



YISHUN INNOVA JUNIOR COLLEGE
JC 2 PRELIMINARY EXAMINATION
Higher 2

CANDIDATE
NAME

CG

INDEX NO

CHEMISTRY

Paper 4 Practical

9476/04

13 August 2026

2 hours 30 minutes

You must answer on the question paper.

You will need: The materials and apparatus listed in the confidential instructions.

INSTRUCTIONS

- Answer **all** questions.
- Use a black or dark blue pen. You may use an HB pencil for any diagrams or graphs.
- Write your name, class and index number in the boxes at the top of the page.
- Write your answer to each question in the space provided.
- Do **not** use an erasable pen. Do **not** use correction fluid or tape.
- You may use an approved calculator.
- Qualitative analysis notes are printed in the question paper.
- Write the details of the shift and laboratory in the boxes provided.

INFORMATION

- The total mark for this paper is 50.
- The number of marks for each question or part question is shown in brackets [].

Shift
Laboratory

For Examiner's Use	
1	/ 16
2	/ 25
3	/ 9
Total	/ 50

This document has **24** pages. Any blank pages are indicated.

Answer **all** the questions in the spaces provided.

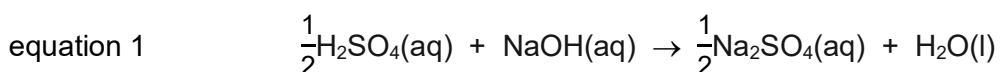
1 Determination of an enthalpy change of neutralisation, ΔH_{neut} , by thermometric titration

FA 1 is 1.80 mol dm⁻³ sulfuric acid, H₂SO₄.

FA 2 is aqueous sodium hydroxide, NaOH.

The enthalpy change of neutralisation between an acid and an alkali can be determined using thermometric titration. This involves using a fixed volume of alkali with progressive addition of small volumes of acid and monitoring the temperature of the resulting solution.

In this experiment, you are to determine the enthalpy change of neutralisation for the reaction as shown in equation 1 using thermometric titration.



(a) Prepare a table in the space provided on page 4 and record, to the appropriate level of precision:

- all volumes of **FA 1** added, V
- the maximum temperature, T , reached after each addition of acid, **FA 1**.

It is important that the volume of **FA 1** recorded is the total volume you have added up to the point when the temperature reading was made.

Note: If you overshoot on an addition, record the actual total volume of **FA 1** added up to that point.

Procedure

1. Place one polystyrene cup inside a second polystyrene cup and place both cups in a 250 cm³ beaker.
2. Use a pipette to transfer 25.0 cm³ of **FA 2** into the first polystyrene cup.
3. Fill a burette with **FA 1**.
4. Stir the **FA 2** in the cup gently with the thermometer. Read and record its temperature.
5. Use the burette to add 5.00 cm³ of **FA 1** into the cup. Stir the mixture gently with the thermometer. Read and record both the maximum temperature and the actual total volume of **FA 1** added.
6. Repeat step 5 until a total of 45.00 cm³ of **FA 1** has been added. For each addition of **FA 1**, read and record both the maximum temperature and the actual total volume of **FA 1** added up to that point.

Question 1 continues on the next page.

Results

Volume of FA 1 added, V / cm^3	Maximum temperature, $T / ^\circ\text{C}$
0.00	29.8
5.00	32.2
10.00	34.0
15.00	35.2
20.00	36.0
25.00	36.8
30.00	35.8
35.00	35.2
40.00	34.6
45.00	34.0

- M1** Table has correct headers and units, all volumes recorded to 0.05 cm^3 , temperatures to nearest $0.1 ^\circ\text{C}$.
- M2** Tabulates 10 sets of temperature and volume of FA1 from 0.00cm^3 to 45.00cm^3
- M3** Temperature increases from 5.00 to 25.00 cm^3 and decreases after 25.00 cm^3 of **FA 1** is added.

[3]

- (b) Plot a graph of temperature, T , on the y -axis, against total volume of **FA 1**, V , on the x -axis on the grid in Figure 1.1. Your scale on the y -axis should allow for extrapolation above the highest temperature recorded.

Draw two lines of best-fit, taking into account the points when the temperature of the mixture was rising and the points when the temperature was falling. Each line should have a shape best suited to its plotted points.

Extrapolate (extend) both lines until they intersect.

