 4.

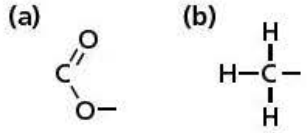


Figure 4

3 Draw the molecular model structures of the molecules in 3D in the results section. This is how we think they look in reality.

Method

1 Which of the molecules **(a)** and **(b)** in Figure 3 do you think has the highest melting and boiling point? [1]

.....

Explain your answer. [2]

.....

.....

2 Why do you think we chose molecules **(a)** and **(b)** in 1? [1]

.....

Results and calculations

1 Molecular model 3D drawing: [8]

(a)

(b)

(c)

(d)

2 Record the bond angles for each bond type **(a)** and **(b)** in Figure 4. [2]

3

Table 1

Molecule	Number of elements present	*Relative formula mass	Total number of atoms present	Number of single covalent bonds	Number of double covalent bonds
a					
b					
c					
d					

*To calculate the relative formula mass (R.F.M) you need to add together all the relative atomic masses (R.A.M.) of the atoms present in a molecule: RAM: H = 1, C = 12, O = 16. [20]

Conclusion

Organic molecules have a ..... structure. Their melting and boiling points depend on this ..... structure, ... and ..... groups present. The bond angles found in organic molecules are the ..... throughout organic chemistry. [5]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Identify the functional group present in each of the molecules. [3]

.....

.....

2 Which homologous series do the molecules **(a)** to **(d)** in Figure 3 belong to? [3]

.....

.....

3 When compound **(b)** and compound **(c)** react together they form an ester. Make a model of the compound created.
What is the name of this compound? [1]

.....

15.2 Nylon rope trick

Aim

To make the polymer nylon.

Theory

There are two types of **polymers** – addition and condensation. An example of an addition polymer is poly(ethene) which is formed by an addition reaction from an alkene, ethene, as the **monomer**. Not all polymers are formed by addition reactions. Condensation polymers are produced, as the name suggests, by a condensation reaction in which a small molecule, such as water or hydrogen chloride gas, is produced alongside the polymer. An example of a condensation polymer is nylon. This was discovered by Wallace Carothers in 1935.

Nylon is made by reacting two different chemicals together, unlike poly(ethene) which is made only from monomer units of ethene. The starting molecules for nylon are more complicated than those for poly(ethene) and are called 1,6-diaminohexane and hexanedioic acid.

1,6-diaminohexane + hexanedioic acid

$H_2N(CH_2)_6NH_2$ $\quad\downarrow\quad$ $HOOC(CH_2)_4COOH$

$H_2N(CH_2)_6\mathbf{NHOC}(CH_2)_4COOH + H_2O$

The polymer chain is made up from the two starting molecules arranged alternately as these molecules react and link up. Each time a reaction takes place, a molecule of water is lost.

Apparatus and chemicals

- ☐ eye protection
- ☐ 50 cm^3 beaker
- ☐ glass rod
- ☐ tweezers
- ☐ 2 × retort stands with clamps and bosses
- ☐ plastic gloves
- ☐ 1,6-diaminohexane (2.2 g dissolved in 50 cm^3 sodium carbonate solution)
- ☐ sebacoyl chloride (1.5 g dissolved in 50 cm^3 hexane or cyclohexane)
- ☐ safety screen
- ☐ paper towels

Safety!

Safety gloves and eye protection should be worn throughout the experiment due to the hazardous nature of the reagents.

1,6-Diaminohexane – when pure is corrosive

Sebacoyl chloride – corrosive

Hexane – highly flammable, harmful

Sodium carbonate – low hazard

Procedure

1 The teacher or demonstrator should wear eye protection and plastic gloves throughout the demonstration experiment.

2 Set up a safety screen.

3 Pour a 20 cm^3 sample of 1,6-diaminohexane solution into a 50 cm^3 beaker. Carefully pour a 20 cm^3 sample of the diacid chloride solution of sebacoyl chloride down the side of the same beaker.

4 A nylon thread can now be pulled from the interface between the two phases using a pair of tweezers.

5 Wrap the thread around the glass rod and continue to pull and turn the glass rod gently.

6 Push the nylon from the glass rod onto a paper towel.

Method

1 Why must plastic gloves be worn? [1]

.....

2 Why is a safety screen required? [1]

.....

3 Why are the solutions poured carefully? [2]

.....

4 Why is it necessary to put the nylon thread on a paper towel? [1]

.....

Results and calculations

1 Draw a diagram of what is happening as the thread is made. [2]

2 Give a description of the nylon thread. [2]

.....

.....

Conclusion

The ..... nylon is made by the mixing of 1,6–diaminohexane and sebacoyl chloride. During this reaction the small molecule ..... is produced (or ..... out). The name of this type of reaction is ..... polymerisation. [4]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 Define the terms in bold in the theory section. [5]

.....

.....

.....

