RIVER VALLEY HIGH SCHOOL

JC 2 PRELIMINARY EXAMINATION

H2 CHEMISTRY 9729 Paper 4

19 AUGUST 2022

2 HOURS 30 MINUTES

NAME ______________________________

CLASS 21J (   )

INDEX NO. ______________________________

INSTRUCTIONS TO CANDIDATES

DO NOT OPEN THIS BOOKLET UNTIL YOU ARE TOLD TO DO SO.

Read these notes carefully.

Write your name, class and index number in the spaces at the top of this page.
Give details of the practical shift and laboratory where appropriate, in the boxes provided.
Write in dark blue or black pen.
You may use a 2B pencil for any diagram or graph.
Do not use staples, paper clips, highlighters, 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.

Shift
Laboratory

For Examiner's Use	
s.f.	
Units	
Total	55

This Question Paper consists of **19** printed pages and **1** blank page.

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

1 Investigation of the oxidising ability of substances

FA 1 is an aqueous solution of iron(II) sulfate.

FA 2 is made by dissolving 0.750 g of $KMnO_4$ in deionised water and made up to 250 cm^3 in a volumetric flask.

FA 3 is 0.500 mol dm^{-3} compound **A**.

FA 4 is 1 mol dm^{-3} sulfuric acid.

You will perform tests to:

- investigate the oxidising ability of three substances
- carry out a titration to determine the concentration of iron(II) sulfate in **FA 1**.

(a) **(i)** Carry out the following tests. Carefully record your observations in Table 1.1. The volumes given below are approximate and should be estimated rather than measured.

In this section, there is only one gas evolved. You only need to carry out the identification test and identify the gas once.

Table 1.1

Test		Observations
1	To a 1 cm depth of **FA 1** in a test-tube, add 1 cm depth of **FA 4**, then gradually add **FA 2** till 1 drop in excess.	Purple **FA 2** decolourised. Pale green/colourless **FA 1** turned pale yellow and finally orange/ pink.
2	To a 1 cm depth of **FA 3** in a test-tube, add 1 cm depth of **FA 2**.	Purple $KMnO_4$ decolourised. A brown solution/ppt formed. Effervescence/ bubbles of gas evolved rapidly. Colourless, odourless gas evolved relight a glowing splint. The gas is oxygen.
3	To a 1 cm depth of **FA 1** in a test-tube, add about 1 cm depth of **FA 4**, followed by 1 cm depth of **FA 3**. To a portion of resulting solution, add aqueous sodium hydroxide till excess.	Pale green/colourless **FA 1** turned yellow. Red-brown ppt formed is insoluble in excess NaOH(aq). Effervescence/ bubbles of gas evolved (rapidly).

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(ii) Explain the chemistry involved for all the observations in Test 1.

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(iii) Write two equations to show the changes to Fe^{2+} ions in Test 3.

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(iv) Using the results in Test 3, state the stronger oxidising agent.

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(b) **(i)** **Titration of FA 1 against FA 2**

1. Fill a burette with **FA 2**.
2. Use the pipette to transfer 25.0 cm^3 of **FA 1** into a 250 cm^3 conical flask.
3. Use a measuring cylinder to add 20.0 cm^3 of **FA 4** into the same conical flask.
4. Titrate the mixture in the conical flask until the end-point is reached.
5. Record your titration results, to an appropriate level of precision, in the space provided below.
6. Repeat points 2 to 5 until consistent results are obtained.

Titration results

Titration number	1	2
Final burette reading /cm^3	24.20	24.25
Initial burette reading /cm^3	0.00	1.00
Volume of **FA 2** (added) /cm^3	24.20	24.25

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Difference in titres

(ii) From your titrations, obtain a suitable volume of **FA 2** to be used in your calculations. Show clearly how you obtained this volume.

volume of **FA 2** used = …………………………

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(c) **(i)** Calculate the amount of MnO_4^- used in the titration.

