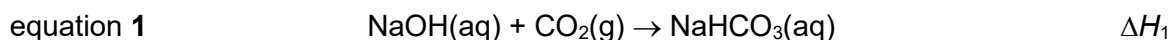
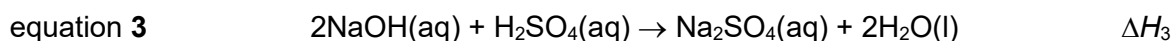
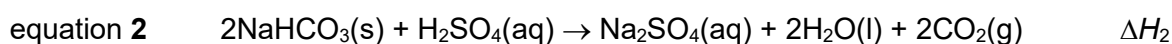


1 Determination of the enthalpy change for the reaction between NaOH and CO₂

In this experiment you will determine the enthalpy change, ΔH_1 , for the reaction shown in equation 1.



Sodium hydrogencarbonate and sodium hydroxide each react with sulfuric acid as shown in equations 2 and 3.



You will determine the values of ΔH_2 and ΔH_3 and use Hess's law to calculate ΔH_1 .

(a) (i) Reaction of sodium hydrogen carbonate with sulfuric acid

FA 1 is solid sodium hydrogen carbonate

FA 2 is 1.0 mol dm⁻³ sulfuric acid, H₂SO₄

In this experiment, you will measure the initial temperature of sulfuric acid in a polystyrene cup. After a sample of sodium hydrogen carbonate is added to sulfuric acid which is in excess, you will measure the minimum temperature reached in the experiment. You will then determine the temperature fall caused by the reaction.

In the space provided on page 3, prepare tables in which to record for the experiment:

- all weighings to an appropriate level of precision,
- all required temperature measurements to an appropriate level of precision.

Procedure:

1. Weigh the capped bottle containing solid **FA 1**.
2. Using a measuring cylinder, transfer 50.0 cm³ of **FA 2** into a polystyrene cup. Place this cup inside a second polystyrene cup, which is placed in a 250 cm³ glass beaker.
3. Place the thermometer in this **FA 2** and record the initial temperature, T_i .
4. Adding small quantities at a time, tip all the **FA 1** into the **FA 2** in the polystyrene cup.
5. Using the thermometer, stir the mixture continuously until it reaches its minimum temperature. Record this temperature, T_{\min} .
6. Calculate and record the temperature fall.
7. Weigh the capped bottle with any residual **FA 1** and record the mass of **FA 1** added.

Results

Mass of capped bottle with FA 1 /g	11.658
Mass of capped bottle with residual FA 1 / g	5.661
Mass of FA 1 added / g	5.997

Initial temperature/ T_i / °C	30.5
Minimum temperature/ T_{\min} / °C	24.2
Temperature fall / °C	6.3

[3]

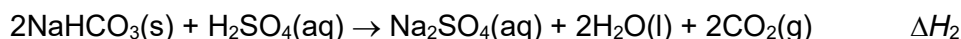
- (a) (ii) Calculate the heat change, q , for your experiment in 1(a)(i).

You should assume that the specific heat capacity of the solution is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$, and that the density of the solution is 1.00 g cm^{-3} .

$$q = 50.0 \times 4.18 \times 1.00 \times 6.3 = \underline{1317 \text{ J} = 1.32 \text{ kJ}}$$

[1]

- (iii) Determine the enthalpy change for the reaction of sodium hydrogen carbonate with sulfuric acid, ΔH_2 . In this reaction, sulfuric acid is used in excess.



Include the sign of ΔH_2 in your answer.

[Ar: C, 12.0; H, 1.0; Na, 23.0; O, 16.0]

$$n_{\text{NaHCO}_3} = \frac{5.997}{84.0} = 0.0714 \text{ mol}$$

$$n_{\text{H}_2\text{SO}_4 \text{ reacted}} = \frac{0.0714}{2} = 0.0357 \text{ mol}$$

$$\Delta H_2 = + \frac{1.32}{0.0357} = \underline{+ 36.9 \text{ kJ mol}^{-1}}$$

[2]

(b) Reaction of sodium hydroxide with sulfuric acid

FA 2 is 1.0 mol dm^{-3} sulfuric acid, H_2SO_4

A student conducted this experiment to determine the enthalpy change for the reaction of sodium hydroxide with sulfuric acid, ΔH_3 .

The student added 40.0 cm^3 of **FA 2** into 30.0 cm^3 of 2.00 mol dm^{-3} NaOH in a polystyrene cup and recorded a temperature rise of 11.0°C .

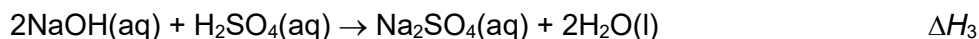
- (i) Calculate the heat change, q , for the student's experiment.

You should assume that the specific heat capacity of the solution is $4.18 \text{ J g}^{-1} \text{ K}^{-1}$, that the density of the solution is 1.00 g cm^{-3} .

$$q = 70.0 \times 4.18 \times 1.00 \times 11.0 = \underline{3219 \text{ J} = 3.22 \text{ kJ}}$$

[1]

- (ii) Determine the enthalpy change for the reaction of sodium hydroxide with sulfuric acid, ΔH_3 . In this reaction, sulfuric acid is used in excess.



