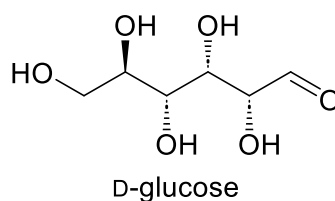




### 1 Investigation of redox reactions involving D-glucose

D-Glucose is a sugar with molecular formula  $C_6H_{12}O_6$ , which can react directly or indirectly with potassium manganate(VII),  $KMnO_4$ .



- (a) **FA 1** is a solution of potassium manganate(VII),  $KMnO_4$ .  
**FA 2** is a  $1.0 \text{ mol dm}^{-3}$  sulfuric acid solution,  $H_2SO_4$ .

You are also provided with a solution of D-glucose and bench reagents.

Perform the tests described in Table 1.1 and record your observations in the table.

**Table 1.1**

test		observations
(i)	Add 1 cm depth of <b>FA 1</b> to a test-tube. Add an equal volume of aqueous sodium hydroxide.  Add 1 drop of glucose solution and shake thoroughly.  Allow to stand for 10 minutes, with occasional shaking. Continue with the remaining parts of Question 1.	<b>Purple FA 1</b> turns <b>green</b> , then to <b>yellow</b> and a <b>brown ppt</b> is formed in <b>colourless</b> solution on standing.
(ii)	Add 1 cm depth of <b>FA 1</b> to a boiling tube. Add 2 cm depth of <b>FA 2</b> .  Add 1 cm depth of glucose solution and warm gently. Shake well to mix.	<b>Purple FA 1</b> turns <b>colourless</b> .

[2]

(b) Manganese species exist in ions of different oxidation states and colours.

Suggest the identity of the **final** manganese-containing species in (a)(i) and in (a)(ii).

final manganese-containing species in (a)(i)  $\text{MnO}_2$  .....

final manganese-containing species in (a)(ii)  $\text{Mn}^{2+}$  .....

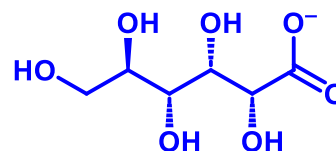
[1]

(c) In (a)(i), D-glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$ , reacts with  $\text{KMnO}_4$  in alkaline medium to give D-gluconate,  $\text{C}_6\text{H}_{11}\text{O}_7^-$ .

(i) Name the type of reaction that D-glucose undergoes in (a).

Oxidation ..... [1]

(ii) Draw the structure of the D-gluconate ion.



[1]

[Total: 5]

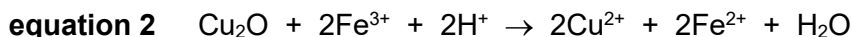
## 2 Determination of mass of D-glucose in a sample

The Bertrand method is a classical redox titration technique used to determine the amount of reducing sugars, such as D-glucose.

In this method, a sample of D-glucose,  $C_6H_{12}O_6$ , is reacted with excess alkaline copper(II) solution,  $Cu^{2+}$ , to form insoluble copper(I) oxide,  $Cu_2O$ .

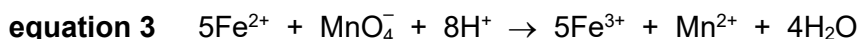


The  $Cu_2O$  is filtered and dissolved in 100 cm<sup>3</sup> acidic iron(III) solution,  $Fe^{3+}$ .



The resulting solution is then diluted to 1 dm<sup>3</sup>. The diluted solution is labelled **FA 3**.

The amount of  $Fe^{2+}$  present in **FA 3** is then determined by titration with potassium manganate(VII),  $KMnO_4$ , in acidic medium. The end-point of the titration occurs when the presence of unreacted  $MnO_4^-$  causes the colour of the solution to become pale pink.



You are to determine the amount of  $Fe^{2+}$  present in **FA 3** and hence determine the mass of D-glucose in the sample.

**FA 1** is 0.0200 mol dm<sup>-3</sup> potassium manganate(VII),  $KMnO_4$ .

**FA 2** is 1.0 mol dm<sup>-3</sup> sulfuric acid solution,  $H_2SO_4$ .

### (a) (i) Procedure

1. Fill the burette with **FA 1**.
2. Pipette 25.0 cm<sup>3</sup> of **FA 3** into a clean 250 cm<sup>3</sup> conical flask.
3. Use a measuring cylinder to add 10.0 cm<sup>3</sup> of **FA 2** to the conical flask.
4. Run **FA 1** from the burette into the conical flask. The end-point is reached when the **first permanent** trace of pale pink colour is seen.
5. Record your titration results, to an appropriate level of precision, in Table 2.1.
6. Repeat points 2 to 5 until consistent titre values are obtained.