[Ar: K, 39.1; Mn, 54.9; O,16.0]

amount of MnO_4^- = ...

14

(ii) Given that 1 mole of MnO_4^- reacts with 5 moles of Fe^{2+}, calculate the molar concentration of iron(II) sulfate in **FA 1**.

molar concentration of iron(II) sulfate = ...

15

(d) A student plans to titrate 25.0 cm^3 of aqueous solution containing iron(II) chloride with **FA 2** as the titrant. This iron(II) chloride solution has similar molar concentration as iron(II) sulfate in **FA 1**.

Without changing the identity of the chemicals used, suggest a modification to improve the experiment. Explain your answer.

You are not required to show any calculations.

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(e) In a series of titrations, a student pipetted 25.0 cm^3 of **FA 1** into the conical flask and added varying volumes of **FA 3** from the burette to form a mixture. This mixture was titrated with **FA 2** till end-point was reached. The results are shown below in Table 1.2.

Table 1.2

experiment	1	2	3	4	5
volume of **FA 3** added /cm^3	5.00	10.00	15.00	20.00	25.00
volume of **FA 2** used /cm^3	15.45	10.55	5.70	1.25	3.45

(i) Explain why the volumes of **FA 2** obtained decrease as V_{FA3} increases from 5.00 cm^3 to 20.00 cm^3.

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(ii) Considering the chemistry involved, deduce whether the volume of **FA 2** used in **experiment 5** is an anomaly.

Explain your answer.

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(iii) Calculate the highest percentage uncertainty in the volume of **FA 2** used as shown in Table 1.2.

percentage uncertainty = ... 20

[Total: 20]

2 **Investigation of the kinetics of the reaction between iron(III) ions and iodide ions**

In acidic solutions, iron(III) ions are reduced by iodide ions to form iron(II) ions. The iodide ions are oxidised to iodine.

$$2Fe^{3+}(aq) + 2I^{-}(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$$

The rate of this reaction can be investigated by using starch indicator, which turns blue-black in the presence of iodine. Sodium thiosulfate is added to the reaction mixture to react with iodine as it is formed. The blue-black colour is seen when all the thiosulfate has reacted.

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^{-}(aq) + S_4O_6^{2-}(aq)$$

You will perform a series of four experiments to investigate how the rate of reaction is affected by changing the concentration of the iodide ions.

For each experiment, you will note the volume of **FA 5** added, $V_{\mathbf{FA\ 5}}$, the volume of water added, V_{H_2O}, and the time taken, t, for the reaction mixture to turn blue-black.

In this series of experiments, the rate equation for the reaction can be simplified to rate = $k'\,[I^-]^{\boldsymbol{m}}$, where $\boldsymbol{m}$ is the rate order with respect to I^- and k' is $k[Fe^{3+}]$.

The simplified rate equation can be further manipulated to derive the following relationship:

$$\lg\left(\frac{3600}{\text{reaction time}}\right) = \boldsymbol{m} \times \lg(V_{\mathbf{FA\ 5}}) + \text{constant}$$

FA 5 is 0.0500 mol dm^{-3} potassium iodide, KI.
FA 6 is 0.0500 mol dm^{-3} iron(III) chloride, $FeCl_3$.
FA 7 is 0.0050 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.
FA 8 is starch indicator.

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

- all values of $V_{\mathbf{FA\ 5}}$ and V_{H_2O} to an appropriate level of precision
- all values of t
- all calculated values of lg($V_{\mathbf{FA\ 5}}$) and lg(rate) to three significant figures.

Experiment 1

1. Use a 25.00 cm^3 measuring cylinder to place 20.00 cm^3 of **FA 5** in a 100 cm^3 beaker.
2. Use appropriate measuring cylinders to add the following to the same beaker.
   - 20.0 cm^3 of **FA 7**
   - 10.0 cm^3 of **FA 8**
3. Use an appropriate measuring cylinder to measure 10.0 cm^3 of **FA 6**.
4. Add this **FA 6** to the same 100 cm^3 beaker and start timing immediately.
5. Stir the mixture and place the beaker on a white tile.
6. Stop timing as soon as the solution turns blue-black.
7. Record this reaction time to the nearest 0.1 second.
8. Wash the beaker and dry it with a paper towel.