.....

2 Use your textbook to:

(a) identify the starting materials for making terylene, another condensation polymer [2]

.....

.....

(b) state the two uses of nylon and two of terylene. [4]

.....

.....

.....

.....

15.3 Properties of dilute ethanoic acid

Aim

To examine the properties of dilute ethanoic acid.

Theory

Dilute ethanoic acid is a typical **weak** acid. It is made by diluting concentrated or glacial ethanoic acid with water. It produces few hydrogen ions when it dissolves in water compared with a strong acid, such as hydrochloric acid, of the same concentration. It is only partially ionised. Its solution has a higher pH than a strong acid, but still less than 7.

$$\text{ethanoic acid} \underset{}{\overset{\text{water}}{\rightleftharpoons}} \text{hydrogen ions} + \text{ethanoate ions}$$

$$CH_3COOH(l) \overset{\text{water}}{\rightleftharpoons} H^+(aq) + CH_3COO^-(aq)$$

The $\rightleftharpoons$ sign means that the reaction is reversible. This means that although the ethanoic acid molecules break down to give hydrogen ions and ethanoate ions, they also react together to re-form the ethanoic acid molecule. The fact that fewer ethanoic acid molecules dissociate compared with a strong acid, and that the reaction is reversible, means that few hydrogen ions are present in the solution.

Ethanoic acid has the typical properties of an acid in that it will react with and be **neutralised** by alkalis, metal oxides, **MAZIT** metals and carbonates. Being an organic acid, it also reacts some organic molecules such as ethanol, C_2H_5OH. This reaction forms a substance called ethyl ethanoate, which belongs to the **homologous series** of esters. You are going to examine some of these properties in this experiment.

Apparatus and chemicals

- ☐ goggles
- ☐ test-tube rack
- ☐ 4 × test tubes
- ☐ test-tube holder
- ☐ thermometer (–10–110 °C)
- ☐ Bunsen burner
- ☐ heat-resistant mat
- ☐ delivery tube
- ☐ teat pipette
- ☐ 2 mol/dm^3 ethanoic acid
- ☐ wooden splints
- ☐ magnesium ribbon
- ☐ sodium carbonate
- ☐ dilute sodium hydroxide
- ☐ ethanol
- ☐ 4 mol/dm^3 sulfuric acid
- ☐ limewater

Safety!

Limewater – irritant

Sodium carbonate – low hazard

2 mol/dm³ Ethanoic acid – irritant

Magnesium ribbon – low hazard

2 mol/dm³ sodium hydroxide – corrosive

Ethanol – highly flammable

4 mol/dm³ sulfuric acid – corrosive

Procedure

1 Put on your goggles and light your Bunsen burner. As you work through the experiments, fill in your observations in Table 1.

2 Place a piece of magnesium ribbon into a test tube. And add a small amount of ethanoic acid to it. Put your thumb over the end to stop any gas escaping and test the collected gas with a lighted splint.

3 Place a small sample of the sodium hydroxide solution into a test tube in the test-tube rack. Insert the thermometer and read the starting temperature. Add a small amount of ethanoic acid to it. Measure the new temperature as the reaction occurs.

4 Place a little sodium carbonate into a test tube and add a small amount of ethanoic acid to it. Put the delivery tube into the neck and pass any gas through limewater.

5 Place a little ethanol into a test tube and add a small amount of ethanoic acid, followed by a few drops of 4 M sulfuric acid, using a teat pipette.

6 After a couple of minutes, waft the fumes from the top of the test tube.

Method

1 Why do you put your thumb over the test tube in the reaction between magnesium and sulfuric acid? [1]

.....

2 Why is a thermometer needed in the reaction between sodium hydroxide solution and sulfuric acid? [2]

.....

.....

3 Why is a delivery tube needed in the reaction between sodium carbonate and sulfuric acid? [1]

.....

4 Why is it necessary to add 4 M sulfuric acid to the reaction between ethanoic acid and ethanol? [1]

.....

Results and calculations

Table 1

Substance	Reaction with dilute ethanoic acid	Inference
magnesium		
sodium hydroxide		
sodium carbonate		
ethanol		

[8]

Conclusion

Dilute ethanoic acid is a ..... acid. It reacts with MAZIT metals such as magnesium to produce a ..... and release ..... gas. When it reacts with sodium hydroxide, the ..... reaction forms sodium ..... and water only. In the reaction with sodium carbonate, it forms sodium ..... and releases the gas ..... . In the final experiment with ethanol, ethyl ethanoate, an ..... , is produced. [8]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

.....

Extension

1 Write the word and chemical equations for the reactions that take place in the four experiments.

(a) with magnesium [3]

.....

.....

(b) with sodium hydroxide [3]

.....

.....

(c) with sodium carbonate [3]

.....

.....

(d) with ethanol [3]

.....

.....