**Table 2.1**

final burette reading / cm <sup>3</sup>	20.20	39.10	21.50		
initial burette reading / cm <sup>3</sup>	0.20	20.20	2.50		
volume of <b>FA 1</b> used / cm <sup>3</sup>	20.00	18.90	19.00	18.30	

[2]

- (ii) From your titration results, obtain a suitable volume of **FA 1** to be used in your calculations. Show clearly how you obtained this volume.

$$\text{volume of FA 1 used} = \frac{18.90 + 19.00}{2} = 18.95 \text{ cm}^3$$

$$\text{volume of FA 1} = \dots\dots\dots 18.95 \dots\dots\dots \text{ cm}^3 \text{ [3]}$$

- (b) (i) Calculate the amount of  $\text{Fe}^{2+}$  in  $1 \text{ dm}^3$  of **FA 3**.

$$\text{amount of MnO}_4^- \text{ used} = \frac{18.95}{1000} \times 0.0200 = 3.79 \times 10^{-4} \text{ mol}$$

$$\text{amount of Fe}^{2+} \text{ in } 25.0 \text{ cm}^3 \text{ of FA 3} = 5 \times 3.79 \times 10^{-4} = 1.895 \times 10^{-3} \text{ mol}$$

$$\text{amount of Fe}^{2+} \text{ in } 1 \text{ dm}^3 \text{ of FA 3} = 1.895 \times 10^{-3} \times \frac{1000}{25} = 0.0758 \text{ mol}$$

$$\text{amount of Fe}^{2+} \text{ in } 1 \text{ dm}^3 \text{ of FA 3} = \dots\dots\dots 0.0758 \text{ mol} \dots\dots\dots \text{ [3]}$$

- (ii) Calculate the mass of D-glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$ , in the sample.  
[ $A_r$ : H, 1.0; C, 12.0; O, 16.0]

$$\text{amount of Cu}_2\text{O formed} = \frac{1}{2} \times 0.0758 = 0.0379 \text{ mol}$$

$$\text{amount of D-glucose} = 0.0379 \text{ mol}$$

$$\text{mass of D-glucose} = 0.0379 \times 180.0 = 6.822 \approx 6.82 \text{ g}$$

$$\text{mass of glucose in sample} = \dots\dots\dots 6.82 \text{ g} \dots\dots\dots \text{ [2]}$$

- (c) A student repeated the procedure in (a) and obtained an average titre value of  $19.50 \text{ cm}^3$ .

The errors (uncertainties) associated with **each reading** using pipette and burette are  $\pm 0.10 \text{ cm}^3$  and  $\pm 0.05 \text{ cm}^3$  respectively.

Calculate the maximum total percentage error (uncertainty) of this average titre volume.

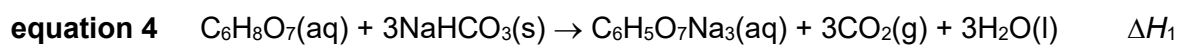
$$\text{maximum \% error} = \pm \left( \frac{0.10}{25.0} + \frac{2 \times 0.05}{19.50} \right) \times 100\% = \underline{\underline{\pm 0.913\%}}$$

[1]

[Total: 11]

### 3 Determination of the enthalpy change when citric acid reacts with sodium hydrogencarbonate

Citric acid,  $\text{C}_6\text{H}_8\text{O}_7$ , is found in citrus fruit such as lemons and limes. It is a triprotic (tribasic) acid – one mole of citric acid reacts with three moles of sodium hydrogencarbonate.



**FA 4** is  $0.080 \text{ mol dm}^{-3}$  citric acid,  $\text{C}_6\text{H}_8\text{O}_7$ .

**FA 5** is solid sodium hydrogencarbonate,  $\text{NaHCO}_3$ .

- (a) Calculate the minimum mass of sodium hydrogencarbonate that will react completely with the citric acid in  $50.0 \text{ cm}^3$  of **FA 4**.  
[Ar: Na, 23.0; H, 1.0; C, 12.0; O, 16.0]

$$\text{amount of citric acid} = \frac{50}{1000} \times 0.08 = 0.00400 \text{ mol}$$

$$\text{amount of NaHCO}_3 = 0.00400 \times 3 = 0.0120 \text{ mol}$$

$$\text{minimum mass of NaHCO}_3 = 0.0120 \times 84.0 = 1.008 \approx 1.01 \text{ g}$$

$$\text{minimum mass of NaHCO}_3 = \dots\dots\dots 1.01 \text{ g} \dots\dots\dots [1]$$

**(b) Determination of the molar enthalpy change of reaction,  $\Delta H_1$** 

In this experiment, you will measure the temperature of the contents of a polystyrene cup at timed intervals, both before and after an excess of **FA 5** is added. You will analyse your results graphically to obtain an accurate value for the temperature change caused by the reaction.

