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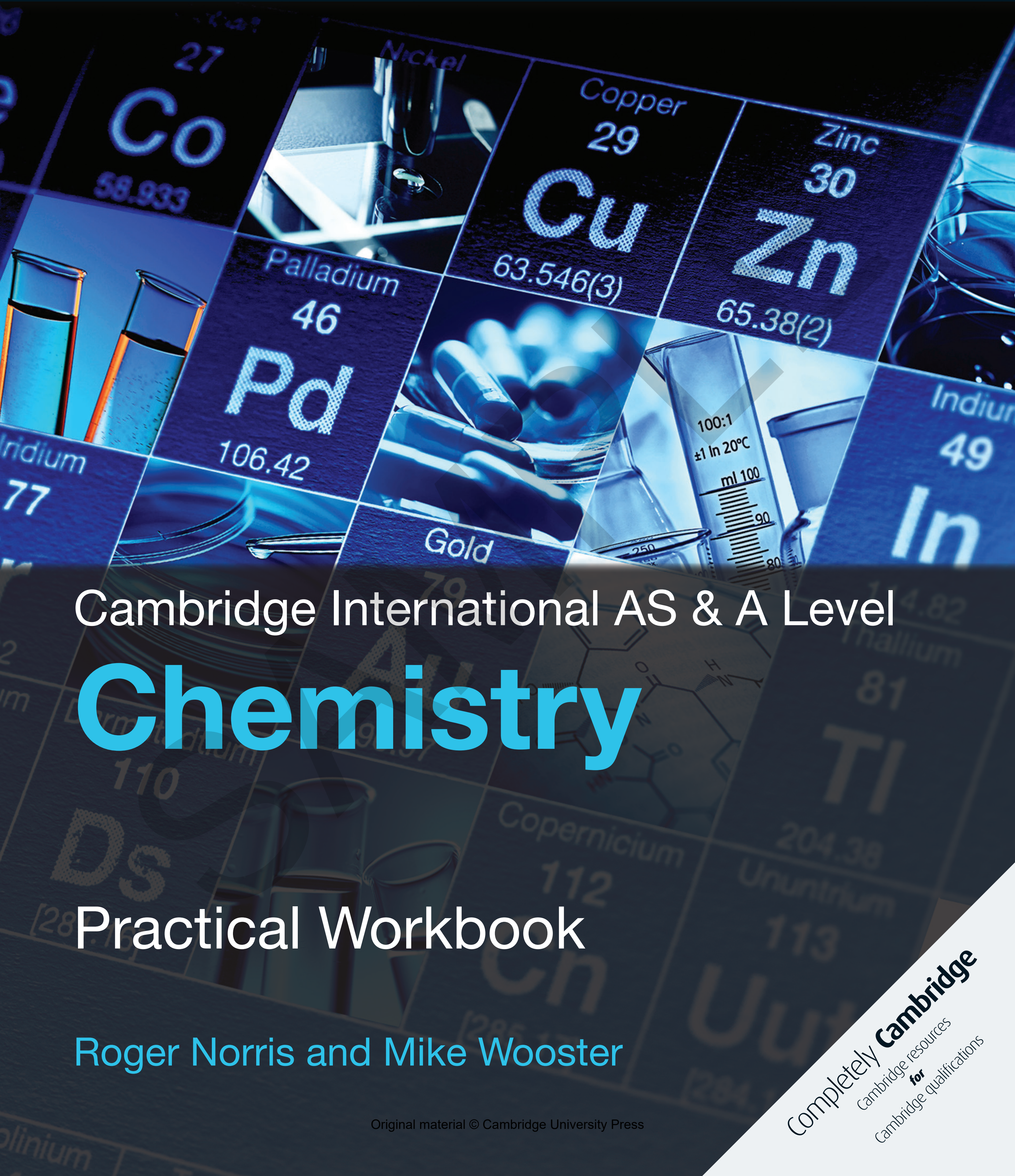

Cambridge International AS & A Level

Chemistry

Practical Workbook

Roger Norris and Mike Wooster

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Introduction

Practical work is an essential part of your advanced Chemistry course. Experimental investigations allow you to gain first-hand experience of the arrangement and names of chemical apparatus and how this apparatus is used to obtain meaningful experimental results. For Cambridge International AS & A Level Chemistry Paper 3 and Paper 5 focus on the assessment of practical skills.

The practical investigations in this workbook have been carefully chosen to allow you to practise and improve your practical skills. The practical work introduced in this workbook emphasises the spirit of enquiry and first-hand experience that helps reinforce your knowledge and helps you apply the results and draw conclusions. It also helps you to test your knowledge and application of theoretical work.

The order of the investigations presented largely follows the order of the topics in the Cambridge International AS & A Level Chemistry coursebook. This does not mean that this is the order that will be chosen by your teacher. Some coursebook chapters involve the use of quantitative techniques and when you carry out these investigations you will need calculators and equipment for drawing graphs. All techniques listed in the practical guidance are covered in the workbook.

There are two parts to this practical guide.

The first part deals with subject matter and practical techniques described in the AS level syllabus. A variety of investigations introduces you to a range of experiments which will provide you with practice in manipulating apparatus and taking measurements. Some investigations also ask you to present and analyse data and observations and/or give you practice in drawing conclusions and evaluating information.

The second part deals with subject matter and practical techniques described in the A Level syllabus. This part gives you practice in planning experiments, analysing results, drawing conclusions and evaluating information. The syllabus stresses that ‘candidates cannot be adequately prepared for Paper 5 without extensive laboratory work’. With this in mind, some of the investigations give you further practice in laboratory work as well as giving you the opportunity to analyse information, draw conclusions and evaluate the experiment. A number of open-ended investigations have also been included, which give you only the basic information to enable you to plan and carry out an experiment in the laboratory. Other investigations are set in areas of Chemistry that may be new to you or are difficult to investigate experimentally in a school laboratory. In these cases, relevant information is always given so that you can complete the investigations successfully.

The various investigations and accompanying questions will help you gain confidence in tackling laboratory work and develop a wide range of skills related to practical chemistry. Apart from the necessary preparation for both practical papers, it is hoped that these investigations will help you understand the importance of laboratory work in development and assessment of theoretical chemistry.

Safety

Practical work has its own set of skills. A number of these are related to working safely. Working safely is essential in getting the maximum advantage from your practical work.

In each investigation involving practical work, you are expected to:

1. Wear eye protection such as safety goggles or safety spectacles (note that goggles give more protection)
2. Tie back long hair and make sure that items of clothing are tucked in
3. Wear gloves when weighing, pouring or filtering hazardous chemicals.

It is also advisable for you to wear a laboratory coat to protect your clothing from chemical splashes.

All chemicals should be treated as hazardous. If they are spilt on the skin you must wash them off immediately using plenty of water. You may not be aware of the dangers of particular chemicals and therefore using them without safety precautions can lead to unforeseen problems. Remember that you should also think about the hazards of all of the substances that are being produced in a chemical reaction, especially when a gas is given off. Chemical reactions which produce hazardous gases should be done in a fume cupboard or well-ventilated room.

As a learner you should take responsibility for working safely and you must learn the meanings of the safety symbols shown in the table below. Table S1 shows the most common hazard symbols found in school science laboratories. An up-to-date list of CLEAPPS hazards can be downloaded from the internet.

Hazard symbol	What does it mean?	Special points
	The substance is **corrosive**. It will damage your skin and tissues if it comes into contact with them.	Always wear safety goggles and if possible gloves when using corrosive substances
	The substance is an **irritant**. If it comes into contact with your skin it can cause blisters and redness.	Always wear safety spectacles when using irritants.
	The substance is **toxic** and can cause death if swallowed, breathed in or absorbed by skin.	Wear gloves and eye protection.
	The substance is **flammable** and catches fire easily.	Keep the substance away from naked flames and when heating reaction mixtures use the hot water from a kettle rather than using Bunsen burners.
	The material is a **biohazard**. Examples are bacteria and fungi.	Seek advice about particular biohazards.
	The substance is an **oxidising agent**. It will liberate oxygen when heated or in the presence of a suitable catalyst.	Keep oxidising agents well away from flammable materials.

Table S1

Skills chapter

Chapter outline

This chapter introduces the key practical methods, processes and procedures that you will use regularly throughout your course. Within the investigations you'll find cross-references to the techniques covered in this chapter so you are encouraged to refer back to the relevant sections whenever you need to.

This chapter covers the following:

- **Preparing a standard solution**
  - **a** Calculating the mass of solute required
  - **b** Making 250cm^3 of a standard solution
- **Carrying out an acid–base titration**
  - **a** preparing the burette
  - **b** pipetting a solution
  - **c** carrying out the titration
  - **d** processing your results
- **Gas collection and measurement**
  - **a** Choosing your apparatus
  - **b** General advice for measuring volumes of gases
- **Qualitative analysis: testing for gases and ions**
  - **a** Tests for gases
  - **b** Tests for ions
- **Strategies for measuring heat changes**
  - **a** Temperature–time graphs
  - **b** Calculating enthalpy (heat) changes
- **Drawing graphs and charts**
- **Calculating errors in your experiments**
  - **a** Calculating systematic errors due to apparatus inaccuracy
    - **i** Temperature readings
    - **ii** Measuring cylinders
    - **iii** Burette readings
    - **iv** Top-pan balance readings
- **Using significant figures**

Preparing a standard solution

Introduction

A standard solution is one that has a **known concentration.** With a standard solution, it is possible to investigate the concentration of other solutions of unknown concentration by titration (see section 2). A standard solution is made by dissolving an accurate mass of solute into a known volume of water. The first step is to calculate the mass of solute required to make up a standard solution. For example, if asked to prepare 250 cm^3 (0.25 dm^3) of a 0.100 mol dm^{-3} sodium carbonate solution you must first calculate what mass of sodium carbonate you need to weigh out.

In the equations used for calculating amounts and concentrations, the symbols refer to the following quantities:

C = concentration (units = mol dm^{-3})

n = number of moles

V = volume (units = dm^3)

m = mass (units = g)

M_r = molar mass (units = g mol^{-1})

Please note that not all substances make good standard solutions. This is due to the fact that some substances can be difficult to obtain in a completely pure form, are unstable in air or not readily soluble in water.

a Calculating the mass of solute required

Before you start to prepare your solution you need to calculate the mass of solute you will need to weigh out using the relationships:

$C = \frac{n}{V}$ and $n = \frac{m}{M_r}$.

The calculations to work out the mass of sodium carbonate (M_r = 106) required to prepare 250 cm^3 of a 0.100 mol dm^{-3} solution are shown in Figure 0.1

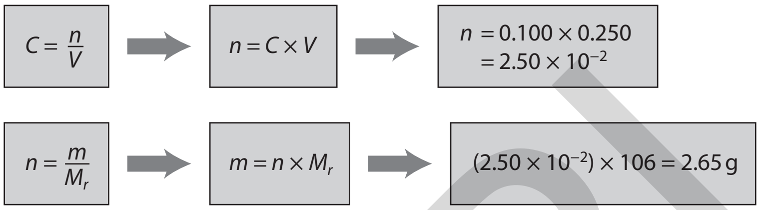

Figure 0.1

b Making 250 cm^3 of a standard solution

Equipment

You will need:

- top-pan balance and weighing boat
- 250 cm^3 beaker
- glass or plastic stirring rod
- filter funnel
- plastic dropper for delivering small volumes
- 250 cm^3 volumetric flask
- eye-protection

Access to:

- distilled water in wash bottle
- solute

HINT

you must make sure the volume of solution in the volumetric flask **does not go over the mark** on the neck of the volumetric flask

Method

1. Use the weighing boat to weigh out the required amount of solute. Empty it into a 250 cm^3 beaker. To ensure there is no solute remaining in the weighing boat, wash the weighing boat twice using distilled water from a wash bottle pouring the washings into the beaker.
2. Add more water to the beaker so that you have about 100 cm^3. Stir the mixture with the stirring rod until all the sodium carbonate has dissolved.
3. Place the filter funnel into the neck of the 250 cm^3 volumetric flask and pour the contents of the beaker into the flask.
4. Using a wash bottle rinse the beaker and pour the washings into the volumetric flask. Repeat this several times. You must also rinse the stirring rod.
5. When the level of the liquid is just a few cm^3 below the mark on the neck of the volumetric flask, take the dropper and **with great care** use it to add distilled water from the wash bottle to the solution one drop at a time until the lowest point of the meniscus is touching the line as shown in Figure 0.2

HINT

If you go over the mark and the level of the liquid is above the line then you must reweigh your solute and repeat the preparation of the solution

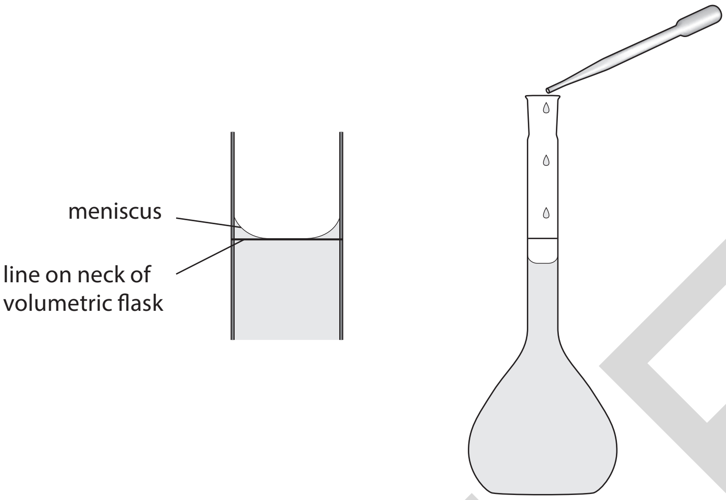

Figure 0.2

6 Place the stopper in the neck of the volumetric flask and **keeping the stopper firmly in the neck using your thumb** mix the solution by turning the flask upside down at least five or six times (see Figure 0.3). If you move the flask and still see swirling currents in the liquid you have not mixed enough – just turn upside down a few more times.

