H

NANYANG JUNIOR COLLEGE
JC 2 PRELIMINARY EXAMINATION
Higher 2

CANDIDATE NAME

CLASS

TUTOR'S NAME

CHEMISTRY

Paper 4 Practical

9729/04

23 August 2021

2 hours 30 minutes

Candidates answer on the Question Paper.

Additional Materials: As listed in the Confidential Instructions

READ THESE INSTRUCTIONS FIRST

Write your name and class on all the work you hand in.
Give details of the practical shift and laboratory where appropriate, in the boxes provided.

Write in dark blue or black pen.
You may use an HB pencil for any diagrams, graphs.
Do not use staples, paper clips, glue or correction fluid.

Answer **all** questions in the spaces provided on the Question Paper.
The use of an approved scientific calculator is expected, where appropriate.
You may lose marks if you do not show your working or if you do not use appropriate units.
Qualitative Analysis Notes are printed on pages 16 and 17.
At the end of the examination, fasten all your work securely together.
The number of marks is given in brackets [ ] at the end of each question or part question.

Shift
Laboratory

For Examiner's Use	
1	**/ 13**
2	**/ 15**
3	**/ 9**
4	**/ 18**
Total	**/ 55**

This document consists of **17** printed pages..

[Turn Over

1 Determination of the concentration of hydrogen peroxide

FA 1 is a solution of hydrogen peroxide, H_2O_2, of unknown concentration.

In this experiment, you will react hydrogen peroxide with excess potassium iodide solution in an acidic condition to produce iodine as shown in equation 1.

$$H_2O_2 + 2I^- + 2H^+ \rightarrow I_2 + 2H_2O \qquad \text{equation 1}$$

The iodine produced is then titrated with a standard solution of sodium thiosulfate as shown in equation 2.

$$I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-} \qquad \text{equation 2}$$

FA 2 is 0.50 mol dm^{-3} sulfuric acid, H_2SO_4.
FA 3 is 0.500 mol dm^{-3} potassium iodide, KI.
FA 4 is 0.200 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.
Starch solution.

You will determine the concentration of hydrogen peroxide by first reacting a fixed amount of hydrogen peroxide with excess potassium iodide and subsequently titrating the iodine produced with sodium thiosulfate.

(a) In a preliminary experiment, a student pipetted 25.0 cm^3 of **FA 1** into a conical flask containing **FA 2** and **FA 3**. The resulting solution was titrated against **FA 4** from a burette and a very large titre value was obtained. He deduced the concentration of **FA 1** was too high and required dilution.

(i) Procedure

1. Using the burette labelled **FA 1**, run between 24.00 cm^3 and 24.50 cm^3 of **FA 1** into a clean 250 cm^3 graduated flask.

Set aside this burette labelled FA 1 and its contents for Question 2.

2. Record your burette readings to an appropriate level of precision in the space provided below.
3. Top up the flask to mark with deionised water. Stopper the flask and mix the contents thoroughly to obtain a homogenous solution. Label this solution as **FA 5**.

Dilution of FA 1

4. Fill another clean burette with **FA 4**.
5. Pipette 25.0 cm^3 of **FA 5** into a 250 cm^3 conical flask.
6. Using separate measuring cylinders, transfer 50.0 cm^3 of **FA 2** and 20.0 cm^3 of **FA 3** into the same conical flask. Leave the solution to stand for **7 minutes.**

While you are waiting, prepare another of this solution by repeating steps 5 and 6.

7. At the 7th minute, titrate the contents in the conical flask against **FA 4** from the burette until a pale-yellow solution is obtained.
8. Add about 5 drops of starch solution and continue titrating until the solution just turns colourless. Ignore any subsequent return of the colour.
9. Repeat the titration until consistent results are obtained and record your readings in the space below. Make certain that your results show the precision of your working.

Results

[4]

(ii) From your titrations, obtain a suitable volume of **FA 4**, $V_{FA\,4}$, to be used in your calculations. Show clearly how you obtained this volume.

$V_{FA\,4}$ = ... cm^3 [1]

(b) (i) Use your answer from **1(a)(ii)** to calculate the concentration of hydrogen peroxide in **FA 5**.

$[H_2O_2]$ in **FA 5** = ... mol dm^{-3} [2]

(ii) Hence, calculate the concentration of hydrogen peroxide in **FA 1**.

[H_2O_2] in **FA 1**= .... mol dm^{-3} [1]

(c) Besides expressing concentration of H_2O_2 in moles per cubic decimetre, mol dm^{-3}, it can also be expressed in *volume strength.*
Volume strength refers to the volume of oxygen formed when 1.00 cm^3 of hydrogen peroxide solution completely decompose at room temperature and pressure.