You will use this value to calculate the heat change,  $q$ , for the experiment and hence determine a value for the molar enthalpy change of the reaction,  $\Delta H_1$ .

In the space provided on page 8, prepare a table in which to record for your experiment:

- all values of temperature,  $T$ , to an appropriate level of precision
- all values of time,  $t$ , recorded to the nearest 0.5 min.

It is important that you measure each temperature at the specified time.

**Procedure**

1. Measure 50.0 cm<sup>3</sup> of **FA 4** using a 50.0 cm<sup>3</sup> measuring cylinder.
2. Place one polystyrene cup inside a glass beaker.
3. Transfer 50.00 cm<sup>3</sup> of **FA 4** into the polystyrene cup.
4. Carefully stir the **FA 4** in the polystyrene cup with the thermometer. Read and record the temperature,  $T$ . Start the stopwatch ( $t = 0.0$  min). The stopwatch must be left to run for the rest of the experiment.
5. Continue to stir **FA 4**. Read and record  $T$  every 0.5 minute for 1.5 minutes.
6. At exactly 2.0 minutes, transfer all the solid **FA 5** (an excess) to the polystyrene cup. Stir the mixture thoroughly but do not read  $T$ .
7. Continue to stir the mixture. Read and record  $T$  at  $t = 2.5$  min and every 0.5 min until the temperature reaches a minimum, and then every minute for a further 4 minutes.

## (i) Results

$t / \text{min}$	$T / ^\circ\text{C}$
0.0	30.2
0.5	30.2
1.0	30.2
1.5	30.2
2.5	28.6
3.0	28.4
3.5	28.3
4.0	28.2
4.5	28.2
5.5	28.3
6.5	28.4
7.5	28.5

[3]



- (ii) Plot a graph of the temperature on the y-axis, against time,  $t$ , on the x-axis on the grid in Fig. 3.1.

Draw a best-fit straight line taking into account all of the points before  $t = 2.0$  min.

Draw another best-fit straight line taking into account all of the points after the temperature of the mixture has started to rise steadily.

Extrapolate (extend) both lines to  $t = 2.0$  min.

[3]