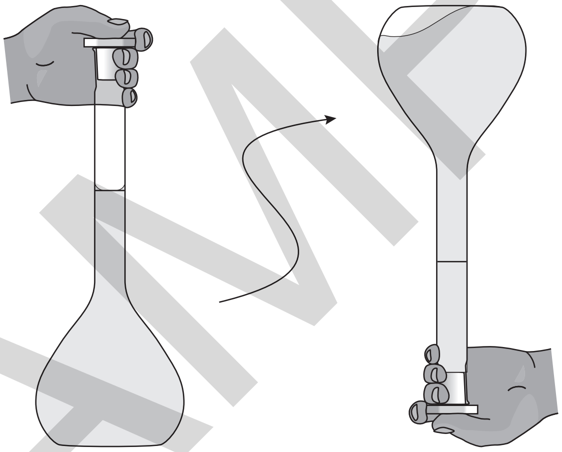

Figure 0.3

Carrying out an acid–base titration

Introduction

Titrations are used to measure the volume of one solution that exactly reacts with another solution. Titration is an analytical technique widely used in industry and is an essential chemical skill. The food industry, for example, uses titrations to determine the amount of salt or sugar in a product or the concentration of beneficial vitamins. Acid–base titrations involve neutralisation between an acid and a base when mixed in solution. An indicator is used to determine the end-point of the titration as it changes colour. This technique is also used in other areas of your syllabus, for example redox reactions.

Additional advice

- When doing acid–base titrations, it is best if the acid is delivered from a burette. This is because alkalis and soluble carbonates can form solids in the taps of burettes and clog them up.
- Burettes should be clamped firmly but not too tightly.
- It is often a good idea to place the burette and clamp stand on a chair or stool. This will make it easier to fill the burette.
- **Never fill a pipette by mouth.** Always use a pipette filler (see Figure 0.5).
- Most pipette fillers have a way of pushing the liquid out of the pipette. Unfortunately, this method is often very difficult to use to the fine level necessary.

Equipment

You will need:

• burette • burette stand or clamp stand • boss and clamp • pipette and pipette filler
• white tile • conical flask • 100 cm^3 and 250 cm^3 beakers • protective gloves • eye protection

Access to:

• distilled water • an indicator (e.g. methyl orange)

a preparing the burette

Method

1 Set up the burette in a burette or clamp stand; it should be clamped firmly but not too tightly.

2 Place a filter funnel in the neck of the burette. Close the tap on the burette.

3 Wearing gloves, add some of the acid to a **dry** 100 cm^3 beaker. Use the beaker to add a few cm^3 of acid to the burette; you are only rinsing the burette at this stage.

4 Open the tap, run out the acid rinse. Ensure the tap is closed, then fill the burette to **above** the zero mark. Remove the funnel from the burette.

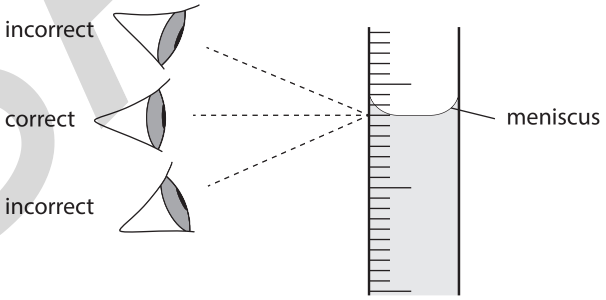

Figure 0.4

HINT
When reading a burette you must always **view it at eye level** as shown in Figure 0.4 so that you accurately record the level of the lowest point of the solution's meniscus

5 Run out some of the acid into the 100 cm^3 beaker to ensure the jet at the bottom of the burette is full and there is no air in it. Remove the funnel from the top of the burette.

6 Read and record the starting level of acid in the burette.

HINT
There is no need to adjust the volume of the solution to exactly 0.00 cm^3 but you must note the starting volume

b pipetting a solution

A pipette and pipette filler is used to measure out an accurate volume of the alkali or soluble carbonate in an acid–base titration.

Method

1 Pour a volume of the alkali/soluble carbonate into a dry 250 cm^3 beaker.

2 Using a pipette filler, fill the pipette (e.g. a 25.0 cm^3 pipette) up a little way above the line on the neck. Quickly remove the pipette from the pipette filler and cover the open end with your index (first) finger as shown in Figure 0.5(b).

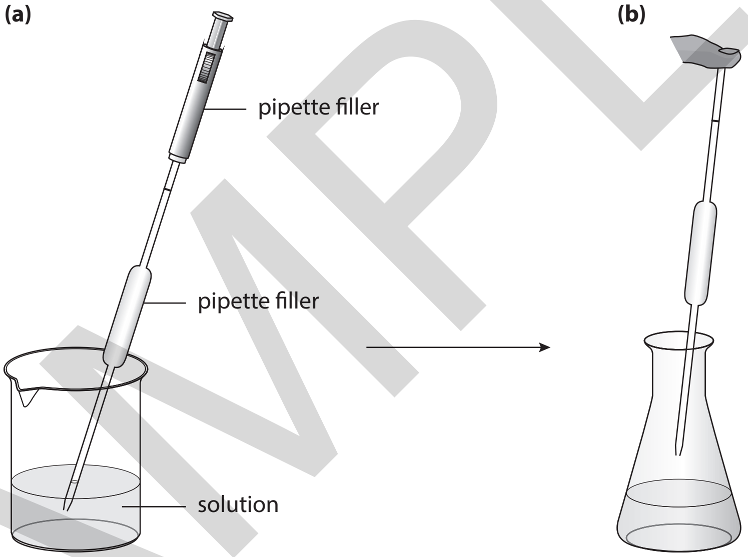

Figure 0.5

3 By releasing your index finger you can now let the solution out of the pipette. When the solution's meniscus is on the line on the neck of the pipette you have exactly 25.0 cm^3 of solution in the pipette.

4 The solution can now be transferred to the conical flask. When the end of the pipette is over the conical flask release your index finger and let the solution run into the flask.

HINT
There will be a very small amount of solution in the end of the pipette. **Do not add this small drop of solution!** The pipette is calibrated to deliver the exact volume with this drop remaining in the pipette

c carrying out the titration

Now that you have your burette prepared and filled with acid and an accurate amount of your alkali/soluble carbonate in a conical flask, you are ready to start your acid/base titration. You will need to repeat the titration at least three times, usually more. The first titration is a rough titration which will help you to be more accurate when you repeat the process. You will need to prepare a results table in which to record your results.

1 Place the conical flask on a white tile directly under the outlet of the burette.

2 Add 2–3 drops of the indicator provided.

3 Remember that the first titration is a **rough** titration. You will add acid from the burette 1.00 cm^3 at a time. After each addition, swirl the flask and if the indicator does not change colour continue adding 1.00 cm^3 at a time until it does.

HINT
Remember that you don't have to adjust the level of acid in the burette to exactly 0.00 cm^3 as this is time consuming and unnecessary; just record the exact starting level in the burette

4 Note the volume of the acid added. What does this result tell you? If the indicator changed colour after 24.00 cm^3 of acid was added then you know that the end-point of the titration was somewhere between 23.00 and 24.00 cm^3. You now know that you can safely run in 23.00 cm^3 of acid from the burette without the indicator changing colour.

5 Wash out the conical flask with plenty of tap water and then rinse with distilled water ready for your second titration.

6 Using the pipette add another 25.00 cm^3 of the solution of the base to the flask and add another 2–3 drops of indicator.

7 At the point when you have added 1 cm^3 less than the volume recorded in your rough titration, you now need to add one drop of acid at a time, swirling the conical flask as you do so. When you are near the end-point, the colour of the indicator will take longer to return to its original colour. As soon as the colour does not change back you know you have added exactly the right amount of acid. Note down the volume.

8 Wash out the conical flask with plenty of tap water and then rinse with distilled water ready for your third titration. If you have enough acid left in your burette to repeat the titration, go ahead; if not you will need to fill the burette up again (taking care to record the starting level) and then repeat.

9 Continue to repeat the method until you have **at least two concordant (consistent) results**. Then you will know that you have accurately estimated the volume of acid required to react with the 25.00 cm^3 of your alkali/soluble carbonate.

d processing your results

Before completing calculations using your results you need to check and process your results to determine the average titre.

Figure 0.6 and Table 0.1 provide sample burette readings and a results table which are reviewed.

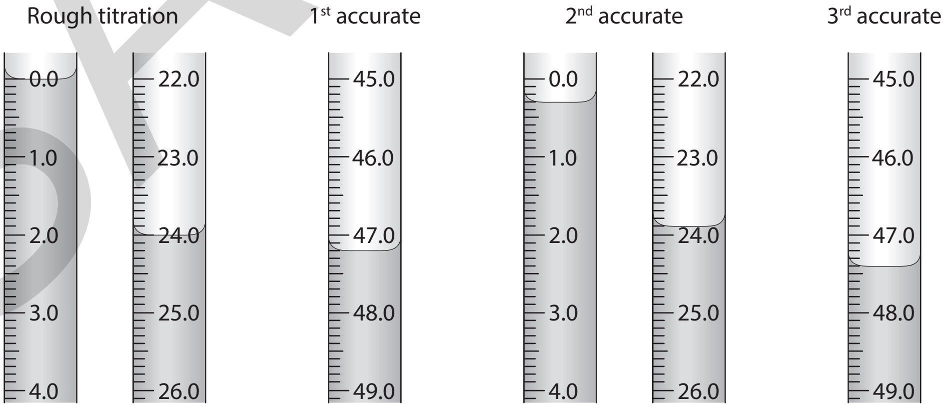

Figure 0.6

	Rough titration/ cm^3	First accurate titration/cm^3	Second accurate titration/cm^3	Third accurate titration/cm^3
Final burette reading	24.00	47.20	23.90	47.40
Initial burette reading	0.00	24.00	0.30	23.90
Titre /cm^3	24.00	23.20	23.60	23.50

Table 0.1

Review of readings and results:

- The results are recorded to two decimal places but this is not because the burette can be read to 0.01 cm^3. The burette is accurate to ±0.05 cm^3.
- The initial burette reading is usually recorded on the second line of the table to aid the calculation of the titre.
- The initial burette reading for the second accurate titration is 0.30 cm^3. Note that no time was wasted in adjusting the volume to 0.00 cm^3. In this titration, acid was added until a reading of 23.30 cm^3 before then being added drop by drop.
- **Three** accurate titrations were necessary because the first two titre values were not close enough in value (concordant).
- The titres for the second and third accurate titrations were concordant so there was no need to do any further titrations.
- The average titre was calculated using the second and third accurate titrations:

$$\frac{23.60 - 23.50}{2} = 23.55\,\text{cm}^3$$

Gas collection and measurement

Introduction

You can investigate a chemical reaction by measuring the volume of gas given off (evolved) at certain time intervals (when investigating reaction rates), or measuring the total volume of gas produced. There are different techniques to collect gas during an experiment and your choice of apparatus depends on the volume of gas produced and the apparatus that is available. The gas produced must be only slightly soluble or insoluble in water.

a Choosing your apparatus

Two common methods for collecting gas are shown in Figure 0.7. In (a) the gas produced is collected in a gas syringe. In (b) the gas is collected by the displacement of water. It is a suitable method for gases that are insoluble in water such as hydrogen.