When 1.00 cm^3 of a 100 volume hydrogen peroxide sample is allowed to completely decompose at room temperature and pressure, 100 cm^3 of oxygen gas is produced.

Calculate the volume strength of the hydrogen peroxide in **FA 1** at room temperature and pressure that you have calculated in **1(b)(ii)**.

Volume strength of **FA 1** =.................................volume [2]

(d) State and explain how the use of a measuring cylinder to measure the volume of **FA 3** in step 6 will affect the accuracy of the titration result.

....

....

.... [1]

(e) Based on the procedure in **1(a)(i)**, the value you obtained in **1(b)(ii)** for concentration of H_2O_2 in **FA 1** differs from the actual concentration.
Explain with evidence how the procedure in **1(a)(i)** will impact the calculated concentration of H_2O_2 in **FA 1**.

Impact: ....

.... [1]

Evidence: ....

....

.... [1]

[Total: 13]

2 Determination of the enthalpy change of decomposition of hydrogen peroxide

Decomposition of hydrogen peroxide occurs spontaneously but slowly at room temperature. The heat evolved from the decomposition is not significant hence is not measurable. By increasing the rate of decomposition, temperature changes of the reaction become more measurable and the enthalpy change of decomposition can be determined experimentally.

FA 6 is 0.250 mol dm^{-3} sodium hydroxide, NaOH.

You will determine the enthalpy change for the decomposition of hydrogen peroxide by mixing **FA 1** and **FA 3** from **Question 1** together with **FA 6** and measuring the temperature of the solution over a period of time.

(a) Procedure

1. Using a clean 50 cm^3 measuring cylinder, measure 48.0 cm^3 of **FA 3** followed by 2.0 cm^3 of **FA 6**. Stir the solution with a glass rod.
2. Using the burette labelled **FA 1** from **Question 1**, transfer 50.00 cm^3 of **FA 1** into a polystyrene cup. Place the cup inside a 250 cm^3 beaker.
3. Stir the solution in the cup with a thermometer and measure its temperature, *T*, and start the stopwatch. This is the temperature when time, t = 0 min.
   Record this temperature in a suitable format in the space below.
   Repeat the measurement at one-minute interval for the next three minutes.
4. At the 4th minute, pour the contents of the measuring cylinder into the polystyrene cup and stir the solution. **Do not** read the temperature at the 4th minute.
5. Stir the solution continuously and record the temperature at the 5th minute and every half–minute interval till the 12th minute. Record all the temperatures measured.

Temperature readings

[2]

(b) Plot a graph of temperature, T, on the y-axis, against time, t, on the x-axis on the grid.
Draw a best–fit straight line taking into account all of the points before t = 4.0 min.
Draw another best–fit straight line taking into account all of the points after the temperature of the mixture has started to fall steadily. Extrapolate both lines to t = 4.0 min.

Temperature change, ΔT = ... °C [4]

(c) Determine the heat change, *q*, for your experiment and hence calculate the enthalpy change, ΔH_{decomp}, for the decomposition of H_2O_2.

You should assume the:

- concentration of H_2O_2 in **FA 1** is 0.950 mol dm^{-3};
- decomposition of H_2O_2 in **FA 1** is complete;
- specific heat capacity of the solution is 4.18 J g^{-1} K^{-1};
- density of the solution is 1.00 g cm^{-3}.

q = ...

ΔH_{decomp} = ...

[3]

(d) The following is a proposed mechanism for the decomposition of H_2O_2.

$I^-(aq) + H_2O_2(aq) \downarrow HOI(aq) + OH^-(aq)$		step 1
$HOI(aq) + OH^-(aq) \downarrow H_2O(l) + OI^-(aq)$		step 2
$OI^-(aq) + H_2O_2(aq) \downarrow I^-(aq) + H_2O(l) + O_2(g)$		step 3

(i) Give the overall equation for the decomposition of H_2O_2, including state symbols.

....[1]

(ii) Deduce the role of iodide ions in the mechanism. Explain your reasoning.

....

....[1]

(iii) A similar decomposition experiment was carried out using aqueous iron(III) ions instead of iodide ions.
By considering the bonds formed and broken between H_2O_2 molecules and iron(III) ions, suggest how will this enthalpy change of decomposition of H_2O_2 compare with that when H_2O_2 molecules and iodide ions were used in **2(a)**. Explain your answer.

....

....

....

....[2]

(e) **(i)** A student decided to use half the volumes of all reagents in **2(a)** so that he can repeat the experiment to obtain two sets of data to increase the reliability of his results. Comment if his method would lead to a more reliable result.

....

....

....

....[1]

(ii) The same student carried out the experiment while the fans were turned on. Comment if his ΔT value obtained will be reliable due to cooling by the draught.