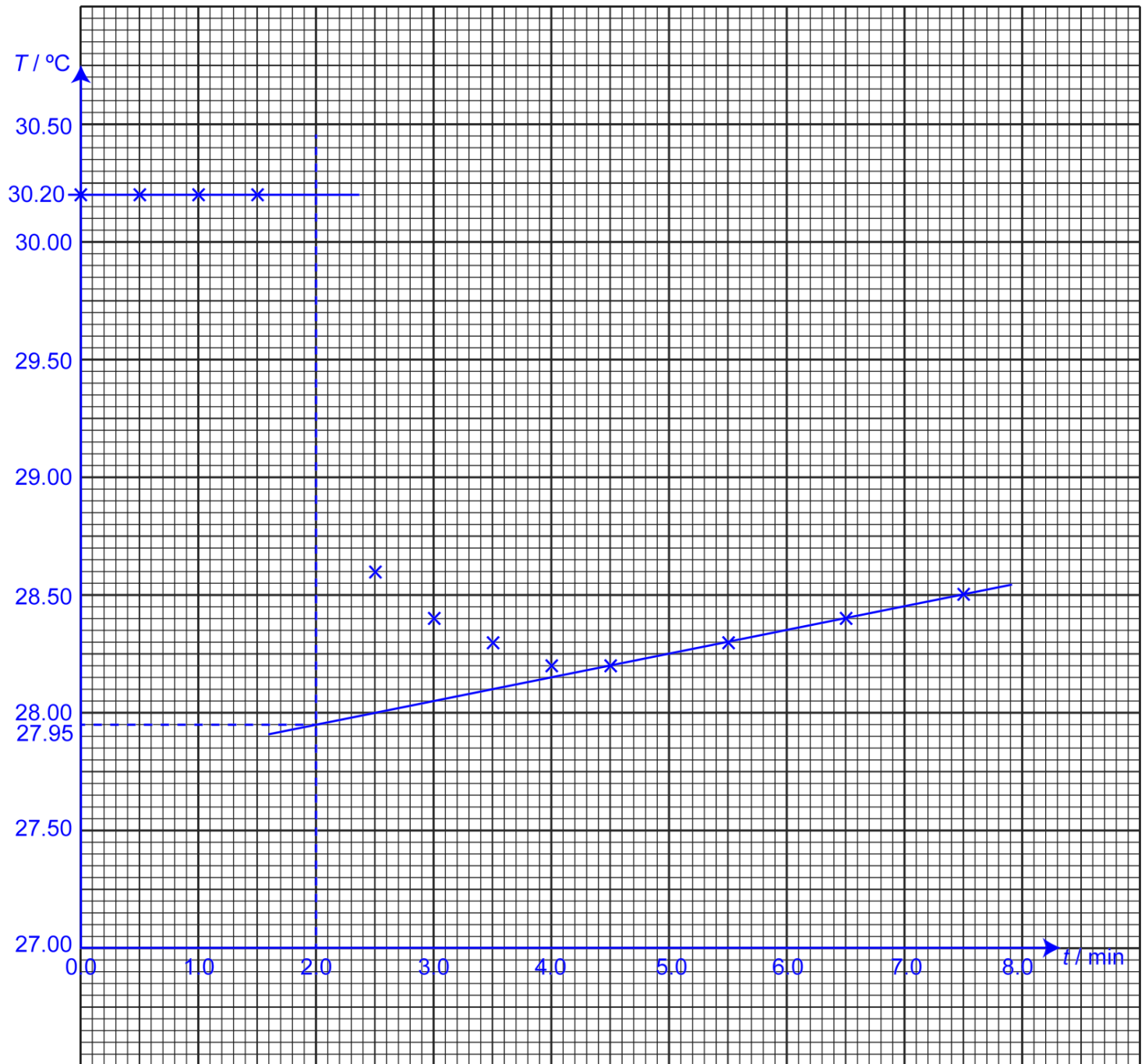


Fig. 3.1

- (iii) From your graph, read the minimum temperature,  $T_{\min}$ , and the maximum temperature,  $T_{\max}$ , at  $t = 2.0$  min. Record these values in the spaces provided.

Deduce the temperature change,  $\Delta T$ , at  $t = 2.0$  min.

$$T_{\min} = \dots\dots\dots 27.95\text{ }^{\circ}\text{C}$$

$$T_{\max} = \dots\dots\dots 30.20\text{ }^{\circ}\text{C}$$

$$\Delta T = \dots\dots\dots -2.25\text{ }^{\circ}\text{C}$$

[1]

- (iv) Calculate the heat change,  $q$ , for your experiment using the value you calculated in (b)(iii).

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

$$\begin{aligned} q &= -mc\Delta T \\ &= -(50.00 \times 1.00) \times 4.18 \times -2.25 \\ &= +470.25\text{ J} \\ &\approx +470\text{ J (3 s.f.)} \end{aligned}$$

$$q = \dots\dots\dots +470\text{ J} \quad [1]$$

- (v) Determine the molar enthalpy change,  $\Delta H_1$ , for the reaction in equation 4. The sodium hydrogencarbonate is in excess.

Include the sign of  $\Delta H_1$  in your answer.

$$\begin{aligned} n_{\text{citric acid}} &= \frac{50.00}{1000} \times 0.080 = 4.00 \times 10^{-3}\text{ mol} \\ \Delta H_1 &= \frac{q}{n} = \frac{+470.25}{4.00 \times 10^{-3}} \\ &= +117562.5\text{ J mol}^{-1} \\ &= +118\text{ kJ mol}^{-1} \end{aligned}$$

$$\Delta H_1 = \dots\dots\dots +118\text{ kJ mol}^{-1} \quad [5]$$

- (c) (i) The maximum temperature change,  $\Delta T_{\max}$ , can also be determined by direct measurement of the initial temperature and the lowest temperature reached only.

Explain why the graphical method is likely to give a more accurate value for  $\Delta T_{\max}$  than the direct measurement method.

The graphical method takes into account of the heat gained from the surrounding, hence  $\Delta T_{\max}$  will be more accurate.

..... [1]

- (ii) What would be the expected change in temperature obtained in (b)(iii) if the volume of citric acid used was halved? Explain your answer.

Since citric acid is the limiting reagent, using half the volume of citric acid means only half the original amount of citric acid is reacted, hence only half the heat will be released. However, the volume of solution to be heated is also halved. Hence the change in temperature will be the same.

[2]

[Total: 17]

#### 4 Qualitative Analysis

**S** is a mixture of two solids: **FA 6**, which is soluble in water and **FA 7**, which is insoluble in water. Each of **FA 6** and **FA 7** contains one cation and one anion.