It is important to have some idea of the volume of gas that will be generated during your experiment so that you can choose an appropriate size of syringe or measuring cylinder. The volume of measuring cylinder chosen should be about 2–3 times the volume of gas.

Remember that the larger the volume of measuring cylinder that you use, the greater the error in measurement (e.g. if the volume to be collected is 12 cm^3 then the ideal size of measuring cylinder is 25 cm^3).

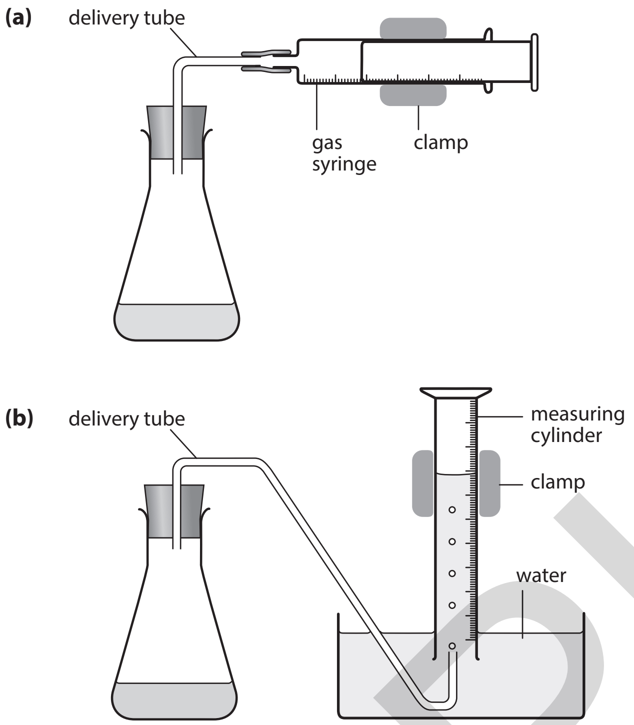

Figure 0.7

b General advice for measuring volumes of gases

- After selecting your apparatus it will be necessary to complete a range of measurements. Trial runs are therefore essential to make sure that your approach will work.

  Example 1
  When asked to investigate the effect of concentration of hydrochloric acid (e.g. 0.100–2.00 mol dm^{-3}) on magnesium ribbon, you would collect gas using the displacement of water (Figure 0.7 (b)).

- It is advisable to complete trial runs using the lowest and highest values of your range that you intend to use in your experiment (e.g. in Example 1, you would use the lowest and highest acid concentrations: 0.100 and 2.00 mol dm^{-3}). This will indicate which size of measuring cylinder is correct to use.

- When you are measuring the volume at different times it can be easily read at the correct time if you start reading the volume a few seconds before the required time and count down.

Qualitative analysis: testing for gases and ions

Introduction

Knowing how to identify different ions and gases is a key skill for all chemists. In particular it is important to understand the chemical basis for each test. The practical exam will test your knowledge of common tests and their expected results.

a Tests for gases.

After collecting gas during an experiment, it will be necessary to complete a test to establish what gas you have. The tests for common gases (and the method used in the testing) are shown in Table 0.2.

Gas	Test and result	Method
Carbon dioxide (CO_2)	Bubble gas through limewater (calcium hydroxide) solution. Turns cloudy in presence of CO_2.	Using dropper, collect gas from above the surface of the reaction mixture; bubble gas through limewater solution.
Hydrogen (H_2)	Use lighted splint. H_2 produces squeaky 'pop' when burnt.	Collect gas in upside down test tube above reaction mixture. Then insert lighted splint into test tube.
Oxygen (O_2)	Use glowing splint. O_2 relights glowing splint.	Collect gas using displacement of water. Insert glowing splint into the measuring cylinder.
Ammonia (NH_3)	Universal Indicator (UI) or red litmus paper. NH_3 turns moist UI or red litmus paper blue/purple.	Hold indicator paper at mouth of test tube. Must be carried out in a fume cupboard.
Chlorine (Cl_2)	Universal Indicator (UI) or blue litmus paper. Chlorine gas bleaches moist UI or blue litmus paper.	Hold indicator paper at mouth of test tube. Must be carried out in a fume cupboard.
Sulfur dioxide (SO_2)	Acidified potassium manganate(VII) solution. SO_2 decolourises acidified potassium manganate(VII) solution on filter paper.	Add 2–3 drops of potassium manganate(VII) solution to filter paper and hold near mouth of test tube. Must be carried out in a fume cupboard.

Table 0.2

b Tests for ions

When presented with a substance or solid to analyse it may be necessary to first prepare a solution of the compound. Here is some general advice regarding the making and testing of solutions.

1. Do not be tempted to make your solution too concentrated
2. When testing for ions it is sensible to add the test solution a drop at a time
3. When testing for anions you must first add the appropriate acid before the testing solution.

Example 1
When testing for halide ions you use silver nitrate solution. But before adding the silver nitrate solution, you add **nitric acid**.

Example 2
When testing for sulfate ions you use barium chloride solution. But before adding the barium chloride solution, you add a drop or two of **hydrochloric acid**, but sulfate ions (SO_4^{2-}) can be distinguished from sulfite (SO_3^{2-}) ions by first adding the barium chloride and then adding the hydrochloric acid.

Strategies for measuring heat changes

Introduction

One of the main challenges when taking measurements of heat change is avoiding heat loss, mainly by conduction and convection. There is also the issue that should the reaction take a long time to go to completion then the maximum temperature may not be reached. Ideally there would be no heat loss; the reaction would take place immediately and it would be possible to measure the temperature immediately.

In order to overcome these challenges it is necessary when investigating heat change to construct a temperature–time graph

a Temperature–time graphs

A typical temperature–time graph is shown in Figure 0.8.

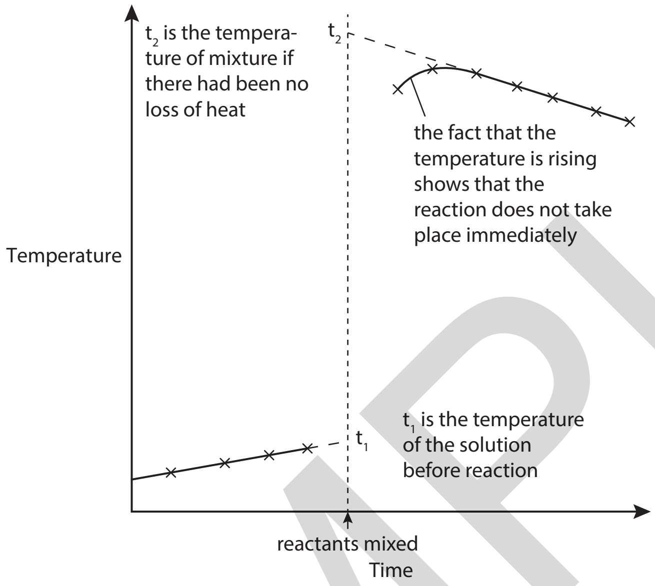

Figure 0.8

The experiment shown in Figure 0.8 was conducted and measurements, which were later plotted, were recorded as follows:

1. The temperature of the reacting solution was measured each minute for the first 4 min.
2. On the fifth minute the temperature was not measured; the reactants were mixed by stirring them together.
3. On the sixth minute the temperature was measured and every minute afterwards for as long as it was deemed necessary (in this case for a further 6 min).

Review Figure 0.8. You can see that the first four measurements are plotted but that the temperature of the reaction mixture at 5 min is obtained by extrapolating the line to the vertical line at 5 min. Similarly, after 5 min the temperature at 5 min is obtained by extrapolating back to the line at 5 min.

The temperature change = $t_2 - t_1$.

HINT
Enthalpy change is the heat change at constant pressure

b Calculating enthalpy (heat) changes

The enthalpy change produced by a reaction in a solution is represented by the following formula:

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

Unit = Joules (J)

Where m = mass of solution being heated up or cooled down, c = specific heat capacity of the solution and ΔT = temperature change.

Usually a solution being heated or cooled is an aqueous solution. The mass of solution (g) is therefore assumed equal to the volume (cm^3).

This is because under the conditions of the experiment it is **assumed** that the density of water = exactly 1.00 g cm^{-3}.

The value of *c* for water = 4.18 J g^{-1} K^{-1} and it is **assumed** that the value of the solution has the same value.

Example 1
If 50 cm^3 of water is heated up by 8 °C (≡ an 8 °K increase) then the enthalpy change is calculated as follows:

$$q = 50 \times 4.18 \times 8\ \text{J} = 1672\ \text{J}$$

$$= 1.67\ \text{kJ (three significant figures)}$$

Drawing graphs and charts

Introduction

Drawing accurate graphs and charts is often an essential part of the analysis of experiments. There are many occasions within this workbook when you are asked to first record experimental data and then produce a graph. Here are some general tips and advice to remember:

1 Bar charts (see Figure 0.9(a)) are used to present categorical variables while line graphs (see Figure 0.9(b)) are used for continuous variables.

Example 1
If you are investigating the effect of surface area on the rate of a chemical reaction you might complete an experiment to compare a solid or powder form of a substance on the volume of gas produced in 1 min.

As lumps or powder are categorical variables you would express the results as a bar chart (see Figure 0.9(a)).

Example 2
If you are investigating the effect of concentration on the rate of a reaction then concentration can have any value and is therefore a continuous variable.

Your results would be presented as a line graph (see Figure 0.9(b)).

1 Whether a bar chart or line graph, make sure you use at least $\frac{3}{4}$ of the available grid provided.

2 When plotting two variables, plot the independent variable on the horizontal axis (*x*-axis) and the dependent variable on the vertical axis (*y*-axis).

3 After plotting the individual points on a line graph, you are often asked to drawn a best-fit line (see Figure 0.9(b)). If the line is obviously a curve then do not draw a 'point-to-point' line. The curve must be a smooth curve through the points.

4 Points lying well away from this best-fit line are 'anomalous' and not taken into consideration.

5 In many investigations, a zero value for your independent variable will obviously give a zero value for your dependent variable. One point you can be certain of is (0,0) and your best-fit line must go through the origin.

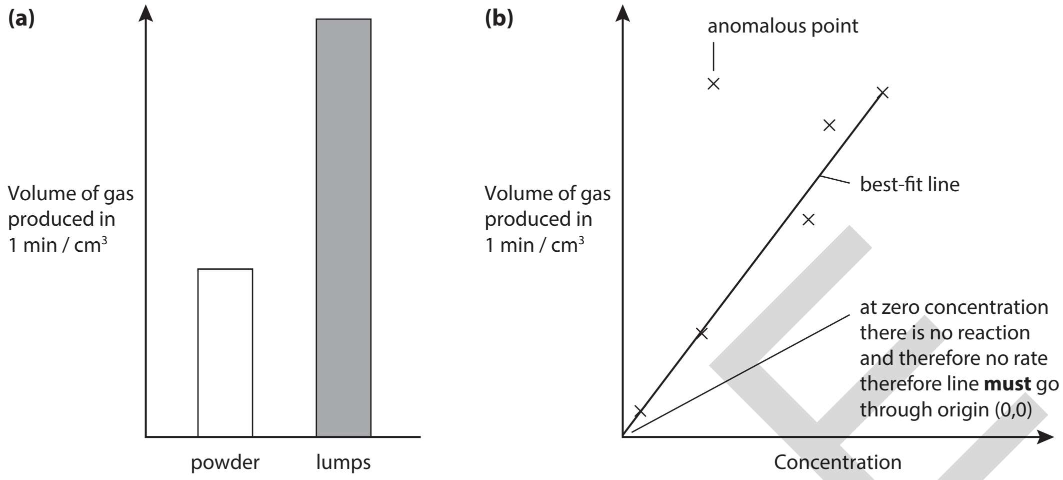

Figure 0.9

Calculating errors in your experiments

Introduction

There are several reasons why the final value that you record during an experiment may be inaccurate. Some of the errors associated with a value may be random (e.g. the substances used may be impure). More commonly though, errors are systematic and are associated with the apparatus that you've used.

For every experiment you complete you must assess and state the total percentage error associated with the values you report.

In some circumstances you can check what the actual value should be. If you know this value then you can calculate the experimental error using the following formula:

$$\text{Percentage error} = \frac{|\text{Actual} - \text{experimental}|}{\text{Actual}} \times 100\,\%$$

The | sign in this equation reflects the fact that the value being calculated is the absolute value and is always positive.