Experiment 2

1. Use the same 25.00 cm^3 measuring cylinder to measure 10.00 cm^3 of **FA 5** and make up the volume to 20.00 cm^3 using deionised water.
2. Place this solution in a 100 cm^3 beaker.
3. Use appropriate measuring cylinders to add the following to the same beaker.
    - 20.0 cm^3 of **FA 7**
    - 10.0 cm^3 of **FA 8**
4. Use an appropriate measuring cylinder to measure 10.0 cm^3 of **FA 6**.
5. Add this **FA 6** to the same 100 cm^3 beaker and start timing immediately.
6. Stir the mixture and place the beaker on a white tile.
7. Stop timing as soon as the solution turns blue-black.
8. Record this reaction time to the nearest 0.1 second.
9. Wash the beaker and dry it with a paper towel.

Experiments 3 and 4

Carry out two further experiments to investigate how the reaction time changes with different volumes of aqueous potassium iodide, **FA 5**.

The combined volume of **FA 5** and deionised water must always be 20.00 cm^3.

Do not use a volume of **FA 5** that is less than 6.00 cm^3.

Results

The rate of the reaction is defined as shown.

$$\text{rate} = \frac{3600}{\text{reaction time}}$$

Expt	$V_{\textbf{FA 5}}$/ cm^3	V_{H_2O} /cm^3	Reaction time, ***t*** /s	lg($V_{\textbf{FA 5}}$)	lg (rate)
1	20.00	0.00	15.1	1.30	2.38
2	10.00	10.00	56.5	1.00	1.80
3	15.00	5.00	26.9	1.18	2.13
4	6.00	14.00	144.5	0.778	1.40

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(b) Plot a graph of lg($\frac{3600}{\text{reaction time}}$), on the y-axis, against the lg($V_{\textbf{FA 5}}$), on the x-axis, on the grid. Draw a straight line of best fit through the points.

Your scale should allow lg($V_{\textbf{FA 5}}$) of 0.700 and lg($\frac{3600}{\text{reaction time}}$) of 1.30 to be plotted.

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(c) **(i)** Calculate the gradient of the line to three significant figures, and hence suggest the rate order, ***m***, showing clearly how you did this.

gradient = ……………………………..

m = ……………………………..

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(ii) Using your graph, calculate the expected reaction time if 7.9 cm^3 of **FA 5** is used**.** Show your working clearly.

reaction time = ……………………………..

30

(d) **(i)** Using data from **Experiments 1** and **2**, show by calculation that the volume of aqueous potassium iodide, **FA 5**, used was directly proportional to the concentration of iodide ions.

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(ii) Explain, by referring to your table of results, how the rate of reaction is affected by an increase in the concentration of aqueous potassium iodide, **FA 5**.

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(e) Thiosulfate ions can reduce iron(III) ions and also react with acid to form sulfur, sulfur dioxide and water.

(i) Write an ionic equation, with state symbols, for the reaction between thiosulfate ions and hydrogen ions in aqueous solution.

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(ii) A student carried out the same investigation as in **(a)** but the solutions were mixed in a different order. The student kept the **FA 5** and an appropriate volume of deionised water in the measuring cylinder and all the other reactants in the 100 cm^3 beaker. The student then transferred the solution from the measuring cylinder into the 100 cm^3 beaker and started timing.

State and explain whether the student's method is

- better than that in **(a)**,
- as good as that in **(a)**, or
- not as good as that in **(a)**.

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[Total: 14]

3 Investigation of an inorganic compound

FA 9 is an aqueous solution of an inorganic compound, MY_2, which contains a cation and an anion (Y^-), both listed in Qualitative Analysis Notes.