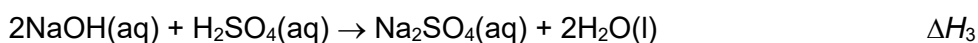
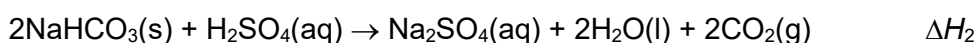
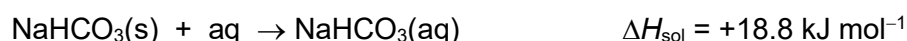
Include the sign of ΔH_3 in your answer.

$$n_{\text{NaOH}} = \frac{30}{1000} \times 2.00 = 0.0600 \text{ mol}$$

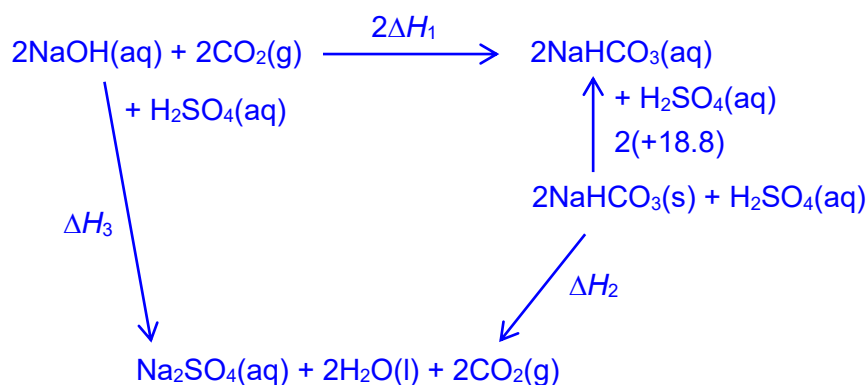
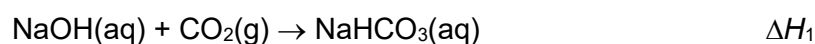
$$n_{\text{H}_2\text{SO}_4 \text{ reacted}} = \frac{0.0600}{2} = 0.0300 \text{ mol}$$

$$\Delta H_3 = -\frac{3.22}{0.03} = \underline{-107 \text{ kJ mol}^{-1}} \quad [2]$$

- (c) The enthalpy change of solution of NaHCO_3 is $+18.8 \text{ kJ mol}^{-1}$.



Using your answer in **1(a)(iii)**, **1(b)(ii)** and ΔH_{sol} given in this question, use Hess's law to calculate ΔH_1 .



By Hess's law,

$$2\Delta H_1 = \Delta H_3 - \Delta H_2 + 2(18.8)$$

$$\Delta H_1 = \frac{-107 - (+36.9) + 2(18.8)}{2} = \underline{-53.2 \text{ kJ mol}^{-1}} \quad [1]$$

- (d) Calculate the percentage error in the temperature fall you obtained in **1(a)(i)**.

$$\text{percentage error} = \pm \frac{2(0.1)}{6.3} \times 100\% = \underline{\pm 3.17\%} \quad [1]$$

[Total: 11]

2 Investigation of the kinetics of the autocatalytic reaction between manganate(VII) ions and ethanedioate ions.

FA 3 is 0.0100 mol dm⁻³ sodium thiosulfate, Na₂S₂O₃

FA 4 is 0.0200 mol dm⁻³ potassium manganate(VII), KMnO₄

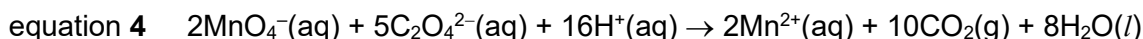
FA 5 is 0.200 mol dm⁻³ ethanedioic acid, H₂C₂O₄

FA 6 is 0.100 mol dm⁻³ potassium iodide, KI

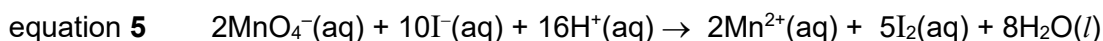
You are also provided with:

- **FA 2**, 1.0 mol dm⁻³ sulfuric acid, H₂SO₄, which you used in Question 1
- Starch Indicator

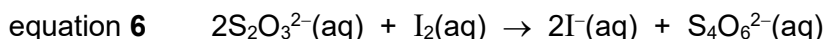
Manganate(VII) ions, MnO₄⁻, react with ethanedioate ions, C₂O₄²⁻, in acidic medium as shown in equation 4. The Mn²⁺ ions produced in equation 4 act as a catalyst for the reaction. This is an example of an autocatalytic reaction.



You will prepare a reaction mixture containing **FA 4**, **FA 5** and **FA 2**. At timed intervals, you will transfer aliquots (portions) of the reaction mixture to excess KI solution which will “quench” the reaction by reacting away manganate(VII) ions as shown in equation 5.



You will then titrate the iodine produced in the resulting solution against sodium thiosulfate solution. The reaction is shown in equation 6.