Example 1
When determining the relative atomic mass of magnesium, the accepted value is $24.3\ \text{g mol}^{-1}$. Your experimental determination gives the value of $26.6\ \text{g mol}^{-1}$.

The percentage error is therefore: $\frac{|24.3 - 26.6|}{24.3} = \frac{2.3}{24.3} \times 100\% = 9.47\,\%$

More commonly you need to calculate the total percentage error by adding up the percentage errors inherent in the apparatus you've used.

The overall **percentage error** will depend on the apparatus which has the greatest percentage error.

a Calculating systematic errors due to apparatus inaccuracy

i Temperature readings

A thermometer can have an uncertainty of ±1 °C.

In the same experiment, if you take two temperature readings and the first one gives you a reading of 21 °C and the second one gives you a reading of 42 °C then the temperature rise is 21 °C.

However, due to the inaccuracy inherent in the apparatus we know that the lower reading is 21±1 °C so could be as low as 20 °C or as high as 22 °C. Similarly, the second reading could be as high as 43 °C or as low as 41 °C.

The maximum difference is therefore 23 °C and the minimum difference is 19 °C.

The true reading is therefore 21±2 °.

The percentage error = $\left(\frac{\text{Maximum error}}{\text{Value of measurement}}\right) \times 100\% = \frac{2}{21} \times 100\% = 9.52\%$.

ii Measuring cylinders

Measuring cylinders are accurate to ±1 cm^3.

It you measure out 50.0 cm^3 of a solution in a 100 cm^3 cylinder

The percentage error = $\frac{1}{50} \times 100\% = 2\%$.

iii Burette readings

Burettes used in schools usually read to ±0.05 cm^3.

If you make two burette readings in a titration then each of them has an error of ±0.05 cm^{-3} and the total error is therefore 0.10 cm^3.

If you run in 22 cm^3 of solution,

the percentage error = $\left(\frac{\text{Maximum error}}{\text{Value of measurement}}\right) \times 100\% = \frac{0.10}{22} \times 100\% = 0.45\%$

iv Top-pan balance readings

The maximum error of a top-pan balance will depend on the quality of the balance.

The error of an electronic device is usually half of the last precision digit**.**

The accuracy of a two decimal place balance is ±0.005 g. Each reading has this error and if you make two readings then the maximum error is ±0.01 g.

Example 2

Mass of weighing boat + solid = 10.34 g	Maximum error = 0.005 g
Mass of weighing boat = 10.00 g	Maximum error = 0.005 g
Mass of solid = 0.34 g	Maximum error = 0.01 g

The percentage error = $\left(\frac{\text{Maximum error}}{\text{Value of measurement}}\right) \times 100\% = \frac{0.01}{0.34} \times 100\% = 2.94\%$

Using significant figures

Introduction

When you note the value of a result you need to be very aware of the amount of significant figures you use. The correct number of significant figures depends on the apparatus you use and the number of significant figures quoted for each measurement.

Example 1

In an investigation involving the use of a top-pan balance, burettes and thermometers:

The mass quoted in the results is 9.76 g

The measurement of solution from the burette was 25.00 cm^3

The temperature is quoted as 7.4 °C.

The measurement quoted to the lowest number of significant figures is the temperature. That is to two significant figures. This means that your final result should also be quoted to two significant figures.

Please note that the final result should be rounded down to two significant figures, but that this should be done right at the very end of your calculation. If you round down too early then you'll introduce a rounding error.

Chapter 1:
Masses, moles and atoms

Chapter outline

This chapter relates to Chapter 1: Moles and equations, Chapter 2: Atomic structure and Chapter 3: Electrons in atoms in the coursebook.

In this chapter you will complete investigations on:

- 1.1 Empirical formula of hydrated copper(II) sulfate crystals
- 1.2 Relative atomic mass of magnesium using molar volumes
- 1.3 Percentage composition of a mixture of sodium hydrogen carbonate and sodium chloride
- 1.4 Relative atomic mass of calcium by two different methods: molar volume and titration

Practical investigation 1.1:
Empirical formula of hydrated copper(II) sulfate crystals

Introduction

In this investigation you will determine the empirical formula (refer to Chapter 1 of the coursebook if required) of hydrated copper(II) sulfate by finding the value of **x** in $CuSO_4.\mathbf{x}H_2O$. You will weigh out some hydrated copper(II) sulfate in an evaporating basin, heat it to constant mass, determine the mass of water present in your sample and then find the molar ratio $CuSO_4 : H_2O$.

Equipment

You will need:

• a pipe-clay triangle • an evaporating basin • Bunsen burner and tripod • tongs • glass stirring rod • two heat-resistant pads • spatula

Access to:

• a supply of gas • a top-pan balance that reads to at least two decimal places

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times in this experiment and tie your hair back if it is long.
- Copper(II) sulfate is an irritant and is harmful if swallowed.
- Carry the evaporating basin to the top-pan balance on a heat-resistant pad. Do not use the tongs to carry it.

Method

1 Weigh an empty evaporating basin and then weigh the mass of crystals that you have been given.

Mass of basin + $CuSO_4.xH_2O$ crystals ……………g

Mass of basin ……………...g

Mass of $CuSO_4.xH_2O$ crystals ………………….…...g

2 Place the pipe-clay triangle and the evaporating basin containing your crystals on the tripod as shown in Figure. 1.1.

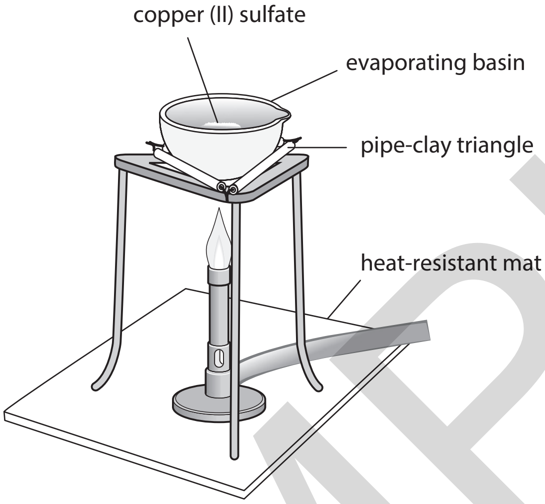

Figure 1.1

3 Copper(II) sulfate decomposes if heated too strongly. Heat the crystals **very gently**. A low, just-blue Bunsen flame should be used for this.

4 While you are heating the crystals, stir them using the glass stirring rod. At the same time grip the evaporating basin using the tongs to prevent the evaporating basin toppling over and spilling the contents.

5 At first, the copper(II) sulfate will be 'sticky' but after a short amount of time it should not cling to the glass rod and will become powdery.

6 The colour of the copper(II) sulfate will change from blue to light blue and then to a very light grey, almost white.

7 When it gets to this stage weigh the evaporating basin and anhydrous copper(II) sulfate.

Mass of basin + copper(II) sulfate = ……………..g

8 Reheat the powder for a short while and then reweigh it. If constant mass is obtained, then all the water of crystallisation will have been driven off from the crystals.

Mass of basin + copper(II) sulfate =……………...g

HINT
Note when cool, anhydrous $CuSO_4$ absorbs water from the air

9 If the mass has decreased, then keep on reheating and reweighing until a constant mass is obtained.

Repeat (1) Mass of basin + copper(II) sulfate =……………...g

Repeat (2) Mass of basin + copper(II) sulfate =……………...g

Repeat (3) Mass of basin + copper(II) sulfate =……………...g

Data analysis

a Calculate the mass of the anhydrous copper(II) sulfate remaining and then the mass of water that has been lost from the crystals on heating. This is the water of crystallisation.

Mass of anhydrous $CuSO_4$ =.……………...g

Mass of water of crystallisation =.……………...g

b Using the grid supplied, draw a set of axes with mass of anhydrous copper(II) sulfate on the horizontal (*x*) axis and mass of the water of crystallisation on the vertical (*y*) axis. Use suitable scales and label the axes.

HINT
Remember your line must go through the origin

- Plot the points on the graph.
- Reject any points that are obviously wrong.
- Draw a best-fit line through the points.

c Use your line to calculate the mass of water that combines with 1.60 g of anhydrous copper(II) sulfate ($CuSO_4$)

Mass of water =.…………………..g

d From your result, calculate the number of moles of water that combine with one mole of $CuSO_4$.

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e Calculate the value of x in the formula $CuSO_4.xH_2O$

x = ………………

Evaluation

f Which point should you be most confident about?

Explain your answer.

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HINT
If the point lies above the line, the ratio of water to anhydrous copper(II) sulfate is greater than it should be

g Explain any points:

i that lie **above** your best-fit line.

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ii that lie **below** your best-fit line.

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HINT
If the point lies below the line, the ratio of water to anhydrous copper(II) sulfate is less than it should be

h Copper(II) sulfate crystals lose their water of crystallisation between 100 °C and 350 °C. They start to decompose at approximately 600 °C.

Briefly describe a better way of heating the copper(II) sulfate crystals in this experiment and explain why it is an improvement on the method you used.

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Practical investigation 1.2:

Relative atomic mass of magnesium using molar volumes

Introduction

The objective of this investigation is to find the relative atomic mass of magnesium using its reaction with dilute hydrochloric acid to give hydrogen gas.

Refer back to Chapter 1 Moles and equations of the coursebook for more details of the theory.

The equation for the reaction between magnesium and hydrochloric acid is:

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

1 mol of any gas occupies 24 000 cm^3 (at room temperature and pressure)

This reaction can be used to find the relative atomic mass of magnesium. By determining the number of moles of hydrogen produced by a known mass of magnesium (*m*), the number of moles (*n*) of magnesium can be determined.

The relative atomic mass of magnesium can then be found using $A_r = \frac{m}{n}$.

As the masses of short lengths of magnesium ribbon are very small and difficult to measure on a top-pan balance, you will measure out a 10 cm length and weigh it. You will then estimate the masses of different shorter lengths and use these for your experiments.

Equipment

You will need:

- apparatus for collection and measurement of gas
- small piece of steel wool
- one 10.0 cm length of magnesium ribbon
- 30 cm ruler
- plastic gloves (see safety considerations)
- scissors

Access to:

- a top-pan balance reading to **at least** two decimal place
- 2 mol dm^{-3} hydrochloric acid

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times in this experiment.
- Magnesium is highly flammable.
- Hydrogen is a flammable gas.
- 2 mol dm^{-3} hydrochloric acid is an irritant.
- Steel wool sometimes splinters so use gloves if you have sensitive skin.
- If you are using a glass measuring cylinder for collecting the gas or a gas syringe then take care when clamping it because over-tightening of the clamp could shatter the glass.

Method

1. Take a 10.0 cm length of magnesium ribbon and **gently** clean it by using the steel wool.
2. Weigh the cleaned magnesium ribbon and record its mass.