2 Use your textbook to help you write ionic equations for the reactions:

(a) with magnesium [2]

.....

(b) with sodium hydroxide [2]

.....

→

(c) with sodium carbonate. [2]

.....

3 Define the terms shown in bold in the theory section. [8]

.....

.....

.....

.....

.....

.....

.....

.....

.....

.....

16.1 Missing labels from reagent bottles: what a problem!

Aim

To use simple qualitative analysis to identify unknown solutions.

Theory

The branch of chemistry that deals with the identification of elements or grouping of elements present in a sample is called **qualitative chemical analysis**, or **qualitative analysis** for short. It does not deal in quantities. The techniques employed in qualitative analysis vary in their complexity, depending on the nature of the sample under investigation. In some cases, it is only necessary to confirm the presence of certain elements or groups for which specific chemical tests, or 'spot' tests may be available.

In this experiment, an everyday problem is investigated – a problem has arisen due to missing labels from reagent bottles A, B and C! These bottles were known to contain solutions of sodium sulfate, sodium chloride and sodium carbonate. Can you identify which is which?

Test for a sulfate: if you take a solution of a suspected sulfate and add it to a solution of a soluble barium salt (such as barium chloride), then a white precipitate of barium sulfate will be produced.

barium ion + sulfate ion → barium sulfate

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

A few drops of dilute hydrochloric acid are also added to this mixture. If the precipitate does not dissolve, then it is barium sulfate and the unknown salt was, in fact, a sulfate.

Test for a chloride: if you take a solution of a suspected chloride and add to it a small volume of dilute nitric acid, to make an aqueous acidic solution, followed by a small amount of a solution of a soluble silver salt (such as silver nitrate), a white precipitate of silver chloride will be produced.

chloride ion + silver ion → silver chloride

$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

Test for a carbonate: If a small amount of an acid is added to some of the suspected carbonate (either solid or in solution), then effervescence occurs. If it is a carbonate, carbon dioxide gas is produced, which will turn limewater 'milky' (a cloudy white precipitate of calcium carbonate forms).

carbonate ions + hydrogen ions → carbon dioxide + water

$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

Apparatus and chemicals

- ☐ eye protection
- ☐ 3 × test tubes
- ☐ test-tube rack
- ☐ dropping pipettes
- ☐ 30 cm^3 of solutions A, B and C
- ☐ 2 mol/dm^3 nitric acid
- ☐ barium chloride solution
- ☐ silver nitrate solution
- ☐ 2 mol/dm^3 hydrochloric acid
- ☐ distilled/deionised water

Safety!

Nitric acid (2 mol/dm^3) – corrosive

Hydrochloric acid (2 mol/dm^3) – irritant

Barium chloride solution – toxic

Silver nitrate solution – corrosive, dangerous for the environment

Procedure

Test a

1. Put on your eye protection.
2. One-third fill a test tube with solution A.
3. To the solution, add a few drops of dilute nitric acid followed by a few drops of silver nitrate solution.
4. Record your observations in the table provided.

Test b

1. Put on your eye protection.
2. One-third fill a test tube with solution A.
3. To the solution, add about 1 cm^3 of dilute hydrochloric acid.
4. Record your observations in the table provided.

Test c

1 Put on your eye protection.

2 One-third fill a test tube with solution A.

3 To the solution, add a few drops of barium chloride solution.

4 Now add a few drops of dilute hydrochloric acid.

5 Record your observations in the table provided.

Wash out your test tubes with distilled/deionised water and then repeat tests **a, b** and **c** with solutions B and C.

Method

What are you testing for with test a, test b and test c? [3]

.....

.....

.....

Results and calculations

Unknown solution	Test	Observation	Inference
A	a		
	b		
	c		
B	a		
	b		
	c		
C	a		
	b		
	c		

[18]

Conclusion

Solution A is ..... . [1]

Solution B is ..... . [1]

Solution C is ..... . [1]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

.....

.....

Extension

1 Write ionic equations for the results you have obtained for A, B and C. [6]

.....

.....

.....

2 If the substances A, B and C were potassium chloride, copper(II) chloride and calcium chloride respectively, how could you modify your approach to identify these substances? [3]

.....

.....

.....

.....

3 In the case of the sulfate test, if the precipitate dissolves what does this tell you about the salt you have present? [1]

.....

.....

16.2 Using flame colours to identify unknown metal ions

Aim

To use flame tests to identify metal ions in solution.

Theory

If a clean nichrome wire is dipped into a metal compound and then held in the hot part of a Bunsen flame, the flame can become coloured. For example, a compound of lithium gives a red colour in a Bunsen flame. Certain metal ions may be detected in their compounds by observing their flame colours. Forensic scientists use this technique on an everyday basis to identify the metal ions in substances left at crime scenes.

You are given seven solutions. Each one has a different metal ion present in the compound found in that solution. Using flame tests you are expected to identify the metal ions present.

