Data Processing Exercises

[Data Processing]: Volumetric Analysis Practical 1

1. Determination of the number of molecules of H_2O of crystallisation in borax

In this experiment, you are to determine the number of moles of H_2O in hydrated borax.

Borax, also known as disodium tetraborate, reacts with hydrochloric acid according to the equation bxelow:

$$Na_2B_4O_7\,(aq) + 2HCl\,(aq) + 5H_2O\,(l) \rightarrow 2NaCl\,(aq) + 4H_3BO_3\,(aq)$$

You are given the following:

FA 1 is 109.30 g dm^{-3} borax, $Na_2B_4O_7.xH_2O$
FA 2 is 0.100 mol dm^{-3} HCl
Indicator

Procedure

1. Using a burette, place between 17.80 cm^3 and 18.20 cm^3 of **FA 1** into a 100 cm^3 volumetric flask. Record your burette reading below.
2. Make up to the mark with deionised water and shake thoroughly. Label this solution **FA 3**.
3. Fill in a second burette with **FA 2**.
4. Pipette 25.0 cm^3 of **FA 3** into a conical flask and add a few drops of indicator provided.
5. Titrate the contents of the conical flask with **FA 2** until the end point indicated by a colour change.
6. Repeat the titration as many times as necessary to obtain consistent results.
7. Tabulate your results in an appropriate manner.

Results

Table 1: Dilution of FA 1

Final burette reading / cm^3	18.00
Initial burette reading / cm^3	0.00
Volume of **FA 1** used / cm^3	

Table 2: Titration of FA 3 against FA 2

	Rough	1	2
Final burette reading / cm^3	25.95	25.80	26.40
Initial burette reading / cm^3	0.00	0.00	0.60
Volume of **FA 2** used / cm^3			

(a) **(i)** Using your titration results, calculate the number of moles of the $Na_2B_4O_7$ in 25.0 cm^3 of **FA 3**.

number of moles of $Na_2B_4O_7$ in 25.0 cm^3 of **FA 3** = ... [1]

(ii) Calculate the number of moles of $Na_2B_4O_7$ in 100 cm^3 of **FA 3**.

number of moles of $Na_2B_4O_7$ in 100 cm^3 of **FA 3** = ... [1]

(iii) Hence calculate the concentration of $Na_2B_4O_7$ in **FA 1** in mol dm^{-3}.

concentration of $Na_2B_4O_7$ in **FA 1** = ... [1]

(iv) Determine the relative molecular mass of $Na_2B_4O_7.\mathbf{x}H_2O$.

M_r of $Na_2B_4O_7.\mathbf{x}H_2O$ = ...

Hence calculate the value of **x** in $Na_2B_4O_7.\mathbf{x}H_2O$. Give your answer to the nearest whole number.
[*A*r: H, 1.0; B, 10.8; O, 16.0; Na, 23.0]

x is ……..…….[2]

(b) (i) Calculate the oxidation state of boron in $Na_2B_4O_7$ and in H_3BO_3.

oxidation state of B in $Na_2B_4O_7$ is …….. [1]

oxidation state of B in H_3BO_3 is …….. [1]

(ii) Hence suggest the type of reaction between $Na_2B_4O_7$ and HC*l*.

……..
[1]

(iii) Hence or otherwise, suggest a identity for the indicator used in this titration.

……..
[1]

(c) Student **A** repeated the experiment, however in step **1**, he transferred **FA 1** to a 250 cm^3 graduated flask by mistake and top up with deionised water to the mark accordingly. He then obtained the titration results as shown below.

	1	2	3
Final burette reading / cm^3	10.00	20.80	30.80
Initial burette reading / cm^3	0.10	10.50	20.40

(i) With reference to his titration results, explain how the mistake affected the reliability of his results.

(ii) Identify any anomalous titre value in the student's result. Suggest an error that could have lead to the anomalous result.

[Date Processing] : Volumetric Analysis Practical 2

2. To calculate the values of a and b in the formula of the ammonium salt $(NH_4)_aFe(SO_4)_b.6H_2O$

Iron(II) ammonium sulfate, $(NH_4)_aFe(SO_4)_b.6H_2O$ dissociates into iron(II) sulfate and ammonium ions when in solution. In this experiment, you will perform a redox titration of $(NH_4)_aFe(SO_4)_b.6H_2O$ with potassium manganate(VII).

You are given the following:

FA 1 is iron(II) ammonium sulfate crystals
FA 3 is 0.0100 mol dm^{-3} potassium manganate (VII)
sulfuric acid solution

You are to record all the mass measurements and titration results in the space provided.

Preparation of FA 2 by dissolving FA 1 in sulfuric acid

1. Weigh accurately about 5.00 g of **FA 1**. Record your readings in an appropriate manner.
2. Using a 50 cm^3 measuring cylinder, transfer 30.0 cm^3 of sulfuric acid to dissolve the crystals in a small beaker.
3. Transfer the solution and washings into a 250 cm^3 graduated flask and make up to the mark with deionised water. Shake the solution thoroughly and label this **FA 2**.

Titration with potassium manganate(VII) solution

4. Pipette 25.0 cm^3 of **FA 2** into a conical flask.
5. Titrate the mixture against **FA 3** until the solution changes from yellow to permanent orange.
6. Repeat the titration as many times as necessary to obtain consistent results.

(a) Results

Table 1: Mass readings

Mass of weighing bottle and iron(II) ammonium salt /g	9.256
Mass of weighing bottle and residual iron(II) ammonium salt /g	4.253
Mass of iron(II) ammonium salt /g	

Table 2: Titration with FA 3

	Rough	1	2
Final burette reading / cm^3	26.50	27.50	25.50
Initial burette reading / cm^3	0.00	2.00	0.00
Volume of $KMnO_4$ used /cm^3			
Consistent results			

(b) **(i)** Give the balanced chemical equation for the reaction taking place during titration.

…….. [1]

(ii) Using your titration results, calculate the amount of $KMnO_4$ used in the reaction.

amount of $KMnO_4$ = ……..…… [1]

(iii) Calculate the amount of Fe^{2+} in 25.0 cm^3 of **FA 2**. Hence, calculate the amount of Fe^{2+} in the mass of **FA 1** used.

amount of Fe^{2+} = ……..…… [2]

(c) Upon futher analysis, the amount of NH_4^+ ion in **FA 2** is found to be 2.55×10^{-2} mol. Using your answer to **(b)(iii)**, calculate the values of **a** and **b** in the formula $(NH_4)_{\mathbf{a}}Fe(SO_4)_{\mathbf{b}}.6H_2O$.

a is …… and **b** is ……... [2]

(d) **(i)** Explain the purpose of using sulfuric acid, instead of deionised water, to dissolve **FA 1** in step **2**.

(ii) A student claims that the amount of NH_4^+ in **FA 2** can be determined by directly titrating 25.0 cm^3 of **FA 2** with a known concentration of NaOH. Explain whether you agree with the student's claim.

(ii) Using knowledge of qualitative analysis, suggest how you can

I. verify that NH_4^+ and Fe^{2+} ions are present in the **FA 2**.

II. hence separate the NH_4^+ and Fe^{2+} ions in **FA 2** in the form of solution.

[Date Processing] : Volumetric Analysis Practical 3

3. Determination of % of mass of H_2O in $CuSO_4.nH_2O$

Copper(II) ions reacts with an excess of iodide ions to form a mixture containing white copper(I) iodide precipitate in brown I_2 solution:

Equation 1: $4I^-$ (aq) + $2Cu^{2+}$ (aq) → I_2 (aq) + 2CuI (s)

The mass of water of crystallisation present in the hydrated copper(II) salt can be calculated by titrating the iodine liberated with sodium thiosulfate.

Equation 2: $2S_2O_3^{2-}$ (aq) + I_2 (aq) → $S_4O_6^{2-}$ (aq) + $2I^-$ (s)

You are provided with

FA 1 is 200 g dm^{-3} hydrated copper(II) sulfate solution
FA 3 is potassium iodide solution
FA 4 is 0.100 mol dm^{-3} sodium thiosulfate solution
Starch indicator

Procedure

1. Using a burette, place exactly 25.00 cm^3 of **FA 1** into a 250 cm^3 volumetric flask. Record your readings in a table.
2. Make up to the mark with distilled water and shake thoroughly. Label this solution **FA 2**.
3. Pipette 25.0 cm^3 of **FA 2** into a conical flask and add about 15.0 cm^3 of **FA 3** using a 50 cm^3 measuring cylinder. This will produce the aforementioned mixture.
4. Titrate the iodine liberated with **FA 4** until the solution turns pale yellow.
5. Add about 1 cm^3 of starch indicator, using a dropper, and continue titration until the blue-black colour is discharged.
6. Repeat the titration as many times as necessary to obtain consistent results.
7. Tabulate your results in an appropriate manner.

Results

Table 1: Dilution of FA 1

Final burette reading / cm^3	25.00
Initial burette reading / cm^3	0.00
Volume of **FA 1** used / cm^3	

Table 2: Titration with FA 4

	Rough	1	2	3
Final burette reading / cm^3	23.55	43.60	20.25	20.30
Initial burette reading / cm^3	0.00	23.55	0.00	0.20
Volume of **FA 4** used / cm^3				
Consistent results				

(a) Calculation

(i) From your titrations, obtain a suitable volume of **FA 4** to be used in your calculations. Show clearly how you obtained this volume.

Volume of **FA 4** = ……..…… [1]

(ii) Calculate the number of moles of $Na_2S_2O_3$ required to react with the iodine liberated by 25.0 cm^3 of **FA 2**.

number of moles of $Na_2S_2O_3$ = ……..…… [1]

(iii) Hence, calculate how many mole of Cu^{2+} were there in 25.0 cm^3 of **FA 2**.

number of moles of Cu^{2+} = ……..…… [1]

(iv) Using your answer in **(iii)**, calculate the number of moles of Cu^{2+} in 250 cm^3 of **FA 2**. Hence, calculate the concentration of $CuSO_4.\mathbf{n}H_2O$ in mol dm^{-3} in **FA 1**.

concentration of $CuSO_4.\mathbf{n}H_2O$ = ……..…… [1]

(v) Determine the relative molecular mass of $CuSO_4.\mathbf{n}H_2O$.

M_r of $CuSO_4.\mathbf{n}H_2O$ = …….

Hence calculate the value of **n** in $CuSO_4.\mathbf{n}H_2O$. Give your answer to the nearest whole number.
[*A*r: H, 1.0; O, 16.0; S, 32.1; Cu, 63.5]

n is ……..……. [2]

(vi) Calculate the percentage mass of water of crystallisation in the salt.

percentage mass of water of crystallisation ……..…….[1]

(b) Calculate the maximum percentage error in the volume of **FA 4** in your first accurate titre.

percentage error = ……..……. [1]

(d) A student carried out this practical and calculated a value of 2 for **n**. The textbook value of **n** is more than 2.

Suggest an error the student could have done in the practical procedure of the experiment that could account for this result and explain why this gives a value of **n** that is too low.

[Date Processing] : Volumetric Analysis Practical 4

4. Determination of change in oxidation number of unknown element X

In this experiment, you are to determine the change in oxidation number of element **X**, when $\mathbf{X}O_3^-$ ions react with Fe^{2+} ions.

There are two reactions taking place.

$\mathbf{X}O_3^-$ is first reacted with excess Fe^{2+} to give the reduced form of **X** and Fe^{3+}. The unreacted Fe^{2+} is subsequently titrated with acidified $KMnO_4$. The reduced form of **X** does not react with $KMnO_4$.

You are given the following:

FA 1: 0.75 mol dm^{-3} acidified Fe^{2+} ions
FA 2: 0.075 mol dm^{-3} $\mathbf{X}O_3^-$
FA 3: 0.020 mol dm^{-3} $KMnO_4$

Procedure

1. Using a burette, place exactly 50.00 cm^3 of **FA 1** into a 250 cm^3 volumetric flask.
2. Make up to the mark with distilled water and shake thoroughly. Label this solution **FA 4**.
3. Pipette 25.0 cm^3 of **FA 4** into a conical flask.
4. Using the other pipette, add 10.0 cm^3 of **FA 2** to the same conical flask.
5. Titrate the mixture with **FA 3**.
6. Repeat the titration as many times as necessary to obtain consistent results.
7. Tabulate your results in an appropriate manner.

Results

	Rough	1	2
Final burette reading / cm^3	22.65	22.45	44.95
Initial burette reading / cm^3	0.00	0.00	22.45
Volume of **FA 3** used /cm^3			

(a) (i) Write the balanced chemical equation for the reaction taking place during the titration.

…….. [1]

(ii) Using your titration results, calculate the amount of Fe^{2+} that reacted with MnO_4^-.

amount of Fe^{2+} = …….. [1]

(iii) Calculate the amount of Fe^{2+} in 25.0 cm^3 of **FA 4** before **FA 2** is added.

amount of Fe^{2+} = …….. [1]

(iv) Hence calculate the amount of Fe^{2+} that reacted with 10.0 cm^3 of **FA 2**.

amount of Fe^{2+} that reacted with 10.0 cm^3 of **FA 2**
= …….. [1]

(v) Calculate how many number of moles of Fe^{2+} ions reacted with one mole of $\mathbf{X}O_3^-$ ions.

number of moles of Fe^{2+} ions = …….. [1]

(vi) Hence, deduce the change in oxidation number of the element **X** when $\mathbf{X}O_3^-$ ion reacts with Fe^{2+} ions.

[1]

(b) Using your answer to **(a)(vi)**, construct a half-equation for the reaction undergoes by $\mathbf{X}O_3^-$ to form $\mathbf{X}O_2^-$.

Hence give a balanced chemical equation for the reaction between Fe^{2+} and $\mathbf{X}O_3^-$.

half equation ……. [1]

chemical equation ..…….. [1]

(c) Explain the importance of performing step **1** and **2** in the procedure, by stating the effect on the titration results if it is not done.

[Date Processing] : Volumetric Analysis Practical 5

5. Determination of concentration of HC*l* and mass of NaOH in solid mixture through double–indicator titration

When a solution containing a mixture of sodium hydroxide and potassium carbonate is titrated against hydrochloric acid, three reactions occur.

1. $NaOH + HCl \rightarrow NaCl + H_2O$ indicator: *thymolphthalein*
2. $K_2CO_3 + HCl \rightarrow KHCO_3 + KCl$ indicator: *thymolphthalein*
3. $KHCO_3 + HCl \rightarrow KCl + CO_2 + H_2O$ indicator: *methyl orange*

If a solution of this mixture is titrated against hydrochloric acid using thymolphthalein indicator, the blue colour of the indicator is discharged when reactions **1** and **2** are complete.

If methyl orange indicator is then added at this point to the mixture and the titration is continued, the end-point corresponds to the completion of reaction **3**.

You are given the following:

FA 1 is prepared by dissolving a known mass of solid sodium hydroxide with 2.20 g solid potassium carbonate in 250 cm^3 of deionised water in a beaker.
FA 2 is hydrochloric acid of unknown concentration

Titration of FA 1 against FA 2

1. Fill the burette with **FA 2**.
2. Using a 50 cm^3 measuring cylinder, transfer 25.0 cm^3 of **FA 1** into a conical flask and titrate with **FA 2**, using thymolphthalein as indicator.
3. Run **FA 2** from the burette into this flask until solution changes from blue to colourless. Record the volume of acid required.
4. Add 1-2 drops of methyl orange indicator into the reaction mixture in the conical flask and continue titrating until the colour changed from yellow to orange. Record the volume of acid required.
5. Record your titration results in the spaces provided below. Make certain that your recorded results show the precision of your instrument used.
6. Repeat steps **2** to **6** until your results are consistent.

Results

Table 1: Titration of FA 1 with FA 2

	1	2
Initial burette reading / cm^3	0.00	0.00
Burette reading at thymolphthalein end-point / cm^3	21.30	21.20
Burette reading at methyl orange end-point / cm^3	29.30	29.30

(a) Calculation

(i) From your titrations, obtain the average volume of **FA 2** at first end-point and second end-point.

average volume of **FA 4** at first end-point = …… [1]

average volume of **FA 4** at second end-point = …… [1]

(ii) Calculate the concentration of K_2CO_3 in **FA 1** in mol dm^{-3}.
[*A*r: C, 12.0; O, 16.0; K, 39.1]

concentration of K_2CO_3 = …… [1]

(iii) Using your answer in **(ii)**, calculate the amount of HC*l* required to react with $KHCO_3$ in reaction **3**.

amount of HC*l* required = …… [1]

(iv) Hence using your titration results and answer in **(iii)**, calculate the concentration of HC*l* in **FA 2**.

concentration of HC*l* in **FA 2** = …… [1]

(v) Hence, calculate the mass of NaOH used in the solid mixture to prepare **FA 1**.
[*A*r: H, 1.0; O, 16.0; Na, 23.0]

mass of NaOH = ……..…… [2]

(b) **(i)** State one **significant** source of error in the procedure used in this experiment and explain the effect on the titration results. Suggest an improvement that could be made to reduce this error.

(ii) A student states that the titration results is inaccurate as solid sodium hydroxide, used in preparation of **FA 1**, tends to absorb CO_2 and moisture from the air readily. Explain whether this error will cause the titration results to be higher or lower than expected.

[Date Processing] : Volumetric Analysis Practical 6

6 Determination of the value of *n* in IO_n^-

IO_n^- is an oxoanion of iodine. It can react with iodide ions, I^-, in an acidic medium, to produce only molecular iodine and water.