   Mass of ribbon…………...g

3 Cut the ribbon up into 2 x 0.5 cm, 2 x 1.0 cm, 2 x 1.5 cm and 2 x 2.0 cm lengths.

4 From your mass for 10.0 cm of ribbon, estimate the mass of the 1.0 cm, 1.5 cm and 2.0 cm lengths.

Estimated mass of 1.0 cm lengths ………………….g

Estimated mass of 1.5 cm lengths………………….g

Estimated mass of 2.0 cm lengths ………………….g

5 Depending on which gas collecting system you are going to use, set up your apparatus as shown in Figure 1.2.

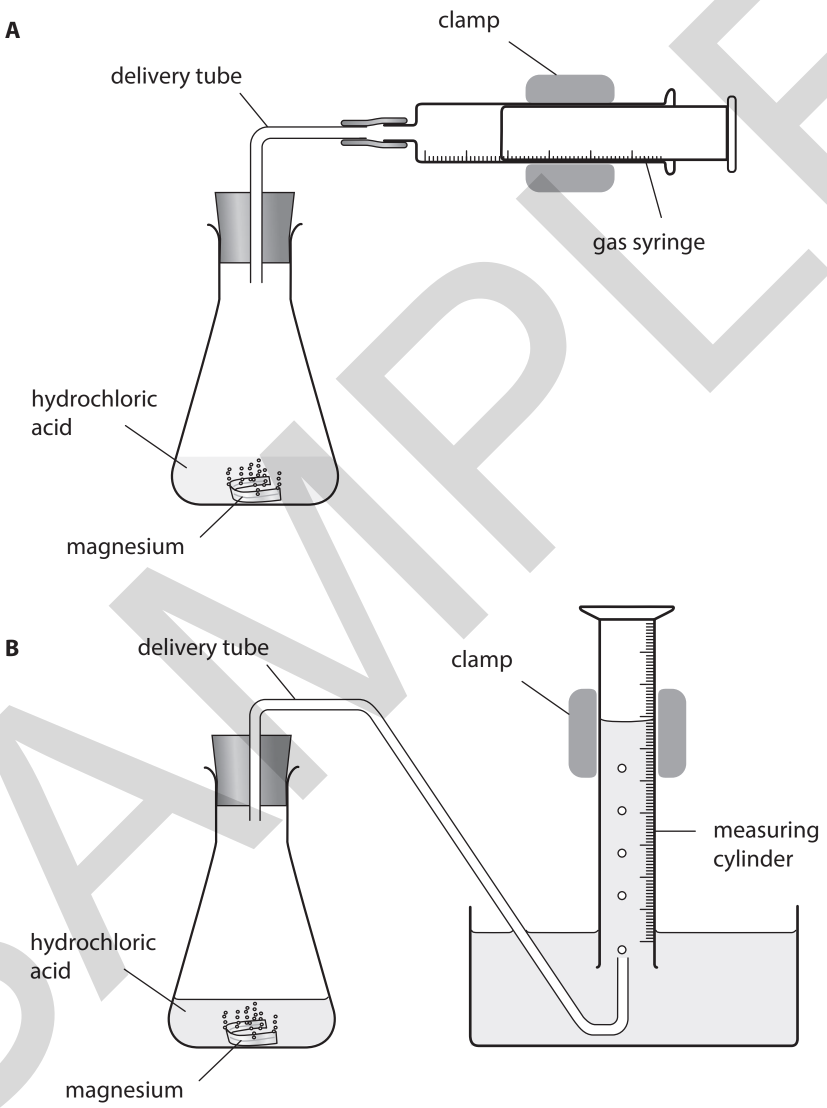

Figure 1.2

6 Measure out 25.0 cm^3 of hydrochloric acid into the conical flask.

a Set up the apparatus ready for measurement of a gas.

b Add one of the 1 cm lengths of magnesium ribbon to the acid, quickly replace the bung and start collecting the gas.

c Continuously swirl the flask because the magnesium will stick to the sides of the flask.

d When the reaction is finished, record the volume of gas produced.

Volume of gas given by a 1.0 cm length of ribbon = ………………………….. cm^3

7 Repeat Step 6 with all the other known lengths of magnesium ribbon.

Volume of gas (from 1.0 cm of ribbon)…………………………….cm^3
Volume of gas (from 1.5 cm of ribbon)…………………………….cm^3
Volume of gas (from 1.5 cm of ribbon)…………………………….cm^3
Volume of gas (from 2.0 cm of ribbon)…………………………….cm^3
Volume of gas (from 2.0 cm of ribbon)…………………………….cm^3

Results

Use Table 1.1 to record the mass of the ribbon used and the volumes of hydrogen produced.

Length of Mg ribbon/cm	**Mass of Mg/g**	**Volume of gas produced/cm^3**		
		Experiment 1	**Experiment 2**	**Average**
0.5 cm				
1.0 cm				
1.5 cm				
2.0 cm				

Table 1.1

Data analysis

a Plot a graph of mass of magnesium along the horizontal axis (*x*-axis) against volume of gas up the vertical axis (*y*-axis). You should use at least $\frac{3}{4}$ of the space available on the graph.

- Discard any results that are obviously wrong.
- Draw a best-fit line through your points.

b Using your graph, calculate the mass of magnesium that gives 24.0 cm^3 of hydrogen gas.

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c From this value, calculate the number of moles of magnesium that give this volume of gas and use $A_r = \frac{\text{mass of magnesium}}{\text{number of moles}}$ to find the relative atomic mass of magnesium. Assume that under the conditions of the experiment, 1 mol of gas occupies 24 dm^3 or 24 000 cm^3.

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d Compare your value for A_r with the value given in your Periodic Table.

Using the following formula, calculate the percentage error in your result.

Percentage error = $\frac{|\text{Actual value} - \text{experimental value}|}{\text{Actual value}} \times 100$

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HINT

Look back at the Skills chapter for how to calculate the percentage error from your readings

e What was the maximum error for the top-pan balance that you used?

The percentage error for your weighing = $\frac{\text{maximum error}}{\text{mass weighed out}} \times 100\%$

f The percentage error from your measurements of lengths of magnesium ribbon.

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HINT

The ruler measures to 1 mm and the maximum error is ± 0.5 mm or 0.05 cm. Therefore, a 2 cm length is really 2.0 ± 0.05 mm and the percentage error = $\frac{0.05}{2.0} \times 100\%$ = 2.5%

g Using this information, calculate the total error from your length measurements. Remember, you did one weighing but several volume and length measurements and these should be added up.

i Calculated error from length measurements:

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ii Possible errors from volume measurements:

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iii Total possible percentage error from apparatus readings:

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Evaluation

h What other factors could limit the accuracy of your results and contribute to the error?

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Practical investigation 1.3:

Percentage composition of a mixture of sodium hydrogen carbonate and sodium chloride

Introduction

In this practical, you will investigate the percentage composition of a mixture of sodium hydrogen carbonate and sodium chloride using an acid–base titration.

Equipment

You will need:

- a 150 cm^3 conical flask
- a 250 cm^3 volumetric flask
- wash bottle of distilled water
- burette stand
- a 25 cm^3 pipette
- white tile
- a 250 cm^3 beaker and 100 cm^3 beaker
- stirring rod
- small dropper
- small filter funnel for burette and larger one for volumetric flask
- a 50 cm^3 burette
- weighing boat

Access to:

- a top-pan balance reading to two or ideally three decimal places
- a mixture of sodium hydrogen carbonate and sodium chloride
- 0.100 mol dm^{-3} hydrochloric acid
- methyl orange indicator and dropper
- distilled water

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- Eye protection must be worn at all times in this investigation.
- The hydrochloric acid is an irritant.
- Methyl orange is poisonous. If you get it on your skin, wash it off immediately.

Method

Part 1: Making up the solution of the mixture

1 Weigh out 1.90–2.10 g of the mixture of sodium hydrogen carbonate and sodium chloride.

Weight of mixture ……………………..g

2 Dissolve this solid sample in distilled water and make up to a total volume of 250 cm^3 in your volumetric flask as described in the Skills chapter.

Part 2: The titrations

1 Titrate 25 cm^3 samples of this solution against the standard 0.100 mol dm^{-3} hydrochloric acid. Use methyl orange as the indicator.

2 You should look back at the Skills chapter for how to do this.

Results

Complete Table 1.2.

	Rough titration/cm^3	First accurate titration/cm^3	Second accurate titration/cm^3	Third accurate titration/cm^3
Final burette reading/cm^3				
Starting burette reading/cm^3				
Titre/cm^3				

Table 1.2

Data analysis

a Identify the concordant titres and give the average of these values.

Concordant titres = ………………….cm^3 and ………………….cm^3

Average of concordant titres = ………………….cm^3

Using the data you've collected, you can calculate the number of moles of the sodium hydrogen carbonate present in your sample. You can then calculate the mass of this compound and from that, the composition of the mixture.

The equation for the reaction between hydrochloric acid and sodium hydrogen carbonate is shown below.

$$NaHCO_3(aq) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

b Calculate:

- **i** The volume of 0.100 mol dm^{-3} hydrochloric acid needed to react completely with the sodium hydrogen carbonate present in 25 cm^3 of the mixture = ……………………cm^3
- **ii** The number of moles of hydrochloric acid reacting =……………… × ……………mol = number of moles of sodium hydrogen carbonate present in 25.00 cm^3 of solution = …………………………..mol
- **iii** Mass of sodium hydrogen carbonate present (Remember $m = n \times M_r$) = ……………………………g
- **iv** Total mass of mixture = ……………………………….g
- **v** Therefore, mass of sodium chloride present in mixture = ………………………….g
- **vi** Percentage of sodium hydrogen carbonate present in mixture = ……………………….%
- **vii** What is the actual percentage composition of the mixture? (Ask your teacher/supervisor.) Answer = ……………………….%

c You should also calculate the percentage error in your results as you did in Investigation 1.2.

$$\text{Percentage error} = \frac{|\text{Actual value} - \text{experimental value}|}{\text{Actual value}} \times 100$$

Evaluation

d Identify and calculate the systematic errors in your experiment from the following apparatus:

- **i** The top-pan balance

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- **ii** The pipette

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- **iii** The burette readings

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e Identify the random errors in your experiment.

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f What was the main contribution (if any) to your percentage error?

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g How could this be overcome?

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HINT

Remember in each titration you take two readings, each with a possible error of ±0.05 cm^3. So, for example, a titre of 20.00 cm^3 has a maximum possible error of ±0.10 cm^3

Practical investigation 1.4:

Relative atomic mass of calcium by two different methods: molar volume and titration

Introduction

The equation for the reaction between calcium and water is shown below:

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

This reaction can be used to find the relative atomic mass of calcium by finding the number of moles of hydrogen produced by a known mass of calcium. The number of moles of calcium (n) can then be calculated using the relative atomic mass calculated using $A_r = m/n$.

The calcium hydroxide formed in the reaction can then be titrated against standard hydrochloric acid.

Equipment

You will need:

• apparatus for measuring gas volumes as used in Investigation 1.2 • small filter funnel for burette • 50.00 cm^3 burette • weighing boat • 150 cm^3 conical flask • wash bottle of distilled water • burette stand • 25.00 cm^3 pipette • white tile • 250 cm^3 beaker • 25.0 cm^3 measuring cylinder • methyl orange indicator in dropper bottle

Access to:

• top-pan balance reading to **at least** two decimal places • 0.200 mol dm^{-3} hydrochloric acid • **fresh** calcium granules • distilled water

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times in this experiment.
- Calcium reacts vigorously with water. Do not handle it with bare hands.
- Hydrogen is a flammable gas.
- 0.2 mol dm^{-3} hydrochloric acid is an irritant.
- If you are using a glass measuring cylinder for collecting the gas or a gas syringe, then take care when clamping it. Over-tightening of the clamp could shatter the glass.
- The calcium hydroxide is an alkali and should be regarded as being corrosive. If you get any on your skin then wash it off immediately.

Part 1: Determination by molar volume

Method

1. Set up your apparatus for reacting the calcium with water and collecting the gas formed during the reaction. Use either of the two arrangements shown in Figure 1.2.
2. Measure out 25 cm^3 of distilled water and add to the conical flask.
3. Weigh out between 0.040 g and 0.080 g of calcium.
4. Make sure that your gas collection apparatus is ready.

5 Add the calcium granules to the conical flask and quickly replace the stopper. Swirl the flask vigorously to make sure that all the calcium has reacted.

6 When the reaction is finished note the volume of gas evolved and record it in Table 1.3.

Results

Mass of calcium/g	**Volume of hydrogen/cm^3**	**Burette reading for hydrochloric acid/cm^3**	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	

Table 1.3

Data analysis

a Assume that 1 mol of gas occupies 24 000 cm^3 at room temperature and pressure.

- **i** Calculate the number of moles of hydrogen formed in your first experiment.
- **ii** From this, calculate the number of moles of calcium.
- **iii** Calculate the relative atomic mass of calcium.

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Evaluation

b Using the value shown on your Periodic Table, calculate the percentage error in your result.

$$\text{Percentage error} = \frac{|\text{Actual value} - \text{experimental value}|}{\text{Actual value}} \times 100$$

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c Systematic errors: calculate the percentage errors in your apparatus.

- **i** The weighing out of the calcium.

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ii The measurement of gas volume.

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iii Random errors: identify the random errors in the method.

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iv Are there any improvements you would make to this method?

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Part 2: Determination by titration

Method

1 Remove the flask from the gas collection apparatus and wash any liquid and white solid on the sides into the solution.

2 **a** Fill up your burette to near the zero mark with 0.200 mol dm^{-3} hydrochloric acid.

b Place a white tile under the burette.

c Add a few drops of methyl orange indicator to the calcium hydroxide in the conical flask. **There are no opportunities for a rough titration**.

3 **a** Add the acid to the flask and after each addition swirl the flask vigorously.

b When the indicator shows signs of colour change to orange red, add the acid more slowly – a drop at a time until an orange colour is obtained.

c Note the new burette reading.

4 **a** Wash the flask thoroughly with tap water and then rinse with distilled water.

b Repeat steps in the two experiments with a new mass of calcium.

HINT
Look back at Skills chapter for full details on carrying out titrations

Results

Complete Table 1.4

Mass of calcium/g	Volume of hydrogen/cm³	Burette reading for hydrochloric acid/cm³	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	
		2nd	
		1st	
		Titre	

Table 1.4

Data analysis

d Calculate the number of moles of hydrochloric acid reacting with the calcium hydroxide.