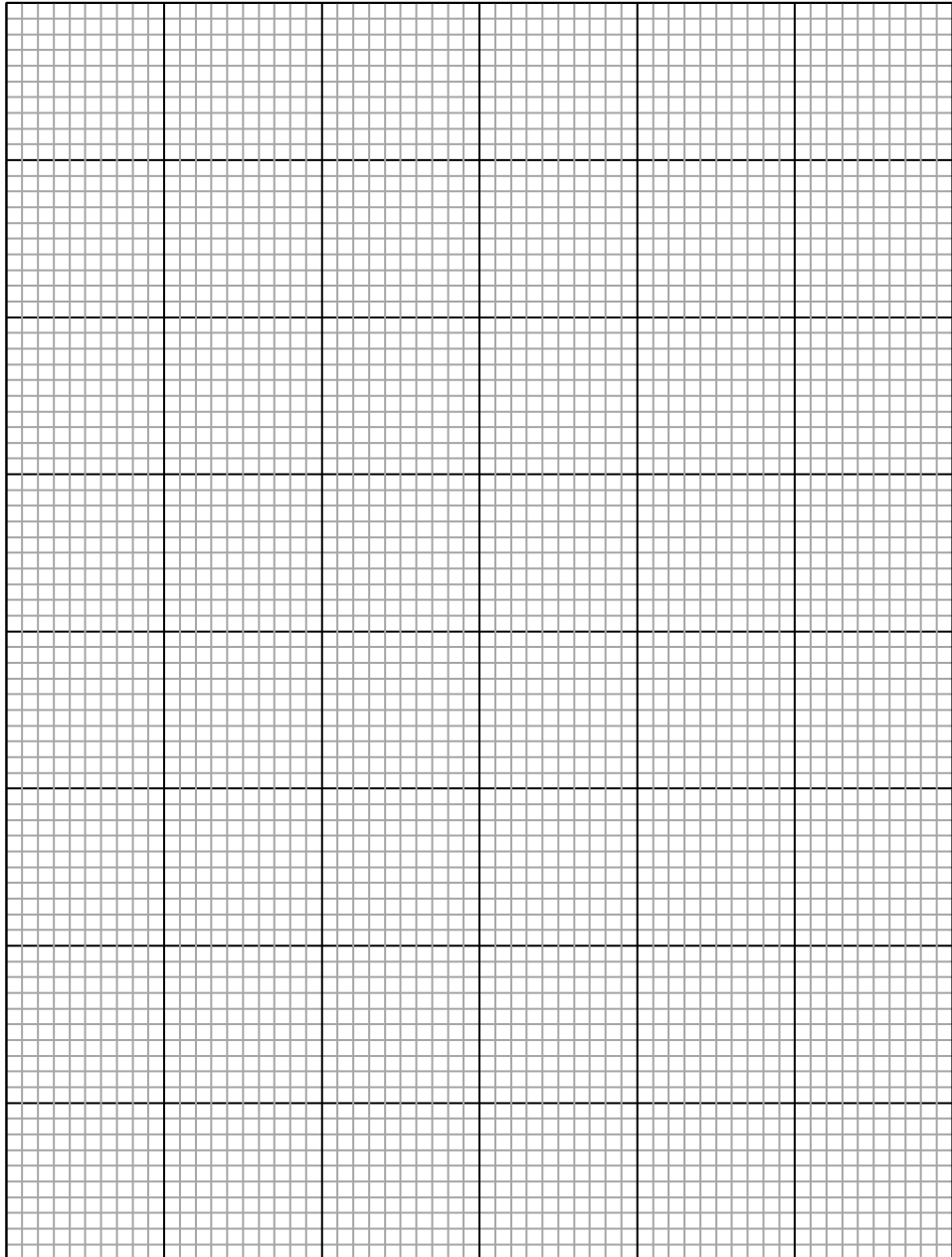
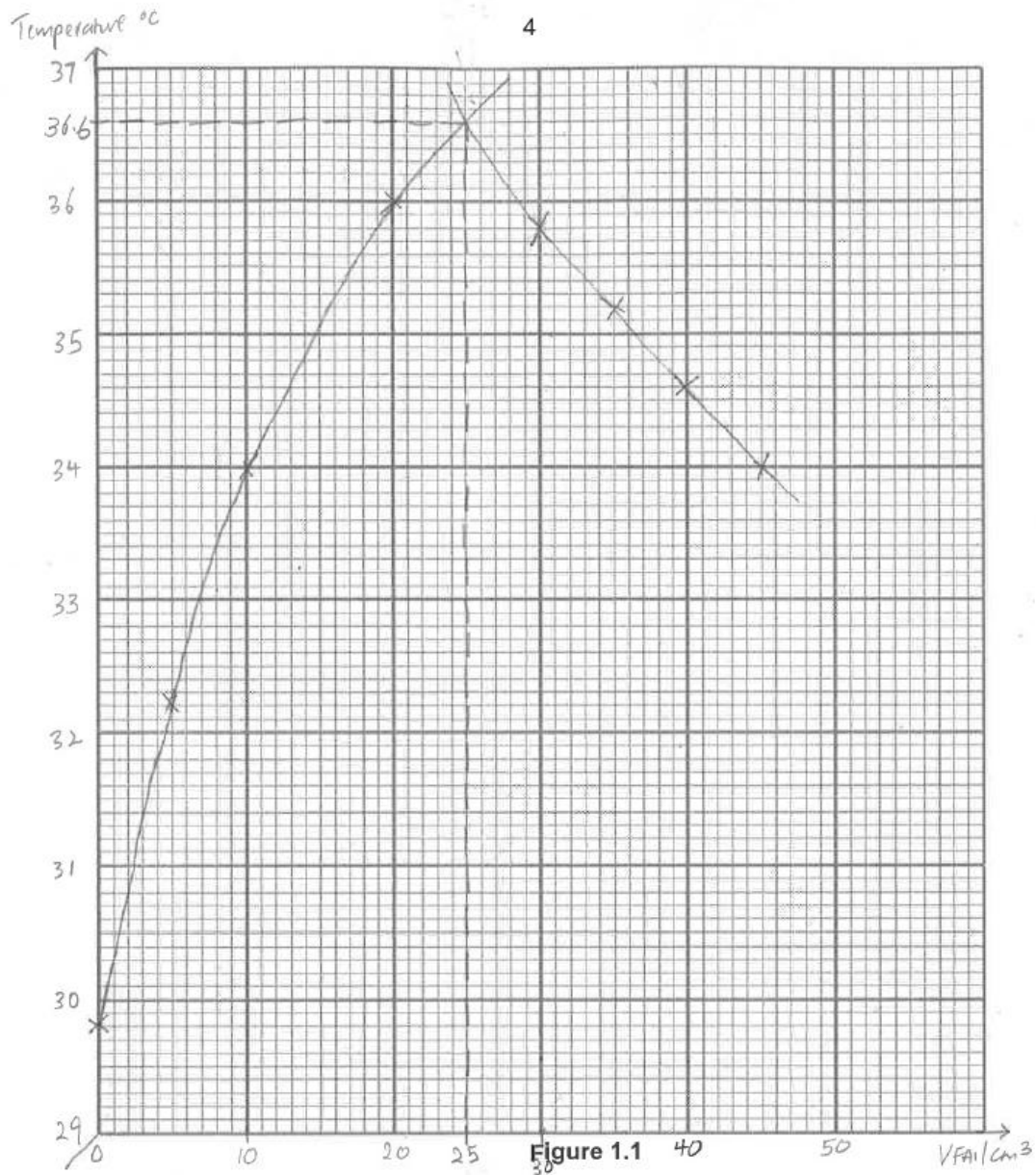