**S** does not contain halide ions.

- (a) Carry out the following tests. Test and identify any gases evolved. Carefully record your observations in Table 4.1.

If there are no observable changes, write **no observable change**.

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

**Table 4.1**

	test	observations
1	Place all of solid <b>S</b> , into a boiling tube. Add 15 cm <sup>3</sup> of deionised water and shake for 2 minutes.  Filter the mixture, collecting the filtrate in a test-tube. Keep the filtrate for tests 2 and 3.  Wash the residue in the filter paper, with a little deionised water. Discard the washings. Keep the residue for test 4.	<b>colourless filtrate</b>  <b>white residue</b>
2	Add 1 cm depth of the filtrate to a test tube.  Add aqueous ammonia.	<b>no observable change</b>
3	Add 1 cm depth of the filtrate to a boiling tube.  Add aqueous sodium hydroxide.  Warm the mixture carefully for 2 minutes.  Add a small piece of aluminium foil and warm the mixture carefully.	<b>white ppt</b> was formed with NaOH(aq) and <b>insoluble in excess</b> .  <b>no observable change</b>  gas evolved turns <b>moist red litmus turns blue</b> .

4	<p>Using a spatula, transfer some of the solid residue from the filter paper into a test-tube.</p> <p>Add 2 cm depth of <b>FA 2</b> (<math>\text{H}_2\text{SO}_4</math>) to the same tube.</p> <p>Keep this solution for test 5.</p>	<p><b>Effervescence</b> observed. Gas forms <b>white ppt with limewater</b></p>
5	<p>To the resultant solution from test 4, add an equal volume of deionised water and shake the test-tube. Divide the liquid into two equal portions.</p> <p>To the first portion, add aqueous sodium hydroxide.</p> <p>To the second portion, add aqueous ammonia.</p>	<p><b>white ppt</b> was formed with <math>\text{NaOH(aq)}</math> and <b>soluble in excess</b> to form a colourless solution</p> <p><b>white ppt</b> was formed with <math>\text{NH}_3(\text{aq})</math> and <b>soluble in excess</b> to form a colourless solution.</p>

[4]

- (b) Identify the ions in **FA 6** and **FA 7** respectively and state the evidence by completing Table 4.2.

Table 4.2

		evidence
(i)	cation in <b>FA 6</b>	<p><b><math>\text{Ca}^{2+}</math></b>  <b>White ppt</b> of <math>\text{Ca(OH)}_2</math> was <b>formed with <math>\text{NaOH(aq)}</math> and not <math>\text{NH}_3(\text{aq})</math></b></p>
(ii)	cation in <b>FA 7</b>	<p><b><math>\text{Zn}^{2+}</math></b>  The <b>white ppt <math>\text{Zn(OH)}_2</math> dissolves in excess <math>\text{NaOH(aq)}</math> and <math>\text{NH}_3(\text{aq})</math> respectively</b> to give complexes <math>[\text{Zn(OH)}_4]^{2-}</math> and <math>[\text{Zn(NH}_3)_4]^{2+}</math>.</p>
(iii)	anion in <b>FA 7</b>	<p><b><math>\text{CO}_3^{2-}</math></b>  Carbonate react with acid to release <b><math>\text{CO}_2</math> gas</b> or which forms <math>\text{CaCO}_3</math> <b>white ppt with limewater</b>.</p>

[3]

- (c) Using your observation in Test 3, suggest the possible identities of the anion in **FA 6**. Hence describe a test to confirm the identity of the anion. You do not need to carry out this test.

$\text{NO}_3^-$  or  $\text{NO}_2^-$

To 1 cm depth of filtrate, add dilute hydrochloric acid (or  $\text{HNO}_3$  or  $\text{H}_2\text{SO}_4$ );. If **brown**

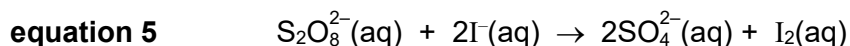
**fumes** are observed, then the anion is  $\text{NO}_2^-$ . If no brown fumes are observed, the

anion is  $\text{NO}_3^-$  [2]

[Total: 9]

## 5 Planning

Iodide ions,  $\text{I}^-$ , react with peroxodisulfate ions,  $\text{S}_2\text{O}_8^{2-}$ , as shown in equation 5.