Your titre values will indicate the concentration of MnO₄⁻ remaining in the reaction mixture at different times. Hence, the rate of reaction between MnO₄⁻ and C₂O₄²⁻ at different times can be determined by graphical analysis of your results.

In an appropriate format in the space provided on **page 8**, prepare a table to record, for each aliquots

- the time of transfer, t , in minutes and seconds,
- the decimal time, t_d , in minutes, to 0.1 min, for example, if $t = 4 \text{ min } 33 \text{ s}$ then $t_d = 4 \text{ min} + 33/60 \text{ min} = 4.6 \text{ min}$,
- the burette readings and the volumes of **FA 3** added.

t_d

You are advised to read all the instructions on this page before you start your experiment.

Note: You will perform each titration once only. Great care must be taken that you do not overshoot the end-point.

Once you have started the stopwatch, it must continue running for the duration of the experiment. You must not stop until you have finished this experiment.

(a) (i) Preparation of FA 6 and FA 3

1. Using a 10 cm³ measuring cylinder, add about 10 cm³ of **FA 6** to each of the labelled boiling tubes, **1** to **5**. Wash this measuring cylinder as it will be used in point 5.
2. Fill a burette with **FA 3**.

Preparation of the reaction mixture and withdrawing aliquots

3. Use appropriate measuring cylinders to add to the beaker labelled **reaction mixture**
 - 50.0 cm³ of **FA 5**,
 - 5.0 cm³ of **FA 2**,
 - 45.0 cm³ of deionised water.
4. Using a measuring cylinder, measure 25.0 cm³ of **FA 4**. Add this **FA 4** to the beaker labelled **reaction mixture**. Start the stopwatch at the instant of mixing and stir the mixture thoroughly using a glass rod.
5. At approximately 1 minute, transfer a 10.0 cm³ aliquot (portion) of the reaction mixture to a 10 cm³ measuring cylinder, using a dropping pipette.
6. **Immediately** transfer this aliquot into the boiling tube **1** and shake the mixture. Read and record the time of transfer in minutes and seconds, to the nearest second, when the aliquot is added.

You are advised to make use of the waiting time and start on your titrations.

7. At approximately 4 minutes, repeat points **5** and **6** but transfer the aliquot into boiling tube **2**.
8. Repeat points **5** and **6** three more times at your chosen times. You should not exceed a maximum reaction time of 15 mins. At each of your chosen times, add each aliquot into the corresponding boiling tube.

Titration of samples in boiling tube

9. Pour all the contents of boiling tube **1** into a clean conical flask. Rinse this boiling tube with deionised water and add the washings to the same conical flask.
10. Titrate the iodine in this solution with **FA 3** until the solution turns pale yellow. Add 10 drops of starch indicator, the solution will turn blue-black. Continue to titrate until the blue-black colour **just** disappears at the end-point. Record your titration results.
11. Wash this conical flask thoroughly with water.
12. Repeat points **9** to **11** as required for each of the remaining boiling tubes.

Results

t	t_d / min	Final burette reading / cm^3	Initial burette reading / cm^3	Volume of FA 3 added / cm^3
1 min 18 s	1.3	19.00	0.00	19.00
4 min 25 s	4.4	17.70	0.00	17.70
7 min 21 s	7.4	30.80	17.90	12.90
10 min 13 s	10.2	35.80	30.80	5.00
13 min 15 s	13.3	37.60	35.80	1.80

[4]