Figure 1.1

[3]



- M4** Axes correct way round + correct labels + units + scale (scale must be chosen to allow for the lines to be extrapolated to intersect)
Graph must occupy more than half the grid for both axes.
- M5** Plotting – all points within $\pm \frac{1}{2}$ small square.
- M6** Both graph lines are best-fit lines **and** lines are correctly extrapolated to intersect.

- (c) From your graph, read the initial temperature of **FA 2**, T_{initial} , and the maximum temperature of the mixture, T_{max} .

Use these values to calculate the temperature change in the reaction, ΔT .

Read the volume of **FA 1** added, V_{neut} , at the maximum temperature of the mixture.

Record all these values below.

T_{initial} =
 T_{max} =
 ΔT =
 V_{neut} = [2]

M7 T_{initial} , T_{max} and V_{neut} **correctly read** from the graph to $\pm 1/2$ small square.
Do not accept if the T_{initial} is taken from the table.

M8 ΔT calculated correctly (T_{initial} may come from graph plot or table)

- (d) (i) Calculate the heat change, q , when **FA 1** has completely neutralised 25.0 cm³ of sodium hydroxide, using the values you calculated in (c).

You should assume that the specific heat capacity of the solution is 4.18 J g⁻¹ K⁻¹, and that the density of the solution is 1.00 g cm⁻³.

$$\text{Total volume} = 25 + 25 = 50 \text{ cm}^3$$

$$\begin{aligned}
 q &= mc\Delta T \\
 &= 50 \times 4.18 \times 6.8 \\
 &= 1421\text{J} \\
 &= 1420\text{J}
 \end{aligned}$$

M9 correct m

M10 correct of values into $mc\Delta T$

Ignore sign on q

q = [2]

- (ii) Determine the enthalpy change of neutralisation, ΔH_{neut} , for the reaction.

Include the sign of ΔH_{neut} in your answer.

$$\begin{aligned}\text{No of moles of H}_2\text{SO}_4 &= 1.80 \times \frac{25}{1000} \\ &= 0.0450 \text{ mol}\end{aligned}$$

$$\text{M11 No. of moles of water} = 2 \times \text{no. of moles of H}_2\text{SO}_4 = 2 \times 0.0450 = 0.0900 \text{ mol}$$

$$\begin{aligned}\Delta H_{\text{neut}} &= -(q \div \text{moles of water}) \\ &= -(1421 \times 10^{-3} \div 0.0900) \\ &= -15.79 \\ &= -15.8 \text{ kJ mol}^{-1}\end{aligned}$$

M12

Allow ecf

If no negative sign, no marks

$$\Delta H_{\text{neut}} = \dots\dots\dots [2]$$

- (e) Explain why the temperature, T , first increases and then decreases as the total volume of **FA 1** added, V , increases.

[2]

M13 As more acid was added, **more water is formed**, T increases, as the neutralisation reaction **is exothermic**.

M14 NaOH has been **completely neutralised/reacted**. T of the mixture decreases as **heat spread over a larger volume**.

- (f) A student carried out the experiment and obtained a value of $-60.0 \text{ kJ mol}^{-1}$ for ΔH_{neut} .

The published value for this enthalpy change is $-57.0 \text{ kJ mol}^{-1}$.

- (i) Express the difference between these two values as a percentage of the published value.

[1]

$$\text{Difference} = |-60.0 - (-57.0)| = 3.0 \text{ kJ mol}^{-1}$$

$$\text{Percentage difference} = \frac{3.0}{57.0} \times 100 = \mathbf{5.26\% \text{ M15}}$$

- (ii) Explain whether the omission of heat loss to surroundings could have been the reason for the discrepancy between the value obtained by the student and the one published.