This amount of iodine can be analysed via titration with sodium thiosulfate, $Na_2S_2O_3$.

$$2S_2O_3^{2-} + I_2 \longrightarrow S_4O_6^{2-} + 2I^-$$

In this question, you will perform an experiment to determine the stoichiometric ratio between IO_n^- and I_2 and the value of *n* in IO_n^-.

FA 1 is solid hydrated sodium thiosulfate, $Na_2S_2O_3 \cdot 5H_2O$
FA 2 is a solution containing 0.00825 mol dm^{-3} KIO_n
FA 3 is 1.0 mol dm^{-3} sulfuric acid, H_2SO_4
FA 4 is 0.25 mol dm^{-3} potassium iodide, KI

You are also provided with starch solution.

(a) Preparation of a standard solution of 0.0484 mol dm^{-3} sodium thiosulfate

(i) Determine the mass of **FA 1** required to prepare 250 cm^3 of sodium thiosulfate solution with a concentration 0.0484 mol dm^{-3}.
[A_r: S, 32.1; O, 16.0; Na, 23.0; H, 1.0]

mass of **FA 1** required = …….. g [1]

You will follow the instructions to perform the preparation of the standard solution.

Record your measurements in Table 1.1.

(ii)
1. Using a mass balance, weigh out accurately the mass of **FA 1** determined in **(a)(i)** into a weighing bottle.
2. Dissolve **FA 1** in about 50 cm^3 of deionised water in a 100 cm^3 beaker.
3. Transfer the solution and washings into a 250 cm^3 volumetric flask, using a filter funnel.
4. Top up to the 250 cm^3 mark with deionised water.
5. Stopper the flask and shake to ensure solution is homogeneous. Label this solution as **FA 5**.

Complete Table 1.1 with your mass measurements and calculate the mass of **FA 1** that you used.

Table 1.1

Mass of weighing bottle and **FA 1** / g	
Mass of empty weighing bottle / g	4.020
Mass of weighing bottle and residual **FA 1** / g	
Mass of **FA 1** / g	3.00

[2]

(b) Preparation and titration of the reaction mixture

1. Fill a burette with **FA 5**.
2. Using a 25.0 cm^3 pipette, add 25 cm^3 of **FA 2** to a 250 cm^3 conical flask.
3. Using a measuring cylinder, add 10.0 cm^3 of **FA 3**, followed by 10.0 cm^3 of **FA 4** to the same conical flask and swirl quickly.
4. **Immediately** titrate the I_2 in the conical flask with **FA 5**. When a pale yellow colour is obtained, add 1 cm^3 of starch solution. The end-point is reached when the blue – black colour of the solution decolourises.
5. Repeat steps **2** to **4** until consistent results are obtained and record your results below.

Results

Final burette reading/cm^3	24.35	23.65
Initial burette reading/cm^3	0.70	0.00
Volume of FA5/ cm^3		

[5]

Obtain, from your titration results, a suitable average titre of **FA 5**. Show clearly the titres you used in calculating this average.

Average volume of **FA 5** = …….. [1]

Calculations

Show your working and appropriate significant figures in all of your calculations.

(c) (i) Calculate the amount, in moles, of $Na_2S_2O_3$ present in the volume of **FA 5** determined in **(b)** and hence the amount of iodine produced.

Amount of I_2 = ……..[2]

(ii) Calculate the amount of **FA 2** used in the titration.

Amount of **FA 2** = …….. [1]

(d) **(i)** Construct a balanced half-equation, in terms of *n*, for the reduction of IO_n^- to I_2.

……
[1]

(ii) Hence construct a balanced equation, in terms of *n*, for the reaction between IO_n^- and I^- to give I_2 and H_2O.

…… [1]

(iii) Determine the amount of I_2 produced by one mole of IO_n^- and determine the value of *n* in IO_n^-.

n =…….. [1]

[Date Processing] : Volumetric Analysis Practical 7

7. Determination of K_{sp} of $M(OH)_2$

$M(OH)_2$ is an ionic compound that has a low solubility in water. The solubility of such a compound can be represented using a type of equilibrium constant called the *solubility product, K_{sp}*.

You will determine the solubility product of $M(OH)_2$ by a titration method using sulfamic acid and the indicator screened methyl orange.

You are given the following:

FA 1 is a **saturated** solution of $M(OH)_2$ in water, it is the filtrate formed by adding excess $M(OH)_2(s)$ to a fixed volume of water followed by filtration.
FA 2 is 0.500 mol dm^{-3} sulfamic acid, NH_2SO_3H, a strong monobasic acid
Screened methyl orange

Procedure

1. Using a burette, place between 30.00 and 31.00 cm^3 of **FA 2** into a 250 cm^3 graduated flask and make up to the mark with distilled water. Shake the mixture and label it **FA 3**. Record your burette readings.
2. Pipette 25.0 cm^3 of **FA 1** into a conical flask. Add a few drops of screened methyl orange and titrate with **FA 3**, until it turns from green to grey.
3. Repeat the titration as many times as necessary.
4. Tabulate your results in an appropriate manner.

Results

Table 1: Dilution of FA 2

Final burette reading / cm^3	30.00
Initial burette reading / cm^3	0.00
Volume of **FA 2** used /cm^3	

Table 2: Titration

	rough	1	2
Final burette reading / cm^3	20.50	40.80	20.20
Initial burette reading / cm^3	0.00	20.50	0.00
Volume of **FA 3** used /cm^3			

(a) (i) Calculate the concentration, in mol dm^{-3}, of sulfamic acid in **FA 3.**

concentration, in mol dm^{-3}, of sulfamic acid in **FA 3** = …… [1]

(ii) Calculate the amount, in moles, of NH_2SO_3H in your mean titre.

amount of NH_2SO_3H = ... [1]

(b) Give the balanced equation for the neutralisation reaction between $M(OH)_2$ and sulfamic acid in the titration.

..... [1]

(c) Calculate the concentration in mol dm^{-3}, of $M(OH)_2$ **FA 1**.

concentration, in mol dm^{-3}, of $M(OH)_2$ **FA 1** = ... [1]

(d) Aqueous $M(OH)_2$ is fully ionised in solution **FA 1**.

(i) Use your answer to **(c)** to deduce the concentration of M^{2+} in **FA 1**.

concentration of M^{2+} = ... [1]

(ii) Deduce the concentration of OH^- in **FA 1.**

concentration of OH^- = ... [1]

(e) **(i)** Give the equation for the solubility product K_{sp}, of $M(OH)_2$, indicating the units.

…….. [1]

(ii) Use your equation and your answers to **(d)(i)** and **(ii)** to calculate the solubility product of $M(OH)_2$.

K_{sp} of $M(OH)_2$ = …….. [1]

(f) Another student performs this experiment using a 50 cm^3 measuring cylinder to measure 30.0 cm^3 of **FA 2** into the graduated flask instead of a burette. State and explain in terms of percentage error (uncertainty), whether this student was wise to choose this measuring cylinder to measure **FA 2.**

Data Processing

[Data Processing] : *Gas Collection Practical 1*

1. Determination of M_r of the ethanedioate salt

Calcium ethanedioate is the major component of the most common type of human kidney stones. It is one of a series of salts formed from ethanedioic acid, $H_2C_2O_4$. Another of these salts can be represented by the formula $X_2C_2O_4$, where **X** is a Group 1 metal.

Solution **Q** contains 43.0 g dm^{-3} of $X_2C_2O_4$ in deionised water. You are not provided with **Q**.

You are provided with the following chemicals.

FA 1 is a diluted solution of **Q**, in which 35.70 cm^3 of **Q** was made up to 250 cm^3 with deionised water in a graduated flask.
FA 2 is 0.0200 mol dm^{-3} potassium manganate(VII), $KMnO_4$.
FA 3 is 1.00 mol dm^{-3} sulfuric acid, H_2SO_4.
FA 4 is a solution containing manganese(II) ions, Mn^{2+}.

In this experiment, you will react the ethanedioate salt with potassium manganate(VII), in the presence of a small amount of manganese(II) ions. You will measure the volume of CO_2 gas produced at timed intervals and determine the maximum volume of CO_2 gas produced.

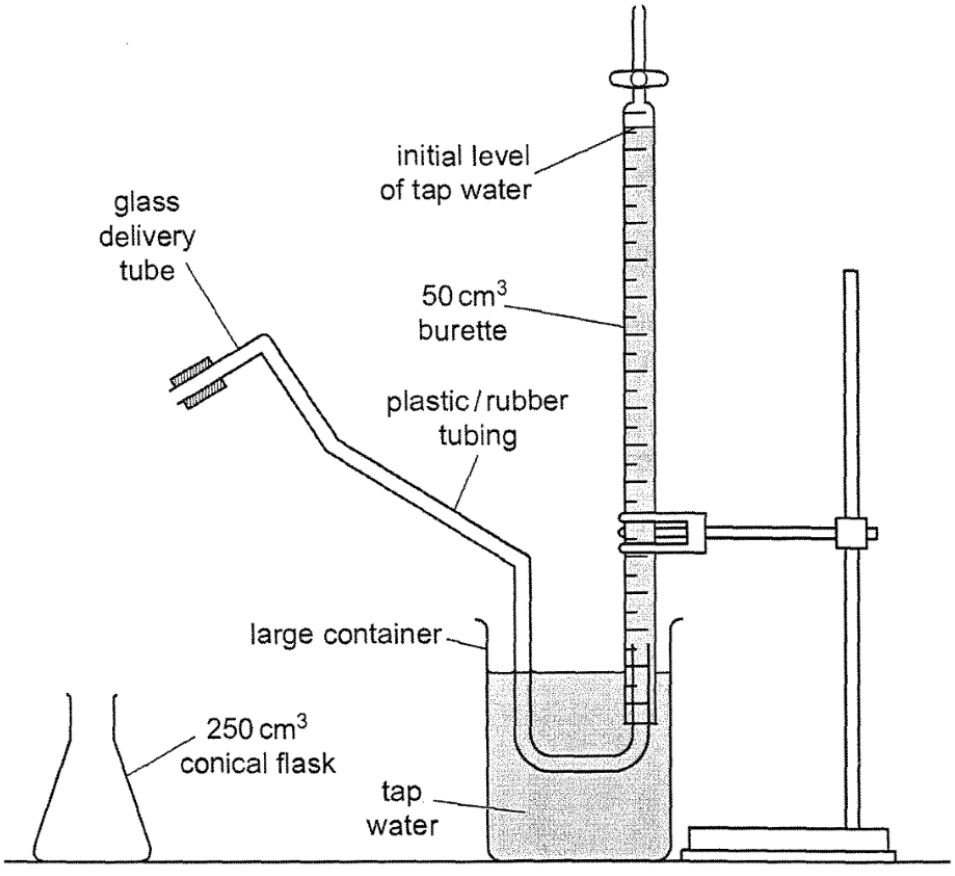


Fig. 1

You will measure the level of water in the burette at timed intervals.

In an appropriate format in the space provided on the next page, prepare a table in which you may record each burette reading and the time it was taken.

In addition, you will need to show the total volume of CO_2 collected up to that time, recorded to one decimal place.

1. Set up the apparatus as shown in Fig 1. You should insert the plastic/rubber tubing to a sufficient depth in the burette so that it will not subsequently shake loose.

2. Adjust the water level in the burette until it is between 48.0 cm^3 and 50.0 cm^3. You may find it helpful to use an empty dropping pipette to introduce small amounts of air from the bottom of the burette tube.

3. Using an appropriate measuring cylinders to add to the 250 cm^3 conical flask
   - 20.0 cm^3 of **FA 1**, dil I
   - 50.0 cm^3 of **FA 3**. H_2SO_4

4. Using a dropping pipette, add about 1 cm^3 of **FA 4** to the conical flask. Mn^{2+}

5. Using an appropriate measuring cylinder, measure out 30.0 cm^3 of **FA 2**. $KMnO_4$

6. Transfer the **FA 2** into the conical flask and insert the bung into the conical flask. $KMnO_4$

7. Allow any bubbles created in the burette when the bung was inserted in the conical flask to rise to the top. Start the stopwatch, read and record the initial water level in the burette.

   Note: Once you have started the stopwatch, allow it to continue running for the duration of the experiment. You **must not** stop the stopwatch until the reaction is complete.

8. Check that the plastic/rubber tubing is securely positioned in the burette.

9. Hold the flask by its neck and gently swirl it continuously.

10. At t = 0.5 min, read and record the water level in the burette, to 1 decimal place, together with the time it was measured in your table.

11. Continue to gently swirl the flask. Read and record the water level in the burette every minute, until the reaction is complete.

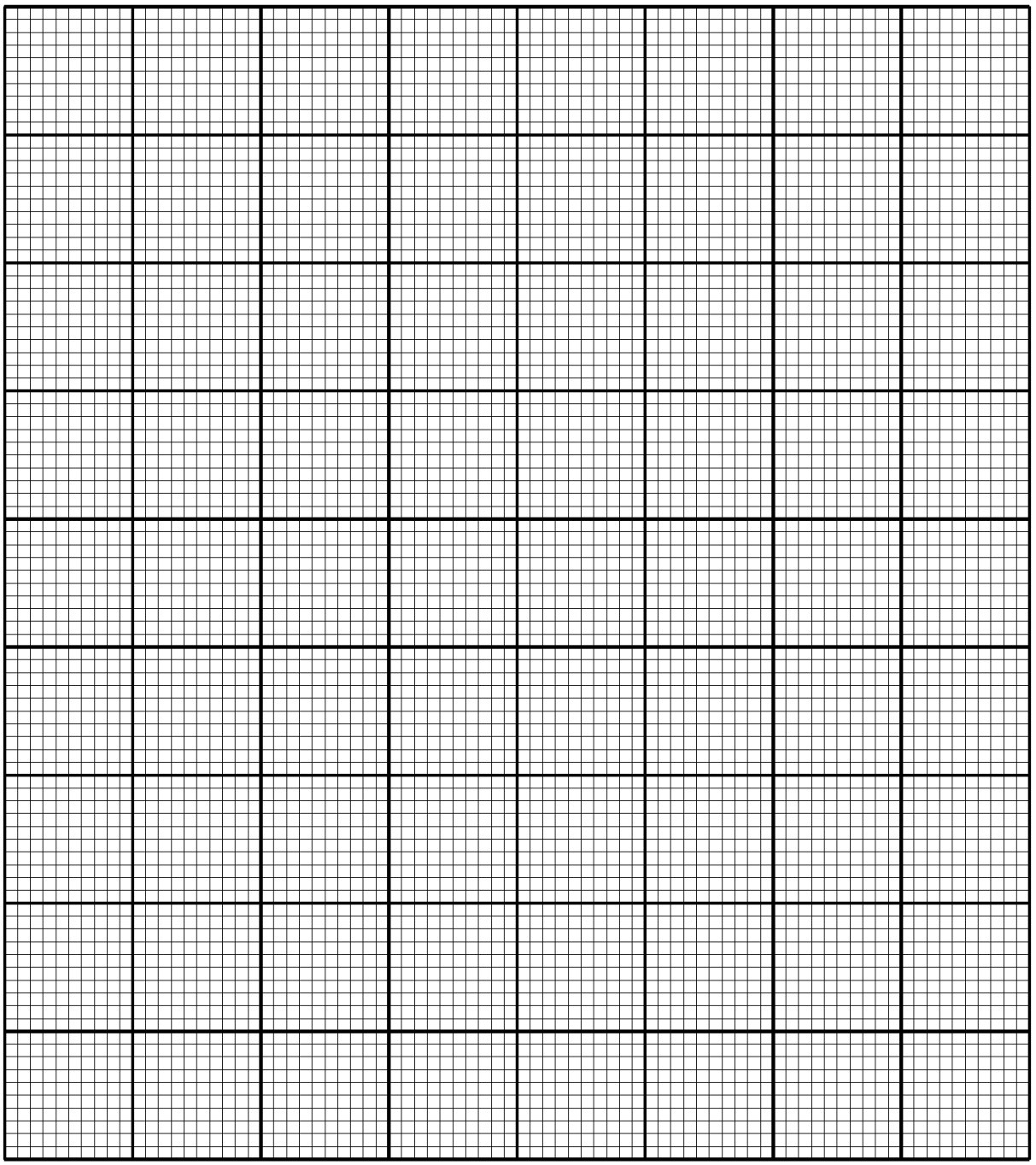
(a) (i) Experimental results

time / min	burette reading / cm^3	vol of CO_2 produced / cm^3
0.0	50.0	
0.5	41.0	
1.5	35.6	
2.5	32.7	
3.5	30.5	
4.5	29.6	
5.5	28.6	
6.5	28.2	
7.5	28.0	
8.5	27.5	
9.5	27.5	
10.5	27.5	

(ii) Plot on the grid below, a graph of volume of CO_2 on the y-axis, against time, *t*, on the x-axis.

Draw the most appropriate line, taking into account all your points.

[4]



(iii) Write the balanced chemical reaction for the reaction between ethanedioate ions and manganate(VII) ions.

…….. [1]

(iv) Using your equation in **(iii)** and appropriate data from your graph, calculate a value for the amount of ethanedioate, $C_2O_4^{2-}$ ions, present in 20.0 cm^3 of **FA 1**.
[molar volume of gas = 24.0 dm^3 mol^{-1} at r.t.p.]

amount of ethanedioate, $C_2O_4^{2-}$ ions = ……. [1]

(v) Determine the concentration, mol dm^{-3}, of $C_2O_4^{2-}$ ions in **Q**.

concentration of $C_2O_4^{2-}$ ions in **Q** = ……. [2]

(vi) Use your answer to **(v)** to calculate the M_r of the ethanedioate salt.

M_r of the ethanedioate salt = …… [1]

Hence deduce the identity of **X**.
Show your working.