FA 10 is a saturated solution of a sodium salt, NaY.

You will perform tests to:
- investigate the effect of reagents on **FA 9**
- identify the anion present in **FA 9**.

(a) Carry out the following tests. Carefully record your observations in Table 3.1. The volumes given below are approximate and should be estimated rather than measured.

Table 3.1

Test		Observations
1	Place a test-tube containing 2 cm depth of **FA 9** in an almost boiling water bath for a few minutes.	Blue/ bluish-green **FA 9** turned green.
2	To a 2 cm depth of **FA 9** in a test-tube, add gradually add 2 cm depth of **FA 10**. Pour half of this mixture into another test-tube and place it in the almost boiling water bath for a few minutes for comparison.	(Blue/ bluish-green) **FA 9** turned green upon adding **FA 10**. The resultant solution turned yellowish-green/ brighter green after warming.
3	To a 1 cm depth of **FA 9** in a boiling-tube, gradually add 3 cm depth of aqueous sodium hydroxide. Filter the resultant mixture and collect the filtrate. Keep the filtrate for Test 4.	Pale blue/ Blue-green/greenish-blue ppt formed is insoluble in excess NaOH(aq). The residue is dark blue. The filtrate is a colourless solution.
4	To a 1 cm depth of filtrate in a test-tube, add 1 cm depth of nitric acid, followed by $AgNO_3$(aq). To a portion of the resultant mixture, add NH_3(aq).	White ppt formed with $AgNO_3$(aq) is soluble in aqueous ammonia to give a colourless solution.

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(b) **(i)** State the identity of the cation in **FA 9**.

Cation ....

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(ii) Table 3.1 Test 4 confirms the identity of the anion in **FA 9**. Explain why the filtrate from Test 3 was used instead of **FA 9**.

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(c) When solid MY_2 is dissolved in water, the cation could exist either as a blue soluble complex or a yellow soluble complex. The complexes are $[M(H_2O)_6]^{2+}$ and $[MY_4]^{n-}$, and they can be converted from one form to the other.

Equation 1 $[M(H_2O)_6]^{2+}(aq) + 4Y^{-}(aq) \rightleftharpoons [MY_4]^{n-}(aq) + 6H_2O(l)$

(i) Consider your observations in Table 3.1 Test 2, identify the yellow complex and explain your answer.

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(ii) Considering the information provided and your answer in **3(b)(i)**, explain the chemistry involved that account for the change(s) observed in Table 3.1 Test 1.

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[Total: 9]

4 Planning

The reaction between an acid and a metal hydroxide is exothermic. It is possible to make use of this fact to determine the equivalence-point of a neutralisation reaction without the use of an indicator. This process is known as thermometric titration, and can be used to calculate the molar concentration of an acid solution.

Portions of acid are progressively added to the hydroxide solution until the equivalence-point is reached and passed. The temperature is monitored throughout the experiment.

The data obtained is plotted and two best-fit graph lines are drawn. One line is drawn using data before the equivalence-point and the second line using the remaining data. These lines are then extrapolated (extended) until they meet.

(a) **(i)** Aqueous hydrochloric acid is neutralised by the addition of aqueous barium hydroxide. The enthalpy change of neutralisation has the symbol ΔH_{neut}.

Write an ionic equation for the reaction for which the enthalpy change is ΔH_{neut}. You should include state symbols in your equation.

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(ii) A student carried out a thermometric titration by adding 1.50 mol dm^{-3} hydrochloric acid to 1.00 mol dm^{-3} barium hydroxide, $Ba(OH)_2$.

Given that the magnitude of ΔH_{neut} is about 57 kJ mol^{-1}, show that 25.0 cm^3 of barium hydroxide used will give a temperature change of 11.7 °C.

You should assume that 4.18 J of heat energy changes the temperature of 1.0 cm^3 of solution by 1.0 °C.