[1]

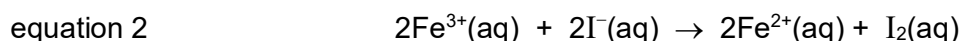
No. Since heat loss will result in **lower recorded maximum temperature** OR **smaller ΔT_{max} value calculated** OR **smaller heat change calculated**, ΔH_{neut} would have been less exothermic. **M16**

[Total:16]

2 Determination of the kinetics of redox reactions

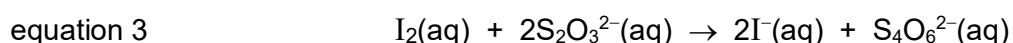
Iron(III) ions can be involved in redox reactions.

In this question, you will study the kinetics of the redox reaction between iron (III) ions and iodide ions, as shown in equation 2.



When starch is added to the reaction mixture, a blue colour is seen immediately due to the formation of an iodine–starch complex.

If a small amount of sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, is also present in the reaction mixture, the formation of the blue colour is delayed as the iodine, I_2 , produced reacts immediately with $\text{Na}_2\text{S}_2\text{O}_3$, as shown in equation 3.



When all the $\text{Na}_2\text{S}_2\text{O}_3$ has reacted, the I_2 which continues to be produced will turn the starch indicator blue.

FA 3 is $0.0200 \text{ mol dm}^{-3}$ aqueous iron(III) chloride, FeCl_3 .

FA 4 is $0.0500 \text{ mol dm}^{-3}$ aqueous potassium iodide, KI .

FA 5 is $0.00500 \text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$.

Solution S is starch solution.

You will perform a series of three experiments. You will add a fixed amount of sodium thiosulfate, **FA 5**, to each of your experiments. You will add an amount of deionised water so that the total volume is 60 cm^3 . The time taken for the blue colour to form allows the reaction rate to be determined.

(a) Fill Table 2.1 with the volume of deionised water needed in each experiment.

Table 2.1

experiment	FA 3 / cm^3	FA 4 / cm^3	FA 5 / cm^3	solution S / cm^3	deionised water / cm^3	time, t / s
1	20.0	10.0	15.0	10.0	5.0	16.6
2	10.0	10.0	15.0	10.0	15.0	31.3
3	25.0	5.0	15.0	10.0	5.0	53.1

[4]

M17	Correct volume of deionised water for all 3 experiments
M18	Volume of deionised water to 1d.p and time to 0.1s.
M19	Accuracy
M20	Accuracy

Experiment 1

The end-point of the reaction is the **first** appearance of a blue colour.

Note: Insufficient stirring of the reaction mixture may lead to a blue colour appearing before the true end-point is reached.

1. Use an appropriate measuring cylinder to transfer 10.0 cm³ of **FA 4** into a 250 cm³ beaker. Place the beaker on a white tile.
2. Using appropriate measuring cylinders, transfer to the beaker
 - 15 cm³ of **FA 5**,
 - 10 cm³ **solution S**,
 - appropriate volume of deionised water recorded in Table 2.1.
3. Stir the contents of the beaker using a glass rod.
4. Use the measuring cylinder labelled **FA 3** to transfer 20.0 cm³ of **FA 3** to the beaker. Start the stopwatch during this addition.
5. Mix the contents in the beaker by **thoroughly** stirring.
6. Stop the stopwatch when the solution **first** turns blue.
7. Record the time taken, t , to the nearest 0.1 s in Table 2.1.
8. Wash out the beaker and stand it upside down on a paper towel to drain.

Experiments 2 and 3

Repeat experiment 1 using the volumes of **FA 3**, **FA 4**, **FA 5**, **solution S** and deionised water given in Table 2.1.

In each case, you will need to ensure that the **same total volume** of reaction mixture is used by adding deionised water as required.

You should alternate the use of two 250 cm³ beakers.

Record all values of t in Table 2.1.

- (b) (i) Calculate the initial concentrations, in mol dm^{-3} , of $\text{S}_2\text{O}_3^{2-}$ ions, Fe^{3+} ions and I^- ions for experiment 1.

$$[\text{S}_2\text{O}_3^{2-}] \text{ in the reaction mixture} = \frac{15}{1000} \times 0.0500 \div \frac{60}{1000} = 1.25 \times 10^{-3} \text{ mol dm}^{-3}$$

$$[\text{Fe}^{3+}] = \frac{20 \times 0.02}{60} = 6.67 \times 10^{-3} \text{ mol dm}^{-3}$$

$$[\text{I}^-] = \frac{10 \times 0.02}{60} = 8.33 \times 10^{-3} \text{ mol dm}^{-3}$$

M21 for all 3 concentrations

initial $[\text{S}_2\text{O}_3^{2-}] =$

initial $[\text{Fe}^{3+}] =$

initial $[\text{I}^-] =$

[1]

- (ii) The expression below shows the experimental rate of this reaction as the change in concentration of $\text{S}_2\text{O}_3^{2-}$ ions per unit time. At the end-point of the reaction, $[\text{S}_2\text{O}_3^{2-}]$ is zero.

$$\text{experimental rate} = \frac{\text{change in } [\text{S}_2\text{O}_3^{2-}]}{\text{time, } t}$$

Complete Table 2.2 by calculating the experimental rates for each experiment.