[A_r: C, 12.0; O, 16.0; H, 1.0; Li, 6.9; Na, 23.0; K, 39.1; Rb, 85.5; Cs, 132.9; Fr, 223.0]

X is ……. [2]

(b) **(i)** Another 25.0 cm^3 **FA 1** sample was analysed by titrating against acidified **FA 2** and the titre volume was found to be 28.50 cm^3.

Based on the results, calculate the concentration of $C_2O_4^{2-}$ in **FA 1**.

concentration of $C_2O_4^{2-}$ in **FA 1** = ……[1]

(ii) Using your answer to **(b)(i)**, determine the maximum volume of CO_2, at r.t.p., that could have been produced from 20.0 cm^3 of **FA 1**.

maximum volume of CO_2 produced from 20.0 cm^3 of **FA 1** = ……[1]

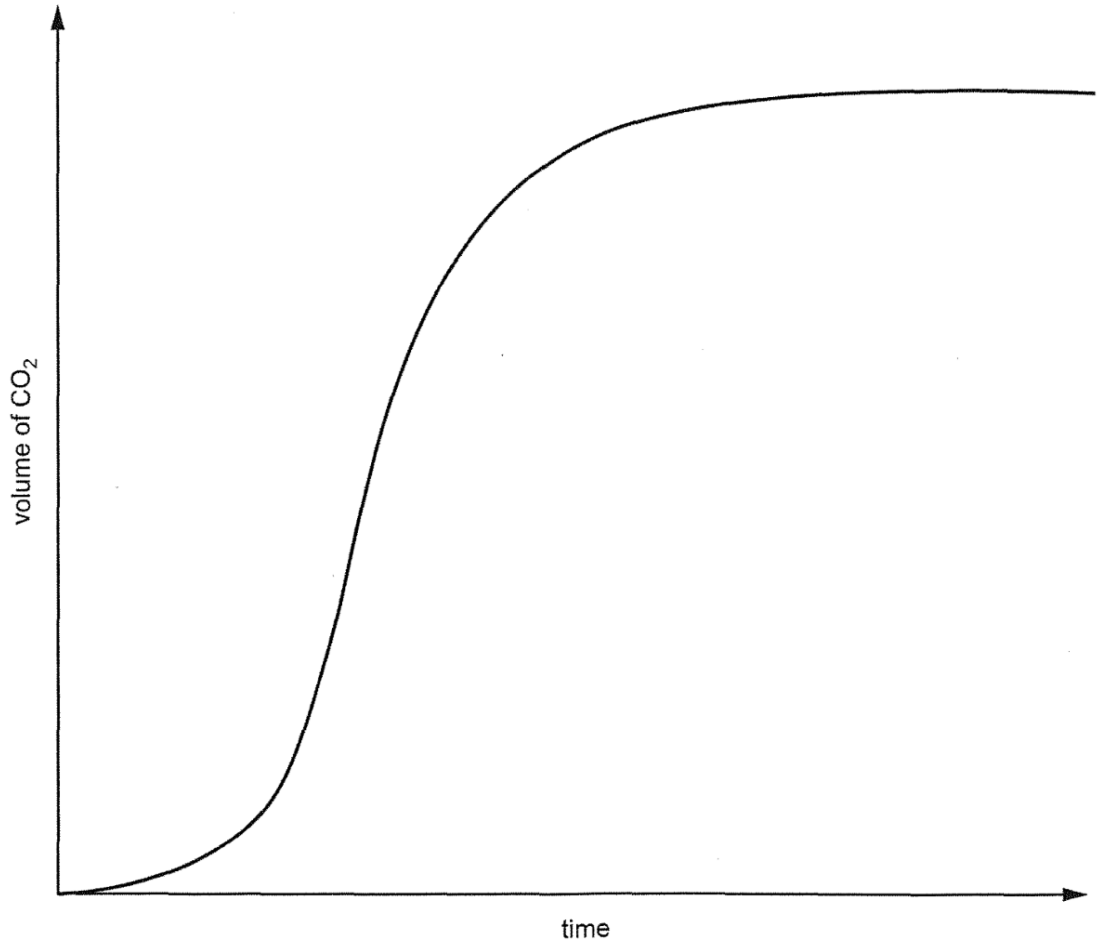
(iii) Suggest a reason for the difference between the total volume of CO_2 you collected and the maximum volume of CO_2 calculated in **2(b)(ii)**.

(iv) Suggest an improvement to this experiment that would overcome this problem.

(c) The presence of Mn^{2+} ions, which are produced in the reaction between MnO_4^- ions and $C_2O_4^{2-}$ ions, is thought to catalyse this reaction.

A student performed the experiment you performed in **2(a)(i)** but forgot to add **FA 4** to the mixture of **FA 1** and **FA 3** before adding the **FA 2**.

The student performed the experiment at the same temperature as your experiment and obtained the graph shown below.



Consider the **shape** of the graph above and your graph in **2(a)(ii)**.

Describe one major difference between their shapes. Suggest an explanation for your answer.

Difference:

Explanation:

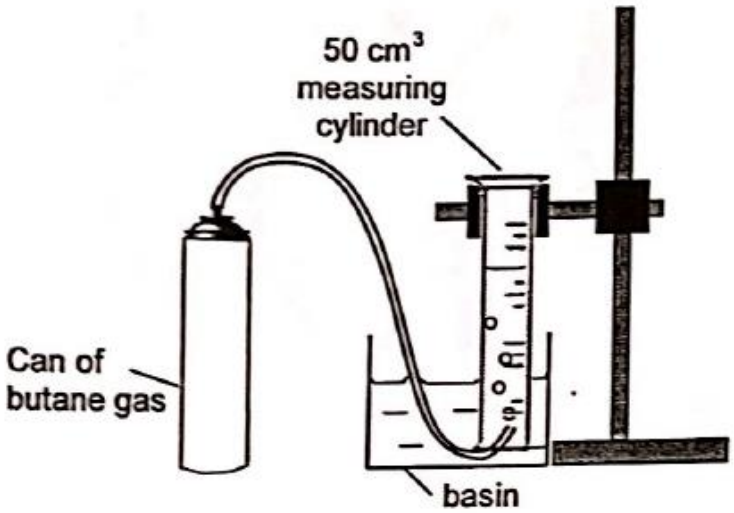
[Data Processing] : Gas Collection Practical 2

2. Determination of M_r of Butane

In this experiment, you are to use the ideal gas equation to find the relative molecular mass of a gas. Butane gas from a pressurised container is collected and the decrease in mass of the container is determined to allow relative molecular mass of butane to be calculated.

Procedure

1. Weigh a can of gas and record its mass in an appropriate manner.
2. Half fill the water container but fill the measuring cylinder with water fully.
3. Set up the retort stand next to the water container. Carefully invert the measuring cylinder full of water into the water container for gas collection. Clamp the measuring cylinder in position.
4. Insert the open end of the rubber tubing into the measuring cylinder. Attach the other end of the tubing with an adaptor to the gas can nozzle.
5. The liquid level inside the measuring cylinder should be on the graduated scale. If it is not, introduce some air bubbles into the cylinder using a dropper. Remember that the scale is now upside down. Record the initial measuring cylinder reading.



6. Lightly depress the gas nozzle to release some gas into the measuring cylinder and collect about 40 cm^3 of gas. Record the final measuring cylinder reading.
7. Reweigh the can of gas, making sure that it is dry.
8. Record the temperature of the laboratory.

Results

Table 1: Mass readings

mass of can of gas before experiment /g	239.409
mass of can of gas after experiment /g	239.304
Mass of butane gas used /g	

Table 2: Volume readings

Initial measuring cylinder reading/ cm^3	8.0
Final measuring cylinder reading /cm^3	50.0
Volume of butane gas / cm^3	

Temperature of laboratory = 30.5 ºC

(a) By using the readings that you have obtained in this experiment and the following equation, determine the relative molecular mass of butane.

$$pV = nRT$$

M_r of butane = …….. [2]

(b) **(i)** Compare your calculated M_r of butane in **(a)** with the theoretical value. Give **one** reason that could account for the difference between your calculated M_r of butane and the theoretical value.

(iii) Identify **one** error in the procedure that affects the reliability of the results. Suggest an improvement that could be made to reduce the error you have identified.

Data Processing

[Data Processing] ; Gravimetric Analysis Practical 1

1. Determination of the formula of magnesium oxide

In this experiment, the formula of magnesium oxide will be determined by heating magnesium wire in air to form magnesium oxide. The increase in mass at the end of the heating process is due to the oxygen that has combined with the magnesium to form magnesium oxide.

(a) You are to follow the practical procedure exactly and record your measurements in a suitable format. You should carefully consider the most effective way of handling the materials and apparatus, so as to obtain the best result.

- 1 Weigh a crucible and lid.
- 2 Clean about 2.5 cm of magnesium ribbon using a sandpaper.
- 3 Place the magnesium ribbon in the crucible and weigh this together with the lid.
- 4 Place the crucible on a wire gauze on a tripod and heat strongly for 2 minutes. Lift the lid very slightly to let more oxygen in but immediately replace it to prevent magnesium oxide from escaping.
- 5 Do this until the magnesium no longer flares up. Remove the lid and heat strongly for five minutes. Allow the crucible to cool on a heat-proof mat and then re-weigh with lid.
- 6 Repeat the heating, cooling and weighing cycle until constant mass of ±0.010 g is achieved.

Results

Table 1: Mass of Magnesium ribbon

Mass of crucible, lid and magnesium ribbon /g	5.225
Mass of crucible and lid /g	5.025
Mass of magnesium ribbon /g	0.200

Table 2: Mass of Magnesium oxide

Mass of crucible, lid and content after 1st heating /g	5.230
Mass of crucible, lid and content after 2nd heating /g	5.335
Mass of crucible, lid and content after 3rd heating /g	5.359
Mass of crucible, lid and content after 4th heating /g	5.367
Mass of crucible and lid /g	5.025
Mass of magnesium oxide /g	

(b) **(i)** Determine the mass of oxygen that has combined with magnesium.

mass of O_2 = …… [1]

(ii) Hence, determine the formula of magnesium oxide.

formula = ……..…… [2]

(c) **(i)** Explain the purpose of **step 2** in the procedure in **(a)**.

(ii) A student repeated the experiment using the procedure in **(a)** but did not carry out **step 6**. Explain the effect on his determination of the formula of magnesium oxide.

[Data Processing] ; Gravimetric Analysis Practical 2

2. Determination of the number of molecules of H_2O in hydrated barium bromide

The water of crystallisation in hydrated barium bromide is lost when heated.

$$BaBr_2.xH_2O\ (s) \rightarrow BaBr_2\ (s) + x\ H_2O\ (g)$$

In this experiment, you are to heat samples of varying mass of barium bromide and to determine the number of molecules of H_2O in hydrated barium bromide by plotting a graph of the mass of barium bromide against the mass of the anhydrous barium bromide left after heating.

(a) Procedure

A student weighed a clean, dry crucible and transferred 0.40 g of hydrated barium bromide to it. The crucible and its contents were heated until a constant mass has been reached (heating-cooling-weighing method). This mass was then recorded.

The student then repeated the experiment using different masses of hydrated barium bromide.

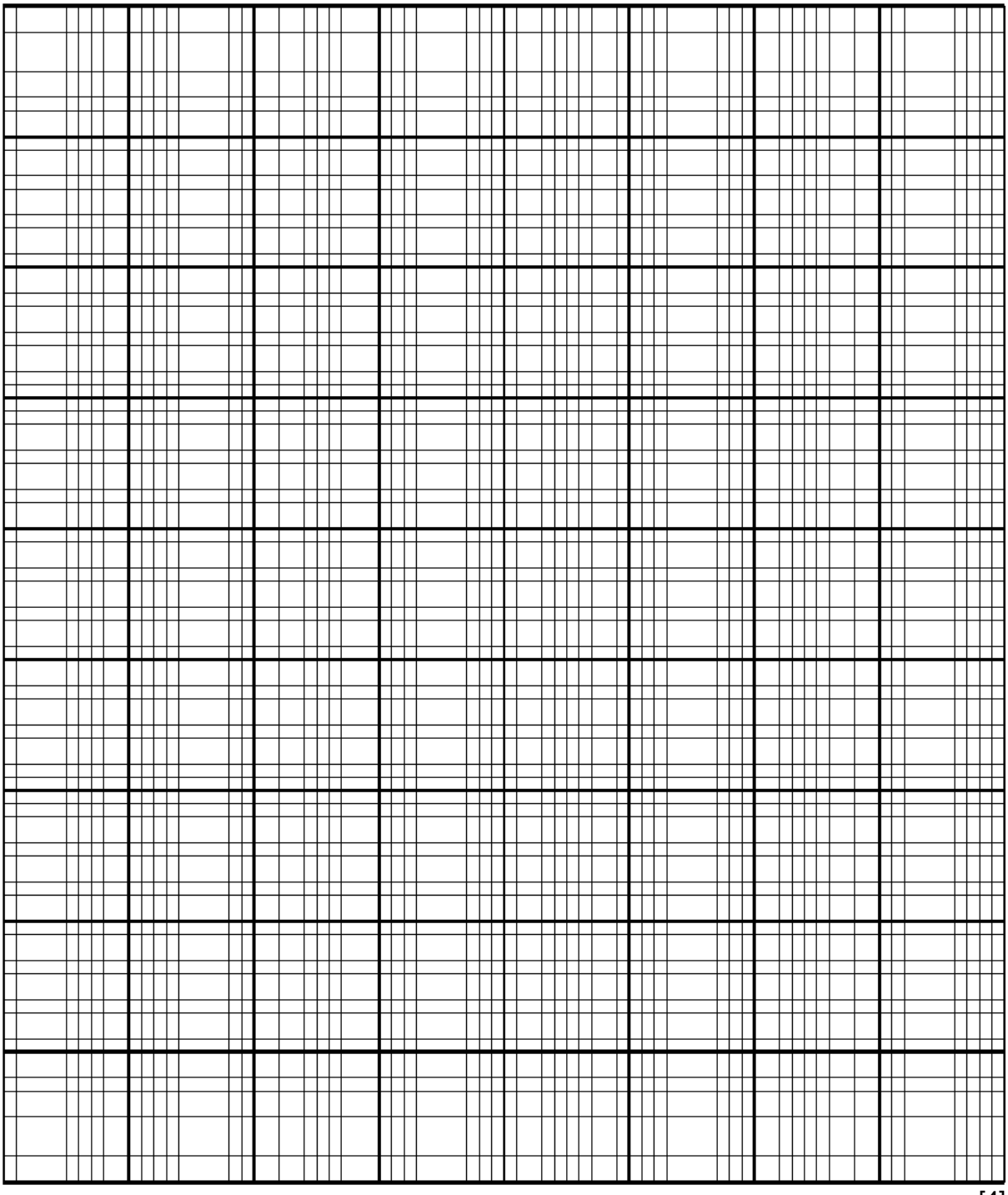
For each experiment the student recorded the original mass of the hydrated barium bromide and the mass of anhydrous barium bromide left after heating. The results are as shown below.

Results

Experiment	mass of $BaBr_2.xH_2O$ / g	mass of $BaBr_2$ / g
1	0.40	0.34
2	0.60	0.58
3	0.70	0.61
4	0.90	0.77
5	1.00	0.86
6	1.30	1.13

(b) **(i)** Plot on the grid below, a graph of the mass of barium bromide on the y-axis, against the mass of the anhydrous barium bromide on the x-axis.

Draw the most appropriate line.



[4]

(ii) Use this graph to determine the mass of hydrated barium bromide which forms 1.00 g of anhydrous barium bromide.

mass = ……..…… [1]

(iii) Calculate the number of moles of $BaBr_2$ present in 1.00 g of anhydrous barium bromide.
[*A*r: Ba, 137.3; Br, 79.9]

No. of moles of $BaBr_2$ = ……..…… [1]

(iv) Using your answer to **(ii)** and **(iii)**, calculate the M_r of hydrated barium bromide.

M_r of $BaBr_2.\mathbf{x}H_2O$ = ……..…… [1]

Hence calculate the value of **x** in $BaBr_2.\mathbf{x}H_2O$. Give your answer to the nearest whole number.
[*A*r: H, 1.0; O, 16.0; Ba, 137.3; Br, 79.9]

x is ……..…….[2]

(c) Considering your answer obtained in **(b)(iv)** above, comment on the results obtained by the student. Is your line of best fit good enough to be used with confidence?

(d) A databook value for the M_r of hydrated barium bromide is 333.1. Calculate the difference between the M_r value obtained from the student's data and the databook value. Express this difference as a percentage of the databook value.

percentage difference = ……..…… [1]

(e) **(i)** Explain why there is a need to heat the crucible and its content to constant mass.

(ii) Suggest an explanation why the hydrated barium bromide should not be stirred during heating.

[Data Processing] ; Gravimetric Analysis Practical 3

3 Determination of the the M_r of an unknown metal carbonate.

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The amount of carbon dioxide evolved when a metal carbonate, X2CO3, reacts with an excess of an acid is directly proportional to the mass of the carbonate used. The relative molecular mass of X2CO3 can hence be determined by conducting a simple experiment.

Procedure

1. The mass of a weighing bottle containing X_2CO_3 is measured.
2. Using a suitable measuring cylinder, 50.0 cm^3 of 2.0 mol dm^{-3} hydrochloric acid (in excess) is transferred to a 250 cm^3 conical flask.
3. A cotton wool plug is placed in the neck of the flask and the flask + acid is weighed.
4. The weighed X_2CO_3 is added to the acid in the flask and the cotton wool plug is quickly replaced in the neck of the flask to prevent any loss of acid spray.
5. The mass of the empty weighing bottle is measured.
6. When the reaction in the flask has stopped, the flask is left for 10 minutes to allow carbon dioxide to diffuse from the flask.
7. The mass of the flask and its contents after the reaction are measured.