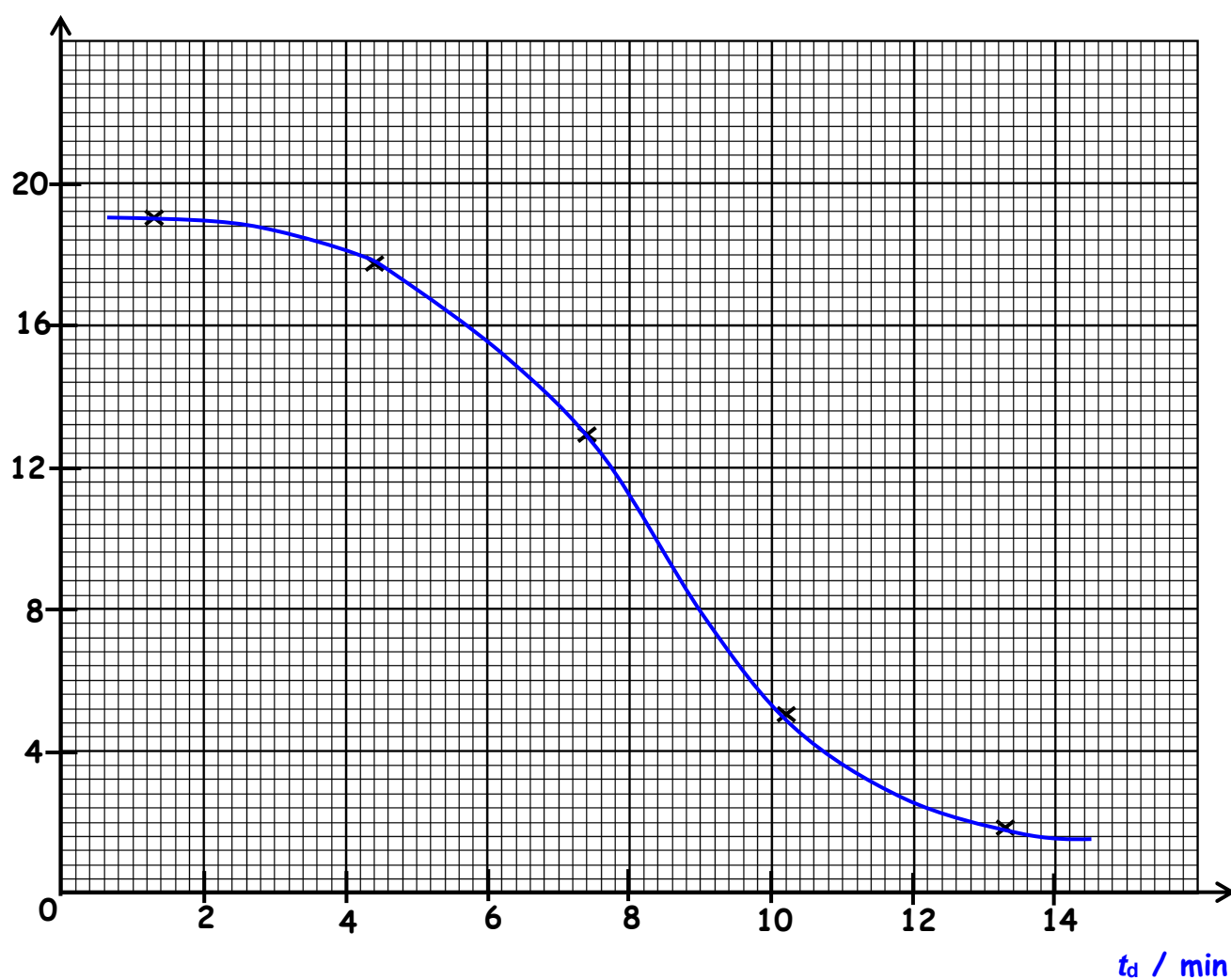


Fig. 2.1

(iii) Rate of reaction at time t is indicated by the gradient of tangent at time t .

Describe and explain how the gradient changes in your graph in **2(a)(ii)** show that the reaction between MnO_4^- and $\text{C}_2\text{O}_4^{2-}$ is an autocatalytic reaction.

At the start, the gradient of tangent is gentle/ slightly negative. This indicates that the rate of reaction is low due to absence of Mn^{2+} /catalyst or low concentration of Mn^{2+} /cataylst.

As the reaction proceeds, the gradient of tangent becomes steeper/more negative or magnitude of gradient increases. This indicates that the rate of reaction increases due to more catalyst/ Mn^{2+} is produced or concentration of Mn^{2+} /catalyst increases.

Towards the end, the gradient of tangent becomes less negative/ magnitude of gradient decreases. This indicates that the rate of reaction decreases due to decrease in concentration of reactants.

[2]

- (b) A student performed the experiment you performed in 2(a)(i) but in point 3, the student also added 5.0 cm³ of manganese(II) sulfate, MnSO₄, a source of Mn²⁺ to catalyse the reaction. The student kept the total volume of the reaction mixture constant by adjusting the volume of deionised water used in the experiment.

The results from the student's experiment are shown on the graph in Fig 2.2.