If you were unable to calculate a value for the initial $[\text{S}_2\text{O}_3^{2-}]$ in **(b)(i)**, you should assume it is $2.00 \times 10^{-3} \text{ mol dm}^{-3}$. Note that this is not the correct value.

Working:

$$\text{rate} = \frac{1.25 \times 10^{-3}}{16.6} = 7.53 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$$

Table 2.2

experiment	FA 3 / cm^3	FA 4 / cm^3	experimental rate / $\text{mol dm}^{-3} \text{ s}^{-1}$
1	20.0	10.0	7.53×10^{-5}
2	10.0	10.0	3.99×10^{-5}
3	25.0	5.0	2.35×10^{-5}

[1]

M22	Correct calculated values for rate
------------	------------------------------------

The rate equation for the reaction shown in equation 2 is

$$\text{rate} = k[\text{Fe}^{3+}]^a[\text{I}^-]^b$$

- (iii) Deduce the order, a , with respect to $[\text{Fe}^{3+}]$. Explain your deduction.

Comparing experiment 1 and 2

$$\frac{7.53 \times 10^{-5}}{3.99 \times 10^{-5}} = \frac{(20.0)^a}{(10.0)^a}$$

$$1.89 = 2^a$$

$$\Rightarrow a = 0.9 \approx 1 \text{ M23}$$

OR

Comparing Experiments 1 and 2,

When $[\text{Fe}^{3+}]$ is **doubled** from Experiment 1 to Experiment 2, rate is **approximately doubled**.

Hence, the rate is directly proportional to $[\text{Fe}^{3+}]$, so the order with respect to $[\text{Fe}^{3+}]$ is 1.

$$a = \dots\dots\dots [1]$$

- (iv) Deduce the order, b , with respect to $[\text{I}^-]$. Explain your deduction.

Comparing Experiment 2 and 3,

$$\frac{3.99 \times 10^{-5}}{2.35 \times 10^{-5}} = \frac{(10.0)^1 (10.0)^b}{(25.0)^1 (5.0)^b}$$

$$4.24 = 2^b$$

$$\Rightarrow b = 2.08 \approx 2 \text{ M24}$$

$$b = \dots\dots\dots [1]$$

- (v) Write the expression for the rate equation.

[1]

$$\text{rate} = k[\text{Fe}^{3+}][\text{I}^-]^2 \text{ M25}$$

- (vi) Hence, calculate the rate constant, k , and state its units.

$$\text{rate} = k[\text{Fe}^{3+}][\text{I}^-]^2$$

$$k = \frac{7.53 \times 10^{-5}}{6.667 \times 10^{-3} \times (8.33 \times 10^{-3})^2} = 163 \text{ mol}^{-2} \text{ dm}^6 \text{ s}^{-1}$$

M26 calculates k , allow ecf on error in rate equation and (e)(ii).

M27 correctly deduces units for k , allow ecf on error in rate equation

Correct significant figures (3 sf) in 1(d)(i), 1(d)(ii), 1(f)(i), 2(b)(i), 2(b)(ii), 2(b)(vi)	M28
Show working in all calculations in 1(d)(i), 1(d)(ii), 1(f)(i), 2(b)(i), 2(b)(ii), 2(b)(iii), 2(b)(iv), 2(b)(vi)	M29
correct units in 1(c), 1(d)(i), (d)(ii)	M30

$$k = \dots\dots\dots [5]$$

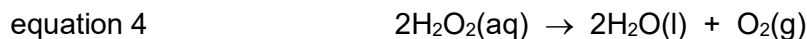
- (c) State, and explain, how the value of t would be affected if 15.0 cm^3 of a sodium thiosulfate solution of lower concentration was used in place of **FA 5**.

[1]

The reaction time should be shorter (or decreases) since less/ smaller amount of thiosulfate ions is present to react with iodine. **M31**

Question 2 continues on the next page.

- (d) Iron(III) ions can also act as a catalyst in the decomposition of hydrogen peroxide, as shown in equation 4.



You will plan an experiment to study the kinetics of this redox reaction by monitoring the volume of oxygen gas produced over time.

The order of reaction with respect to hydrogen peroxide may be confirmed to be first order graphically by plotting $V_{\text{final}} - V_t$ against time, where:

- V_{final} is the final volume of oxygen gas measured when the decomposition is complete
- V_t is the volume of oxygen gas collected at time = t .

You may assume that you are provided with the following materials:

- 0.100 mol dm⁻³ H₂O₂(aq) solution
 - **FA 3** solution
 - 100 cm³ gas syringe
 - the equipment normally found in a school or college laboratory.
- (i) Write a plan for a single experiment to obtain sufficient data for a graph of $V_{\text{final}} - V_t$ against time to be plotted. You may draw a labelled diagram of the set-up that could be used to carry out this experiment.