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(b) Using the information provided, you are required to write a plan for a thermometric titration in which hydrochloric acid is added to aqueous barium hydroxide.

You may assume that you are provided with:

- 1.00 mol dm^{-3} $Ba(OH)_2(aq)$,
- hydrochloric acid of approximate concentration 1.50 mol dm^{-3},
- the equipment normally found in a school or college laboratory.

Your plan should include:

- brief, but specific, details of the apparatus you would use, bearing in mind the levels of precision they offer;
- an outline of how accurate results would be obtained;
- how you would recognise that the equivalence-point had been passed.

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(c) **(i)** On the axes provided, sketch the graph you would expect to obtain.

You should consider your answer in **(a)(ii)** and label the graph.

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(ii) Explain the shape of your graph in **(c)(i)**.

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(iii) The exact concentration of HC*l* provided is higher than 1.50 mol dm$^{`3}$.

On the same axes provided in **(c)(i)**, sketch the graph you would expect to obtain.

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[Total: 12]

2022 H2 Chemistry Preliminary Examination Paper 4

Suggested Answers

1 Investigation of the oxidising ability of substances

(a) (i) Table 1.1

	Test	Observations
1	To a 1 cm depth of **FA 1** in a test-tube, add 1 cm depth of **FA 4**, then gradually add **FA 2** till 1 drop in excess.	Purple **FA 2** decolourised. Pale green/colourless **FA 1** turned pale yellow and finally orange/ pink.
2	To a 1 cm depth of **FA 3** in a test-tube, add 1 cm depth of **FA 2**.	Purple $KMnO_4$ decolourised. A brown solution/ppt formed. Effervescence/ bubbles of gas evolved rapidly. Colourless, odourless gas evolved relight a glowing splint. The gas is oxygen.
3	To a 1 cm depth of **FA 1** in a test-tube, add about 1 cm depth of **FA 4**, followed by 1 cm depth of **FA 3**. To a portion of resulting solution, add aqueous sodium hydroxide till excess.	Pale green/colourless **FA 1** turned yellow. Red-brown ppt formed is insoluble in excess NaOH(aq). Effervescence/ bubbles of gas evolved (rapidly).

(ii) Purple MnO_4^- oxidised (pale green) iron(II) to (yellow) iron(III) ions.

Itself is reduced to colourless Mn^{2+}. (reject pale Mn2+)

(iii) $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

(iv) Compound **A** / **FA 3**

(b) (i) Titration results

Titration number	1	2
Final burette reading /cm^3	24.20	24.25
Initial burette reading /cm^3	0.00	1.00

Volume of **FA 2** (added) /cm^3	24.20	24.25

(ii) average volume of **FA 2** used = $\frac{24.20+24.25}{2}$

$= 24.23\ \text{cm}^3$

(c) **(i)** $[\text{KMnO}_4] = \frac{0.750}{39.1+54.9+(4\times16.0)}\times\frac{1000}{250} = 0.01899\ \text{mol dm}^{-3}$

amount of MnO_4^- = $0.01899\times\frac{V_{FA2}}{1000}$ mol

$= 0.000460$ mol

(ii) amount of Fe^{2+} in 25.0 cm^3 = $(c)(i)\times\frac{5}{1}$ mol

$[Fe^{2+}] = (c)(i)\times\frac{5}{1}\times\frac{1000}{25.0}$ mol dm^{-3}

$= 0.0920$ mol dm^{-3}

(d) Identify the cause: Chloride is oxidised by/ reacts with MnO_4^-

And any one of the following modification : **M16**

- The titration needs to be carried out in the fumehood. Chloride is oxidised by MnO_4^- to give toxic chlorine gas.
- Iron(II) chloride needs to be diluted prior to titration. The titre will exceed 50.00 cm^3 if iron(II) chloride is not diluted.
- Prepare higher concentration of MnO_4^- for used. The titre will exceed 50.00 cm^3 if **FA 2** with original concentration is used.
- Using a smaller pipette/ burette, measure a smaller volume of iron(II) chloride for titration, so that the titre will not exceed 50.00 cm^3.