If  $\text{I}^-(\text{aq})$  is used in large excess in the reaction, the simplified rate equation becomes

$$\text{rate} = k'[\text{S}_2\text{O}_8^{2-}]^a \quad \text{where } k' = k[\text{I}^-]$$

The order of reaction with respect to peroxodisulfate ions can be determined experimentally using the continuous method. Since iodine ( $\text{I}_2$ ) is formed during the reaction, the concentration of iodine can be measured continuously at various time intervals by sampling, quenching using cold water, followed by titration with aqueous sodium thiosulfate.

(a) You will prepare a reaction mixture consisting of

- 70  $\text{cm}^3$  of 0.400  $\text{mol dm}^{-3}$  aqueous potassium iodide, KI
- 30  $\text{cm}^3$  of 0.010  $\text{mol dm}^{-3}$  aqueous peroxodisulfate ions,  $\text{S}_2\text{O}_8^{2-}$  □

You are also provided with

- 0.003  $\text{mol dm}^{-3}$  sodium thiosulfate,  $\text{Na}_2\text{S}_2\text{O}_3$
- starch
- the equipment normally found in a school or college laboratory.

Plan an investigation to determine the order of reaction with respect to peroxodisulfate,  $\text{S}_2\text{O}_8^{2-}$ , using the continuous method.

In your plan you should include brief details of:

- the quantities of all reagents you would use,
- the apparatus and the procedure you would follow,
- the measurements you would take to allow for a suitable volume-time graph to be drawn from time,  $t = 0$  min to  $t = 20$  min.

1. Fill a **50.00  $\text{cm}^3$  burette** with sodium thiosulfate. ....
2. Using another **50.00  $\text{cm}^3$  burette**, add **30.0  $\text{cm}^3$  of peroxodisulfate ions** into a **250  $\text{cm}^3$  conical flask** labelled reaction mixture. Place the beaker in a thermostatically-controlled water bath kept at 30 °C. ....
3. Using a **100  $\text{cm}^3$  measuring cylinder**, measure **70.0  $\text{cm}^3$  of KI** to the same conical flask and place the measuring cylinder in the thermostatically-controlled water bath. ....

4. After 5 min, add KI into the 250 cm<sup>3</sup> beaker in the thermostatically-controlled water bath.

Start the stopwatch immediately and stir the mixture thoroughly using a glass rod to ensure even mixing.

5. Using a 10.0 cm<sup>3</sup> pipette, transfer a 10.0 cm<sup>3</sup> aliquot (portion) of the reaction mixture into the second 250 cm<sup>3</sup> conical flask.

6. At  $t = 2$  minute, add about 50 cm<sup>3</sup> of cold deionised water, using a 50 cm<sup>3</sup> measuring cylinder, to the second conical flask, and vigorously swirl the mixture.

Or Read and record the time of transfer in minutes and seconds, to the nearest second, when the quenching agent is added.

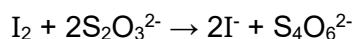
7. Immediately titrate the iodine in the second conical flask with sodium thiosulfate..When the colour of the solution turns pale yellow, use a dropping pipette to add about 1 cm<sup>3</sup> of starch indicator and continue titrating until the point when the solution just turns colourless. Record your titration results.

8. Repeat steps 4 to 6 (at 5<sup>th</sup>, 8<sup>th</sup>, 11<sup>th</sup>, 14<sup>th</sup> and 17<sup>th</sup> minute) until a total of six aliquots have been titrated and their results recorded.

[5]



- (b) (i)  $V_{\max}$  represents the volume of sodium thiosulfate required to react with the iodine produced when the reaction as shown by equation 5 has completed, in a  $10.0 \text{ cm}^3$  aliquot of the reaction mixture.



Show that  $V_{\max}$  is  $20.00 \text{ cm}^3$ .

total amount of  $\text{I}_2$  formed = amount of  $\text{S}_2\text{O}_8^{2-}$  used

$$= \frac{30}{1000} \times 0.010 = 3.00 \times 10^{-4} \text{ mol}$$

$$\text{amount of } \text{I}_2 \text{ in } 10 \text{ cm}^3 = 3.00 \times 10^{-4} \times \frac{10}{100} = 3.00 \times 10^{-5} \text{ mol}$$

amount of  $\text{S}_2\text{O}_3^{2-}$  needed =  $2 \times$  amount of  $\text{I}_2$

$$= 2 \times 3.00 \times 10^{-5} = 6.00 \times 10^{-5} \text{ mol}$$

$$\text{volume of } \text{S}_2\text{O}_3^{2-}, V_{\max} = \frac{6.00 \times 10^{-5}}{0.0030} = 0.0200 \text{ dm}^3 = \underline{\underline{20.0 \text{ cm}^3}}$$

[2]

- (ii) Using information from (b) (i), sketch on Fig. 5.1 an appropriately labelled graph to show that the reaction is first order with respect to peroxodisulfate ions.