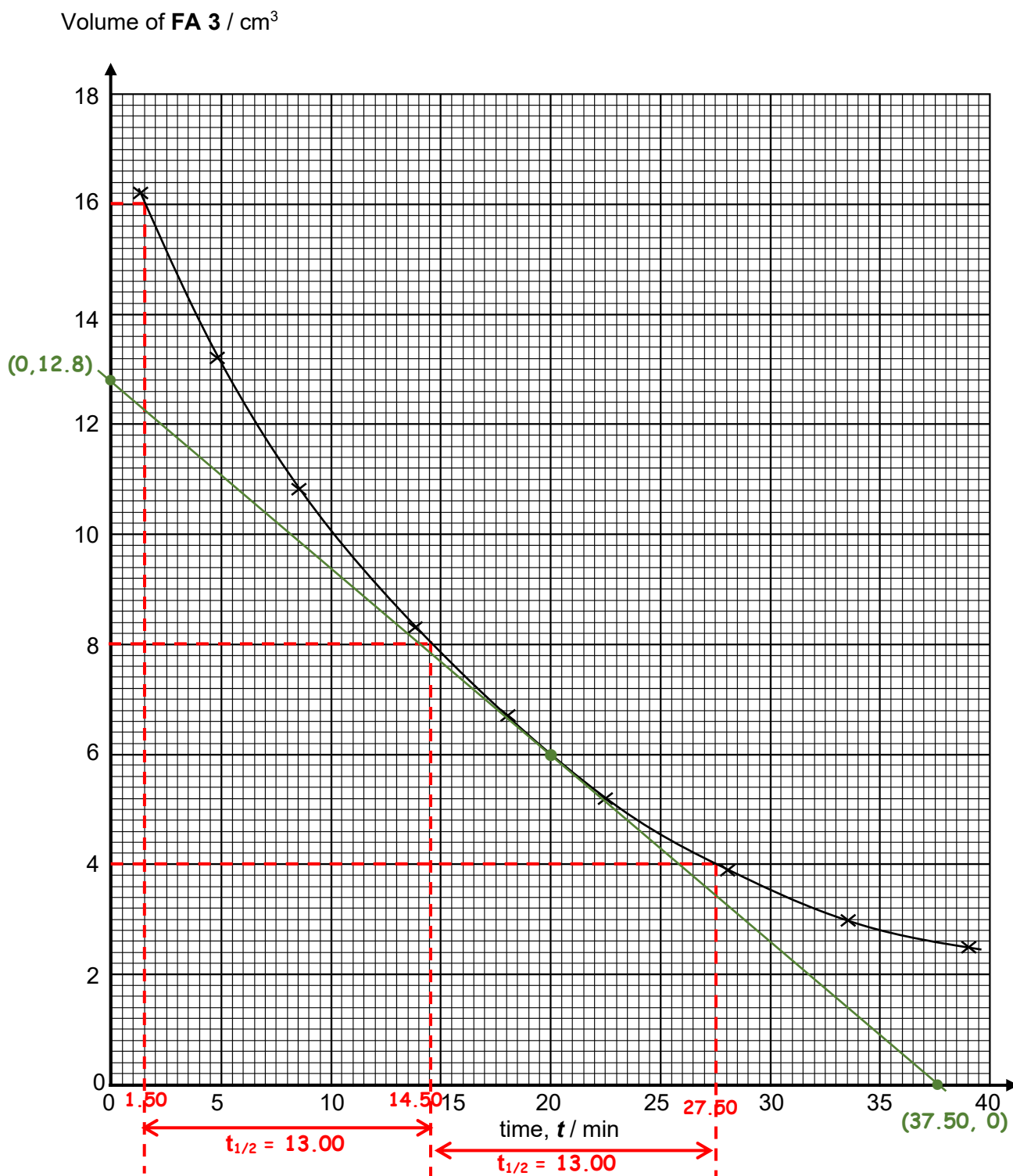


Fig 2.2

- (b) (i) Use the graph in Fig. 2.2 to determine the order of reaction with respect to $[\text{MnO}_4^-]$ for the reaction in **equation 4**. Draw clearly any construction lines on the graph. Explain your reasoning clearly.

From the graph, first $t_{1/2} = 13.00 \text{ min}$,
second $t_{1/2} = 13.00 \text{ min}$,

Volume of FA 3 $\propto [\text{MnO}_4^-]$

Since half-lives are (approximately) constant, the reaction is first order with respect to $[\text{MnO}_4^-]$ in reaction 1.

[2]

- (ii) Draw a tangent to the graph in Fig. 2.2 at time $t = 20 \text{ min}$. Determine the gradient of this line, showing clearly how you did this.

$$\text{Gradient} = \frac{12.8 - 0}{0 - 37.50} = -0.341 \text{ cm}^3 \text{ min}^{-1}$$

$$\text{gradient} = \underline{-0.341} \text{ cm}^3 \text{ min}^{-1} \quad [3]$$

- (iii) Use your gradient in **2(b)(ii)** to determine the rate of change of the amount of $\text{S}_2\text{O}_3^{2-}$ ions at time $t = 20 \text{ min}$ in mol min^{-1} .

$$\begin{aligned} \text{rate of change in amount of } \text{S}_2\text{O}_3^{2-} \text{ ions} &= -0.341 \times 10^{-3} \times 0.0100 \\ &= \underline{-3.41 \times 10^{-6} \text{ mol min}^{-1}} \quad [1] \end{aligned}$$

- (iv) equation 5 $2\text{MnO}_4^- + 10\text{I}^- + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 5\text{I}_2 + 8\text{H}_2\text{O}$
 equation 6 $2\text{S}_2\text{O}_3^{2-} + \text{I}_2 \rightarrow 2\text{I}^- + \text{S}_4\text{O}_6^{2-}$

Using your answer in **2(b)(iii)** and equations 5 and 6, determine the rate of change of the amount of MnO_4^- reacted at time $t = 20 \text{ min}$ in mol min^{-1} .

$$\begin{aligned} \text{rate of change in amount of } \text{MnO}_4^- \text{ ions reacted} \\ = -3.41 \times 10^{-6} \times 1/2 \times 2/5 = \underline{-6.82 \times 10^{-7} \text{ mol min}^{-1}} \quad [1] \end{aligned}$$

- (v) Hence, calculate the rate of change of $[\text{MnO}_4^-]$ in the 10 cm^3 reaction mixture at $t = 20 \text{ min}$, in $\text{mol dm}^{-3} \text{ min}^{-1}$.

$$\begin{aligned} \text{Rate of change of the } [\text{MnO}_4^-] \\ = -6.82 \times 10^{-7} / 0.010 \\ = \underline{-6.82 \times 10^{-5} \text{ mol dm}^{-3} \text{ min}^{-1}} \quad [3] \end{aligned}$$

[Total: 19]

3 Inorganic and Organic Analysis

- (a) **FA 7** is a mixture containing one cation and two anions which are listed in the Qualitative Analysis notes.