- **i** From this value, calculate the number of moles of calcium hydroxide and therefore the number of moles of calcium.
- **ii** Calculate the relative atomic mass of calcium.

  Repeat these calculations if you have more than one set of results.

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Evaluation

e Using the value shown on your Periodic Table, calculate the percentage error in your results for:

- **i** The weighing out of the calcium.

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  ……………………………………………………………………………………

- **ii** The titrations.

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- **iii** Systematic errors: calculate the total percentage errors in your measurements.

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  ……………………………………………………………………………………

- **iv** Random errors: identify the random errors in the method.

  ……………………………………………………………………………………

  ……………………………………………………………………………………

f Are there any improvements you would make to this method?

……………………………………………………………………………………

……………………………………………………………………………………

Chapter 2:
Structure and bonding

Chapter outline

This chapter relates to Chapter 4: Chemical bonding and Chapter 5: States of matter in the coursebook.

In this chapter, you will complete investigations on:

- 2.1 Physical properties of three different types of chemical structure
- 2.2 Effect of temperature on the volume of a fixed mass of gas
- 2.3 Effect of pressure on the volume of a fixed mass of gas

Practical investigation 2.1:
Physical properties of three different types of chemical structure

Introduction

In this investigation, you will carry out some simple tests on substances that are examples of different types of chemical structure. You will then use your knowledge of the different chemical structural types to explain your observations.

Equipment

You will need:

• Bunsen burner, tripod, gauze and heatproof mat • 12 dry test tubes and a test-tube rack • eight stoppers to fit test tubes • two graphite rods in a holder • three spatulas • three leads and two crocodile clips • 12 V bulb • power pack • wash bottle filled with distilled water • small evaporating basin • tongs

Access to:

• cyclohexane • wax • white sand • potassium iodide

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times and tie back long hair.
- Cyclohexane is flammable and when using it you must turn off your Bunsen burner.
- Cyclohexane should be disposed of by pouring the mixture into a large glass bottle in the fume cupboard.
- If the test tube is very hot after heating, place it on the heatproof mat to allow it to cool.

Method

The three materials you are going to test are wax, silicon dioxide (sand) and potassium iodide.

Follow the methods explained in Table 2.1 and complete your observations as you proceed.

Test	Observations		
	Wax	**Silicon dioxide**	**potassium iodide**
1 Place a small sample of each material in a dry test tube and slowly increase the heat until it is very strong. Heat until there is no further change.			
2 Place a small amount of the substance in a dry test tube. Add some cyclohexane to the solid. Stopper the tube and shake it.			
3 Place a small amount of the substance in a dry test tube. Add some water to the solid. Stopper the tube and shake it.			
4 Place a small amount of the substance in an evaporating basin and test its electrical conductivity as a solid and then after the addition of the liquid in which it dissolved.			

Table 2.1

Summarise your findings in Table 2.2. Decide which of these three substances has a structure which is either giant covalent, simple molecular or giant ionic.

Substance	Type of chemical structure	Summary of observations

Table 2.2

Data analysis

a Explain your observations for each of the **three** substances

i Wax

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……………………………………………………………………………………………………

……………………………………………………………………………………………………

ii potassium iodide

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………………………………………………………………………………………………………

………………………………………………………………………………………………………

iii Silicon dioxide

………………………………………………………………………………………………………

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Practical investigation 2.2:

Effect of temperature on the volume of a fixed mass of gas

Introduction

In this investigation, you will determine how the volume of a fixed mass of gas varies with temperature at constant pressure.

Equipment

You will need:

• Bunsen burner, tripod, gauze and heatproof mat • 100 cm^3 round-bottomed flask • stopper for flask attached to short length of plastic or rubber tubing • 100 cm^3 measuring cylinder • permanent marker pen • dropper • 100 cm^3 gas syringe • metal container for heating water • thermometer reading to 110 °C • either a stirring rod or a small 'paddle' for stirring water in metal container • water supply

Safety considerations

- Long hair must be tied back securely.
- Eye protection must be worn at all times in this investigation.
- When you clamp the gas syringe do not over tighten the clamp as this could stop the piston from moving easily or even worse, break the glass.
- Do not use the thermometer for stirring. The thermometer bulb is only thin glass and is easily broken.
- When you stir the water in the metal container hold the container so that your stirring does not move it.
- Take special care when you are carrying out measurements at higher temperatures.

Method

1 In this first step you will determine the **true starting volume** of the gas. This needs to be done before the actual practical because the flask needs to be dry when doing the volume determinations. If this measurement is done the day before, the flask should be placed in an oven to dry it.

- Insert the stopper into the neck of the round-bottomed flask and mark the level of the bottom of the stopper.
- Pour water into the flask up to the mark.
- Measure the volume of the water using the measuring cylinder.
- Your teacher will give you the volume of the tubing.
- What is the total volume of the flask and tubing?

  Total volume of the flask and tubing = ……………… cm^3

2 Set up apparatus as shown in Figure 2.1.

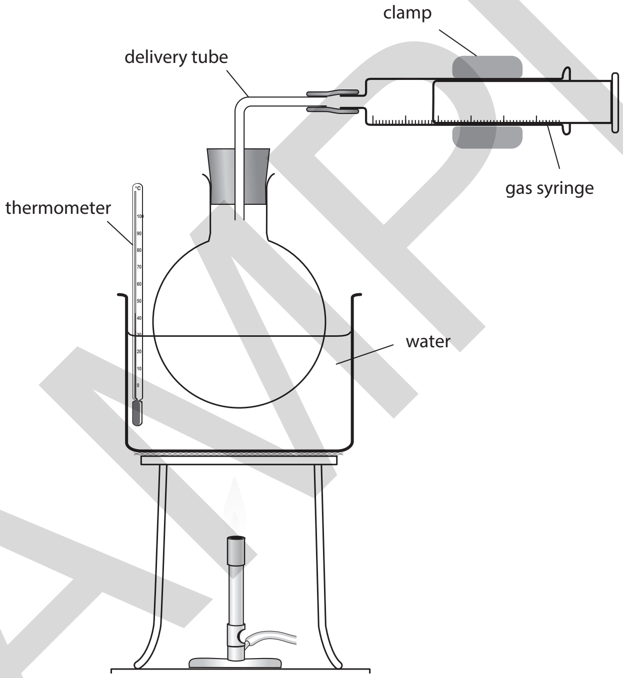

Figure 2.1

3 Measure the volume of the gas at room temperature.

4 Gently heat the water for a few seconds and at the same time stir the water thoroughly.

5 Remove the heat and measure the temperature.

6 Measure the volume of the gas in the syringe and add this to the volume of the flask and the tubing to give the total volume. If the temperature has risen too much then add some cold water to the container and stir again.

7 Repeat steps 5 and 6 until you have made measurements of the volume at several temperatures between room temperature and 90 °C.

8 Record your results in Table 2.3.

Temp/°C											
Reading on syringe/ml											
Total volume of gas/cm^3											

Table 2.3

HINT

Your *x*-axis should start at −300 °C and end at 100 °C.

Data analysis

a Plot the temperature (horizontal axis) against total volume (vertical axis) on the graph paper provided.

- Plot a best-fit line through your points. You should omit any obviously anomalous points.
- Produce your line back to the value where the volume is **zero** and find the temperature at this point.

Temperature where volume is zero = …… …… °C

Evaluation

b Calculate the error in your answer for the absolute temperature. Look back at the Skills chapter for the formula, if required.

……………………………………………………………………………………………………

……………………………………………………………………………………………………

c Apart from the errors due to equipment, what were the other sources of error in your experiment?

……

……

d What is the name given to the temperature where the volume of the gas is zero?

……

e Using your results, write a law which can be applied to all gases and which defines the relationship between the volume and the temperature of an ideal gas.

……

……

Practical investigation 2.3:
Effect of pressure on the volume of a fixed mass of gas

Introduction

In this investigation, you will measure the pressure of a gas as its volume decreases and try to deduce the relationship between pressure and volume at constant temperature.

You will be using your pressure data logger to take readings at precise values of gas volume – this is described by some data logging systems as 'single-step mode'. If your software allows it you can export your results to Microsoft Excel for analysis.

Equipment

This experiment may be done as a demonstration by the teacher.

You will need:

- a 60 cm^3 plastic syringe attached to a short length of plastic tubing that will fit the pressure data logger
- a laptop or tablet that will interface with the data logger and run the software required
- a pressure data logger with any software required

HINT

At the lower volumes, the pressure will be relatively high, and someone will need to hold the syringe piston firmly while another records the pressure. Do not press down too forcefully on the piston!

Safety considerations

- Eye protection must be worn at all times in this practical.
- Make sure your laptop is away from any water.
- Be careful that the tube connecting the pressure monitor to the syringe does not come off.

Method

1 Connect your syringe and pressure data logger. Each data logging system will have its own procedure for recording the separate values.

2 Starting at 60 cm^3, measure the pressure of the gas at that volume.

3 Decrease the volume by pressing down the piston of the syringe and hold the piston at that point whilst the data logger records the pressure of the gas.

4 Measure the pressure at 5 cm^3 intervals.

5 Record your results in Table 2.4.

Volume of gas/cm^3	60	55	50	45	40	35	30	27
$\frac{\mathbf{1}}{\textit{volume}}$								
Pressure of gas/kPa								

Table 2.4

Data analysis

a From your results deduce the relationship between the **volume (*V*)** of a fixed mass of gas and its **pressure (*P*)**.

……………………………………………………………………………………………………

……………………………………………………………………………………………………

b Use a calculator or data processing package (e.g. Microsoft Excel) to calculate the values of $\frac{1}{V}$

c On the graph paper provided draw a graph of $\frac{1}{V}$ (horizontal axis) against P (vertical axis)

d Draw a best-fit line though the points.

e Explain why your graph shows that PV = constant.

……

……

……

f Take your first set of results.

Value of P = ……..kPa; Value of V = ………..m^3; (Note 1 kPa = $1 \times 10_3$ Pa)

The universal gas equation states that $PV = nRT$;

P is in Pa; V is in m^3 and 1 cm^3 = 1×10^{-6} m^3; n = number of moles of gas; T = absolute temperature (K). R is the Universal Gas Constant and its units are J mol^{-1} K^{-1}

If the initial volume = 60 cm^3 then we can use the fact that $n = \frac{60}{24\,000}$ mol because this relationship is true at room temperature and pressure.

g Calculate the value of R by substituting your values into the ideal gas equation.

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h Research the true value of R and calculate the percentage error in your result.

……

……

……

i Using the equation for R, explain why the units of R = J mol^{-1} K^{-1}

……

……

……

Chapter 3: Enthalpy changes

Chapter outline

This chapter relates to Chapter 6: Enthalpy changes in the coursebook.

In this chapter you will complete investigations on:

- 3.1 Enthalpy change for the reaction between zinc and aqueous copper(II) sulfate solution
- 3.2 Enthalpy change of combustion of alcohols
- 3.3 Enthalpy change of thermal decomposition
- 3.4 Change in enthalpy of hydration of copper (II) sulfate

Practical investigation 3.1:

Enthalpy change for the reaction between zinc and aqueous copper(II) sulfate

Introduction

In this practical, you will find the enthalpy change of reaction between zinc and copper(II) sulfate solution.

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

You will carry out the practical at least twice. The first experiment will have the zinc as the limiting reactant and the second will have the copper(II) sulfate as the limiting reactant. You can complete both of the practical investigations before answering the questions.

For both experiments you will construct a temperature–time graph. The reason for this is explained in the Skills chapter.

Equipment

You will need:

• two small polystyrene beakers • glass beaker large enough to hold the polystyrene beakers • −10 to 110 °C thermometer • 25 cm^3 measuring cylinder • plastic covers for polystyrene beakers • a small spatula • two weighing boats

Access to:

• 1 mol dm^{-3} copper(II) sulfate solution • zinc powder • a top-pan balance that reads to at least two decimal places

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times.

- The zinc powder is flammable.
- Copper(II) sulfate solution is harmful to you and to the environment.

Part 1: Enthalpy change for the reaction with zinc as the limiting reactant.