Apparatus and chemicals

- ☐ eye protection
- ☐ Bunsen burner
- ☐ test-tube rack
- ☐ test tube containing 4 M hydrochloric acid
- ☐ flame-test wire
- ☐ access to seven solutions to investigate, labelled A, B, C, D, E, F, G

Safety!
Hydrochloric acid (4 mol/dm^3) – irritant

Procedure

1. Put on your eye protection.
2. You have seven solutions labelled A, B, C, D, E, F, G.
3. Collect a sample of solution A.

4 Clean the nichrome flame-test wire. This is done by placing it in a roaring Bunsen flame (just above the blue cone) and then into the quite concentrated hydrochloric acid (extreme caution). The wire should then be put back into the Bunsen flame to ensure that it is completely free from contaminating substances.

5 Dip the clean nichrome flame-test wire into solution A. The colour of the flame gives you an indication of the metal ion present in the solution.

6 Put the nichrome flame-test wire into the Bunsen flame. Record the colour of the flame in the results table provided.

7 Repeat steps 3–6 with samples B, C, D, E, F, G. Record your results in the results table.

8 Use your textbook to help you identify the metal ions present.

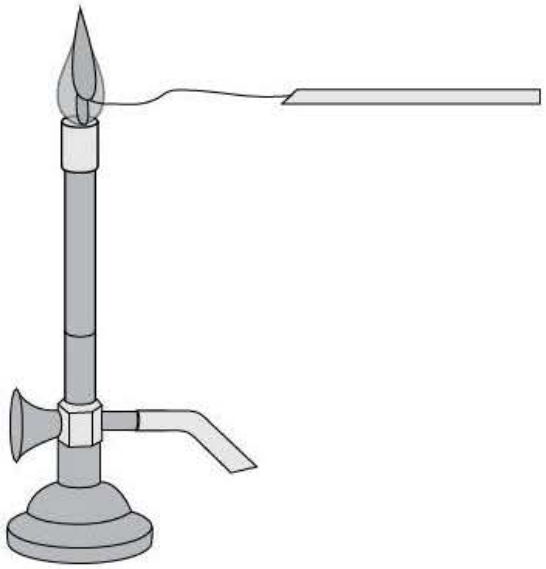


Figure 1

Method

1 Why is the nichrome wire cleaned in the position just above the blue cone of a Bunsen flame? [1]

.....

.....

2 Why have you been made aware of the fact that the acid is quite concentrated? [2]

.....

.....

.....

.....

3 Why do you use a solution rather than solid samples? [2]

.....

.....

Results and calculations

Table 1

Solution	Flame colour	Metal ion present
A		
B		
C		
D		
E		
F		
G		

[14]

Conclusion

..... ions can be identified by carrying out flame tests. These tests are

..... in nature. [2]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

Describe how you might confirm the presence of the metal ions you have identified. [2]

.....

.....

16.3 How pure is your water supply?

Aim

To use simple qualitative analysis procedures to identify potentially harmful substances in drinking water.

Theory

Drinking water has to be very pure. That means it has to be free from harmful or poisonous chemicals. For this reason, chemical analysts working for your water utility company have to check samples of your water supply regularly, if not every day.

Two of the main tests are quite simple. They involve the addition of dilute hydrochloric acid and observation of the result. The second test involves adding some dilute sodium hydroxide to excess and observing the result.

If a carbonate is present, then there will be bubbles of carbon dioxide produced:

carbonate ions + hydrogen ions → carbon dioxide + water

$CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

If a metal ion is present, then a precipitate may be seen when sodium hydroxide is added.

metal ions + hydroxide ions → metal hydroxide

$M^{2+}(aq) + 2OH^-(aq) \rightarrow M(OH)_2(s)$

The precipitate may dissolve when sodium hydroxide is added to excess.

Apparatus and chemicals

- ☐ goggles
- ☐ teat pipettes
- ☐ 10 × test tubes
- ☐ test-tube rack
- ☐ samples of water A, B, C, D, E
- ☐ 2 mol/dm^3 hydrochloric acid
- ☐ 2 mol/dm^3 sodium hydroxide

Safety!

Sodium hydroxide (2 mol/dm^3) – corrosive

Hydrochloric acid (2 mol/dm^3) – irritant

Procedure

1 Put on your goggles.

2 One-third fill a test tube with sample A.

3 Add dilute hydrochloric acid a drop at a time to this sample until no further changes take place.

4 Record your observations carefully in the results table.

5 Repeat steps 2–4 with samples B, C, D and E.

6 One-third fill a test tube with sample A.

7 Add dilute sodium hydroxide a drop at a time to this sample until no further changes take place.

8 Record your observations carefully in the results table.

9 Repeat steps 6–8 with samples B, C, D and E.

10 Use your textbook to help you identify the types of substances present in the different solutions.

Results and calculations

Table 1

Appearance of sample at start	Effect of the addition of dilute hydrochloric acid	Effect of the addition of dilute sodium hydroxide to excess
A		
B		
C		
D		
E		

[10]

Method

1 Why are the reagents added dropwise to the samples? [1]

.....