($5FeCl_2(aq) + 3MnO_4^-(aq) + 24H^+(aq) \rightarrow 5Fe^{3+}(aq) + 5Cl_2(g) + 3Mn^{2+}(aq) + 12H_2O(l)$)

(e) **(i)** As V_{FA3} increases, more compound **A**/**FA 3** was added to oxidise Fe^{2+} in **FA 1**. This leaves less Fe^{2+} to be oxidised by MnO_4^- in **FA 2**.

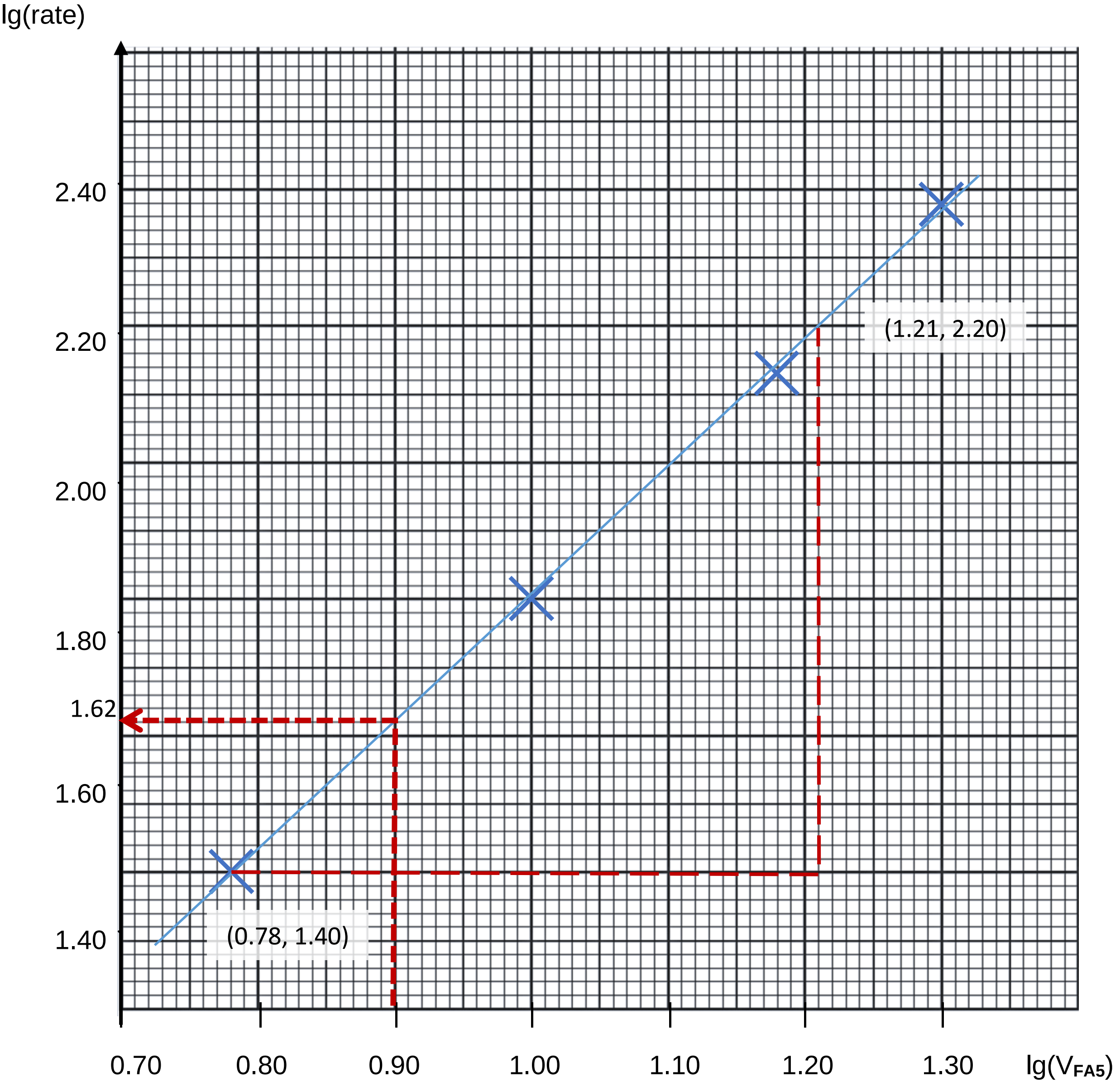
(ii)
- Not an anomaly.
- (Compound **A** in **FA 3** is both an oxidising and reducing agent.)
- In experiment 5, compound **A** is in excess/ Fe^{2+} is limiting.
- The (excess) compound **A** is oxidised by MnO_4^-.