Method

HINT
Enthalpy is the heat change at constant pressure

1 Measure out:

- 25.0 cm^3 of copper(II) sulfate solution into one of your polystyrene beakers. Place the polystyrene beaker into the glass beaker to support it.
- Between 0.64 g to 0.66 g of zinc powder.
  Mass of zinc powder = ……………….g

2 Measure the temperature of the copper(II) sulfate solution for the next three minutes. Swirl the solution regularly to make sure its temperature is uniform. Record your results in Table 3.1

3 On the third minute **do not measure the temperature**. Add the zinc powder to the copper(II) sulfate solution and for the next minute swirl the beaker so that the reactants are well mixed.

4 On the fourth minute resume the measurement of the temperature and continue measuring the temperature until the tenth minute. Record your results in Table 3.1

5 Between temperature measurements it is very important to swirl the beaker so that you have good mixing of the reactants and the solution.

Results

Time/min	**0**	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Temperature/°C				X							

Table 3.1

Data analysis

a What is the maximum temperature change (ΔT) in the first experiment?

………

The heat change (q) is calculated using the formula $q = m \times c \times \Delta T$ where m = mass of solution; c = specific heat capacity of solution and ΔT is the temperature change in the reaction

b Calculate the enthalpy change for the reaction.

Assume that the density of the copper(II) sulfate solution is exactly the same as pure water (1.00 $g\,cm^{-3}$). Therefore, m = …………………… g

The specific heat capacity of the solution is assumed to be the same as that of pure water. $c = 4.18\ J g^{-1} K^{-1}$

………

………

………

c Calculate the number of moles of $CuSO_4$ present in 25.0 cm^3 of 1.00 mol dm^{-3} solution.

……………………………………………………………………………………………………

……………………………………………………………………………………………………

d Calculate the number of moles of zinc present in your sample of zinc (A_r Zn = 65.4)

……………………………………………………………………………………………………

……………………………………………………………………………………………………

e Using the equation for the reaction and your answers to questions c and d above, explain why zinc is the limiting reactant in this experiment.

……………………………………………………………………………………………………

……………………………………………………………………………………………………

f Calculate the standard enthalpy change in kJ mol^{-1}

……………………………………………………………………………………………………

……………………………………………………………………………………………………

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Part 2: Enthalpy change for the reaction with copper(II) sulfate as the limiting reactant.

Method

1 Measure out:

- 25.0 cm^3 of copper(II) sulfate solution into one of your polystyrene beakers
- Between 6.40 g to 6.60 g of zinc powder

  Mass of zinc powder = ……………g

2 Repeat Method Part 1, Steps 2–5 for the second mass of zinc and record your results in Table 3.2**.**

Results

Time/min	**0**	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Temperature/°C				X							

Table 3.2

Data analysis

a Calculate the number of moles of copper(II) sulfate in 25.0 cm^3 of 1.00 mol dm^{-3} solution. Number of moles of $CuSO_4$ = ………………… g

b Calculate the number of moles of zinc you weighed out. (A_r Zn) = 65.00 g.

c On the graph paper provided, plot the results of Part 1 and Part 2 of the investigation as follows:

- Plot a graph of temperature (vertical axis) against time (horizontal axis).
- Use the largest scale possible.
- For each experiment you should get two lines that look something like those shown in the Skills chapter.

d What is the maximum temperature change (ΔT) in the **second** experiment?

………………………………………………………………………………………………………

HINT
m = mass of solution; c = specific heat capacity of solution

e Calculate the enthalpy change(q)for the reaction. Make the same assumptions you did for Part 1.

………………………………………………………………………………………………………

………………………………………………………………………………………………………

………………………………………………………………………………………………………

f Calculate the number of moles of $CuSO_4$ present in 25.0 cm^3 of 1.00 mol dm^{-3} solution.

………………………………………………………………………………………………………

………………………………………………………………………………………………………

g Calculate the number of moles of zinc present in your sample of zinc (A_r Zn = 65)

……………………………………………………………………………………………………

……………………………………………………………………………………………………

h Using the equation for the reaction and your answers to questions c and d above, explain why copper(II) sulfate is the limiting reactant in this experiment.

……………………………………………………………………………………………………

……………………………………………………………………………………………………

i Calculate the standard enthalpy change for the reaction in kJ mol^{-1}

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

j Explain why your two values should be either identical or very close in value.

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……………………………………………………………………………………………………

Evaluation

k The accepted value for the enthalpy change of reaction is -219 kJ mol^{-1}. Calculate your percentage error using the average of your two results.

……………………………………………………………………………………………………

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l Calculate the maximum percentage errors arising from your **mass**, **temperature** and **volume** measurements.

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m Identify the sources of the **non-**systematic errors in your measurement.

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Practical investigation 3.2:
Enthalpy change of combustion of alcohols

Introduction

Enthalpy change means the heat change at constant pressure.

In this practical you will investigate the enthalpy of combustion of the straight chain alcohols methanol, ethanol, propan-1-ol and butan-1-ol.

You will burn the alcohols using spirit burners. To make it a fair test you must make sure the enthalpy change measured is the same each time. Therefore, you will raise the temperature of a measured volume of water by the same temperature for each alcohol.

In this method, we are going to use a copper calorimeter as a container for the water being heated. When we heat up the water we are also heating up the calorimeter as well as the water.

The formula used for calculating heat change is $q = m \times c \times \Delta T$.

HINT
m = mass;
c = specific heat capacity;
ΔT = change in temperature

In this experiment:

Enthalpy change = $(m_{\text{water}} \times c_{\text{water}} \times \Delta T) + (m_{\text{calorimeter}} \times c_{\text{calorimeter}} \times \Delta T)$ J

The specific heat capacity for water = 4.18 J g^{-1} K^{-1}

The specific heat capacity for copper = 0.385 J g^{-1} K^{-1}

If you are using glass as a container then c = 0.840 J g^{-1} K^{-1}

Equipment

You will need:

• spirit burners containing the four alcohols • copper wire stirrer • clamp stand, boss and clamp • at least two heat-resistant pads • thermometer • 100 cm^3 measuring cylinder • cap/cover for spirit burner • wooden splint

Access to:

• a top-pan balance reading to at least two decimal places • a supply of water • a Bunsen burner

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- You must wear eye protection at all times.
- All the alcohols are flammable.
- All the alcohols should be treated as harmful.

Method

Part 1: Preliminary planning

1 Weigh the copper calorimeter and stirrer.

2 Take a spirit burner containing methanol and position it where it will be placed in the actual experiment.

- **a** Using a lighted splint, light the wick to get an idea of the height of the flame. The flame should be no more than 2 cm high.
- **b** Clamp the calorimeter so that the flame just touches the base of the calorimeter (see Figure 3.1).

3 For meaningful measurements, the flame has to be the same for all four experiments and so has the distance between the flame and the base of the calorimeter. Extinguish the flame.

4 Add 100 cm^3 of water to the calorimeter and see if the bulb of the thermometer is covered. If it isn't then you will need to increase the volume of water.

5 The experimental set-up should resemble that shown in Figure 3.1.

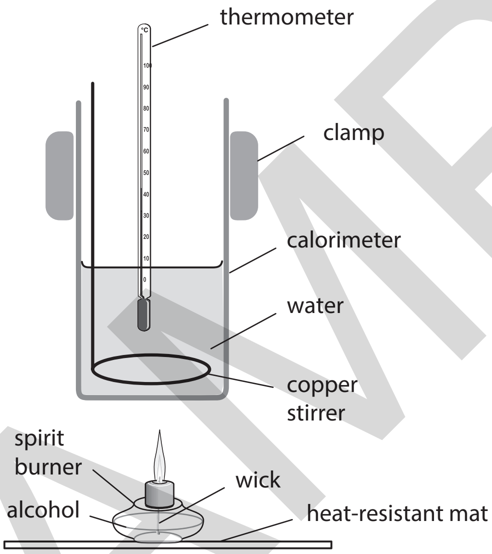

Figure 3.1

Part 2: Procedure

1 Measure out 100 cm^3 of water into the calorimeter.

2 Make sure that the water covers the bulb of the thermometer. If it does not then the volume of water will have to be increased.

3 Clamp the calorimeter in the position and height that you decided upon in the preliminary planning.

4 Weigh the spirit burner plus the cap/cover if one is available.

Mass of burner + methanol = ……………………..g

5 Stir the water thoroughly and measure its temperature.

Initial temperature of water = ……………………...°C

6 **a** Remove the cap from the burner and place it underneath the calorimeter.

b Light the wick using a lighted splint.

7 Stir the water thoroughly until the temperature has risen by exactly 20 °C.

Final temperature of water ……………………………..°C

8 **a** Blow out the flame and cover with the cap/cover if one is provided.

b Carry the burner over to the top-pan balance using a heat-resistant pad and weigh it.

Mass of burner + methanol = ………………………...g

Mass of methanol burned = ………………………...g

9 Repeat Steps 1–8 using the other three alcohols and record your results in Table 3.3

Methanol	Mass of burner + methanol before burning	g
	Mass of burner + methanol after burning	g
	Mass of **methanol** burned	g
Ethanol	Mass of burner + ethanol before burning	g
	Mass of burner + ethanol after burning	g
	Mass of **ethanol** burned	g
Propan-1-ol	Mass of burner + propan-1-ol before burning	g
	Mass of burner + propan-1-ol after burning	g
	Mass of **propan**-1-ol burned	g
Butan-1-ol	Mass of burner + butan-1-ol before burning	g
	Mass of burner + butan-1-ol after burning	g
	Mass of **butan-1-ol** burned	g

Table 3.3

Data analysis

The enthalpy change is the same for all four alcohols because you heated up the same mass of water and the same apparatus by the same temperature. Remember this value is in J and standard enthalpy changes are usually expressed in kJ.

If the mass of methanol burned is *m* g then the number of moles of methanol (*n*) burned in the experiment

$$= \frac{m}{Mr} = \frac{m}{32}$$

If the enthalpy change is q then the standard enthalpy change of combustion (ΔH_c) can be calculated as follows: $\Delta H_c = \frac{q}{n} \div 1000\ \text{kJ mol}^{-1}$

a Calculate the standard enthalpy changes of combustion for all four alcohols and record your results in Table 3.4.

Name of alcohol	Relative molecular mass	No. of moles burned/mol	Experimental value for ΔH_c /kJ mol^{-1}	Literature value for ΔH_c/ kJ mol^{-1}
Methanol(CH_3OH)				−726
Ethanol (C_2H_5OH)				−1367
Propan-1-ol (C_3H_7OH)				−2021
Butan-1-ol (C_4H_9OH)				−2676

Table 3.4

b Calculate the percentage error in your results for each alcohol.

i Methanol

....

....

ii Ethanol

....

....

iii Propan-1-ol

....

....

....

iv Butan-1-ol

....

....

....

c What was the maximum percentage error from your apparatus?

....

....

i The measurement of temperature (note for each temperature change there were two readings taken)

....

....

ii The measurement of the water volume

……………………………………………………………………………………………………

……………………………………………………………………………………………………

d The measurement of mass of alcohol burned for each alcohol:

i Methanol

……………………………………………………………………………………………………

……………………………………………………………………………………………………

ii Ethanol

……………………………………………………………………………………………………

……………………………………………………………………………………………………

iii Propan-1-ol

……………………………………………………………………………………………………

……………………………………………………………………………………………………

iv Butan-1-ol

……………………………………………………………………………………………………

……………………………………………………………………………………………………

e Choose **one** alcohol and calculate the maximum percentage error due to the apparatus used.

……………………………………………………………………………………………………

……………………………………………………………………………………………………

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f What other sources of error could lead to inaccuracies in your results?

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Practical investigation 3.3:

Enthalpy change of thermal decomposition

Introduction

When potassium hydrogen carbonate is heated it decomposes into potassium carbonate, carbon dioxide and water.

$$2KHCO_3(s) \rightarrow K_2CO_3(s) + CO_2(g) + H_2O(l)$$

Because this is a thermal decomposition and is an endothermic reaction, it is impossible to find the enthalpy change directly. To overcome this problem we use Hess' Law to find the enthalpy change indirectly.