2 Why is excess dilute sodium hydroxide added? [1]

.....

Conclusion

Water samples may be tested for the presence of metal ..... by adding dilute hydrochloric acid. If a carbonate is present then bubbles of the gas ..... will be seen. The samples may also be tested for the presence of metal ions which may be harmful using ..... . White or a coloured ..... may be produced, from which an indication of the metal ion present can be obtained. If ..... of dilute sodium hydroxide is used then this will also help in the identification of specific metal ions. [5]

Evaluation

Outline how this experiment could be improved, or made more reliable. [2]

.....

.....

Extension

1 How could the use of aqueous ammonia help in the identification of metal ions? [2]

.....

.....

2 How could the presence of carbon dioxide gas be confirmed? [2]

.....

.....

Practical Test past exam questions

1 You are going to investigate the reaction between dilute sulfuric acid and three aqueous solutions of sodium hydroxide of different concentrations, labelled **A**, **B** and **C**.

Read all the instructions below carefully before starting the experiments.

Instructions

You are going to carry out three experiments.

(a) *Experiment 1*

Fill the burette with the dilute sulfuric acid provided to the 0.0 cm^3 mark.

Use a measuring cylinder to pour 20 cm^3 of solution **A** into a conical flask. Add a few drops of phenolphthalein indicator to the flask.

Add the sulfuric acid from the burette 1 cm^3 at a time, while shaking the flask, until the colour of the phenolphthalein changes. Record the burette readings in the table.

(b) *Experiment 2*

Fill the burette with dilute sulfuric acid to the 0.0 cm^3 mark.

Empty the conical flask and rinse it with water. Use a measuring cylinder to pour 20 cm^3 of solution **B** into the conical flask. Add a few drops of phenolphthalein to the flask.

Add the sulfuric acid from the burette 1 cm^3 at a time, while shaking the flask, until the colour of the phenolphthalein changes. Record the burette readings in the table.

(c) *Experiment 3*

Repeat Experiment 2, using solution **C** instead of solution **B**. Record your burette readings in the table and complete the table.

	experiment 1	**experiment 2**	**experiment 3**
final reading/cm^3			
initial reading/cm^3			
difference/cm^3			

[6]

(d) What colour change was observed after the sulfuric acid was added to the flask?

from ..... to ..... [2]

(e) What type of chemical reaction occurs when sulfuric acid reacts with sodium hydroxide? [1]

.....

(f) (i) Complete the sentences below.

Aqueous sodium hydroxide labelled ..... needed the smallest volume of sulfuric acid to change the colour of the phenolphthalein.

Aqueous sodium hydroxide labelled ..... needed the largest volume of sulfuric acid to change the colour of the phenolphthalein. [1]

(ii) The order of concentration of the solutions of sodium hydroxide is

least concentrated .....

↓ .....

most concentrated ..... [2]

(g) Compare the volumes of sulfuric acid used in Experiments 1 and 2. [1]

.....

(h) If Experiment 3 was repeated using 40 cm^3 of solution **C**, what volume of sulfuric acid would be used? [2]

.....

(i) What would be a more accurate method of measuring the volume of the aqueous sodium hydroxide? [1]

.....

(j) What would be the effect on the results if the solutions of sodium hydroxide were warmed before adding the sulfuric acid? Give a reason for your answer. [2]

effect on results .....

reason .....

.....

(k) Suggest a different method of finding the order of concentrations of the solutions of sodium hydroxide. [3]

.....

.....

.....

.....

[Total: 21]

(Cambridge IGCSE Chemistry 0620, Paper 53 Q1 November 2011)

2 You are provided with two salt solutions, **J** and **K**.

Carry out the following tests on **J** and **K**, recording all of your observations in the table.

Conclusions must **not** be written in the table.

tests	observations
tests on solution **J** **(a)** Describe the appearance of **J**.	[1]
(b) To about 1 cm^3 of the solution, add an equal volume of aqueous sodium hydroxide. Leave to stand for 5 minutes. Note any changes.	[2]
(c) To about 1 cm^3 of the solution, add an equal volume of hydrogen peroxide. Test the gas given off.	[3]
(d) To about 1 cm^3 of the solution, add about 1 cm^3 of aqueous ammonia.	[1]
(e) To about 1 cm^3 of the solution, add a few drops of dilute nitric acid followed by aqueous silver nitrate.	[1]
(f) To about 1 cm^3 of the solution, add a few drops of dilute nitric acid followed by barium nitrate solution.	[2]
tests on solution **K** **(g)** Describe the appearance of **K**.	[1]
(h) To about 1 cm^3 of the solution, add 5 drops of aqueous sodium hydroxide. Now add excess aqueous sodium hydroxide.	[3]
(i) To about 1 cm^3 of the solution, add about 2 cm^3 of aqueous sodium hydroxide and one spatula measure of aluminium powder. Heat the mixture gently. Test the gas given off.	[2]

(j) What conclusions can you draw about solution **J**? [3]

.....