Results

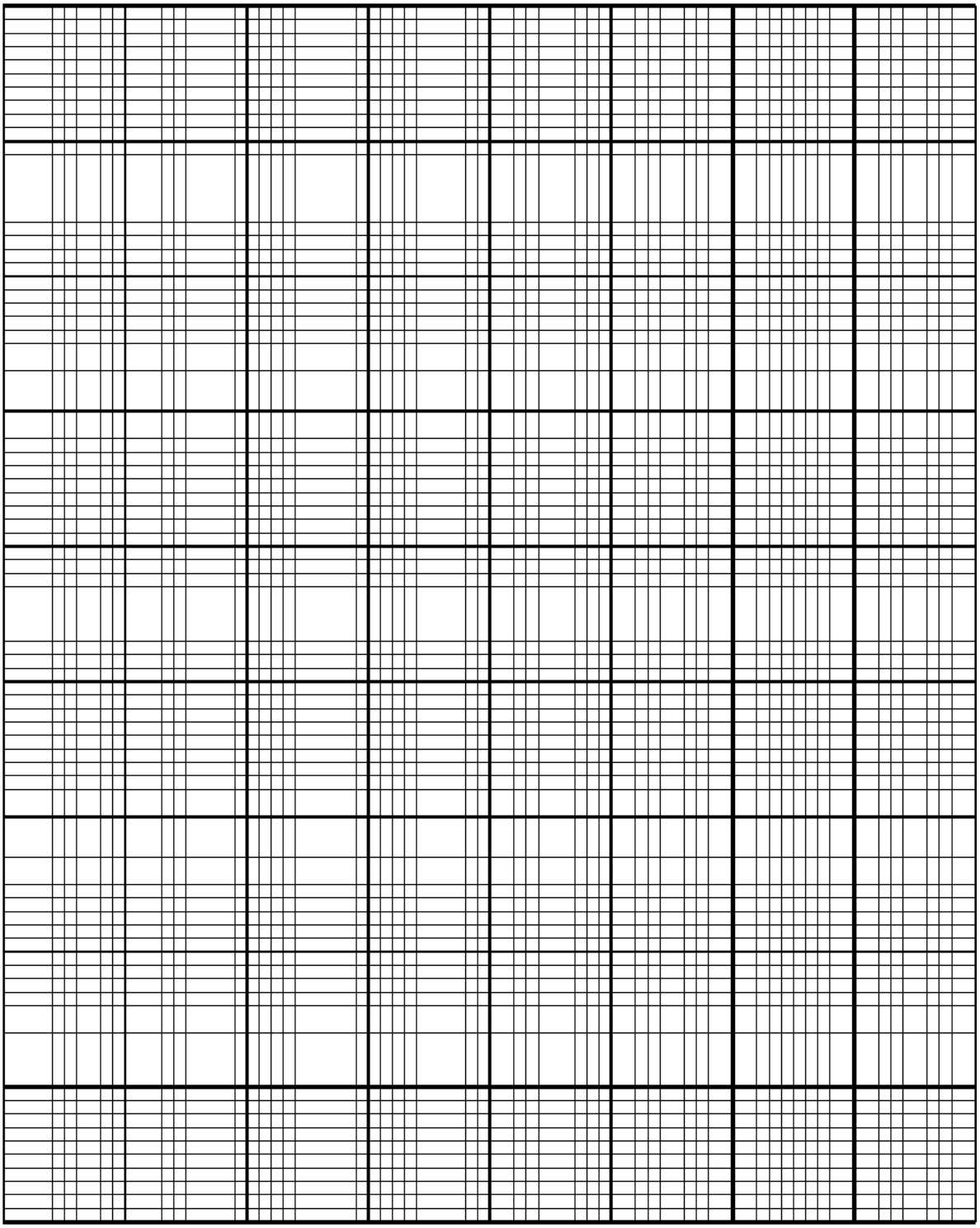
A	**B**	**C**	**D**				
mass of weighing bottle + X2CO3 / g	mass of empty weighing bottle / g	mass of flask + acid before the reaction / g	mass of flask + solution after the reaction / g				
14.29	11.48	221.35	223.21				
16.41	11.76	209.71	212.91				
12.24	11.34	210.45	210.97				
16.77	11.27	214.38	217.80				
16.48	10.68	211.63	215.59				
14.85	11.15	217.18	219.68				

(a) Process the results in the table to calculate both the mass of X_2CO_3 used and the mass of CO_2 given off. Record these values in the additional columns of the table. You may use some or all of these columns.

Label the columns you use, including the units and an equation to show how the value is calculated. You may use the column headings **A** to **H** in these equations, e.g. = **C** – **B**.(Remember that the gas is lost from the flask after the weighed carbonate is added to the weighed acid.) [3]

(b) Plot a graph of mass of CO_2 / g against mass of X_2CO_3 / g.

Draw the line of best fit.



(c) Circle **one** clearly anomalous point on your graph. By reference to the description of the experiment, suggest an explanation for this anomaly.

(d) **(i)** Clearly mark two points on your line of best fit suitable to be used to determine the gradient of the line. Write the coordinates of the points you have selected below.
Use these points to determine the gradient of the line.

gradient = ……..…… [1]

(ii) Use the gradient you have calculated and the information below to calculate the M_r for X_2CO_3.

M_r for X_2CO_3 = ……..…… [1]

(f) Explain how each of the following changes to the experiment would affect the gradient of the graph obtained from the experiment.

(i) The experiment is carried out at a higher temperature.

(ii) The experiment is carried out with a different carbonate, $\mathbf{Y}_2CO_3$, which has a lower M_r than $\mathbf{X}_2CO_3$.

(g) Use your knowledge of acid-base chemistry to suggest a more appropriate way by which the M_r of a soluble metal carbonate might be determined.

[Data Processing] ; Gravimetric Analysis Practical 4

4. Determination of formula of a hydrated basic carbonate

The approximate formula for a hydrated basic carbonate is $CuCO_3.Cu(OH)_2.H_2O$. When heated, the basic copper(II) carbonate decomposes.

$$CuCO_3.Cu(OH)_2.H_2O(s) \rightarrow 2CuO(s) + CO_2(g) + 2H_2O(g)$$

FA 1 is powdered basic copper(II) carbonate, a hydrated mixture of copper(II) carbonate and copper(II) hydroxide.

Procedure

Record all weighings in an appropriate format in the space below.

1. Weigh and record the mass of an empty boiling-tube.
2. Tip the contents of the tube labelled **FA 1** into the weighed boiling-tube. Reweigh and record the total mass of the boiling-tube and **FA 1**.
3. Heat **FA 1** in the boiling-tube **very gently** until the vigorous decomposition of the copper carbonate has stopped; then heat more strongly for 1 to 2 minutes. **Take care not to lose any solid from the tube during the initial heating.**
4. Warm the upper parts of the boiling-tube to remove any water that may have condensed while heating the carbonate.
5. Place the hot tube on a heat-proof mat and leave to cool.
6. Reweigh the boiling-tube and the residual copper(II) oxide.
7. Repeat the heating, cooling and weighing process until a constant mass of ±0.010 g is achieved.

Results

Mass of boiling tube and content after 1st heating /g Mass of boiling tube and content after 2nd heating /g Mass of boiling tube and content after 3rd heating /g Mass of boiling tube and content after 4th heating /g	6.951 6.912 6.869 6.860
Mass of empty boiling tube /g	6.631
Mass of boiling tube + **FA 1** / g	6.984
Mass of **FA 1** / g	
Mass of residual copper oxide /g	

(a) Calculate the loss in mass during the experiment as a percentage of the mass of solid heated.

percentage loss in mass = ……..…… [1]

(b) The theoretical loss in mass is 33.5 %.

The proportions of $CuCO_3$ and $Cu(OH)_2$ in the basic carbonate can vary from the 1:1 ratio given in the formula.

Make use of the following information to account for the difference between the value you have calculated in **(a)** and the theoretical percentage loss in mass.

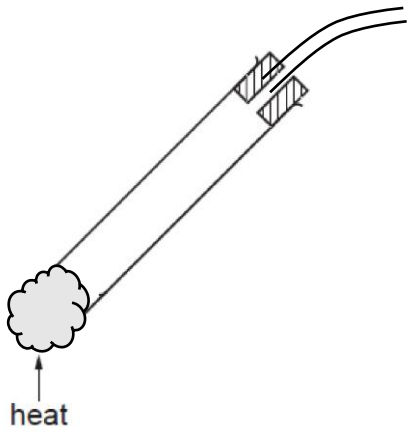
1 mol $CuCO_3(s) \rightarrow$ 1 mol $CO_2(g)$

1 mol $Cu(OH)_2(s) \rightarrow$ 1 mol $H_2O(g)$

Assume that 1 mol of any sample of the solid basic carbonate contains 1 mol H_2O.
[*M*r: CO_2, 44.0; H_2O, 18.0]

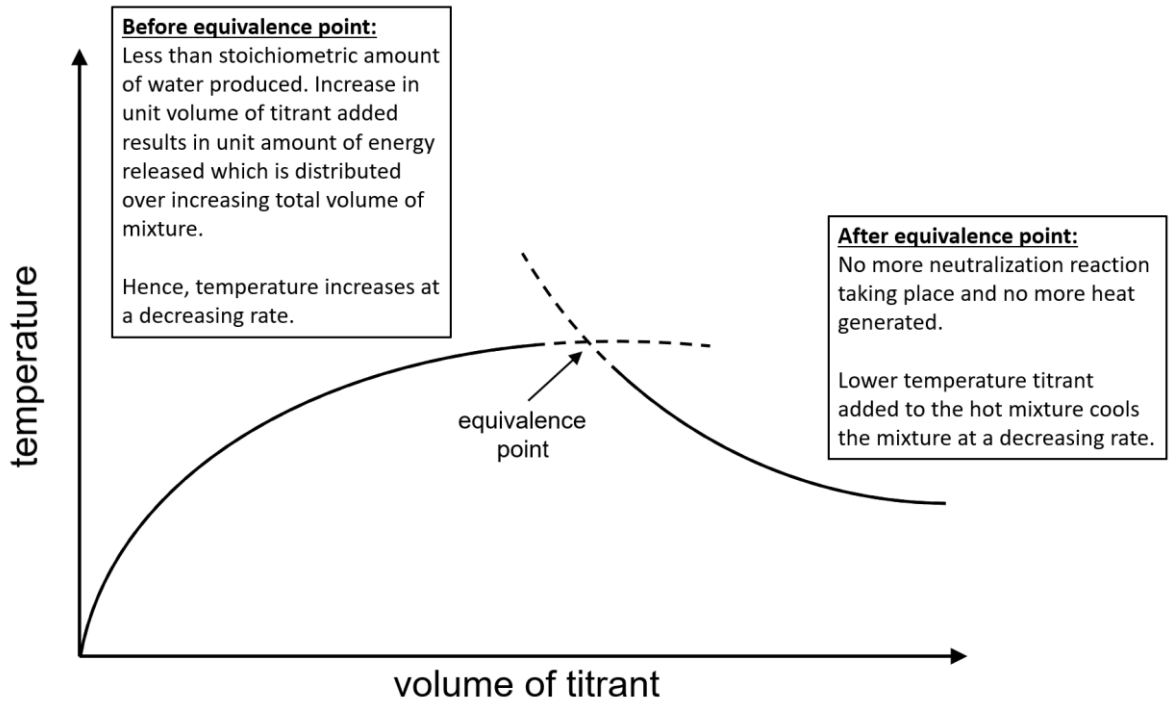
(c) Add to the diagram below additional standard laboratory apparatus that would enable you to collect and measure the volume of carbon dioxide evolved in the experiment.

Ensure that your apparatus does not also collect and measure any of the water vapour evolved.



[2]

Typical graph of a thermometric titration (eg. Neutralization reaction) is as follows:



Data Processing

[Data Processing] : Thermochemistry Practical 1

1. Determination of the enthalpy change of combustion, $\Delta H_{combustion}$, of a fuel.

In this experiment, you are required to determine the enthalpy change of combustion of a fuel **A**, by first calibrating the calorimeter using a tea-light candle.

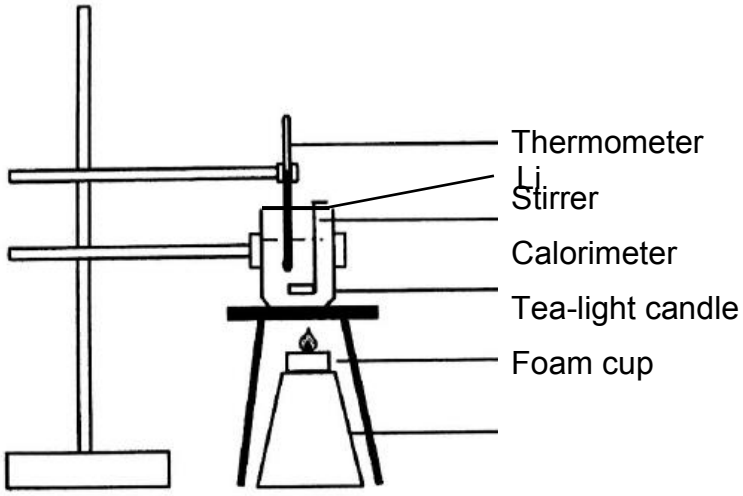
The main compound that is burnt during combustion of tea-light candles is paraffin wax. The paraffin wax or paraffinic hydrocarbons used usually consists of straight chained or branched alkanes with formula C_nH_{2n+2} where n ranges from 18 to 30.

In the first set up, the heat evolved from the combustion of paraffin wax is used to heat up a fixed volume of water.

In the second set-up, an unknown fuel **A** undergoes combustion in a fuel lamp. This process is exothermic and heats up the water in the calorimeter. By measuring the temperature rise and using the heat capacity of the calorimeter, the enthalpy change of combustion of fuel **A** can be calculated.

(a) Calibration of calorimeter using tea-light candles

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



2. Measure the mass of the candle and record its mass.
3. Fill the calorimeter with 100 cm^3 of water using a 100 cm^3 measuring cylinder. Record the initial temperature of the water using a 1 ^oC division thermometer.
4. Light the wick of the candle and place it under the calorimeter. Stir the contents gently.
   *Note: The candle should be placed under the calorimeter such that the tip of the wick is approximately 2 cm from the base of the calorimeter.
5. Extinguish the flame when a rise of about 5 ^oC is reached and record the maximum temperature reached after the flame has been extinguished.
6. Allow the candle to cool sufficiently and reweigh the burnt candle, recording the new mass of the candle. Record your results in the table below.

Table 1: Calibration of calorimeter using tea-light candle

Mass of tea-light candle before combustion /g	12.950
Mass of tea-light candle after combustion /g	12.602
Mass of wax burnt /g	
Initial temperature of water /ºC	28.0
Maximum temperature reached by water / ºC	33.0
ΔT / ºC	

(b) Finding the enthalpy change of combustion of fuel A

7. For the same setup shown above, use the fuel lamp containing fuel **A** in place of tea-light candle.
8. Repeat steps **2** to **6** using the lamp.
9. Record your results in the table below.

Table 2: Combustion of fuel A

Mass of fuel lamp before combustion /g	171.387
Mass of fuel lamp after combustion /g	170.527
Mass of fuel burnt /g	
Initial temperature of water /ºC	28.0
Maximum temperature reached by water / ºC	33.5
ΔT / ºC	

(c) (i) Given that heat energy, q = CΔT, calculate C, the heat capacity of the calorimeter.

Relevant data:

- $\Delta H_{combustion}$ of paraffin wax (n = 22) = −11.0 kJ mol^{-1}
- M_r of fuel **A** = 46.0

heat capacity of calorimeter = …….. [1]

(ii) Calculate the enthalpy change of combustion of fuel **A**

enthalpy change of combustion of fuel **A** = …….. [1]

(d) Identify **two** sources of error in the procedure and measurement that affect the reliability of the results. Explain how each error affects the results obtained and suggest a way to reduce each error.

Error 1:

Error 2:

(e) Assuming the same mass of fuel **A** is used, explain the effect on ΔT measured if the following change is made to the procedure.

(i) The calorimeter is filled with 50 cm^3 of water instead of 100 cm^3.

(ii) The calorimeter is changed from a copper container to a glass beaker.

[Data Processing] ; Thermochemistry Practical 2

2. Determination of ΔH_{rxn} using Hess's Law

Copper is an unreactive metal that does not react directly with dilute acids. It is therefore necessary to find the enthalpy change of reaction for two reactions that do occur so that Hess' Law can be used to determine the enthalpy of reaction.

Reaction 1: $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

Reaction 2: $Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

In this experiment, you are required to determine the enthalpy change of reaction, ΔH, for the reaction shown below.

$$Cu(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2(g)$$

You are given the following:

FA 1 is 1.00 mol dm^{-3} sulfuric acid, H_2SO_4
FA 2 is magnesium powder, Mg in a stoppered tube
FA 3 is 1.00 mol dm^{-3} copper(II) sulfate, $CuSO_4$.
FA 4 and **FA 5** are magnesium powder, Mg in stoppered tubes

(a) Determining the enthalpy change for Reaction 1

Read through the method before you start any practical work and prepare suitable tables for your results.