....

....

....

....[1]

[Total: 15]

3 (a) Planning

The order of decomposition reaction with respect to hydrogen peroxide cannot be determined from the mechanism in **2(d)** unless the slow step is known. A student suggested the order of decomposition could be first order.

You are to plan a series of experiments to verify the decomposition reaction is first order with respect to hydrogen peroxide using a continuous sampling method.

In this method, decomposition of H_2O_2 will run for approximately 25 min once it is mixed with solid MnO_2. A fixed volume of the reaction mixture is then sampled shortly after mixing and diluted with deionised water while noting the time interval from the start of the reaction.
The sample is then analysed to determine the amount of H_2O_2 remaining.

Further samples are withdrawn at timed intervals and analysed in the same way to obtain sufficient data for a graph to be drawn. You will use the graph to verify if the reaction is first order with respect to H_2O_2.

You may assume you are provided with:
- solid manganate(IV) oxide, MnO_2,
- aqueous potassium manganate(VII), $KMnO_4$,
- aqueous sulfuric acid, H_2SO_4,
- apparatus normally found in a school or college laboratory.

Your plan should include brief details of:
- the quantities of reagents you would use (concentration of reagents is not required),
- the apparatus you would use,
- the procedure to sample the reaction mixture at timed intervals,
- the method to determine the amount of H_2O_2 remaining in each sample.

....

....

....

....

....

....

....

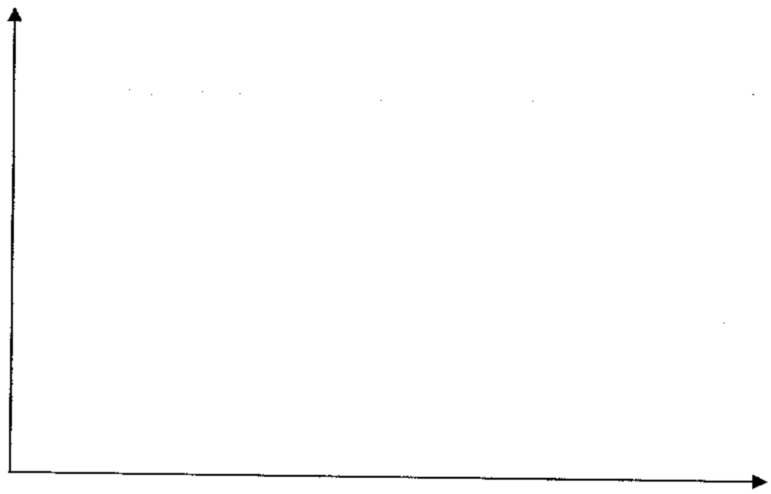
....

....

....

[7]

(b) In the space below, sketch a clearly labelled graph you expect to obtain from carrying out your plan. Indicate clearly on your graph and explain how you would verify if the decomposition of hydrogen peroxide is first order.



[2]

[Total: 9]

4 **Qualitative Analysis**

(a) **Solid FA 7** is a salt containing two cations and one anion.

You are provided with an aqueous solution of **FA 7**, labelled **FA 7 solution**.

You will devise a series of simple tests, based on the Qualitative Analysis Notes on pages **16–17**, to identify the anion present in **FA 7 solution**.

FA 7 does not contain sulfite ions, SO_3^{2-}, nitrite ions, NO_2^-, or nitrate ions, NO_3^-.

(i) Describe **three different tests**, using only the bench reagents provided, which will allow you to identify the anion. State how you will decide if the test result is positive.

test 1 ....

....

....

test 2 ....

....

....

test 3 ....

....

....[3]

(ii) Perform the tests you described in **4(a)(i)**, using the **FA 7** solution provided.

Record your observations and hence deduce the identity of the anion in **solid FA 7**.

Any test requiring heating MUST be performed in a boiling tube.

test 1 ....

....

test 2 ....

....

test 3 ....

....

identity of anion: .... [4]

(iii) Perform the tests in **Table 4.1** and record your observations. If there is no observable change, write **no observable change**. Hence deduce the cations present and the identity of **solid FA 7**.

Table 4.1

Test	Observation
Add 2 cm^3 of **FA 7** into a boiling tube followed by aq NaOH dropwise. Warm the mixture.	

[2]

Cations present in **aqueous FA 7**: .... and ....

Identity of **solid FA 7**: ....

[2]

(b) (i) **FA 8** and **FA 9** are aqueous solutions of covalently bonded compounds. Perform the tests for **FA 8** and **FA 9** described in **Table 4.2** and record your observations. If there is no observable change, write **no observable change**.

All heating should be carried out in a hot water bath.