.....

(k) What conclusions can you draw about solution **K**? [2]

.....

.....

[Total: 21]

(Cambridge IGCSE Chemistry 0620, Paper 53 Q2 November 2012)

3 You are going to investigate what happens when two different solids, **C** and **D**, react with excess dilute hydrochloric acid.

Read all the instructions below carefully before starting the experiments.

Instructions

You are going to carry out five experiments.

(a) *Experiment 1*

Use a measuring cylinder to pour 30 cm^3 of dilute hydrochloric acid into the polystyrene cup supported in the beaker provided. Measure the temperature of the dilute hydrochloric acid and record it in the table below. Add 1 g of solid **C** to the dilute hydrochloric acid and stir the mixture with the thermometer.

Measure the maximum temperature reached by the liquid mixture. Record your result in the table.

(b) *Experiment 2*

Empty the polystyrene cup and rinse it with water.

Repeat Experiment 1 using 2 g of solid **C**.

Record your results in the table.

(c) *Experiments 3 and 4*

Repeat Experiment 2 using 3 g and then 5 g of solid **C**.

Record your results in the table.

Complete the final column in the table.

Experiment	mass of solid C/g	initial temperature of acid/°C	maximum temperature reached/°C	temperature change/°C
1				
2				
3				
4				

[4]

(d) *Experiment 5*

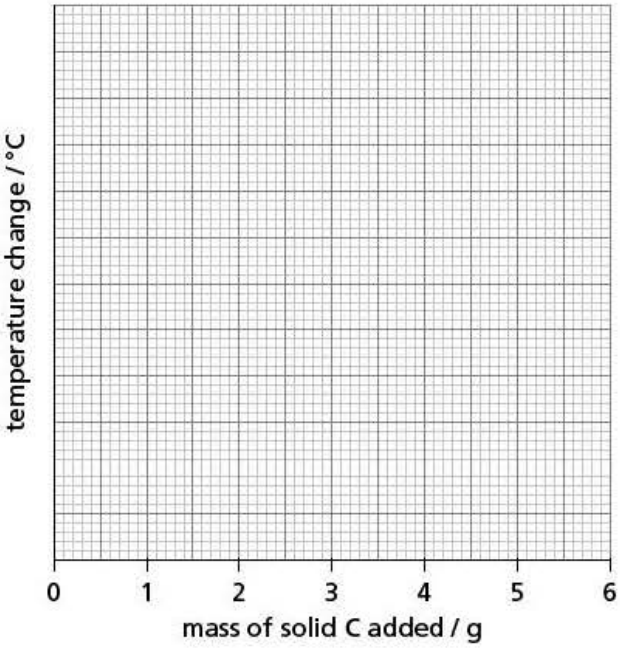
Repeat Experiment 1 using all of the solid **D** provided. Measure the minimum temperature reached by the liquid mixture. Record your results in the spaces below.

Initial temperature of dilute hydrochloric acid = ...°C

Final temperature of liquid mixture = ...°C

Temperature change = ...°C [2]

(e) Plot the results for Experiments 1, 2, 3 and 4 on the grid and draw a straight line graph.



[4]

(f) (i) **From your graph**, deduce the temperature change of the solution when 6g of solid **C** is added to 30cm^3 of dilute hydrochloric acid.

Show clearly **on the grid** how you worked out your answer.

..... °C [2]

(ii) **From your graph**, deduce the mass of solid **C** that would give a temperature rise of 9°C when added to 30cm^3 of dilute hydrochloric acid.

Show clearly **on the grid** how you worked out your answer. [2]

.....

(g) What type of chemical process occurs when solid **D** reacts with dilute hydrochloric acid? [1]

.....

(h) Suggest the effect on the results if Experiment 3 was repeated using 60 cm^3 of dilute hydrochloric acid. [2]

.....

.....

(i) Predict the temperature of the solution in Experiment 4 after 1 hour. Explain your answer. [2]

.....

.....

(j) When carrying out the experiments, what would be **one** advantage and **one** disadvantage of taking the temperature readings after exactly 1 minute? [2]

advantage .....

.....

disadvantage .....

.....

[Total: 21]

(Cambridge IGCSE Chemistry 0620, Paper 51 Q1 June 2013)

4 You are going to investigate what happens when aqueous sodium hydroxide reacts with acid **K**.

Read all the instructions below carefully before starting the experiments.

Instructions

You are going to carry out two experiments.