Both potassium hydrogen carbonate and potassium carbonate react with hydrochloric acid and the enthalpy changes are measurable.

$$KHCO_3(s) + HCl(aq) \rightarrow KCl(aq) + H_2O(l) + CO_2(g) \qquad \text{Enthalpy change} = \Delta H_1$$

$$K_2CO_3(s) + 2HCl(aq) \rightarrow 2KCl(aq) + H_2O(l) + CO_2(g) \qquad \text{Enthalpy change} = \Delta H_2$$

We can construct a Hess' cycle for this reaction (see Figure 3.2)

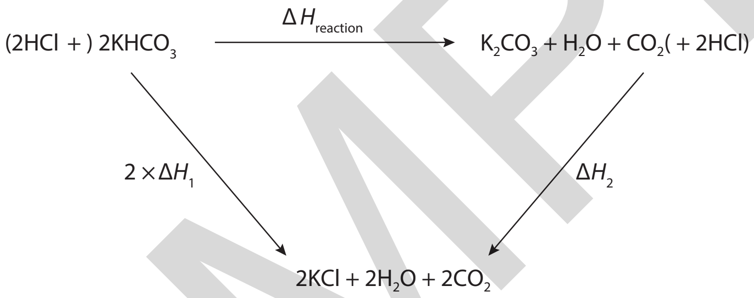

Figure 3.2

Using Hess' Law $\Delta H_r + \Delta H_2 = 2\Delta H_1$

Therefore, $\Delta H_r = 2\Delta H_1 - \Delta H_2$ and by determining the values of ΔH_1 and ΔH_2 we will be able to calculate the enthalpy change for the reaction.

Equipment

You will need:

- polystyrene beaker and cap with hole for thermometer
- glass beaker to hold the polystyrene beaker
- thermometer: one reading from −10 to 50 °C with 0.2 °C divisions is preferable
- weighing boat
- 50 cm^3 measuring cylinder
- cotton wool to act as extra insulation

Access to:

- potassium carbonate solid
- potassium hydrogen carbonate solid
- 2 $mol\,dm^{-3}$ hydrochloric acid
- a top-pan balance reading to at least two decimal places

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- Wear eye protection at all times.
- The hydrochloric acid is an irritant at this concentration.

Method

Part 1: Determining the enthalpy change for Reaction 1.

1 Weigh out 0.025 mol of potassium hydrogen carbonate.

[*Ar* values: K = 39.1; H = 1; C = 12; O = 16]

2 Formula mass of potassium hydrogen carbonate = ………………………g mol^{-1}

3 Mass of 0.025 mol = ……………………g

4 Mass of potassium hydrogen carbonate weighed out = ……………………g

5 **a** Place the polystyrene beaker inside the glass beaker and fit the cotton wool round it to improve the insulation (as shown in Figure 3.3).

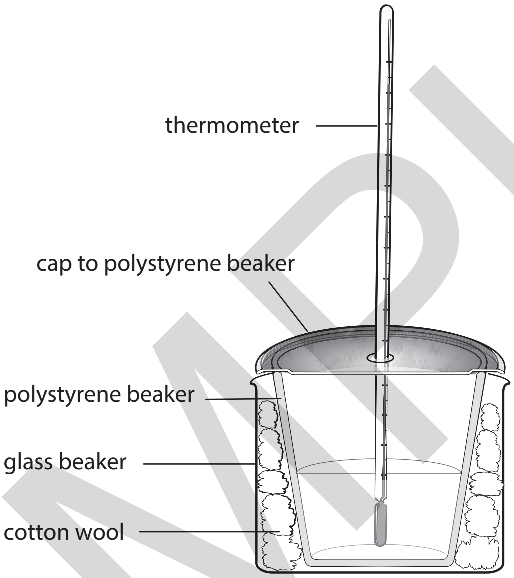

Figure 3.3

b Measure 50.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid and pour it into the polystyrene beaker.

6 Measure the initial temperature of the acid.

Initial temperature of acid = ……………………°C

7 Add the potassium hydrogen carbonate to the acid. There will be a rapid effervescence so make sure you have the cap ready to prevent any spillages.

8 Swirl the beaker and contents to ensure that there is thorough mixing of the contents.

9 When the reaction is complete, record the minimum temperature.

Final temperature of the Reaction 1 mixture = ………………………°C

Part 2: Determining the enthalpy change for Reaction 2.

1 Weigh out 0.025 mol of potassium carbonate.

[*Ar* values: K = 39.1; C = 12; O = 16]

2 Formula mass of potassium carbonate = ………………………g mol^{-1}

3 Mass of 0.025 mol = …………………………g

4 Mass of potassium carbonate weighed out = ……………………g

5 Place the polystyrene beaker inside the glass beaker and fit the cotton wool round it to improve the insulation

6 Measure 50 cm^3 of 2 mol dm^{-3} hydrochloric acid and pour it into the polystyrene beaker.

7 Measure the initial temperature of the acid.

Initial temperature of acid = ……………………..°C

8 **a** Add the potassium carbonate to the acid. Once again there will be a rapid effervescence so make sure you have the cap ready to prevent any spillages.

b Swirl the beaker and contents to ensure that there is thorough mixing of the contents.

9 When the reaction is complete, record the maximum temperature attained.

Final temperature of the Reaction 2 mixture = …………………..°C

Data analysis

a Complete Table 3.5

Temperature	Reaction 1	Reaction 2
Final temperature/ °C		
Initial temperature/ °C		
Temperature change (ΔT)/ °C		

Table 3.5

The enthalpy change for each reaction (q) = m x specific heat capacity x ΔT

Assume the density of the hydrochloric acid is 1.00 g dm^{-3} and its specific heat capacity is 4.18 J g^{-1} K^{-1}.

b Calculate the value of q for Reaction 1

………

………

………

c Calculate the enthalpy change of reaction for Reaction 1

………

………

………

d Calculate the value of q for Reaction 2

………

………

………

e Calculate the enthalpy change of reaction for Reaction 2

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

f Use these two results to find the enthalpy change for the thermal decomposition of potassium hydrogen carbonate.

Evaluation

The standard enthalpies of formation (in kJ mol^{-1}), relevant to this reaction is follows:

$\Delta H_f^{\ominus}(KHCO_3) = -959.4$; $\Delta H_f^{\ominus}(K_2CO_3) = -1146$; $\Delta H_f^{\ominus}(CO_2) = -393.5$; $\Delta H_f^{\ominus}(H_2O) = -285.9$

g Using these values calculate the standard enthalpy change for the reaction.

……………………………………………………………………………………………………

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h Calculate the percentage error in your results.

……………………………………………………………………………………………………

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……………………………………………………………………………………………………

i Calculate the maximum percentage error due to the apparatus used.

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Practical investigation 3.4:

Change in enthalpy of hydration of copper (II) sulfate

Introduction

This practical brings together techniques and theory used in previous investigations. The reaction being studied is as follows:

$$CuSO_4\ (s) + 5H_2O(l) \rightarrow CuSO_4.\ 5H_2O(s)$$

It is impossible to determine the enthalpy change of this reaction directly. Therefore, we have to use Hess's law. Hess's cycle is shown in Figure 3.4.

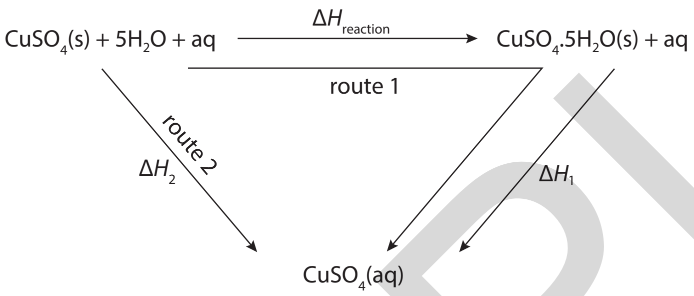

Figure 3.4

Using the cycle, you can see that the arrows meet at $CuSO_4(aq)$ (i.e. an aqueous solution of copper(II) sulfate). However, to get to this solution, the enthalpy change must be the same.

This means that: $\Delta H_{reaction} + \Delta H_1 = \Delta H_2$

In one of the reactions, the temperature change is reasonably high and for this reason you are going to collect results for a temperature–time graph.

Equipment

You will need:

- two polystyrene beakers plus caps
- thermometer which reads from -10 °C to 50 °C in 0.2 °C divisions
- spatula
- wash bottle containing distilled water
- glass beaker large enough to hold the polystyrene beakers
- cotton wool to improve the insulation of the polystyrene beakers
- a 50 cm^3 measuring cylinder
- weighing boats x 2

Access to:

- a top-pan balance which reads at least to two decimal places
- anhydrous copper(II) sulfate
- hydrated copper(II) sulfate crystals
- distilled water
- paper towels

Safety considerations

- Make sure you have read the advice in the Safety section at the beginning of this book and listen to any advice from your teacher before carrying out this investigation.
- Eye protection must be worn at all times during this experiment.
- The copper(II) sulfate solution is an irritant and copper(II) sulfate is harmful to the environment; any solution should be poured into a bottle and re-used.

Part 1: Determination of ΔH_2

Method

1 Weigh out 0.025 moles of anhydrous copper(II) sulfate.

(Relative atomic masses: Cu = 63.5; S = 32.1 and O = 16)

Relative formula mass of copper(II) sulfate = ……………………..g mol^{-1}

Mass of anhydrous copper(II) sulfate = …………………………...g

Therefore, number of moles of anhydrous copper(II) sulfate = …………………………...mol

HINT
When measuring the temperature make sure that the bulb of the thermometer is completely covered by the water

2 Measure out 50 cm^3 of distilled water into one of the polystyrene beakers. Place the polystyrene beaker into the glass beaker and surround it with cotton wool to improve insulation.

3 **a** Measure the temperature of the distilled water every minute for the next three minutes.

b Record your measurements in the Table 3.6.

4 On the fourth minute do not measure the temperature but add the anhydrous copper(II) sulfate to the distilled water and for the next minute swirl the glass beaker and polystyrene beaker vigorously in order to help the dissolving of the anhydrous copper(II) sulfate.

5 On the fifth minute continue with the measurement of the temperature and do so every minute until ten minutes has elapsed.

Results

Time/min	**0**	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**	**9**	**10**
Temperature/°C					X						

Table 3.6

Data analysis

a Draw a graph of time (horizontal axis) against temperature (vertical axis) for the determination of ΔH_2.

b From your graph determine the initial and final temperatures.

Initial temperature = …………………….°C; final temperature =…………………….°C

c Temperature change = …………………….°C

d Enthalpy change for _____mol = ……………………. J

e Therefore, standard enthalpy change of reaction (ΔH_2) = …………………….kJ mol^{-1}

Part 2: Determination of ΔH_1

Method

1 Weigh out 0.025 moles of $CuSO_4.5H_2O$ crystals.

(Relative atomic masses: Cu = 63.5; S = 32.1 and O = 16)

Record:

a Relative formula mass of $CuSO_4.5H_2O$ crystals = …………………………...g mol^{-1}

b Mass of $CuSO_4.5H_2O$ crystals = …………………………...g

c Therefore number of moles of $CuSO_4.5H_2O$ crystals = ……………………..mol

This mass of $CuSO_4.5H_2O$ crystals already contains some water because of the water of crystallisation and this has to be taken into consideration when measuring out the water in which the copper(II) sulfate crystals are to be dissolved.

For example, if you weigh out 0.0250 mol of $CuSO_4.5H_2O$ then the number of moles of water = 5 x 0.0250 = 0.125 mol and the mass of water present = 0.125 x 18 g = 2.25 g. Therefore, you weigh out 50 – 2.25 g of water = 47.75 g

d Number of moles of water in _______mol of copper(II) sulfate crystals = …………………….. mol

e Mass of water present in the crystals = …………………………...g

Therefore, this mass of water (m) has to be subtracted from the mass of water we will measure out for the 2nd enthalpy determination.

f Mass of water to be weighed out = ……………..g of water

2 **a** Take the other polystyrene beaker and place it on the top-pan balance.

b Zero (Tare) the balance and weigh out ………...g of water.

3 Replace the polystyrene beaker from the previous experiment with the new one in the glass beaker.

4 Wash the thermometer thoroughly with distilled water and wipe dry with a paper towel.

5 Add the 0.025 mol of copper(II) sulfate crystals to the distilled water.

6 Swirl the glass and polystyrene beakers until all the copper sulfate has dissolved.

7 Measure the following temperatures obtained for this solution:

a Minimum temperature = ……………………..°C

b Initial temperature = ……………………..°C

c Final temperature =……………………..°C

Data analysis

a Calculate the temperature change = ……………………..°C

b Calculate enthalpy change for _____mol = …………………….. J

= ……………………..mol

c Calculate the standard enthalpy change of reaction (ΔH_2) = ……………………..kJ mol^{-1}

d Calculate the enthalpy change for the following chemical change:

$CuSO_4\ (s) + 5H_2O(l) \rightarrow CuSO_4.\ 5H_2O(s)$

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

……………………………………………………………………………………………………

Evaluation

The accepted value for the enthalpy change is = $-78.2\ \mathrm{kJ\,mol^{-1}}$

e Calculate the percentage error in your experiment.

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f Calculate the maximum expected percentage errors from your all your items of apparatus and account for the percentage error in your experiment.

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