1. Weigh the stoppered tube containing **FA 2**. Record the mass in an appropriate manner.
2. Support the plastic cup in the 250 cm^3 beaker.
3. Use a 50 cm^3 measuring cylinder to transfer 50.0 cm^3 of **FA 1** into the plastic cup.
4. Measure the temperature of **FA 1** in the plastic cup, using a 0.2 ^oC division thermometer and start the stop clock. Record this temperature as being the temperature at time = 0.
5. Measure, and record, the temperature of this **FA 1** every half minute for 2 minutes.
6. At time = 2½ minutes, add the **FA 2** to the acid and stir carefully to reduce acid spray.
7. Measure the temperature of the mixture in the cup at time = 3 minutes and then every half a minute up to time = 7 minutes. Continue stirring occasionally throughout this time.
8. Weigh the stoppered tube that had contained **FA 2**. Record the mass.
9. Calculate and record the mass of **FA 2** added to the sulfuric acid.
10. Rinse the plastic cup with water and shake to dry.

Results

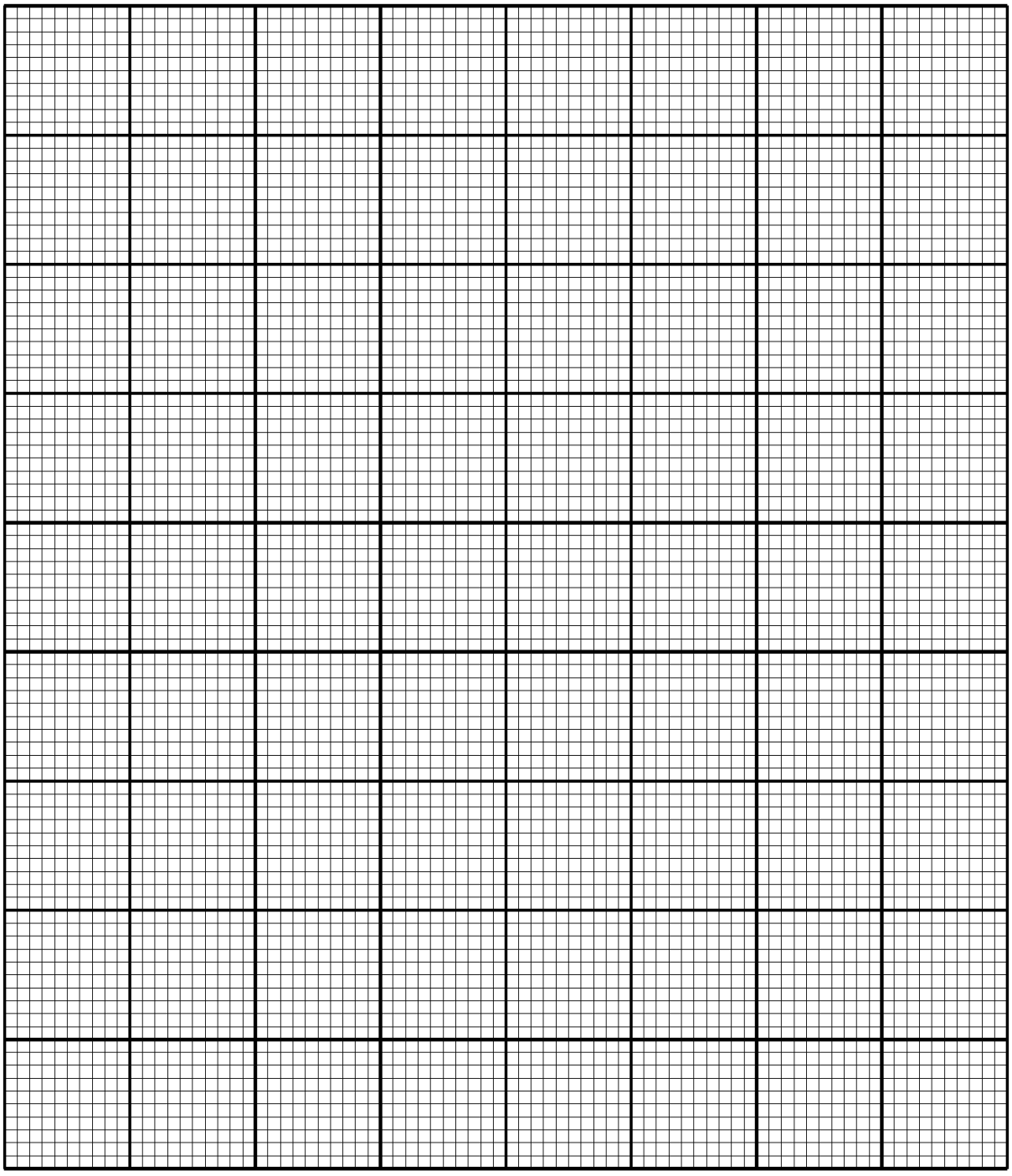
Table 1: Mass of FA2

Mass of stoppered tube and **FA 2** /g	10.655
Mass of stoppered tube and residual **FA 2** /g	10.548
Mass of **FA 2** used /g	

Table 2: Temperature recordings

Time/ min	Temperature/ ºC	Time/ min	Temperature/ ºC
0.0	29.5	**4.0**	37.3
0.5	29.5	**4.5**	36.8
1.0	29.5	**5.0**	36.0
1.5	29.5	**5.5**	35.5
2.0	29.5	**6.0**	34.9
2.5	-	**6.5**	34.1
3.0	38.5	**7.0**	33.7
3.5	37.8		

(i) On the grid below plot a graph of temperature (*y*-axis) against time (*x*-axis).



[3]

(ii) Complete the graph by inserting **two**, **straight** lines of best fit:
- one to show the temperature up to time = 2½ minutes,
- one to show the temperature after time = 2½ minutes

(iii) From your graph, use the two straight lines of best fit to calculate the change in temperature at time = 2½ minutes.

temperature change = ... [1]

(b) **(i)** In the reaction in **(a)**, sulfuric acid was in excess. Without carrying out any additional tests, what observation could you have made during your experiment to confirm this?

(ii) Calculate the heat change that occurred during the reaction in **(a)**.
[Assume that 4.2 J is needed to raise the temperature of 1.0 cm^3 of solution by 1.0 °C.]

heat change = …….. [1]

(iii) Use your answer to **(ii)** to calculate the enthalpy change, in kJ mol^{-1}, for the reaction between sulfuric acid and magnesium.
[*A*r : Mg, 24.3]

Reaction 1: $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

enthalpy change = …….. [2]

(c) Determining the enthalpy change for Reaction 2

Read through the method before you start any practical work and prepare suitable tables for your results.

1. Weigh the stoppered tube containing **FA 4**. Record the mass in an appropriate format.
2. Support the plastic cup in the 250 cm^3 beaker.
3. Use a 50 cm^3 measuring cylinder to transfer 50.0 cm^3 of **FA 3** into the plastic cup.
4. Measure the temperature of **FA 3** in the plastic cup, using a 0.2 ^oC division thermometer and record the temperature.
5. Add the **FA 4** to the **FA 3** in the cup and stir the mixture constantly.
6. Measure and record the maximum temperature reached during the reaction.
7. Calculate and record the maximum temperature change that occurred during the reaction.
8. Weigh the stoppered tube that had contained **FA 4**. Record the mass.
9. Calculate and record the mass of **FA 4** added to the copper(II) sulfate.
10. Discard the contents of the plastic cup
11. Rinse the plastic cup and shake to dry.
12. Repeat this experiment using **FA 5** in place of **FA 4**.

Table 3: Reaction 2 - experiment 1

Mass of stoppered tube and **FA 4** /g	10.895
Mass of stoppered tube and residual **FA 4** /g	10.694
Mass of **FA 4** used /g	
Initial temperature of **FA 3**/ ºC	29.7
Maximum temperature reached/ ºC	37.3
Change in temperature/ ºC	

Table 4: Reaction 2 - experiment 2

Mass of stoppered tube and **FA 5** /g	10.764
Mass of stoppered tube and residual **FA 5** /g	10.567
Mass of **FA 5** used /g	
Initial temperature of **FA 3**/ ºC	29.6
Maximum temperature reached/ ºC	36.8
Change in temperature/ ºC	

(d) (i) Using data from **experiment 1**, show, using a suitable calculation, that the copper(II) sulfate was in excess in these reactions.

[1]

(ii) Using your results, calculate the enthalpy change, in kJ mol^{-1}, for the reaction in experiment 1 and 2 respectively. Hence calculate the average enthalpy change, in kJ mol^{-1}, for the reaction between magnesium and copper(II) sulfate.
[Assume that 4.2 J is needed to raise the temperature of 1.0 cm^3 of solution by 1.0 °C.]
[*A*r : Mg, 24.3]

average enthalpy change = …… [3]

(e) **Reaction 3:** $Cu(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2(g)$

Use your values for the enthalpy changes for **Reactions 1** and **2** to calculate the enthalpy change for **Reaction 3**.

Reaction 1: $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
Reaction 2: $Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$

Show clearly how you obtained your answer.

enthalpy change = …… [1]

(f) **(i)** The method you used to determine the enthalpy change for **Reaction 1** was more accurate than the method you used to determine the enthalpy change for **Reaction 2**.

Suggest **two** reasons why the method used for **Reaction 2** was less accurate. Explain your answers.

(ii) A student suggested that the accuracy of the method used for **Reaction 2** could be improved by using a larger volume of copper (II) sulfate. Explain if you agree with the student's suggestion. Give a reason for your answer.

[Data Processing] ; Thermochemistry Practical 3

3. Determination of ΔH_{rxn} of decomposition of H_2O_2

Hydrogen peroxide decomposes into water and oxygen. This is a slow spontaneous process that can be sped up using a catalyst such as manganese(IV) oxide.

$H_2O_2(aq) \rightarrow H_2O(l) + ½\ O_2(g)$

In this experiment, you are required to determine the enthalpy change, ΔH, for the catalytic decomposition of hydrogen peroxide.

You are given the following:

FA 1 is 0.10 mol dm^{-3} hydrogen peroxide, H_2O_2.
FA 2 is manganese(IV) oxide, MnO_2, the catalyst for the decomposition.

Procedure

Read through the method before starting any practical work and prepare a table for your results in the space below.

Experiment 1

1. Support the plastic cup inside a 250 cm^3 beaker.
2. Use a 50 cm^3 measuring cylinder to transfer 40.0 cm^3 of distilled water into the plastic cup.
3. Use another 50 cm^3 measuring cylinder to add 10.0 cm^3 of **FA 1** into the plastic cup.
4. Measure and record the initial temperature of the mixture using a 0.2 ^oC division thermometer.
5. Add a spatula measure of **FA 2** to the mixture in the plastic cup.
6. Stir constantly until the maximum temperature is reached and record this temperature.
7. Calculate and record the temperature rise.
8. Wash and wipe out your plastic cup and rinse the thermometer.

Experiment 2

Repeat steps 1 to 6 using 30.0 cm^3 of distilled water and 20.0 cm^3 of **FA 1.**

Experiment 3

Repeat steps 1 to 6 using 20.0 cm^3 of distilled water and 30.0 cm^3 of **FA 1.**

Experiment 4

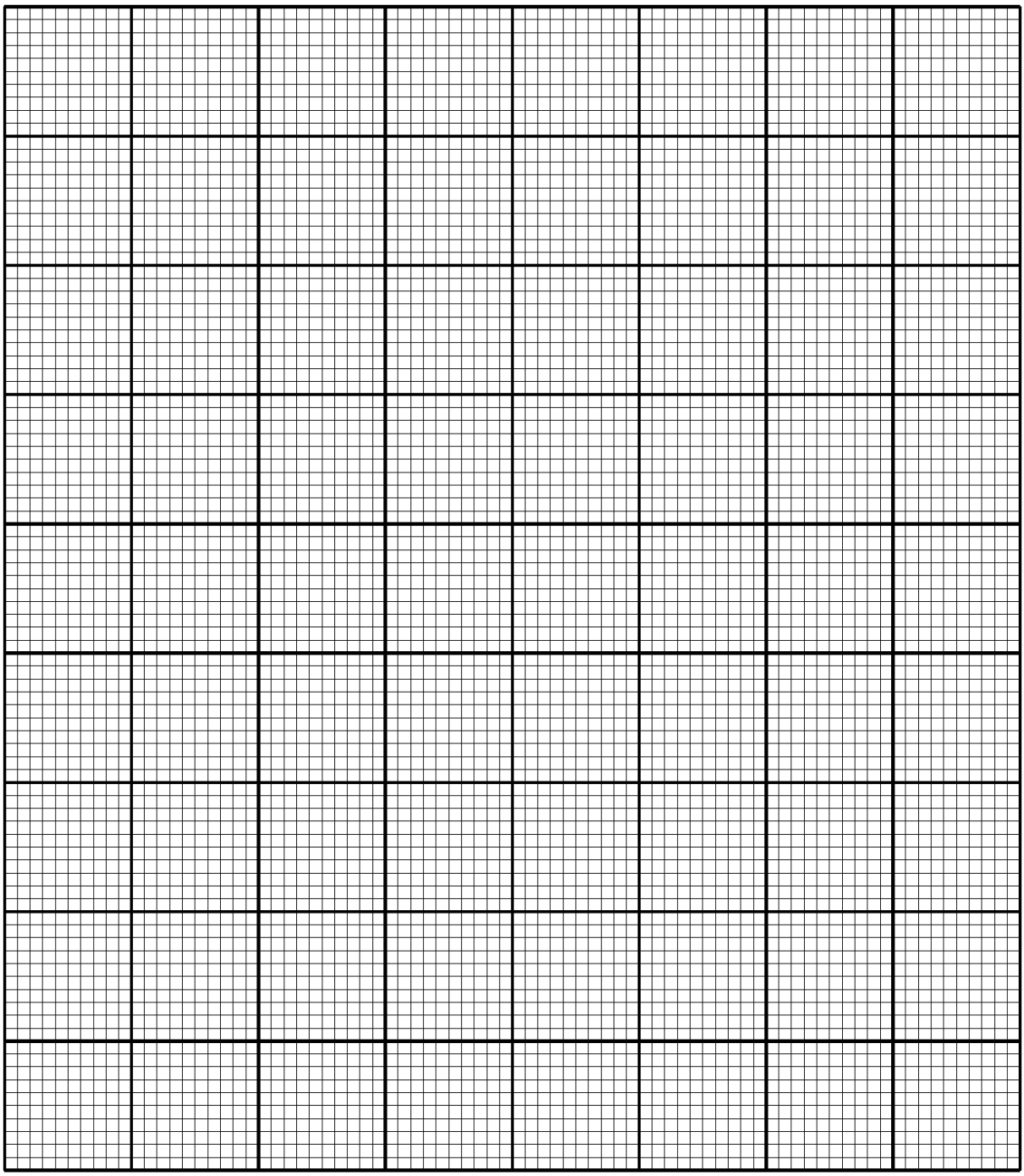
Repeat steps 1 to 6 using 10.0 cm^3 of distilled water and 40.0 cm^3 of **FA 1.**

Results

Expt	Vol of FA 1/ cm^3	Vol of distilled water/ cm^3	Initial T/ ºC	Final T/ ºC	T rise/ ºC
1	10.0		30.1	31.0	
2	20.0		30.1	31.8	
3	30.0		30.1	32.6	
4	40.0		30.1	33.5	

(a) (i) Using the grid below, plot a graph of the temperature rise (*y*-axis) against the volume of **FA 1** (*x*-axis).

Draw the line of best fit.



[4]

(b) (i) Use your graph to extrapolate the temperature rise for 45.0 cm^3 of **FA 1** used. Show your working clearly on the graph.

temperature rise = ……. [1]

(ii) Calculate the energy released when 45.0 cm^3 of **FA 1** is used.
(Assume that 4.2 J are needed to raise the temperature of 1.0 cm^3 of solution by 1.0 °C.)

energy released = …… [1]

(iii) Hence calculate the enthalpy change, in kJ mol^{-1}, for the reaction below.

$$H_2O_2(aq) \rightarrow H_2O(l) + 1/2\ O_2(g)$$

enthalpy change = …… [1]

(c) Which **one** of the four experiments that you carried out is likely to be the least accurate? Explain your choice.

(d) **(i)** Explain what the gradient of your graph represents.

(ii) Hence suggest the effect on the gradient of the graph, if **FA 2** is not added in step **5**.

[Data Processing] ; Thermochemistry Practical 4

4. Determination of ΔH_{sol} of NH_4Cl

When an exothermic reaction takes place in a container such as a beaker, some of the evolved heat energy is absorbed by the beaker.

When an endothermic reaction takes place some of the required heat energy is supplied by the beaker.

The amount of heat energy evolved or supplied for a 1°C change in temperature is known as the heat capacity of the beaker.

In this experiment, three reactions are involved.

NH_4Cl (s) + NaOH (aq) → NH_3 (aq) + $NaCl$ (aq) + H_2O (l) $\quad \Delta H_{reaction1}$
NH_4Cl (s) + aq → NH_4Cl (aq) $\quad \Delta H_{solution} = +14\ kJmol^{-1}$
NH_4Cl (aq) + NaOH(aq) → NH_3 (aq) + $NaCl$ (aq) + $H_2O(l)$ $\quad \Delta H_{reaction2}$

This experiment requires you to determine the following

- The heat capacity of the calorimeter.
- $\Delta H_{reaction1}$

$\Delta H_{reaction2}$ will be determined using Hess' law.

You are given the following:

FA 1 is NH_4Cl solid
FA 2 is 2.00 mol dm^{-3} sodium hydroxide, NaOH.