(a) *Experiment 1*

Use a measuring cylinder to pour 25cm^3 of acid **K** into a conical flask. Add five drops of phenolphthalein to the flask.

Fill the burette with the aqueous sodium hydroxide to the 0.0cm^3 mark.

Slowly add the aqueous sodium hydroxide to acid **K** in the flask and shake the mixture.

Continue to add aqueous sodium hydroxide to the flask until the solution shows a permanent colour change.

Measure and record the volume in the table. Complete the table.

Pour the solution away and rinse the conical flask.

	burette reading
final volume/cm^3	
initial volume/cm^3	
difference/cm^3	

[3]

(b) *Experiment 2*

Use a measuring cylinder to pour 50 cm^3 of acid **K** into a conical flask. Add the 0.3 g of powdered calcium carbonate to the flask and shake the flask until no further reaction is observed.

Add five drops of phenolphthalein to the mixture in the flask.

Fill the burette with aqueous sodium hydroxide and record the burette reading. Slowly add aqueous sodium hydroxide from the burette to the flask and shake the mixture. Continue to add aqueous sodium hydroxide to the flask until the solution shows a permanent colour change.

Measure and record the volume in the table. Complete the table.

	burette reading
final volume/cm^3	
initial volume/cm^3	
difference/cm^3	

[3]

(c) What colour change was observed after the sodium hydroxide solution was added to the flask? [2]

from ..... to .....

(d) What type of chemical reaction occurs when acid **K** reacts with sodium hydroxide? [1]

.....

(e) If Experiment 1 was repeated using 50 cm^3 of acid **K**, what volume of sodium hydroxide would be required to change the colour of the indicator? [2]

.....

(f) (i) What is the effect of adding 0.3 g of powdered calcium carbonate to acid **K**? [2]

.....

.....

(ii) Use your answers from **(b)** and **(e)** to work out the difference in the volume of sodium hydroxide added when 0.3 g of calcium carbonate is mixed with 50 cm^3 of acid **K** in Experiment 2. [2]

.....

.....

(iii) Estimate the mass of calcium carbonate that would need to be added to 50 cm^3 of acid **K** to require 0.0 cm^3 of sodium hydroxide. [1]

.....

(g) What would be the effect on the results if the solutions of acid **K** were warmed before adding the sodium hydroxide? Give a reason for your answer. [2]

effect on results .....

reason .....

(h) Suggest the advantage, if any, of

(i) using a pipette to measure the volume of acid **K**. [2]

.....

.....

(ii) using a polystyrene cup instead of a flask. [2]

.....

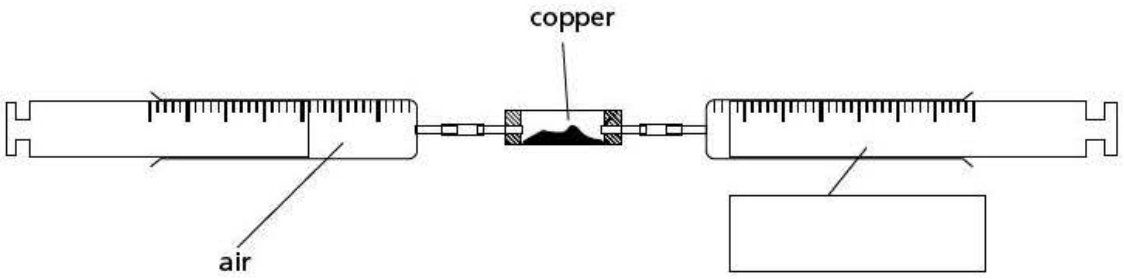
.....

[Total: 22]

(Cambridge IGCSE Chemistry 0620, Paper 52 Q1 November 2013)

Alternative to Practical past exam questions

1 A student investigated the reaction of air with copper. 100 cm^3 of air was passed continuously over heated copper using the apparatus below. When the volume remained constant, the apparatus was left to cool and the volume of gas was measured.



(a) (i) Complete the box to show the apparatus labelled. [1]

(ii) Indicate on the diagram, with an arrow, where heat is applied. [1]

(b) What should be used to transfer the copper from a bottle to the apparatus? [1]

.....

(c) The copper changed colour from brown to ..... [1]

(d) Why was the apparatus left to cool before measuring the final volume of gas? [2]

.....

.....

[Total: 6]

(Cambridge IGCSE Chemistry 0620, Paper 61 Q1 November 2011)

2 Coffee beans contain caffeine and other compounds. Caffeine is soluble in water and in trichloromethane, an organic solvent.

A student obtained crystals of caffeine by the following method.

Stage 1 Some coffee beans were crushed into small pieces.

Stage 2 Hot water was added to the crushed beans to dissolve the soluble substances.

Stage 3 The crushed beans were separated from the liquid solution.

Stage 4 The liquid was allowed to cool and shaken with trichloromethane to extract the caffeine from the water.