In your plan, you should include brief details of:

- calculations to justify the volume of H₂O₂(aq) you will use, assuming that all gases are measured at room temperature with molar gas volume of 24 dm³ mol⁻¹
- the apparatus you would use
- the quantities you would use
- the procedure you would follow
- the measurements you would make.

[5]

Pre-calculations:

Maximum volume of O₂ = 100 cm³

amount of O₂ = $\frac{100}{24000} = 0.0041667$ mol

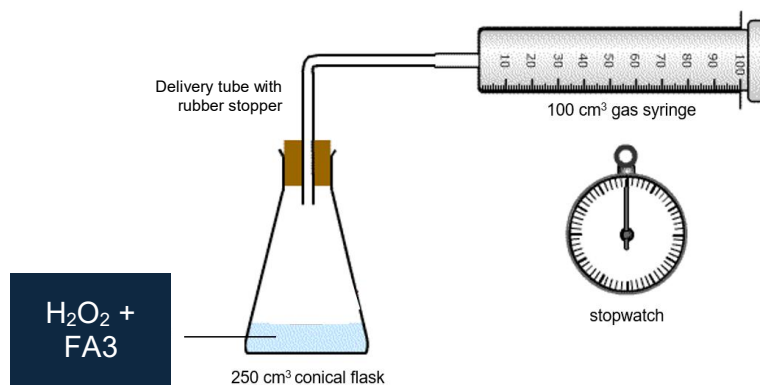
amount of H₂O₂ = 2 × O₂ = **0.00833333 mol**

Volume of H₂O₂ = $\frac{0.00833333}{0.100} \times 1000 = \underline{\underline{83.3 \text{ cm}^3}}$

Suitable volume of H₂O₂ to use = 80.0 cm³

Procedures:

1. Using a 100 cm³ measuring cylinder, transfer 80.0 cm³ of 0.100 mol dm⁻³ hydrogen peroxide into a 250 cm³ conical flask.
2. Assemble the set-up as shown in the diagram.
3. Record the initial reading of the gas syringe (V_0).
4. Using a teat pipette, add 2 cm³ of **FA 3** into the flask as quickly as possible. Immediately stopper the flask, swirl the mixture and start the stopwatch simultaneously.
5. Record the gas syringe reading at 0.5 min intervals (or other suitable regular intervals = V_t) until the gas syringe reading remains constant (V_{final}).
6. Stop the stopwatch and record the time.

Diagram:

Mark Requirements**Pre-calculations**

- M32**
- selecting an appropriate volume of $O_2(g)$ to collect, that must not exceed 100 cm^3 capacity of gas syringe
 - correct calculation of volume of H_2O_2 to use

calculations **MUST** be shown.

Alternative approach: Proposes a suitable volume of $H_2O_2(aq)$ and shows, by calculation, that the expected volume of $O_2(g)$ produced is appropriate for collection using a 100 cm^3 gas syringe.

M33 Appropriate apparatus and capacity

- delivery tube with stopper or rubber bung
- gas syringe
- stopwatch
- 50 cm^3 measuring cylinder / burette to measure H_2O_2
- 250 cm^3 conical flask
- teat pipette / measuring cylinder to measure **FA 3** (BOD this if measuring cylinder was mentioned one time minimally)

Well-labelled drawing of diagrams are acceptable.

M34 Appropriate quantities of chemicals

- specify volume of H_2O_2 to be used (should not exceed 83.3 cm^3 of H_2O_2 which will produce more than 100 cm^3 of O_2)*
- specify a reasonable volume of **FA 3** to be used (quantity should be lower than volume of $H_2O_2 = 83.3\text{ cm}^3$)

ecf is given for students who did not obtain M32.

M35 General Procedures

- **add** the **FA 3** to the reaction mixture as quickly as possible and
- **stopper** the flask immediately.
- **shake/swirl** the mixture
- **starting stopwatch immediately** upon adding the catalyst **FA 3**
- **stop stopwatch** upon **completion of reaction** (eg. until volume of gas collected stays constant) or **indicate end time** (until 10 minutes)

M36 Measurements

Record the **volume of gas produced** at **regular time intervals (V_t)** until there are no further changes in the gas syringe reading

- (ii) Sketch the graph that you would expect to obtain on the axes in Figure 2.1.

Explain how you would use your graph to show that the reaction is first order with respect to hydrogen peroxide.