(a) Determining the approximate heat capacity of the 250 cm^3 beaker

When samples of hot and cold water are mixed in the 250 cm^3 beaker, some heat is lost to the beaker in raising its temperature. To determine the approximate heat capacity of your 250 cm^3 beaker, you will determine the maximum temperature rise when a sample of hot water is added to tap water in the beaker.

1. Use a 50 cm^3 measuring cylinder to transfer 50.0 cm^3 of tap water into the 250 cm^3 beaker.
2. Use the 50 cm^3 measuring cylinder to transfer 50.0 cm^3 of tap water into a 100 cm^3 beaker. Note the temperature of the water in this 100 cm^3 beaker.
3. Heat the content of 100 cm^3 beaker **carefully and gently** until the temperature of the water has increased by 45–50 °C then stop heating. *e.g. if the water is at 20.0 °C, you should warm it to 65–70 °C*.
4. Stir the tap water in the 250 cm^3 beaker with the thermometer.
5. Record the temperature of the tap water (this is the temperature at t = 0 min).
6. Record the temperature each minute for 3 minutes.
7. After you have taken the reading at t = 3 min, use the thermometer to stir the hot water in the 100 cm^3 beaker.
8. At t = 4 min, measure the temperature of the hot water and record this value in the blank below.

9. **Immediately** add the hot water from the 100 cm^3 beaker to the tap water in the 250 cm^3 beaker. Stir with the thermometer but do **not** record the temperature.
10. Continue to stir the water throughout the experiment.
11. Record the temperature at t = 5 min, and then every **½ minute** until t = 8 min.
12. Empty and rinse the 250 cm^3 beaker. Dry it using a paper towel.
13. Record all measurements of time and temperature obtained.

The temperature, T_1, of the hot water at t = 4 min is 95.0 °C

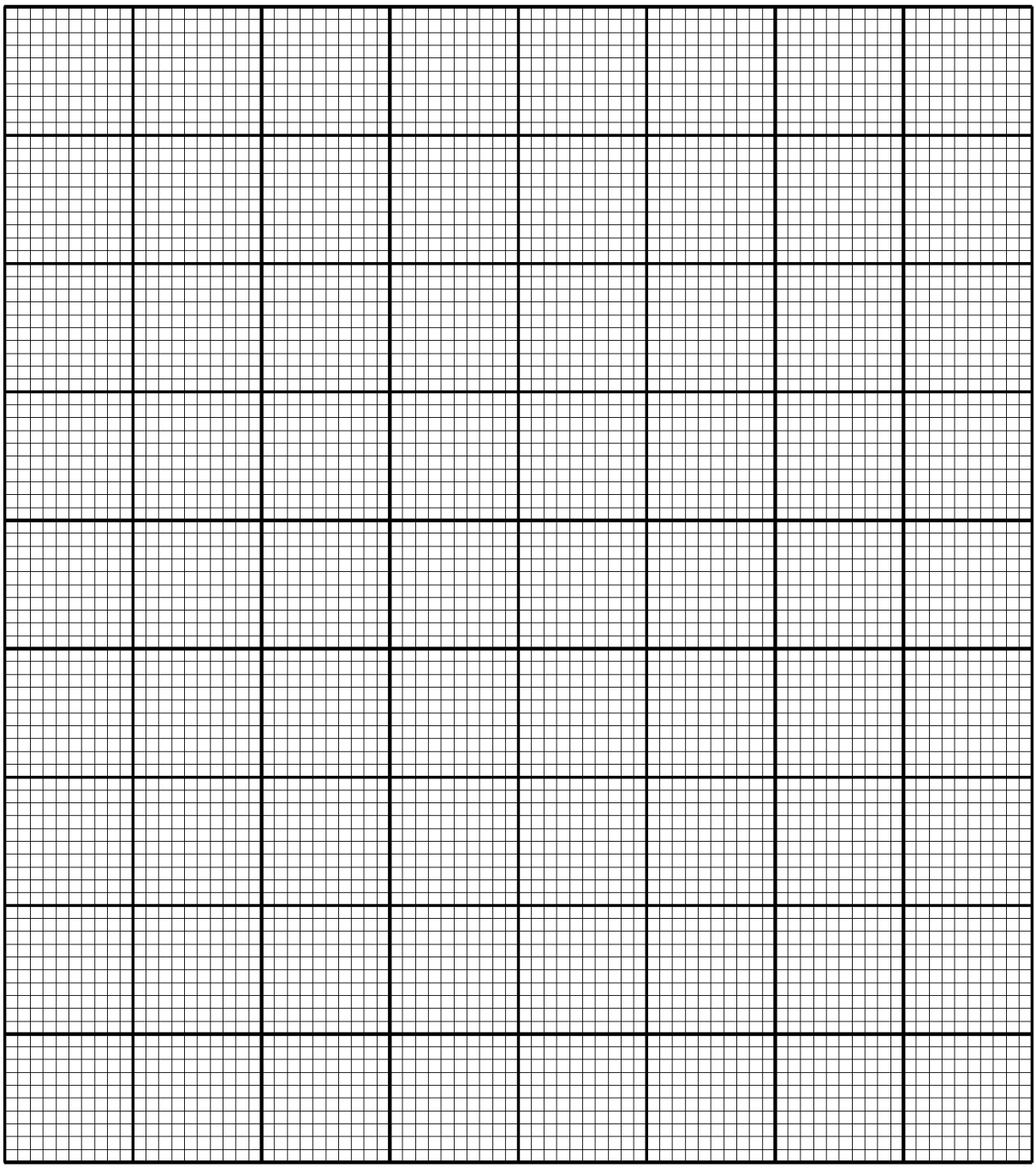
Results

Temperature of tap water in 100 cm^3 beaker : 30.2 °C

Time/ min	Temperature/ °C	Time/ min	Temperature/ °C
0.0	30.2	**5.5**	55.0
1.0	30.2	**6.0**	53.2
2.0	30.2	**6.5**	51.6
3.0	30.2	**7.0**	49.8
4.0	-	**7.5**	48.2
5.0	56.6	**8.0**	46.5

(b) Graph plotting

1. Plot a graph of the temperature of the water in the 250 cm^3 beaker (y-axis) against time (x-axis) on the grid below. Do **not** plot the temperature, T_1, of the hot water at t = 4 min.
2. Draw two straight lines of best fit; one through the points up to t = 3 min; the second through the points from t = 5 min to t = 8 min. Extrapolate both lines to t = 4 min.
3. From the extrapolated lines read the minimum and the maximum temperatures at t = 4 min. Record these values in the spaces provided below.
4. Determine the values for the two temperature changes at t = 4 min.



Minimum temperature, T_2, at t = 4 min is ..... °C.

Maximum temperature, T_3, at t = 4 min is ..... °C.

Temperature rise for 50.0 cm^3 of tap water in the 250 cm^3 beaker, ($T_3 - T_2$) is ..... °C.

Temperature fall for 50.0 cm^3 of hot water from the 100 cm^3 beaker, ($T_1 - T_3$) is ..... °C.

(c) **(i)** Calculate the heat energy gained by the 50.0 cm^3 of tap water in the 250 cm^3 beaker.
[**4.2 J** are absorbed or released when the temperature of 1.0 cm^3 of **water** changes by 1.0 °C.]

heat gained = …… [1]

(ii) Calculate the heat energy lost by the 50.0 cm^3 of hot water from the 100 cm^3 beaker.

heat lost = …… [1]

(iii) The difference between the values calculated in **(i)** and **(ii)** is an approximate value for the total heat energy absorbed by the 250 cm^3 beaker during the experiment.

The heat capacity of the beaker is the amount of heat energy absorbed for a 1 °C change in temperature.

approximate heat capacity of the 250 cm^3 beaker

$$= \frac{(\text{heat energy lost}) - (\text{heat energy gained})}{T_3 - T_2} \text{ J °C}^{-1}$$

Use your answers to **(i)** and **(ii)** and the temperature rise from **(b)** to calculate the approximate heat capacity of the 250 cm^3 beaker.

heat capacity of the 250 cm^3 beaker = …… [1]

(d) Determining the enthalpy change of solution for ammonium chloride

Record all your mass and temperature measurements in the table below.

1. Weigh the stoppered tube containing ammonium chloride, which is labelled **NH_4Cl**.
2. Use the 100 cm^3 measuring cylinder to transfer 100.0 cm^3 of tap water into the rinsed and dried 250 cm^3 beaker used in **(a)**.
3. Stir the water in the beaker with the thermometer and record the temperature.
4. Add the solid from the weighed tube to the water. Stir the mixture constantly with the thermometer.
5. Record the minimum temperature obtained in the solution.
6. Reweigh the tube labelled **NH_4Cl**, its stopper and any residual ammonium chloride.
7. Empty and rinse the beaker and dry it using a paper towel.

Results

mass of tube + ammonium chloride / g	10.007
mass of tube + residual ammonium chloride / g	5.012
mass of ammonium chloride / g	
initial temperature of water / °C	30.1
minimum temperature obtained / °C	27.7
temperature fall, **ΔT** / °C	

(e) (i) Use the temperature fall from the experiment, to calculate the change in heat energy of the solution.
[**4.2 J** are absorbed or released when the temperature of 1.0 cm^3 of **solution** changes by 1.0 °C.]

heat change = …….. [1]

(ii) To calculate the total change in heat energy as ammonium chloride dissolves in water, the change in heat energy of the 250 cm^3 beaker has to be added to the change in heat energy of the solution.

Explain why these two changes in heat energy have to be added together.

(iii) Use your answer in **(i)** above and the approximate heat capacity of the 250 cm^3 beaker calculated in **(c)(iii)** to calculate the combined change in heat energy of the beaker and solution.

change in heat energy = …… [1]

(iv) Calculate how many moles of **FA 1**, NH_4Cl, were used in the experiment. [*A*r; C*l*, 35.5; H, 1.0; N, 14.0]

moles of **FA 1** = …… [1]

(v) Calculate the enthalpy change of solution, $\Delta H_{solution}$ (NH_4Cl).

$\Delta H_{solution}$ (NH_4Cl) = …… [1]

(vi) Explain the significance of the sign you have given in **(v)** and how it is related to your experimental results.

(f) **(i)** A data book value for the molar enthalpy change of solution, $\Delta H_{solution}$ (NH_4Cl), is +15.2 kJ mol^{-1}. The value you have obtained may be significantly different from this value.

Calculate the difference between your value of $\Delta H_{solution}$ (NH_4Cl) and the data book value. Record this difference below. Express this difference as a percentage of the data book value.

percentage difference = ……..[1]

(ii) Describe **one** major source of error in this experiment. Suggest an improvement which would significantly increase the accuracy of the experiment. Explain why your suggestion would produce a more accurate value.

Data Processing

[Data Processing] : *Kinetics Practical 1*

1. Determination of the relationship between rate of decomposition of H_2O_2 and its concentration.

Two experiments on the decomposition of H_2O_2 in the presence of a catalyst, I^- will be carried out.

The reactions involved are:

$$H_2O_2\ (aq) + I^-\ (aq) \rightarrow IO^-\ (aq) + H_2O\ (l)$$
$$2IO^-\ (aq) \rightarrow 2I^-\ (aq) + O_2\ (g)$$

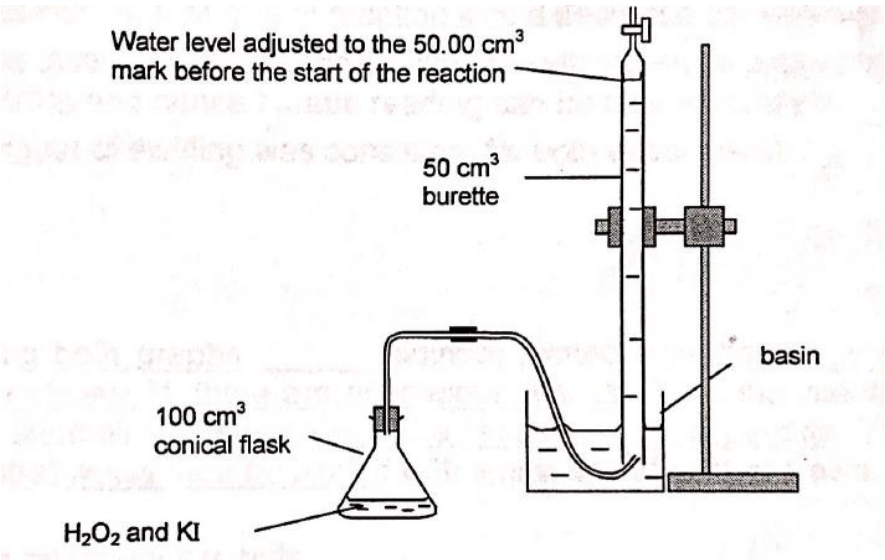
Overall: $2H_2O_2\ (aq) \rightarrow 2H_2O\ (l) + O_2\ (g)$

This reaction can be followed by measuring the volume of oxygen gas evolved as the reaction proceeds.

You are provided with the following:
Aqueous hydrogen peroxide
Aqueous potassium iodide

Procedure:

1. Fill a burette with water. By holding a thumb tightly over the mouth of the burette, invert it over a beaker of water.
2. By controlling the tap of the burette, slowly allow some air to enter until water level corresponds to the 50.00 cm^3 mark. (or until you are able to read off a volume from the burette)
3. Record the initial burette reading.
4. Place the rubber tubing such that the gas evolved will be collected in the burette via downward displacement of water as shown in the diagram.



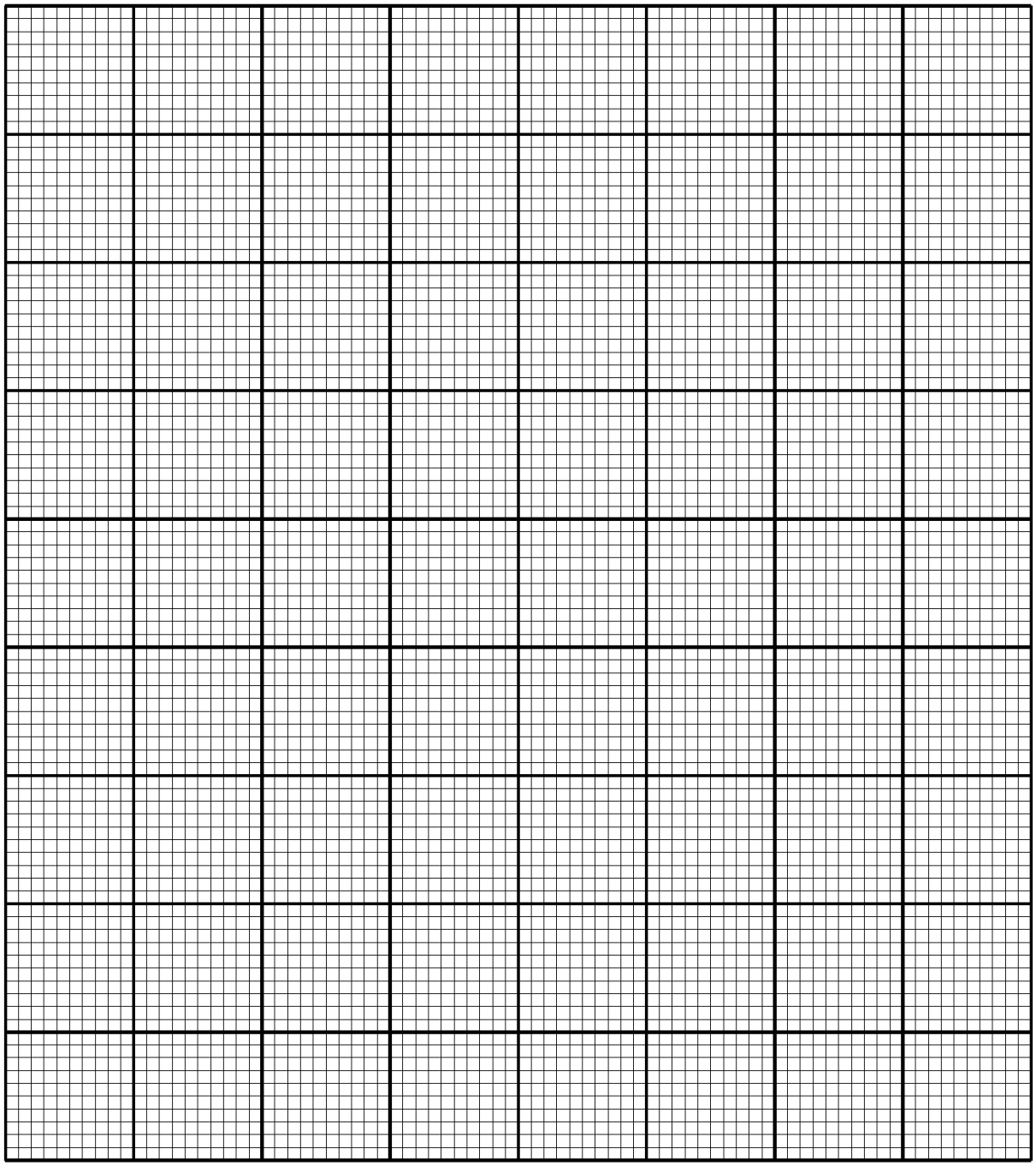
5. Pipette 20.0 cm^3 of hydrogen peroxide into the conical flask.
6. Measure 10 cm^3 of potassium iodide in a measuring cylinder.
7. Remove the stopper from the conical flask. Start the stopwatch the instant when potassium iodide is added to the conical flask and immediately replace the stopper.
8. Swirl the conical flask once immediately to ensure even mixing.
9. Do not swirl the conical flask throughout the experiment.
10. Record the burette reading every 30 seconds for the first 3 minutes and continue to record for every 1 minute for a total duration of 8 minutes.
11. For the second experiment, rinse out the conical flask and add 10 cm^3 of distilled water into a second measuring cylinder.
12. Pipette 10.0 cm^3 of hydrogen peroxide into the conical flask and add 10 cm^3 of distilled water from the second measuring cylinder.
13. Repeat steps 6 to 9.