Stage 5 The caffeine was crystallised from the trichloromethane solution.

Stage 6 The caffeine crystals were checked for purity.

(a) What apparatus should be used to crush the beans in Stage 1? [2]

.....

(b) How could the dissolving process in Stage 2 be speeded up? [1]

.....

(c) Draw a diagram of the apparatus used in Stage 3. [2]

(d) How should Stage 5 be carried out? [2]

.....

.....

(e) What method could be used to check the purity of the crystals in Stage 6? [1]

.....

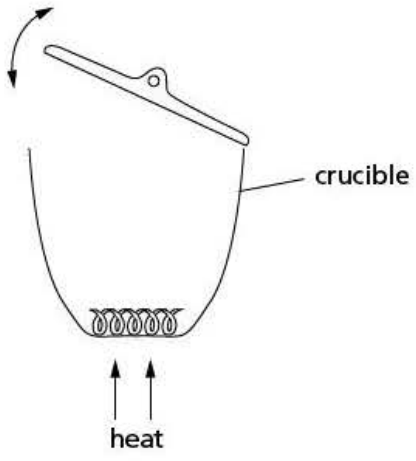
[Total: 8]

(Cambridge IGCSE Chemistry 0620, Paper 61 Q3 June 2012)

3 A student carried out an experiment to find the mass of magnesium oxide formed when magnesium burns in air.

A strip of magnesium ribbon was loosely coiled and placed in a weighed crucible, which was then reweighed.

The crucible was heated strongly for several minutes. During the heating, the crucible lid was lifted and replaced several times as in the diagram.



The magnesium was converted into magnesium oxide. After cooling, the crucible and contents were reweighed.

(a) Describe the appearance of the

(i) magnesium ..... [1]

(ii) magnesium oxide. ..... [1]

(b) Name the element that reacted with the magnesium. [1]

.....

(c) Why was the lid lifted during heating? [1]

.....

(d) Suggest why the mass of the magnesium oxide was found to be **lower** than expected. [2]

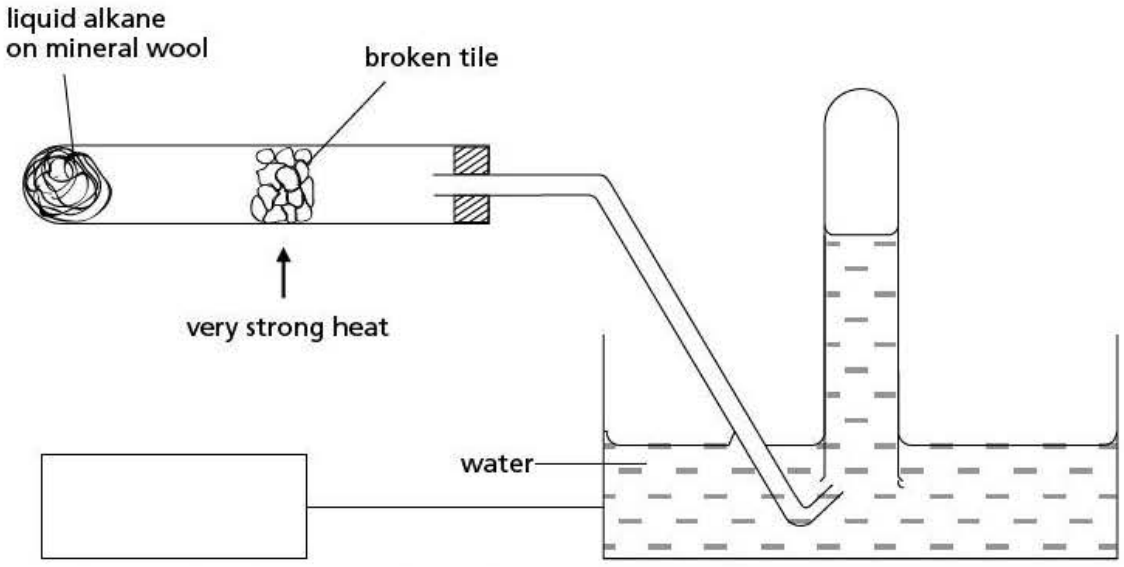
.....

.....

[Total: 6]

(Cambridge IGCSE Chemistry 0620, Paper 61 Q3 November 2012)

4 Alkenes can be made by cracking long chain alkanes. A student used the apparatus below to demonstrate cracking.



(a) Complete the box to show the apparatus used. [1]

(b) Indicate with an arrow where the alkenes are collected. [1]

(c) Suggest the purpose of the mineral wool. [1]

.....

(d) Why are **small** pieces of broken tile used? [1]

.....

.....

(e) Describe a test to show that alkenes have been made. [2]

test .....

result .....

[Total: 6]

(Cambridge IGCSE Chemistry 0620, Paper 62 Q1 November 2013)