Table 4.2

Test	Observation	
	FA 8	**FA 9**
1. Add 2 cm^3 of aqueous sulfuric acid and **FA** solution into a test–tube followed by 1 cm^3 of potassium manganate(VII) solution.		
Warm the test–tube.		
2. Add 1 cm^3 of aqueous sulfuric acid and **FA 8** solution into a test–tube followed by 1 cm^3 of **FA 7**.		
Add aqueous ammonia.		
3. Add 1 cm^3 of **FA** solution into a test–tube followed by 2 cm^3 of aqueous iodine. Add aqueous sodium hydroxide dropwise until the yellow colour just disappears. Warm the test–tube.		
4. Add 1 cm^3 of **FA 9** into a test–tube followed by 1 cm^3 of 2,4–DNPH and warm.		
5. To 1 cm^3 of aqueous silver nitrate, add 2 drops of aqueous sodium hydroxide. Add aqueous ammonia dropwise until the precipitate just dissolves. Add 1 cm^3 of **FA 9** to the resulting mixture and warm.		

[5]

(ii) Using your results in **4(b)(i)**, state and explain the redox nature of **FA 8**.

....

....[1]

(iii) **FA 9** contains a 4–carbon organic compound. Suggest two possible structures for this compound.

Structure 1	**Structure 2**

[1]

[Total: 18]

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

anion	*reaction*
carbonate, CO_3^{2-}	CO_2 liberated by dilute acids
chloride, Cl^-(aq)	gives white ppt. with Ag^+(aq) (soluble in NH_3(aq))
bromide, Br^-(aq)	gives pale cream ppt. with Ag^+(aq) (partially soluble in NH_3(aq))
iodide, I^-(aq)	gives yellow ppt. with Ag^+(aq) (insoluble in NH_3(aq))
nitrate, NO_3^-(aq)	NH_3 liberated on heating with OH^-(aq) and Al foil
nitrite, NO_2^-(aq)	NH_3 liberated on heating with OH^-(aq) and Al foil; NO liberated by dilute acids (colourless NO $\rightarrow$ (pale) brown NO_2 in air)
sulfate, SO_4^{2-}(aq)	gives white ppt. with Ba^{2+}(aq) (insoluble in excess dilute strong acids)
sulfite, SO_3^{2-}(aq)	SO_2 liberated with dilute acids; gives white ppt. with Ba^{2+}(aq) (soluble in dilute strong acids)

(c) Tests for gases

gas	***test and test result***
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

halogen	***colour of element***	***colour in aqueous solution***	***colour in hexane***
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

Chemicals

Chemical Name	Qty per student	Mass / Concentration / Preparation method	Label
Question 1			
Hydrogen peroxide, H_2O_2	120 cm^3	0.800 mol dm^{-3} Using 19% H_2O_2, dilute 129 cm^3 with deionised water to 1 dm^3 Using 3% H_2O_2, dilute 817 cm^3 with deionised water to 1 dm^3	**FA 1**
Sulfuric acid, H_2SO_4	180 cm^3	0.50 mol dm^{-3}	**FA 2**
Potassium iodide, KI	150 cm^3	0.500 mol dm^{-3}	**FA 3**
Sodium thiosulfate, $Na_2S_2O_3$	120 cm^3	0.200 mol dm^{-3}	**FA 4**
Starch solution	10 cm^3	1%	Starch solution
Question 2			
aq NaOH	5 cm^3	0.250 mol dm^{-3} sodium hydroxide,	**FA 6**
Aqueous ammonium iron(II) sulfate	15 cm^3	0.2 mol dm^{-3} ammonium iron(II) sulfate Dissolve 78.4 g of $(NH_4)_2Fe(SO_4)_2.6H_2O$ in each dm^3 of solution	**FA 7**
Question 4			
H_2O_2 (**FA 1 from Qn 1**)	10 cm^3	0.800 mol dm^{-3}	**FA 8**
Ethanol + Propanone	10 cm^3	5 cm^3 ethanol + 5 cm^3 propanone	**FA 9**
Bench reagents			
sodium hydroxide	10 cm^3	2.0 mol dm^{-3} sodium hydroxide	
aqueous silver nitrate		0.05 mol dm^{-3} silver nitrate	silver nitrate
barium chloride or barium nitrate		(approximately 0.2 mol dm^{-3})	barium chloride or barium nitrate
aqueous iron(II) sulfate	5 cm^3	0.2 mol dm–3 iron(II) sulfate Dissolve 55.6 g FeSO4.7H2O in each dm^3 of 1 mol dm^{-3} sulfuric acid. This solution should be prepared shortly before use.	aqueous iron(II) sulfate
limewater		(a saturated solution of calcium hydroxide)	
aqueous sodium hydroxide		(approximately 2.0 mol dm^{-3})	
aqueous ammonia		(approximately 2.0 mol dm^{-3})	Aq NH_3