Results

Time/s	Expt 2 10.0 cm^3 H_2O_2	Expt 1 20.0 cm^3 H_2O_2
	Burette reading/ cm^3	Burette reading/ cm^3
0	50.00	50.00
30	46.80	43.00
60	43.70	39.20
90	40.90	34.90
120	38.70	31.60
150	36.40	28.50
180	34.00	25.70
240	29.80	21.10
300	26.90	17.70
360	24.90	15.40
420	23.60	13.70
480	22.30	12.10

(a) For each of the experiment, plot a graph of burette reading against time on the same graph grid provided.

The rate of this reaction can be investigated by analysing the graph of burette reading against time.



[4]

(i) Using the graph plotted, calculate the initial relative rate of the two graphs. Show your working clearly.

(ii) State and explain the relationship between the volume of H_2O_2 (aq), V_{H2O2} and the initial concentration of H_2O_2, $[H_2O_2]_{mixture}$ in the reaction mixture.

(iii) By comparing the gradient of the two graphs that correspond to different initial concentrations of H_2O_2, deduce the order of the reaction with respect to H_2O_2.

order with respect to H_2O_2 = …… [2]

(iv) Suggest how you can carry out a third experiment to determine the order of reaction with respect to I^-.

(b) Explain the importance of step **8** in the procedure.

[Data Processing] ***: Kinetics Practical 2***

The kinetics of the reaction between Na_2CO_3 and HC*l* is studied. The order of reaction with respect to HC*l* can be investigated using the following set-up by varying the concentration of HC*l*.

Method

(a) Results

1. 100.0 cm^3 of 4 mol dm^{-3} Na_2CO_3 is pipetted into one side of the divided flask.

2. 20.0 cm^3 of 5.0 mol dm^{-3} HC*l* is pipetted into the other side of the divided flask.

3. The apparatus is assembled as shown and the reaction started by swirling the flask to mix the solutions.

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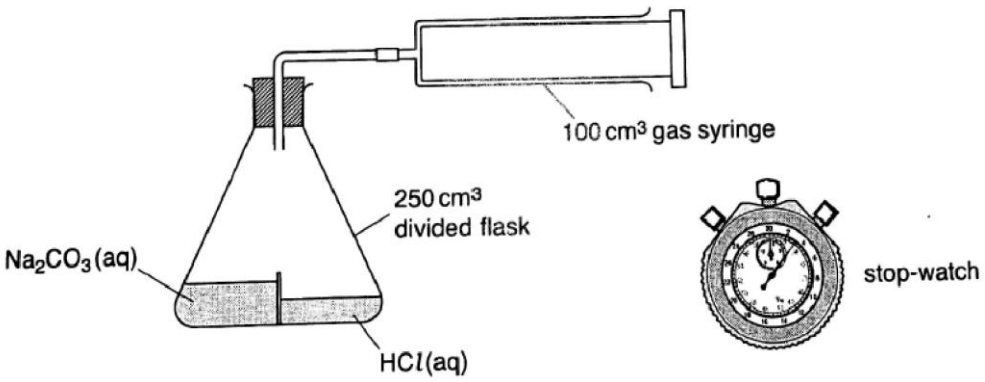


4. The stop watch is started at the moment the reaction commences.
5. The stop watch is stopped and the time recorded when 20 cm^3 of gas have been collected.
6. The apparatus is rinsed out and the experiment repeated using volumes of acid and water shown in **Table 1**.

The acid and water volumes are measured using a graduated pipette.

[3]

Results:

Complete the following table using the following expressions:

- **Initial rate of reaction =** $\frac{20}{time\ taken\ for\ gas\ collection}$

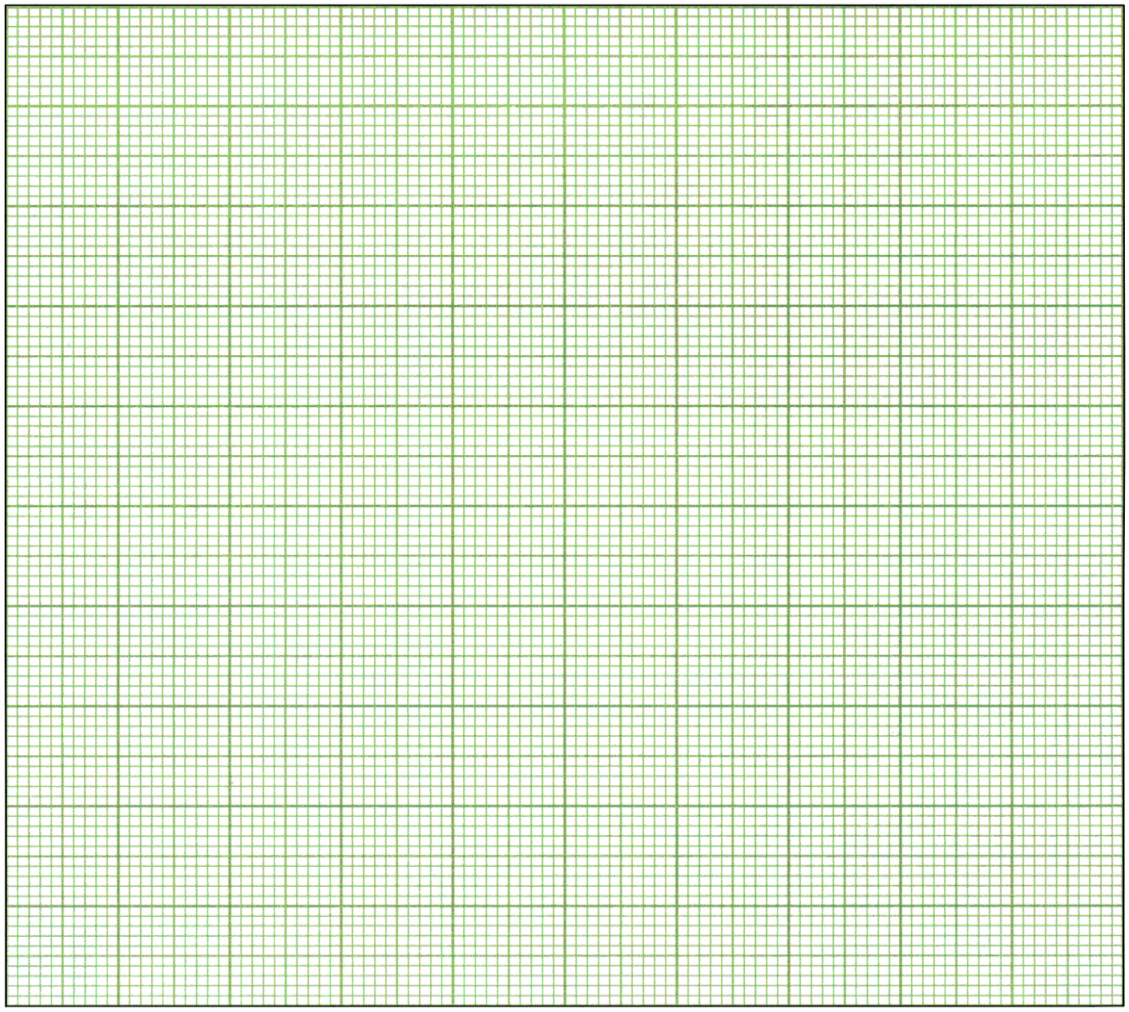
Relative [HC*l*] = $\frac{volume\ of\ acid\ in\ numbered\ experiment}{volume\ of\ acid\ in\ experiment\ 1}$

Record all values of initial rate of reaction and relative [HC*l*] to 2 significant figures.

Expt	Volume of 5.0 mol dm^{-3} HC*l* added / cm^3	Volume of water added / cm^3	Volume of gas collected / cm^3	Time taken to collect 20 cm^3 gas / s	Initial rate of reaction / $cm^3\ s^{-1}$	Relative [HC*l*]
1	20.0	0	20	8		
2	17.0	3.0	20	9		
3	14.0	6.0	20	12		
4	11.0	9.0	20	14		
5	8.0	12.0	20	20		
6	5.0	15.0	20	32		
7	2.0	18.0	20	80		

Table 1

(b) Plot the initial rate of reaction against the relative concentration of the acid. Start each scale at zero. Draw the line of best-fit through the points.



(c) What can you conclude from your plot about the order of reaction with respect to hydrochloric acid? Explain your answer.

(d) A student carrying out the experiment made a mistake and only pipetted 50 cm^3 rather than 100 cm^3 of 0.4 mol dm^{-3} sodium carbonate into the flask.
How will this affect the rate of reaction? Explain your answer.

2 Alternative manipulation of experimental results to determine order of reaction with respect to HC*l*.

Calculate the values of (volume of **HC*l*** × reaction time) in the table below.

Expt	Volume of 5.0 mol dm^{-3} HC*l* added / cm^3	Volume of water added / cm^3	Volume of gas collected / cm^3	Time taken to collect 20 cm^3 gas / s	vol $_{HCl}$ x t / cm^3 s
1	20.0	0	20	8	
2	17.0	3.0	20	9	
3	14.0	6.0	20	12	
4	11.0	9.0	20	14	
5	8.0	12.0	20	20	
6	5.0	15.0	20	32	
7	2.0	18.0	20	80	

(a) Explain how the values of (vol_{HCl} × t) of these experiments show that the order of reaction with respect to HC*l* is in agreement to your answer determined in **(c)** on pg 50**.**

(b) The 1st order of reaction with respect to HCl in this reaction can also be verified to by monitoring the volume of carbon dioxide gas during the course of reaction in experiment **7** of table 1.

In the space below, sketch the graph you would expect to obtain when the reaction is allowed to proceed to completion. You may assume that the total volume of carbon dioxide gas collected in the gas syringe is $\mathbf{V_{max}}$.

Explain your answer.

[Data Processing] ***: Kinetics Practical 3***

2. To investigate the reaction between substance X, iodine and hydrogen ions.

You are provided with the following:

FA 4 is 1.00 mol dm^{-3} sulfuric acid
FA 5 is an aqueous solution of substance **X**
FA 6 is 0.00375 mol dm^{-3} iodine, I_2

Procedure:

1. Using a measuring cylinder, measure out 20 cm^3 of **FA 4** and 20 cm^3 of **FA 5**, as shown in column 1 of **Table 1**, into a conical flask. It is not necessary to rinse the measuring cylinder between solutions.
2. Measure out 4.00 cm^3 of **FA 6** from the burette into a test-tube.
3. Start the reaction by tipping the **FA 6** from the test-tube into the conical flask. Start the stop-clock with one hand and swirl the contents of the flask with the other. Place the flask on a white tile and stop the clock as soon as the yellow colour disappears.
4. Record the time in seconds (to the nearest second) in **Table 1**.
5. Repeat the experiment using the different volumes of **FA 4**, **FA 5** and **FA 6** as shown in **Table 1**. Where water is required, use the measuring cylinder to add the water to the other solutions in the conical flask.

The "rate of reaction" can be calculated by using the relationship:

$$\text{rate} \propto \frac{\text{volume of } \mathbf{FA\ 6} \text{ in cm}^3}{\text{time in seconds for colour to disappear}}$$

(a) Results

Calculate each rate and complete **Table 1**.

Table 1

Experiment	**1**	**2**	**3**	**4**
Volume of **FA 4** /cm^3	20.0	20.0	10.0	20.0
Volume of **FA 5** /cm^3	20.0	20.0	20.0	10.0
Volume of water /cm^3	0.0	2.0	10.0	10.0
Volume of **FA 6** /cm^3	4.0	2.0	4.0	4.0
Time for colour to disappear /s	24	24	48	48
"rate of reaction"				

[3]

(b) Explain the importance of adding 10 cm^3 of deionized water for **experiment 3** and **4** and the effect it would have on the dependent variable if not adhered to.

(c) **(i)** From **Table 1**, determine the order of reaction with respect to

- [**X**]

order of reaction with respect to [**X**] = ……

- $[H^+]$

order of reaction with respect to $[H^+]$ = ……

- $[I_2]$

order of reaction with respect to $[I_2]$ = ……

(ii) Hence, give the rate equation for the reaction and the units of the rate constant.

[2]

(d) Suggest a safety precaution in this experiment.

…….... [1]

[Data Processing] :*Kinetics Practical 4*

3. To investigate the effect of temperature on rate and E_a

This experiment investigates how the rate of the reaction between hydrochloric acid and sodium thiosulfate, $Na_2S_2O_3$, depends on the temperature and hence calculate a value for the activation energy of the following reaction.

$$S_2O_3^{2-}(aq) + 2H^+(aq) \rightarrow S(s) + SO_2(g) + H_2O(l)$$

Since the rate constant k is inversely proportional to time, t, the Arrhenius equation:

$$\lg k = \frac{-E_a}{RT} + C$$

may be written as $$\lg t = \frac{E_a}{2.303RT} + C$$

where E_a is activation energy, T is temperature in Kelvin and C is a constant.

Therefore, a plot of lg t against $\frac{1}{T}$ will be a straight line and the slope of the line is $\frac{E_a}{2.303R}$ where R = 8.31 $Jmol^{-1}K^{-1}$. Thus, the activation energy, E_a, for the reaction can be calcuated.

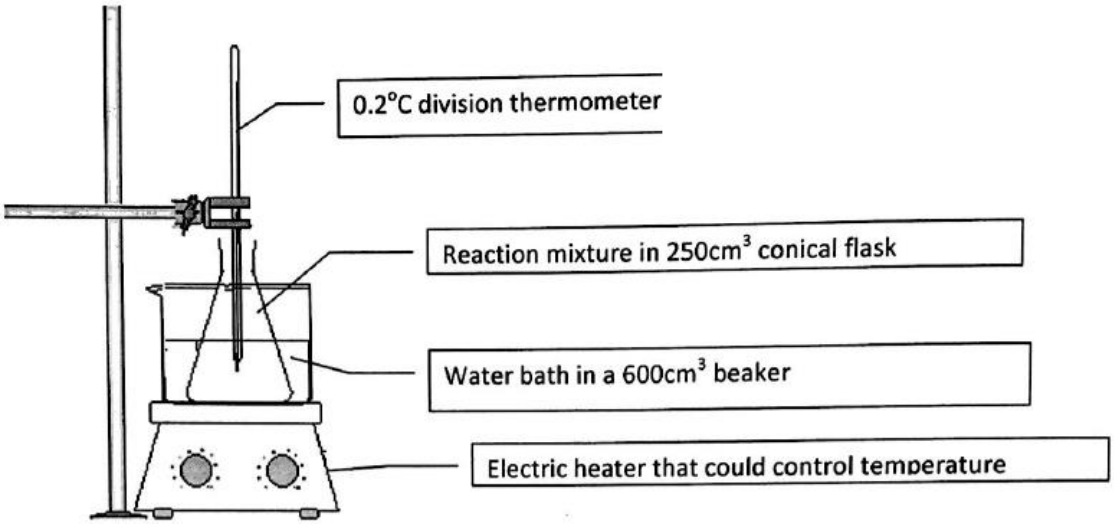
You are provided with the following:

0.100 mol dm^{-3} aqueous sodium thiosulfate
1.00 mol dm^{-3} hydrochloric acid

Procedure

1. Draw a large "X" mark using a marker on the surface of the hot plate.
2. Prepare a water bath using a 600 cm^3 beaker on the hotplate.
3. Add 100 cm^3 of 0.100 mol dm^{-3} $Na_2S_2O_3$ into a 250 cm^3 conical flask using a 100 cm^3 measuring cylinder.
4. Place the 250 cm^3 conical flask from step **3** into the water bath and carefully stir the solution until its temperature is between 70 ºC and 80 ºC.
5. Prepare 10 cm^3 of 1.00 mol dm^{-3} HC*l* using a 10 cm^3 measuring cylinder.
6. Pour the 10 cm^3 HC*l* solution rapidly into the conical flask containing $Na_2S_2O_3$ and start the stopwatch.
7. Stir the mixture using a glass rod and note the initial temperature, T_i of the reaction mixture using a 0.2 ºC division thermometer.
8. Stop the time when the "X" mark is completely obscured.
9. Note the final temperature T_f of the reaction mixture.
10. Note the time to the nearest second.
11. Wash the beaker and thermometer and dry them.
12. Repeat the experiment at temperatures ranging from 80 ºC to 40 ºC.

The experimental set-up is shown on the next page.