You are also provided with **FA 3**, sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, which you used in Question 2.

You will carry out tests to identify three ions in **FA 7**.

Unless otherwise stated, the volumes given in Tables 3.1 and 3.2 are approximate and should be estimated rather than measured.

If there is no observable change, write **no observable change**.

- (i) Carry out the following tests and record your observations in Table 3.1.

Table 3.1

	<i>test</i>	<i>observations</i>
1	Place a small spatula measure of FA 7 in a dry hard-glass test-tube. Heat gently at first, then heat more strongly until no further change is seen. Allow the tube to cool on a heat-proof mat.	Condensation on test-tube Purple gas/vapour produced solid turns yellow on heating yellow solid pales/ turns white/ cream/ off-white on cooling. CO ₂ gas forms white ppt. in limewater.
2	Add 2 cm depth of dilute nitric acid to a test-tube. Add a spatula measure of FA 7 .	Effervescence observed. CO ₂ gas forms white ppt. in limewater. Colourless/pale yellow solution formed.

Dilute the solution from test 2 with an equal volume of deionised water.

Keep the resulting solution for use in **3(a)(ii)**. This solution is **FA 8**.

[3]

- (ii) Carry out the following tests using the **FA 8** that you have prepared in **3(a)(ii)** and record your observations in Table 3.2.

Table 3.2

	<i>test</i>	<i>observations</i>
1	Add 1 cm depth of FA 8 to a test-tube. Add a few drops of potassium manganate(VII). Add FA 3 dropwise to the mixture, with shaking, until there is no further change.	Purple KMnO_4 decolourises. Solution turns brown/ red-brown/orange/yellow. brown/ red-brown/orange/yellow solution decolourises/ turns colourless
2	Add 1 cm depth of FA 8 to a test-tube. Add aqueous sodium hydroxide until there is no further change.	White ppt. formed soluble in excess NaOH to give a colourless solution.
3	Add 1 cm depth of FA 8 to a test-tube. Add aqueous ammonia until there is no further change.	White ppt. formed soluble in excess NH_3 to give a colourless solution.
4	Add 1 cm depth of FA 8 to a test-tube. Add a few drops of aqueous silver nitrate. Add aqueous ammonia to the mixture, with shaking, until the aqueous ammonia is in excess.	Yellow ppt. formed. Ppt. insoluble in excess NH_3

[3]

- (iii) Use your observations in Tables 3.1 and 3.2 to identify three of the ions present in **FA 7**.

Ions present in **FA 7** are Zn^{2+} and CO_3^{2-} and I^-

[1]

- (b) **Before starting this analysis, ensure your Bunsen burner is turned off.**

In this part of the question you will deduce the identities of two organic compounds.

FA 9 $\text{C}_2\text{H}_2\text{O}_4$

FA 10 $\text{C}_4\text{H}_{10}\text{O}_2$

You are provided with aqueous samples of these compounds, **FA 9** and **FA 10**.

You will perform some of the tests described in Table 3.3.

Using the observations in Table 3.3, and the given molecular formulae, you will deduce the identities of **FA 9** and **FA 10**.

- (i) Carry out the following tests. Some of the observations have been completed for you. There is no need to carry out those tests. Record your observations in Table 3.3.

Unless otherwise stated, the volumes given below are approximate and should be estimated rather than measured.

Test and identify any gases evolved.

If there is no observable change, write **no observable change**.

Use a fresh sample of each solution in each test. It is essential that you replace the lid as soon as you have removed the sample for each test.

Table 3.3

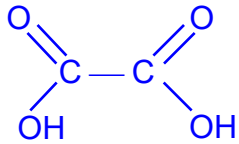
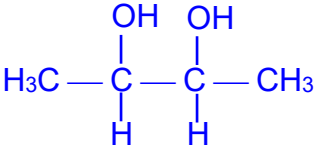
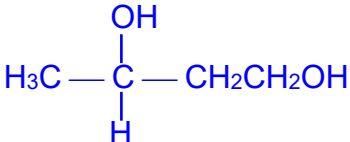
tests	Observations with FA 9 , $C_2H_2O_4$	Observations with 10 , $C_4H_{10}O_2$	FA
<p>1. Add 1 cm depth of the FA solution into a test-tube.</p> <p>Add 1 cm depth of dilute sulfuric acid followed by 2-3 drops of potassium dichromate solution to the test-tube.</p> <p>Leave the test-tube to stand in the hot water-bath</p>	No observable change	Orange $K_2Cr_2O_7$ turns green	
<p>2. Add 1 cm depth of the FA solution into a test-tube.</p> <p>Add, using the tip of a spatula, a small portion of magnesium turnings to this test-tube.</p>	<p>Effervescence observed.</p> <p>H_2 gas "pops" with a lighted splint.</p>	No observable change/ no gas evolved	
<p>3. Add 1 cm depth of the FA solution into a test-tube.</p> <p>To this test-tube add 8 drops of sodium hydroxide solution followed by iodine solution, dropwise, until a permanent orange/red colour is present.</p> <p>Warm the mixture in the hot water bath for two minutes.</p>	<p>solution turns yellow/orange/red or no yellow ppt. or No observable change</p>	Pale yellow ppt. formed	

[2]

- (b) (ii) Complete Table 3.4, using the observations in Table 3.3 to suggest possible identities of compounds **FA 9** and **FA 10**.