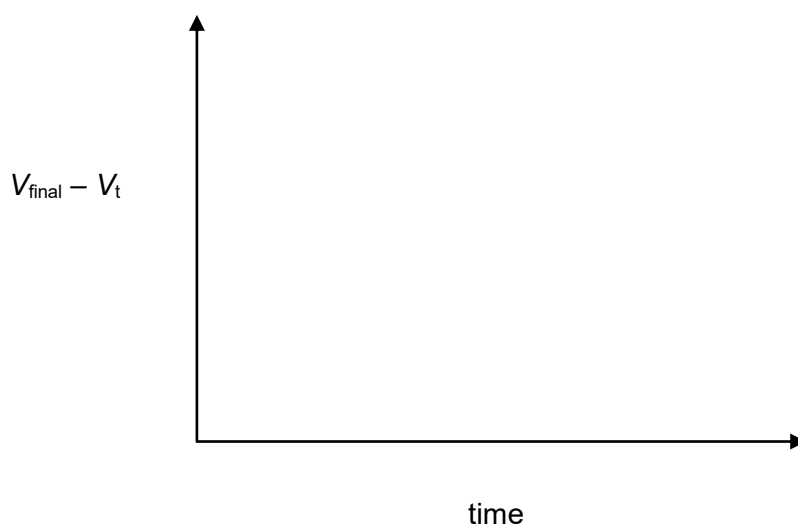
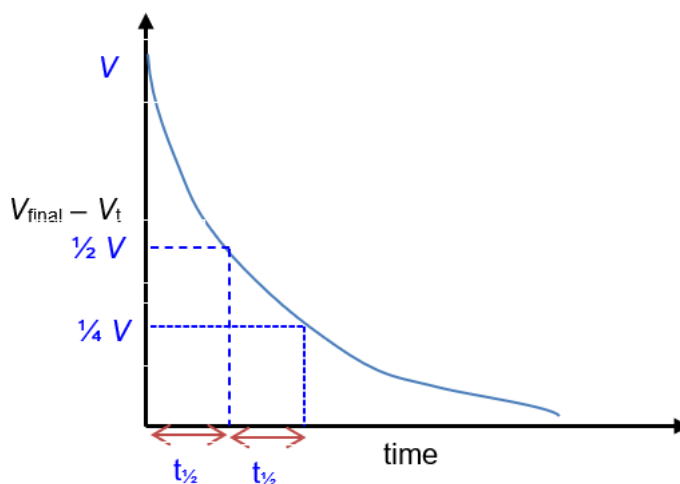


Figure 2.1

[2]



Graph (includes labels for finding $t_{1/2}$) **M37**

Find the half-life from the graph.

A **constant half-life** would indicate that the order with respect to $[\text{H}_2\text{O}_2]$ is one. **M38**

Most students do not give labelled diagrams (explanations in words were often lacking and unable to clearly to explain how half-life is derived) – no credit for M37 despite correct shape of graph. Students are awarded M38 as long as ‘constant half-life’ is mentioned except when they contradict themselves, ie. first order as it is a straight line graph and half life is constant.

- (iii) The oxygen gas evolved can also be collected by displacement of water.

State **one** advantage of collecting and measuring the gas in a gas syringe instead of over water.

.....
..... [1]

No loss of oxygen gas due to dissolving in water. **M39**

Alternative answer:

Oxygen is partially soluble in water hence, this caused volume of oxygen collected to be lower than expected while collection using a gas syringe gives **reliable/accurate** results (or allows actual volume of oxygen to be collected) = full credit

- (iv) With reference to the electrode potentials given in Table 2.3, suggest a two-step reaction sequence which describes the role of $\text{Fe}^{3+}(\text{aq})$ in the decomposition of hydrogen peroxide.

Table 2.3

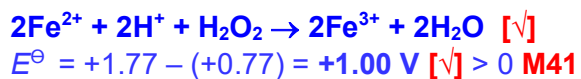
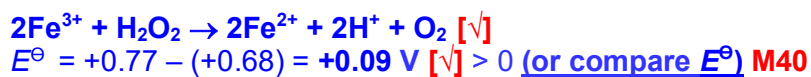
electrode reaction	$E^\ominus / \text{V}$
$\text{Fe}^{3+} + 3\text{e}^- \rightleftharpoons \text{Fe}$	-0.04
$\text{Fe}^{3+} + \text{e}^- \rightleftharpoons \text{Fe}^{2+}$	+0.77
$\text{H}_2\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$	+1.77
$\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}_2$	+0.68

.....

.....

.....

..... [2]



[Total: 25]

3 Identification of two cations and two anions

FA 6 and **FA 7** each contain one cation and one anion. The ions in the two compounds are different.

The volumes given in Table 3.1 are approximate and should be estimated rather than measured.

Test any gases evolved during the tests.

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

(a) Carry out the tests in Table 3.1.

Carefully record your observations **at each stage** of the tests in Table 3.1.

Table 3.1

tests		observations
1	Transfer the FA 6 provided into a boiling tube and add 3 cm depth of dilute nitric acid. Filter the mixture into another clean test-tube. Discard the filter paper with the residue.	- Effervescence observed [✓] and gas gives a white precipitate with limewater. [✓] - Filtrate: colourless filtrate [✓]
2	Add 1 cm depth of filtrate from test 1 into a clean test-tube. Add aqueous ammonia slowly with shaking, until the aqueous ammonia is in excess. Leave to stand for several minutes.	- Off-white precipitate, insoluble in excess $\text{NH}_3(\text{aq})$. [✓] - Off-white precipitate turns brown [✓] on contact with air.
3	Add 1 cm depth of FA 7 into a clean 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.	- White precipitate is formed. [✓] - White precipitate is soluble in excess $\text{NH}_3(\text{aq})$. [✓]
4	Add 1 cm depth of FA 7 into a clean test-tube. Add aqueous sodium hydroxide slowly with shaking, until the aqueous sodium hydroxide is in excess.	White precipitate, soluble in excess aqueous sodium hydroxide. [✓]
5	Add 1 cm depth of FA 7 into a clean test-tube. Add aqueous ammonia slowly with shaking, until the aqueous ammonia is in excess. Add dilute nitric acid dropwise with shaking, until the dilute nitric acid is in excess.	- White precipitate, soluble in excess $\text{NH}_3(\text{aq})$. [✓] - White precipitate formed upon adding HNO_3 dropwise, soluble in excess HNO_3. [✓]

[5]

M42 to M46

2 [✓] – 1 mark

- (b) Use your observations in Table 3.1 to identify the ions in **FA 6** and **FA 7** and state the evidence of each by completing Table 3.2.