Results

Expt	T_i / ºC	T_f / ºC	T_{ave} / ºC	T_k / K	t / s		
1	42.0	48.0			180		
2	50.0	56.0			73		
3	64.0	68.0			16		
4	73.0	77.0			6		

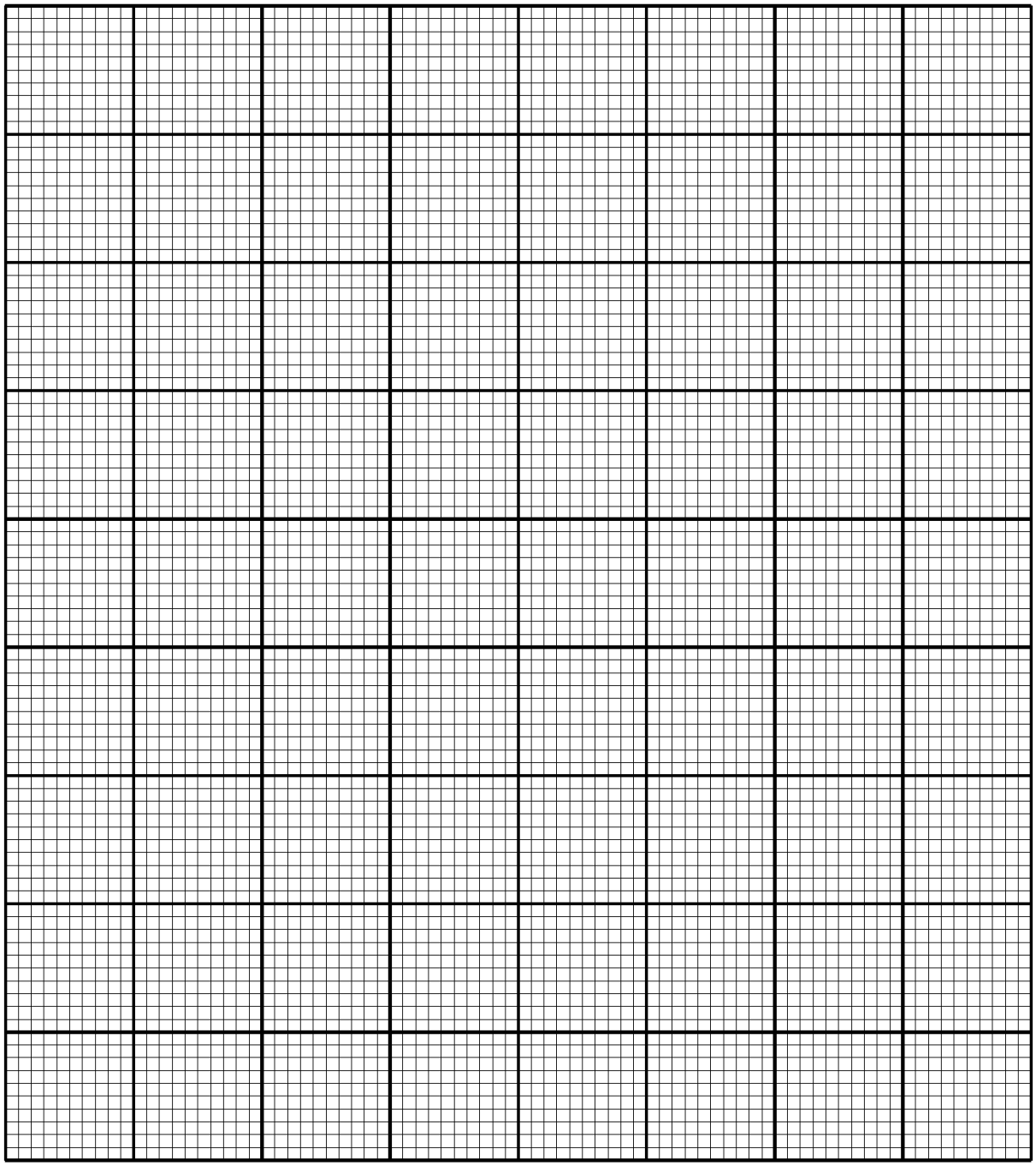
(a) Process the results in the table above to calculate both average temperature, T_{ave}, and temperature in Kelvin, T_K.

Include calculated values in the additional columns of the table that enable you to plot lg t against $\frac{1}{T_K}$. You may use some or all of these columns.

Label the columns you use clearly, including the units.

[3]

(b) Use the grid below to plot a graph of lg t against $\frac{1}{T_K}$.



[4]

(c) Determine the gradient of the graph plotted in **(b)** and hence calculate a value for the activation energy of the reaction.

activation energy = …….. [2]

(d) **(i)** Draw the Boltzmann distribution curve for the reaction between hydrochloric acid and sodium thiosulfate at 45 ^{o}C and 53 ^{o}C.

[2]

(ii) Hence, explain why the rate of reaction is faster at a higher temperature.

(e) **(i)** Calculate the percentage uncertainty for the measurement of hydrochloric acid using a 10 cm^3 measuring cylinder.

percentage error = …….. [1]

(ii) Suggest an improvement to reduce the percentage error obtained from **e(i)**.

(f) State a hazard that must be considered when planning the experiment and describe precautions that should be taken to keep risks to a minimum.

Data Processing

[Data Processing] : Inorganic *Qualitative Analysis Practical 1*

Instructions

For this experiment, you are to watch the video recordings on Google Classroom in order to record your observations.

1. At each stage of any test you are to record details of the following.
- colour changes seen
- the formation of any precipitate
- the solubility of such precipitates in an excess of the reagent added

Where reagents are selected for use in a test, the name or correct formula of the element or compound must be given.

(a) You are provided with a sample labelled **FA 5**. This sample **a mixture of 2 salts**, one of which contains Zn^{2+}. By carrying out the following tests you are able to suggest the identities of the ions present in **FA 5**.

Test	Observations
(i) Transfer the solid **FA 5** into a test tube and add a 3 cm depth of dilute nitric acid to make solution **FA 6**. Use this solution in tests **(ii)** to **(v)**. **Hint:** as 2 salts are present, 2 cations and 2 anions may be present. Both or either of the anions can produce gas with acid.	White solid dissolves in nitric acid to form a …… Gas evolved forms a white precipitate when passed into calcium hydroxide. Gas does not decolourise purple $KMnO_4$.
(ii) To a 1 cm depth of **FA 6** in a test-tube, add aqueous sodium hydroxide. **Hint:** Since **FA6** is a colourless solution, what are the other possible cations? See QA notes on pg **76-77**. Think of how you can distinguish them.	
(iii) To a 1 cm depth of **FA 6** in a test-tube, add aqueous ammonia.	
Test	**Observations**
(iv) To a 1 cm depth of **FA 6** in a test-tube, add a 1 cm depth of aqueous silver nitrate,	

(v) To a 1 cm depth of **FA 6** in a test-tube, add a 1 cm depth of aqueous barium chloride or barium nitrate.	

[5]

(vi) Suggest the identity of the cation (other than Zn^{2+}) and anions present in **FA 5**. Explain your choice.

Cation:	
Anions:	

[3]

(b) You are provided with solid **FA 7**, which is a **mixture of two salts**. Use separate portions of **FA 7** to perform the experiments below.

	Observation
(i) Heat a spatula measure of **FA 7** in a test-tube gently at first, then **Hint:** note the colour of the solid, may indicate presence of transition metal cation.	
heat more strongly and test the gas evolved with litmus paper. **Hint:** Decomposition of polyatomic ion in solid only applies to NH_4^+, CO_3^{2-} and NO_3^-.	
Dissolve a spatula measure of **FA 7** in a test-tube with a 5 cm depth of deionised water for use in tests **(ii)** and **(iii)**. **Hint:** note the colour of the solution formed, may indicate presence of transition metal cation.	
(ii) To a 1 cm depth of **FA 7**(aq) in a test-tube, add aqueous silver nitrate,	
followed by aqueous ammonia.	

(iii) To a 1 cm depth of **FA 7**(aq) add barium chloride or barium nitrate,	
followed by dilute nitric acid.	

[5]

(iv) Use your observations and QA notes on pg **76-77** to identify **three** of the ions present. Give evidence for your choice of ions.

Ion	Evidence

[3]

[Data Processing] : Inorganic *Qualitative Analysis Practical 2*

Instructions

For this experiment, you are to watch the video recording on Google Classroom in order to record your observations.

2. At each stage of any test you are to record details of the following.
- colour changes seen
- the formation of any precipitate
- the solubility of such precipitates in an excess of the reagent added

Where reagents are selected for use in a test, the name or correct formula of the element or compound must be given.

(a) Some cations interfere with tests for anions and have to be removed from the solution before the tests for anions present can be performed. One way in which this can be carried out is to precipitate the cation in the form of its insoluble carbonate.

Test	Observation	
	FA 3	**FA 5**
To 5 cm depth of solution in a test tube, add all of the sodium carbonate, Na_2CO_3, from one of the tubes provided. Stir the mixture.		

Retain the mixture from **FA 3** for use in **(b)**.

(b) Filter the mixture from **FA 3** from **(a)** into another test tube. Ignore any colour in the filtered solution. Add 5 cm depth of dilute nitric acid to the filtrate obtained. This removes any excess of carbonate ions.

Carry out the following tests on the acidified filtrate from **FA 3**.

Test	Observation
To 2 cm depth of the acidified filtrate from **FA 3** in a test-tube, add 1 cm depth of aqueous silver nitrate	
add an excess of aqueous ammonia.	

[1]

(c) Carry out the following test on **FA 4**.

Test	Observation
To 1 cm depth of **FA 4** in a test-tube, add 1 cm depth of **FA 3**. **Hint:** if the ppt is formed in a coloured solution, allow time for the solid/ ppt to settle at the bottom so that colour of the solution can be observed more accurately.	
followed by adding a few drops of starch.	

[1]

(d) Carry out the following tests.

	Observations		
	FA 3	**FA 4**	**FA 5**
To 1 cm depth of solution in a test-tube, add aqueous sodium hydroxide. Next, add aluminium foil and warm cautiously. Test for any gas with a moist red litmus paper			
To 1 cm depth of solution in a test-tube, add aqueous ammonia.			

[6]

(e) The results from tests in **(d)** should enable you to identify a cation in a solution.

solution	cation(s)
FA 3	
FA 4	
FA 5	

[2]

(f) By considering the results of **all** your tests, enter one of the following responses in each of the boxes below.

- chloride
- iodide
- NO_3^-
- NO_2^-
- Insufficient test

FA 3	
FA 4	
FA 5	

[2]

Note: colourless iodide, I^- is typically used to test for ions with oxidising property, a positive test is indicated by the brown solution ($I_2(aq)$)/ blue black coloration due to iodine-starch complex formed with starch indicator

(g) Where you were not able to identify a single anion present in **(f)**, with reference to QA notes on pg **76-77**, suggest a suitable reagent which would allow you to identify which anion is present in the solution from the list in **(f)**.

Do not carry out this test.

(h) State the expected observation for the test in **(g)** to identify the presence of each of the anions.

Anion 1 ……..... [1]

Anion 2 ……..... [1]

(i) When the test for **(a)** was repeated with Fe^{3+}, effervescence was observed. With the aid of two relevant equations with state symbols, explain why Fe^{3+} exhibits acidic properties and the gas produced upon addition of CO_3^{2-}(aq).

Data Processing

[Data Processing] : Organic *Qualitative Analysis Practical 1*

Instructions

For this experiment, you are to watch the video recording on Google Classroom in order to record your observations.

1. Before starting part **(a)** and **(b)**, half-fill a 250cm^3 beaker with water and heat with a Bunsen burner to approximately 60 ^{o}C. You will use this as a hot water bath.

(a) **FB 5**, **FB 6, FB 7** and **FB 8** are solutions each containing a single compound which could be ethanol, propanal, propanone and 1-bromobutane. To identify each compound you will react the samples with Tollens' reagent and with acidified potassium manganate(VII).

(i) **Preparation of Tollens' reagent**

- To approximately 2 cm depth of aqueous silver nitrate in a test tube, add approximately 0.5 cm depth of aqueous sodium hydroxide.
- Add aqueous ammonia a little at a time with continuous shaking until the brown precipitate **just** dissolves. Do **not** add an excess of ammonia.

Test	*observations*			
	FB 5	**FB 6**	**FB 7**	**FB 8**
To a 1 cm depth of each solution in a clean, dry test-tube add a few drops of the Tollens' reagent that you have prepared. **Do not shake the tube**. If no reaction is seen, warm the tube in the hot water bath if necessary.				
To a 1 cm depth of each solution in a test-tube add a 1 cm depth of dilute sulfuric acid. Then add a few drops of aqueous potassium manganate (VII). If no reaction is seen, warm the tube in the hot water bath.	**Do not carry out this test**			
To a 1 cm depth of each solution in a test tube, add 1cm^3 of aq NaOH. Warm in water bath for 5 mins. Cool the mixture and add an excess of HNO_3. Add aq $AgNO_3$ dropwise.	**Do not carry out this test**			

Add 2-3 drops of each solution in a test tube. Add 2-3 drops of 2, 4-dinitrophenylhydrazine.	**Do not carry out this test**			

[5]

(ii) Identify each compound.

..... contains ethanal.
..... contains propanone.
..... contains ethanol.
..... contains 1-bromobutane.

[4]

(b) **(i)** Halogenoalkanes can undergo either S_N1 or S_N2 reaction with NaOH (aq); Primary halogenoalkanes undergo S_N2 while tertiary halogenoalkanes undergo S_N1.
Suggest reason why tertiary halogenoalkanes undergo S_N1.

……

……

……

Given that the rate equation for the reaction between 1-phenyl-1-bromoethane, $C_6H_5CHBrCH_3$ and NaOH is

$$\text{Rate} = k\,[C_6H_5CHBrCH_3]$$

(ii) Draw the mechanism of the S_N1 reaction of $C_6H_5CHBrCH_3$ with NaOH. Show all charges and relevant lone pairs and the movement of electron pairs by using curly arrows.

[3]

(ii) With reference to the intermediate, explain why 1-phenyl-1-bromoethane undergoes S_N1 reaction instead of S_N2.

……

……

……

……

(c) **FB 9** is an aqueous solution of an organic compound with a secondary alcohol and carboxylic acid functional group. Fill in the expected observations in the table below. You do not need to identify **FB 9.**

test	*observations*
To a 1 cm depth of **FB 9** in a test-tube add a 1 cm depth of dilute sulfuric acid. Then add a few drops of aqueous potassium manganate(VII). If no reaction is seen, place the test-tube in the hot water bath and leave to stand.	
To a 1 cm depth of **FB 9** in a test-tube carefully add a small spatula measure of sodium hydrogen carbonate.	

[3]

[Data Processing] : Organic *Qualitative Analysis Practical 2*

Instructions

For this experiment, you are to watch the video recording on Google Classroom in order to record your observations.

2. In this question, you will deduce the identities of four organic compounds, **FA 5**, **FA 6**, **FA 7**, and **FA 8**. You will perform a series of test-tubes experiments and use the observations to help you distinguish between the compounds. Observations for **FA 7** have already been recorded in Table 4.1.

(a) liquid samples **FA 5**, **FA 6**, **FA 7** and **FA 8** are provided. Each of which could be one of the following.

propanal hex-1-ene propan-2-ol propanone

In addition to excess to the usual bench reagents, you are also provided with the following

- aqueous bromine, Br_2 (aq)
- iodine solution

Perform the test described in Table 4.1 and record your observations in the table. Test and identify any gases evolved. If there is no observable change, write no observable change.

Use a fresh sample of each liquid in each test

Samples **FA 5**, **FA 6**, **FA 7** and **FA 8** are provided in sample tubes. It is essential that you replace the lid as soon as you have removed the sample for each test.

Table 4.1

		Observations with **FA 5**	Observations with **FA 6**	Observations with **FA 7**	Observations with **FA 8**
a (i)	1. Label a test-tube **FA 5**. Place 1 cm depth of deionised water into the test tube. To this test-tube, add 5 drops of **FA 5**. To this test-tube, add about 1 cm depth of aqueous bromine. Stopper the test-tube and shake vigourously. **Repeat using FA 6, FA 7 and FA 8 in place of FA 5.**			No observable change.	Yellow Br_2 (aq) decolourised.
	2. Label a test-tube **FA 5**. Place 1 cm depth of sulfuric acid in this test-tube. To this test-tube, add 5 drops of **FA 5**, followed by 5 drops of $KMnO_4$. Warm the test-tube in a water bath **Repeat using FA 6, FA 7 and FA 8 in place of FA 5.**			Purple $KMnO_4$ decolourised	

		Observations with **FA 5**	Observations with **FA 6**	Observations with **FA 7**	Observations with **FA 8**
	3. Label the test-tube **FA 5**. Place about 1 cm depth of deionised water in this test-tube. To this test-tube, add 5 drops of **FA 5** and 6 drops of aqueous sodium hydroxide. Now add iodine solution, dropwise, until a permanent yellow/orange is present. Prepare a hot water bath using the hot water provided. Warm the mixture in the water bath for two minutes. **Repeat using FA 6, FA 7 and FA 8 in place of FA 5.**		Pale yellow precipitate formed.	No observable change	

[4]

(ii) Complete Table 4.2, using the observations in Table 4.1 to identify the compounds **FA 5**, **FA 6**, **FA 7** and **FA 8**.

In each case, give your evidence to support your conclusion.

Table 4.2

	Compound	Evidence
FA 5		
FA 6		
FA 7		
FA 8		

[Data Processing] : Organic *Qualitative Analysis Practical 3*

3. **(a)** **FA 7**, **FA 8**, **FA 9** and **FA 10** are organic compounds. Each contains one of the following different functional groups.

The identities have been given and you are required to fill in the relevant observations.

- **FA 10** – primary alcohol
- **FA 7** – tertiary alcohol
- **FA 8** – alkene
- **FA 9** – amide

The identities have been given.

- **FA 10** – primary alcohol
- **FA 7** – tertiary alcohol
- **FA 8** – alkene
- **FA 9** – amide

Some of the following reagents are provided.
- acidified aqueous potassium manganate (VII)
- sodium hydroxide
- bromine

You are required to fill in the relevant observations for each test.

Table 1

reagent	*observations*			
	FA 7	**FA 8**	**FA 9**	**FA 10**
H_2SO_4 (aq), $KMnO_4$ and heat				
NaOH and warm, followed by test for gas with moist red litmus paper				
Aq bromine				