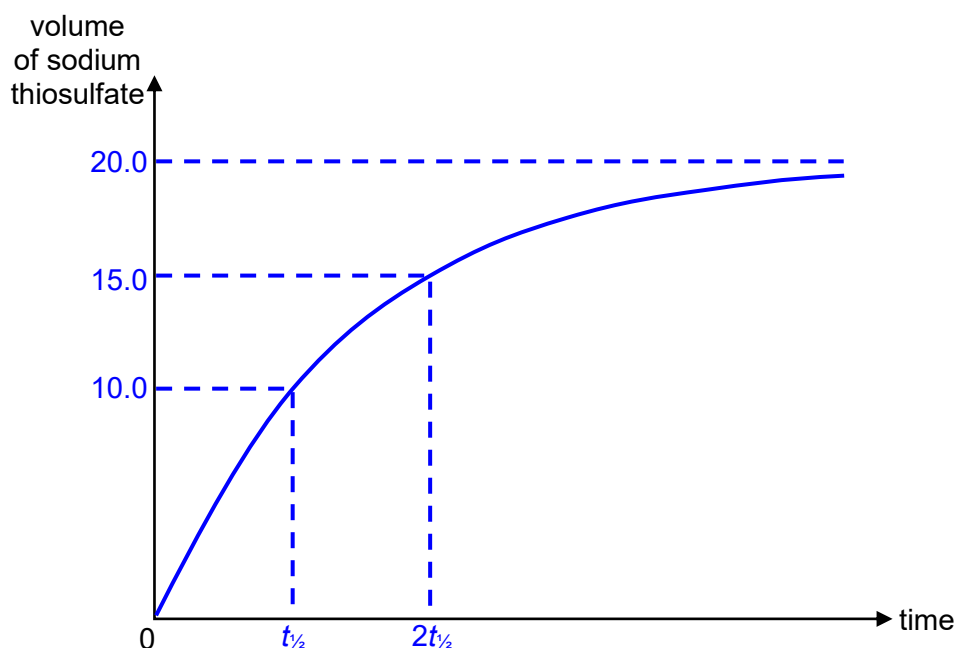


Fig. 5.1

[2]

(c) The reaction between peroxodisulfate ions and iodide ions as shown in equation 5 is slow. The reaction can be sped up by the addition of iron(II),  $\text{Fe}^{2+}$  catalyst.

(i) Suggest why the reaction between peroxodisulfate ions and iodide ions is slow at room temperature.

The activation energy for the reaction is high as it involves the collision between two anions, which repel each other electrostatically.

..... [1]

(ii) The experimental procedure in (a) is modified to determine the order of reaction with respect to the iron(II) catalyst in the reaction between peroxodisulfate and iodide ions.

Two additional experiments were conducted with iron(II) catalyst added at the beginning.

You are provided with  $50 \text{ cm}^3$  of  $0.01 \text{ mol dm}^{-3}$  iron(II) nitrate solution,  $\text{Fe}(\text{NO}_3)_2$ .

- Complete Table 5.1 with suitable volumes of reagents used.
- Describe how you would make use of the data collected to determine the order of reaction with respect to the iron(II) catalyst.

**Table 5.1**

expt	volume of $\text{S}_2\text{O}_8^{2-}$ / $\text{cm}^3$	volume of $\text{I}^-$ / $\text{cm}^3$	volume of $\text{H}_2\text{O}$ / $\text{cm}^3$	volume of $\text{Fe}^{2+}$ / $\text{cm}^3$
1	30	70	0	10
2	30	70	5	5

Order of reaction can be determined by plotting the graph of volume of  $\text{S}_2\text{O}_3^{2-}$  against time. Draw a tangent at  $t = 0$  and calculate the gradient to find the initial rate of reaction for each experiment.

$$\frac{\text{rate}_1}{\text{rate}_2} = \left( \frac{[\text{Fe}^{2+}]_{\text{expt 1}}}{[\text{Fe}^{2+}]_{\text{expt 2}}} \right)^n \Rightarrow \left( \frac{\text{gradient}_1}{\text{gradient}_2} \right) = \left( \frac{V_{\text{Fe}^{2+}, \text{expt 1}}}{V_{\text{Fe}^{2+}, \text{expt 2}}} \right)^n$$

Solve for  $n$ . or state comparison of initial rate

.....

.....

.....

.....

.....