Table 3.2

	identity	evidence
cation in FA 6	Mn^{2+}	In test 2, addition of $\text{NH}_3(\text{aq})$ results in formation of off-white ppt that was insoluble in excess $\text{NH}_3(\text{aq})$. The ppt rapidly turned brown upon contact with air. M47
anion in FA 6	CO_3^{2-}	In test 1, addition of nitric acid results in effervescence of CO_2 gas which gave a white ppt with $\text{Ca}(\text{OH})_2$. M48
cation in FA 7	Zn^{2+}	In test 4, addition of addition of $\text{NaOH}(\text{aq})$ results in formation of white ppt that was soluble in excess $\text{NaOH}(\text{aq})$, AND In test 5, addition of addition of $\text{NH}_3(\text{aq})$ results in formation of white ppt that was soluble in excess $\text{NH}_3(\text{aq})$. M49
anion in FA 7	Cl^-	In test 3, addition of AgNO_3 , gives a white precipitate which is soluble in excess $\text{NH}_3(\text{aq})$. M50

[1] for each identity + evidence

[4]

[Total: 9]

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Qualitative Analysis Notes

[ppt. = precipitate]

(a) Reactions of aqueous cations

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

(b) Reactions of anions

<i>anion</i>	<i>reaction</i>
carbonate, CO_3^{2-}	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, $\text{NO}_3^-(\text{aq})$	NH_3 liberated on heating with $\text{OH}^-(\text{aq})$ and Al foil
nitrite, $\text{NO}_2^-(\text{aq})$	NH_3 liberated on heating with $\text{OH}^-(\text{aq})$ and Al foil; NO liberated by dilute acids (colourless $\text{NO} \rightarrow$ (pale) brown 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, $\text{SO}_3^{2-}(\text{aq})$	SO_2 liberated with dilute acids; gives white ppt. with $\text{Ba}^{2+}(\text{aq})$ (soluble in dilute strong acids)

(c) Tests for gases

<i>gas</i>	<i>test and test result</i>
ammonia, NH_3	turns damp red litmus paper blue
carbon dioxide, CO_2	gives a white ppt. with limewater (ppt. dissolves with excess CO_2)
chlorine, Cl_2	bleaches damp litmus paper
hydrogen, H_2	"pops" with a lighted splint
oxygen, O_2	relights a glowing splint
sulfur dioxide, SO_2	turns aqueous acidified potassium manganate(VII) from purple to colourless

(d) Colour of halogens

<i>halogen</i>	<i>colour of element</i>	<i>colour in aqueous solution</i>	<i>colour in hexane</i>
chlorine, Cl_2	greenish yellow gas	pale yellow	pale yellow
bromine, Br_2	reddish brown gas / liquid	orange	orange-red
iodine, I_2	black solid / purple gas	brown	purple

(e) Tests for organic compounds

<i>organic compounds</i>	<i>reactions</i>
alkene	decolourises orange $\text{Br}_2(\text{aq})$
chloroalkane	heat with $\text{NaOH}(\text{aq})$, white ppt. formed on adding dilute HNO_3 , followed by $\text{Ag}^+(\text{aq})$
bromoalkane	heat with $\text{NaOH}(\text{aq})$, pale cream ppt. formed on adding dilute HNO_3 , followed by $\text{Ag}^+(\text{aq})$
iodoalkane	heat with $\text{NaOH}(\text{aq})$, yellow ppt. formed on adding dilute HNO_3 , followed by $\text{Ag}^+(\text{aq})$
alcohol	- forms white fumes with solid PCl_5 - decolourises purple acidified $\text{KMnO}_4(\text{aq})$ on heating (for primary and secondary alcohols) - gives pale yellow ppt. with alkaline $\text{I}_2(\text{aq})$ on warming (for alcohols with $\text{CH}_3\text{CH}(\text{OH})-$ group)
phenol	decolourises orange $\text{Br}_2(\text{aq})$ and forms a white ppt.
carbonyl compounds (aldehydes and ketones)	- gives orange ppt. with 2,4-dinitrophenylhydrazine - gives pale yellow ppt. with alkaline $\text{I}_2(\text{aq})$ on warming (for carbonyl compounds with $\text{CH}_3\text{CO}-$ group) - gives red-brown ppt. with Fehling's solution on warming (for aliphatic aldehydes) - gives silver mirror with Tollens' reagent on warming (for aldehydes)
carboxylic acid	CO_2 liberated with $\text{Na}_2\text{CO}_3(\text{aq})$ - gives white fumes with solid PCl_5
phenylamine	decolourises orange $\text{Br}_2(\text{aq})$ and form a white ppt.
primary amide	NH_3 liberated on heating with $\text{NaOH}(\text{aq})$