[Turn Over

hydrochloric acid		(approximately 2.0 mol dm^{-3})	
nitric acid		(approximately 2.0 mol dm^{-3})	
sulfuric acid		(approximately 1.0 mol dm^{-3})	
aq iodine			

Apparatus

Per student		
1 × burette labelled **FA 1**	Sticker labels	1 x Lighter
1 × burette (unlabelled)	1 × wash bottle with deionised water	1 x Tripod stand and wire gauze
1 x retort stand and burette clamp	5 plastic dropping pipettes / droppers	1 x metal spatula
1 x 250 cm^3 graduated flask	1 x safety goggles	1 x 250 cm^3 beaker (for water bath)
2 × 250 cm^3 conical flask	1 x Polystyrene cup	
1 x 100 cm^3 waste beaker	1 x 250 cm^3 beaker for polystyrene cup	
3 × plastic funnel	1 x Thermometer 0.2 °C div	
1 × 25.0 cm^3 pipette	1 × white tile	
1 × pipette filler	1 x stopwatch	
1 × 50 cm^3 measuring cylinder	1 x glass rod	
1 × 25 cm^3 measuring cylinder	1 x test–tube holder	
6 x medium test-tubes	1 x test–tube rack	
Wooden splint	2 x boiling tubes	
red and blue litmus paper or Universal Indicator paper	1 x Bunsen burner	

NYJC J2 H2 Chemistry Prelim Answers

Paper 4 Answers

1 (a) (i)

Dilution of FA 1

Final burette reading for **FA 1** / cm^3	
Initial burette reading for **FA 1** / cm^3	
Volume of **FA 1** used / cm^3	24.20
	2 d.p.

Titration Results

Final burette reading for **FA 4** / cm^3		
Initial burette reading for **FA 4** / cm^3		
Volume of **FA 4** used / cm^3	19.15	19.05
	2 d.p.	2 d.p.

Correct header and units for both tables [1]
Correct evaluation and precision for both tables [1]

$$\text{scaled mean titre} = \frac{24.25}{\text{volume diluted}} \times \text{mean titre}$$

Scaled mean titre diff $\leq \pm 0.5$ cm^3 [2] or $\leq \pm 1.0$ cm^3 [1]

(ii) Ave titre value = 19.10 (2 d.p.)

Choice of titre values are consistent and average correctly evaluated **[1]**

(b) (i)

$n_{S_2O_3^{2-}}$ reacted $= 0.200 \times \frac{19.10}{1000} = 3.82 \times 10^{-3}$ mol

$n_{H_2O_2}$ in 25.0 cm^3 **FA 5** reacted $= n_{I_2}$ formed

$= \frac{1}{2} \times 3.82 \times 10^{-3}$ mol

$= 1.91 \times 10^{-3}$ mol [1]

$$[H_2O_2]_{FA5} = \frac{1.91 \times 10^{-3}}{\frac{25.0}{1000}} = 7.64 \times 10^{-2} \text{ mol dm}^{-3} \text{ [1]}$$

(ii)

$$7.64 \times 10^{-2} \times \frac{250}{1000} = [H_2O_2]_{FA1} \times \frac{24.20}{1000}$$

ecf for $[H_2O_2]_{FA5}$ from 2(b)(i)

$$[H_2O_2]_{FA1} = \frac{7.64 \times 10^{-2} \times \frac{250}{1000}}{\frac{24.20}{1000}} = 0.789 \text{ mol dm}^{-3} \text{ [1]}$$

(c) $2H_2O_2 \rightarrow 2H_2O + O_2$

n_{O_2} evolved from 1 dm^3 of **FA 1** $= \frac{1}{2} \times 0.789$

$= 0.394$ mol

n_{O_2} evolved from 1 cm^3 of **FA 1** at r.t.p. $= \frac{0.3946}{1000} = 3.946 \times 10^{-4}$ mol [1]

V_{O_2} evolved from 1 cm^3 of **FA 1** at r.t.p. $= 3.946 \times 10^{-4}$ mol x 24000 cm^3 mol^{-1}

$= 9.47$ cm^3 [1]

Hence 1 cm^3 of **FA 1** gives 9.47 cm^3 of O_2 $\Rightarrow$ Vol strength of **FA 1** is 9.47

(d) H_2O_2 in **FA 5** is the limiting reagent while **FA 3** is measured in excess. Using a measuring cylinder to measure **FA 3** will not affect the amount of I_2 produced hence will not affect the volume / accuracy of **FA 4** used.