..... [3]

[Total: 13]



## Qualitative Analysis Notes

[ppt. = precipitate]

## (a) Reactions of aqueous cations

<b>cation</b>	<b>reaction with</b>	
	NaOH(aq)	NH <sub>3</sub> (aq)
aluminium, Al <sup>3+</sup> (aq)	white ppt. soluble in excess	white ppt. insoluble in excess
ammonium, NH <sub>4</sub> <sup>+</sup> (aq)	ammonia produced on heating	–
barium, Ba <sup>2+</sup> (aq)	no ppt. (if reagents are pure)	no ppt.
calcium, Ca <sup>2+</sup> (aq)	white. ppt. with high [Ca <sup>2+</sup> (aq)]	no ppt.
chromium(III), Cr <sup>3+</sup> (aq)	grey-green ppt. soluble in excess giving dark green solution	grey-green ppt. insoluble in excess
copper(II), Cu <sup>2+</sup> (aq),	pale blue ppt. insoluble in excess	blue ppt. soluble in excess giving dark blue solution
iron(II), Fe <sup>2+</sup> (aq)	green ppt., turning brown on contact with air insoluble in excess	green ppt., turning brown on contact with air insoluble in excess
iron(III), Fe <sup>3+</sup> (aq)	red-brown ppt. insoluble in excess	red-brown ppt. insoluble in excess
magnesium, Mg <sup>2+</sup> (aq)	white ppt. insoluble in excess	white ppt. insoluble in excess
manganese(II), Mn <sup>2+</sup> (aq)	off-white ppt., rapidly turning brown on contact with air insoluble in excess	off-white ppt., rapidly turning brown on contact with air insoluble in excess
zinc, Zn <sup>2+</sup> (aq)	white ppt. soluble in excess	white ppt. soluble in excess

**(b) Reactions of anions**

<b><i>anion</i></b>	<b><i>reaction</i></b>
carbonate, $\text{CO}_3^{2-}$	$\text{CO}_2$ liberated by dilute acids
chloride, $\text{Cl}^-(\text{aq})$	gives white ppt. with $\text{Ag}^+(\text{aq})$ (soluble in $\text{NH}_3(\text{aq})$ )
bromide, $\text{Br}^-(\text{aq})$	gives pale cream ppt. with $\text{Ag}^+(\text{aq})$ (partially soluble in $\text{NH}_3(\text{aq})$ )
iodide, $\text{I}^-(\text{aq})$	gives yellow ppt. with $\text{Ag}^+(\text{aq})$ (insoluble in $\text{NH}_3(\text{aq})$ )
nitrate, $\text{NO}_3^-(\text{aq})$	$\text{NH}_3$ liberated on heating with $\text{OH}^-(\text{aq})$ and $\text{Al}$ foil
nitrite, $\text{NO}_2^-(\text{aq})$	$\text{NH}_3$ liberated on heating with $\text{OH}^-(\text{aq})$ and $\text{Al}$ foil; $\text{NO}$ liberated by dilute acids (colourless $\text{NO} \rightarrow$ (pale) brown $\text{NO}_2$ in air)
sulfate, $\text{SO}_4^{2-}(\text{aq})$	gives white ppt. with $\text{Ba}^{2+}(\text{aq})$ (insoluble in excess dilute strong acids)
sulfite, $\text{SO}_3^{2-}(\text{aq})$	$\text{SO}_2$ liberated with dilute acids; gives white ppt. with $\text{Ba}^{2+}(\text{aq})$ (soluble in dilute strong acids)

**(c) Tests for gases**

<b><i>gas</i></b>	<b><i>test and test result</i></b>
ammonia, $\text{NH}_3$	turns damp red litmus paper blue
carbon dioxide, $\text{CO}_2$	gives a white ppt. with limewater (ppt. dissolves with excess $\text{CO}_2$ )
chlorine, $\text{Cl}_2$	bleaches damp litmus paper
hydrogen, $\text{H}_2$	"pops" with a lighted splint
oxygen, $\text{O}_2$	relights a glowing splint
sulfur dioxide, $\text{SO}_2$	turns aqueous acidified potassium manganate(VII) from purple to colourless

**(d) Colour of halogens**

<b><i>halogen</i></b>	<b><i>colour of element</i></b>	<b><i>colour in aqueous solution</i></b>	<b><i>colour in hexane</i></b>
chlorine, $\text{Cl}_2$	greenish yellow gas	pale yellow	pale yellow
bromine, $\text{Br}_2$	reddish brown gas / liquid	orange	orange-red
iodine, $\text{I}_2$	black solid / purple gas	brown	purple