In each case, give evidence to support your conclusion.

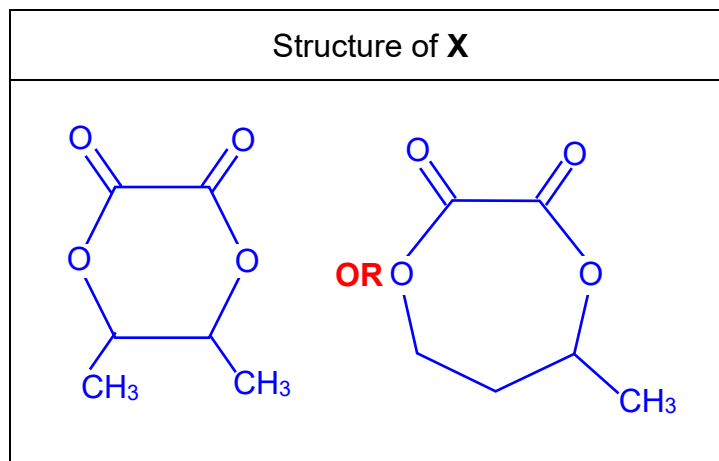
Table 3.4

compound	evidence
FA 9 $\text{C}_2\text{H}_2\text{O}_4$ 	FA 9 is a carboxylic acid since it undergoes redox reaction with Mg to give H_2 gas.
FA 10 $\text{C}_4\text{H}_{10}\text{O}_2$	FA 10 is an alcohol with $-\text{CH}(\text{OH})\text{CH}_3$ since it undergoes a positive iodoform test/ oxidation with alkaline $\text{I}_2(\text{aq})$ to give pale yellow ppt. It undergoes oxidation with acidified $\text{K}_2\text{Cr}_2\text{O}_7$ which turns orange $\text{K}_2\text{Cr}_2\text{O}_7$ green.
 OR 	

[3]

- (iii) **X**, $\text{C}_6\text{H}_8\text{O}_4$, is an organic compound which contains only one functional group. When **X** is heated with aqueous sulfuric acid, **FA 9** and **FA 10** are formed.

Use your answer in Table 3.4 to suggest a possible structure of **X**.

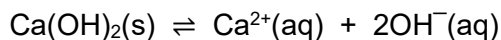


[1]

[Total: 13]

4 Planning

When a sparingly soluble salt, calcium hydroxide, Ca(OH)_2 , is added to water, an equilibrium is established between the undissolved salt and a saturated solution of the salt.



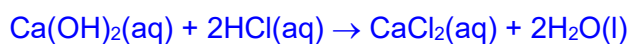
The solubility product can be found by determining the equilibrium concentration of the hydroxide ions in a saturated solution of calcium hydroxide. This equilibrium concentration can be found by titration with a standard solution of hydrochloric acid.

- (a) (i) The solubility of calcium hydroxide, Ca(OH)_2 , at 30 °C, is approximately 1.8 g dm⁻³.

Calculate the concentration of 25 cm³ of HCl required to react completely with 25 cm³ of saturated Ca(OH)_2 solution.

[Ar: Ca, 40.1; H, 1.0; O, 16.0]

$$\text{Solubility of Ca(OH)}_2 = \frac{1.8}{74.1} = 0.0243 \text{ mol dm}^{-3}$$



$$[\text{HCl}] = 2 \times 0.0243 = \underline{0.0486} \text{ mol dm}^{-3} \quad [1]$$

- (ii) Hence, calculate the volume of 0.500 mol dm⁻³ stock solution of hydrochloric acid, HCl, required to prepare 250.0 cm³ of diluted HCl of the concentration you determined in 4(a)(i).

$$\text{Vol of stock HCl} = \frac{(0.0486)(250)}{(0.500)} = \underline{24.3 \text{ cm}^3} \quad [1]$$

- (iii) Plan an experiment to determine the solubility product of calcium hydroxide at 30 °C.

You may assume that you are provided with:

- solid calcium hydroxide, Ca(OH)_2 ,
- 0.500 mol dm⁻³ stock solution of hydrochloric acid, HCl,
- methyl orange indicator,
- the equipment normally found in a school or college laboratory

Your plan should include details of, including quantities:

- the preparation of 250.0 cm³ of diluted HCl of the concentration you determined in 4(a)(i) from the given stock solution.
- the preparation of about 100 cm³ of a saturated solution of calcium hydroxide, Ca(OH)_2 at 30 °C,
- the essential details of the titration process;
- an outline of how you would use your mean titre value to determine the solubility product of Ca(OH)_2 .

To prepare 250.0 cm³ of 0.0486 mol dm⁻³ HCl(aq):

1. Using a 50.00 cm³ burette, run 24.3 cm³ of 0.500 mol dm⁻³ stock HCl solution into a 250 cm³ graduated flask.
2. Make up to the graduated mark with deionised water.
3. Stopper and shake the flask well to obtain a homogeneous solution.