(e) Impact: The oxidation of I^- by H_2O_2 is incomplete when titration is carried out as the oxidation reaction is slow hence the concentration of H_2O_2 calculated is smaller than expected.[1]

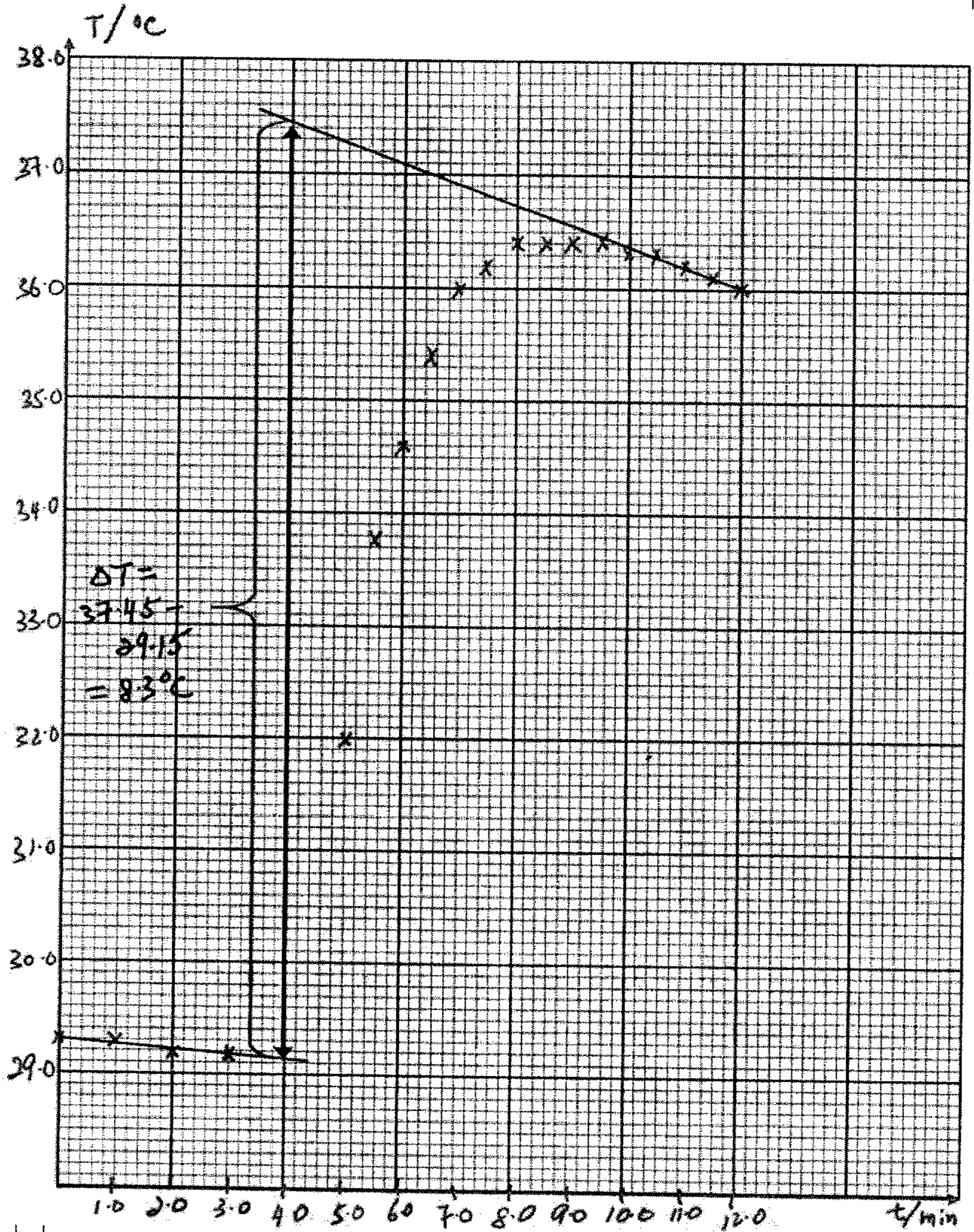
Evidence: The reaction mixture needs to stand for at least 7 minutes before titration. **OR** slow return of the blue black colour at end point suggest oxidation of I^- still continues.[1]

2 (a) Temperature readings

Time / min	Temperature / °C		Time / min	Temperature / °C
0	29.3		7 ½	36.2
1	29.3		8	36.4
2	29.2		8 ½	36.4
3	29.2		9	36.4
4	—		9 ½	36.4
5	32.0		10	36.3
5 ½	33.8		10 ½	36.3
6	34.6		11	36.2
6 ½	35.4		11½	36.1
7	36.0		12	36.0

Correct heading and units [1]

Record time to nearest half minute and temp to 1 d.p except at 4th min.[1]

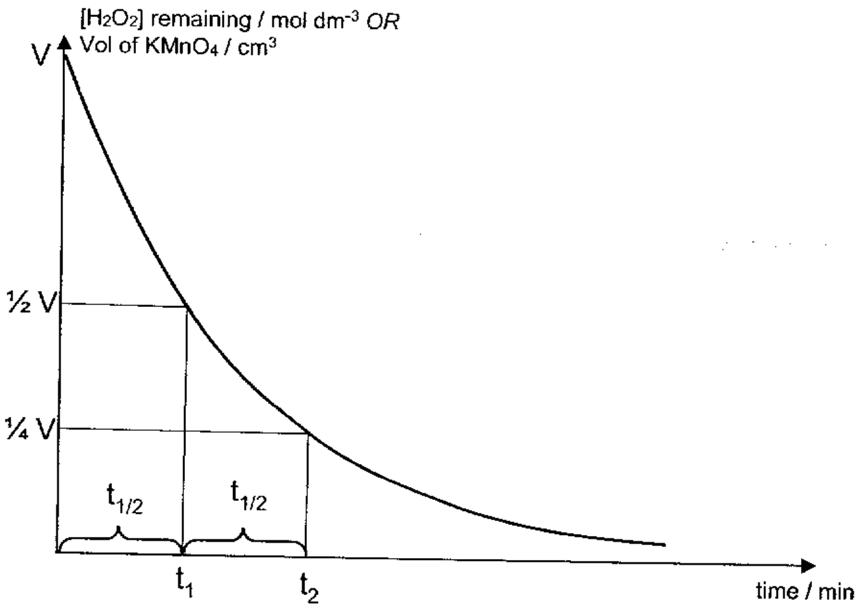

T/°C
38.0
37.0
36.0
35.0
34.0
33.0
32.0
31.0
30.0
29.0
ΔT =
37.45 −
29.15
= 8.3°C
1.0 2.0 3.0 4.0 5.0 6.0 7.0 8.0 9.0 10.0 11.0 12.0
t/min

		Temperature change, ΔT = ... °C [4] Correctly labelled axes and units, no odd scale used [1] Correctly plotted all points within ½ sq except at 4th min [1] Drew and extrapolated 2 best–fit lines to the 4th min, allowing 1 anomalous pt after mixing [1] Correctly determined and labelled ΔT clearly [1]
(c)		$q = mc\Delta T$ = 100 x 4.18 x Ans fr **2(b)** J [1] including units $\Delta H_{decomp} = -\frac{mc\Delta T}{n_{H_2O_2}}$ $= -\frac{100 \times 4.18 \times \text{Ans fr } \mathbf{2(b)}}{\frac{50}{1000} \times 0.950}$ [1] = - 8800 x Ans fr **2(b)** J mol^{-1} [1] with sign and units
(d)	**(i)**	$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$
	(ii)	Iodide is acting as a catalyst. It is reacted in step 1 but regenerated in step 3 hence is not involved in the overall decomposition reaction.[1]
	(iii)	The enthalpy changes for both experiments will be the same / very similar.[1] Energy evolved in forming bonds between H_2O_2 and the ions will require the same amount of energy to eventually break it up hence both catalysts are not involved in the overall reaction.[1]
(e)	**(i)**	The results would be equally reliable. Enthalpy change of reaction, ΔH_{rxn}, and specific heat capacity, *c*, remain the same, while the mass of reactants, m, and the amount of reactants, n, used are half the original quantity hence ΔT would remain the same.[1] $\Delta H_{rxn} = \frac{\frac{1}{2}(100) \times 4.18 \times \Delta T}{\frac{1}{2}n}$
	(ii)	The ΔT value would still be reliable as cooling happened throughout the experiment at a constant rate hence extrapolating the cooling portion of the graph to the point of mixing would compensate for the heat loss due to cooling.[1]

3 (a) Planning

1. Fill a burette with aq $KMnO_4$ solution.
2. Using a 100 cm^3 measuring cylinder, transfer 100 cm^3 of H_2O_2 into a 250 cm^3 conical flask (allow beaker). [1]
3. Add a small spatula of MnO_2 into the conical flask, swirl the mixture and start the stopwatch at the same time. Allow the mixture to stand. [1]
4. Before 2 minutes, use a 10.0 cm^3 pipette to transfer a sample into a clean 250 cm^3 conical flask. step 1 + step 4 [1]
5. At 2 min, use a clean beaker to transfer 100 (to 150) cm^3 of deionised water into the conical flask (containing the 10 cm^3 sample) and record the timing. [1]
6. Using a 25.0 cm^3 measuring cylinder, add 20.0 (to 25) cm^3 of H_2SO_4 into the conical flask. [1]
7. Titrate the resulting solution against $KMnO_4$ from the burette until 1 drop gives a permanent pink colour. [1]
8. Repeat steps 4 to 7 for another 4 (or 5) times but at time intervals of approximately 4 (to 5) min. [1]
9. Calculate the amount of H_2O_2 remaining at each timing using the volume of $KMnO_4$.