To prepare 100 cm³ of Ca(OH)₂ saturated solution at 30 °C:

1. Use a 100 cm³ measuring cylinder to transfer 100 cm³ of deionised water into a 250 cm³ conical flask.
2. Using a spatula, add a few tips of solid Ca(OH)₂ into the conical flask. Swirl the conical flask to dissolve all the solids. Keep adding more solids, with swirling after each addition, until some solids are left undissolved.
3. To ensure that the solution is saturated, leave the conical flask containing the solution to stand in a thermostatically controlled water bath set at 30°C, for some time. There must be some solids left undissolved.
4. To remove undissolved solids, filter the saturated solution into a clean, dry conical flask using a dry filter funnel and a piece of dry filter paper.

Titration procedure:

1. Fill a burette with 0.0486 mol dm⁻³ HCl(aq).
2. Pipette 25.0 cm³ of the saturated solution into a conical flask.
3. Add a few drops of methyl orange into the flask.
4. Titrate the solution in the flask with HCl from the burette, adding dropwise towards the end-point, until the solution changes from yellow to orange at the end-point.
5. Repeat the titration to obtain two consistent titres within 0.10 cm³ in difference. Record all the titration results and average the two consistent titres.

To calculate K_{sp} of Ca(OH)₂:

let V cm³ be the mean titre.

$$\text{Amount of HCl required} = \frac{V}{1000} \times 0.0486 = 4.86 \times 10^{-5} V \text{ mol}$$

$$\begin{aligned} \text{Amount of OH}^- \text{ in } 25.0 \text{ cm}^3 \text{ saturated solution} \\ = 4.86 \times 10^{-5} V \text{ mol} \end{aligned}$$

$$[\text{OH}^-] \text{ in saturated solution} = 4.86 \times 10^{-5} V \times \frac{1000}{25.0} =$$

$$1.94 \times 10^{-3} V = 2s \text{ mol dm}^{-3}$$

[7]

OR

$$\text{Amount of HCl required} = \frac{V}{1000} \times 0.0486 = 4.86 \times 10^{-5} \text{ V mol}$$

$$\text{Amount of Ca(OH)}_2(\text{aq}) \text{ in } 25.0 \text{ cm}^3 \text{ saturated solution} \\ = 2.43 \times 10^{-5} \text{ V mol}$$

$$\text{Conc. of saturated Ca(OH)}_2 = 2.43 \times 10^{-5} \text{ V} \times \frac{1000}{25.0} = s \text{ mol dm}^{-3}$$

$$K_{\text{sp}} (\text{Ca(OH)}_2) = [\text{Ca}^{2+}] [\text{OH}^-]^2 \\ = (s)(2s)^2 = 4s^3 \text{ mol}^3 \text{ dm}^{-9}$$

- (b) The experiment can be modified to determine the ΔH and ΔS for the dissolution of solid Ca(OH)_2 . The K_{sp} values of Ca(OH)_2 can be determined at different temperatures and the enthalpy change, ΔH , and entropy change, ΔS , of the reaction can be graphically determined.

The following expression can be used.

$$\ln K_{\text{sp}} = -\frac{\Delta H}{RT} + \frac{\Delta S}{R}$$

T is the reaction temperature in kelvin.

R is the molar gas constant.

- (i) Sketch on Fig 4.1 the graph you would expect to obtain.

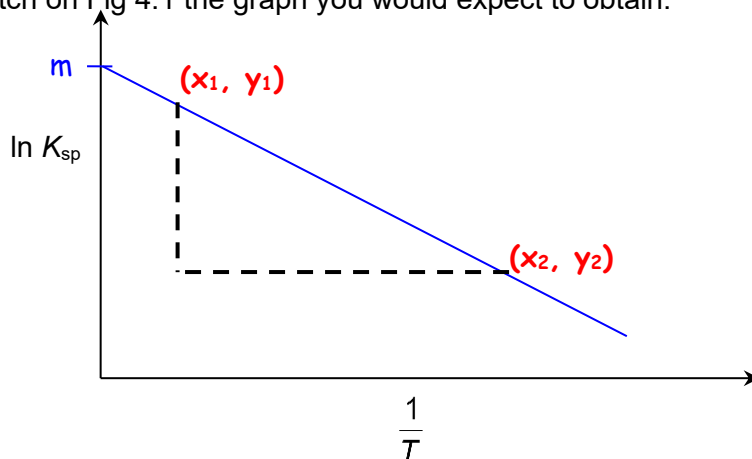


Fig. 4.1

[1]

- (ii) Describe how you would use your graph to determine values for ΔH and ΔS .

$$\text{gradient} = \frac{y_2 - y_1}{x_2 - x_1} = m$$

$$\text{Gradient} = -\frac{\Delta H}{R} \quad \underline{\Delta H = -\text{gradient} \times R}$$

$$\text{y-intercept, } m = \frac{\Delta S}{R} \quad \underline{\Delta S = m \times R}$$

[2]

[Total: 12]