(b)



Sketch labelled graph of V_{KMnO4} against time with units, downward sloping curve [1]

If the decomposition is first order with respect to H_2O_2, the graph would show constant half–life when volume of $KMnO_4$ decreases by half.

Note: V_{KMnO4} reacted α n_{KMnO4} reacted α n_{H2O2} reacted α $[H_2O_2]$

Explains constant half–life and marked out on graph [1]

4 (a) (i)

test 1 To 1 cm^3 of **FA 7** in a test–tube, add about 2–4 cm^3 of H_2SO_4 / HC*l* / HNO_3. If effervescence observed and white ppt is formed when gas evolved is bubbled into limewater, carbonate ions are present.

test 2 To 1 cm^3 of **FA 7 in** a test–tube, (add 1 cm^3 of HNO_3 and) add 1 cm^3 of aq $AgNO_3$ followed by aq NH_3 dropwise till in excess. If white ppt is formed and is soluble in excess aq NH_3, Cl^- ions are present. If cream ppt is formed and is partially soluble in excess aq NH_3, Br^- ions are present. If yellow ppt is formed and is insoluble in excess aq NH_3, I^- ions are present.

test 3 To 1 cm^3 of **FA 7** in a test–tube, add 1 cm^3 of aq $BaCl_2$ followed by 2 cm^3 of aq HNO_3 / HC*l*. If white ppt is formed and is insoluble in acid, SO_4^{2-} ions are present.

(ii)

test 1: No effervescence observed

test 2: No ppt observed.

Note: There are no halides present, but **FA 7** contains Fe^{2+}. When aq NH_3 is added, green ppt is formed and is insoluble in excess aq NH_3. (Ppt turns brown upon standing in air.) The presence of $AgNO_3$ and aq NH_3 may also cause a black ppt to form.

test 3: White ppt formed is insoluble in excess acid.

identity of anion: SO_4^{2-}

(iii)

Table 4.1

Test	Observation
Add 2 cm^3 of **FA 7** into a boiling tube followed by aq NaOH dropwise.	Green ppt formed and is insoluble in excess aq NaOH. Ppt turns brown upon standing in air. [1]
Warm the mixture.	Gas evolved turns damp red litmus paper blue.[1]

Cations present in **aqueous FA 7**: NH_4^+ and Fe^{2+} [1]

Identity of **solid FA 7**: $(NH_4)_2Fe(SO_4)_2$ [1]

(b) (i)

Table 4.2

test	observation	
	FA 8	**FA 9**
1. Add 2 cm^3 of aqueous sulfuric acid and **FA** solution into a test–tube followed by 1 cm^3 of potassium manganate(VII) solution.	Purple $KMnO_4$ decolourised. Effervescence. Gas evolved relights glowing splinter.	No observable change

Warm the test–tube.	No further / observable change	Purple $KMnO_4$ decolourised.
2. Add 1 cm^3 of aqueous sulfuric acid and **FA 8** solution into a test–tube followed by 1 cm^3 of **FA 7.**	Pale green **FA 7** turned (light) yellow / brown.	
Add aqueous ammonia.	Red–brown ppt formed is insoluble in excess $NH_3(aq)$	
3. Add 1 cm^3 of **FA** solution into a test–tube followed by 2 cm^3 of aqueous iodine. Add aqueous sodium hydroxide dropwise until the yellow colour just disappears. Warm the test–tube.	Yellow / colourless solution formed.	(Pale) Yellow ppt formed.
4. Add **1** cm^3 of **FA 9** into a test–tube followed by 1 cm^3 of 2,4–DNPH and warm.		Orange ppt formed
5. To 1 cm^3 of aqueous silver nitrate, add 2 drops of aqueous sodium hydroxide. Add aqueous ammonia dropwise until the precipitate just dissolves. Add 1 cm^3 of **FA 9** to the resulting mixture and warm.		No grey ppt / silver mirror formed.

(ii) FA 8 is both an oxidising and reducing agent. It reduces $KMnO_4$ in test 1 and oxidises (Fe^{2+} in) FA 7 in test 2.

(iii)

Structure 1	Structure 2	Structure 3
$CH_3COCH(OH)CH_3$	$CH_3COCH_2CH_2OH$	$CH_3COCH=CH_2$

Any 2 structures with 4 carbon atoms and matches the correct observations [1]