(iii) percentage uncertainty = $\frac{2 \times 0.05}{1.25} \times 100 = 8.0\ \%$

2 **Results**

Expt	$V_{FA\ 5}$/ cm^3	V_{H_2O} /cm^3	Reaction time, ***t*** /s	lg($V_{FA\ 5}$)	lg (rate)
1	20.00	0.00	15.1	1.30	2.38
2	10.00	10.00	56.5	1.00	1.80
3	15.00	5.00	26.9	1.18	2.13
4	6.00	14.00	144.5	0.778	1.40

(b)



(c) **(i)** Gradient of line = $\frac{1.40-2.20}{0.78-1.21}$

$= 1.86$ (3 s.f.)

$\boldsymbol{m} = 2$ (nearest integer)

(ii) From graph, when $\lg(7.9) = 0.898$,

$$\lg\left(\frac{3600}{\text{reaction time}}\right) = 1.62$$

reaction time = $\frac{3600}{10^{1.62}} = 86.4$ s

(d) **(i)** Experiment 1: [KI] = $0.0500 \times \frac{20.00}{60.00}$ = 0.016667 mol dm^{-3}

Experiment 2: [KI] = $0.0500 \times \frac{10.00}{60.00}$ =0.008333 mol dm^{-3}

$$\frac{20.00}{10.00} = \frac{0.016667}{0.008333} = 2$$

(ii) Using results of experiment 1 and 2,

$$\frac{\text{rate in experiment 1}}{\text{rate in experiment 2}} = \frac{\frac{3600}{t_1}}{\frac{3600}{t_2}} = \frac{56.5}{15.1} = 3.74 \approx 4$$

When [KI] doubled, rate of experiment 2 is 4 times that of experiment 1.

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

(ii) Not as good as that in **(a)**. With any of the following reasons:

- There is <u>less thiosulfate left</u> in the reaction mixture <u>to react with iodine formed/ so shorter time recorded</u>. (or words to the effect)
- More time is needed to transfer the content in measuring cylinder/ 20.00 cm^3 of solution into the beaker for mixing

Or

As good as that in **(a)**. With any of the following reasons:

- [$S_2O_3^{2-}$] is very small, so reaction with iron(III) ions/H$^+$/acid will be very slow and negligible.
- [$S_2O_3^{2-}$] decrease by similar extent for each experiment, the relative rate of experiments is not affected.

3 **(a)** **Table 3.1**

	Test	Observations
1	Place a test-tube containing 2 cm depth of **FA 9** in an almost boiling water bath for a few minutes.	Blue/ bluish-green **FA 9** turned green.
2	To a 2 cm depth of **FA 9** in a test-tube, add gradually add 2 cm depth of **FA 10**. Pour half of this mixture into another test-tube and place it in the almost boiling water bath for a few minutes for comparison	(Blue/ bluish-green) **FA 9** turned green upon adding **FA 10**. The resultant solution turned yellowish-green/ brighter green after warming.
3	To a 1 cm depth of **FA 9** in a boiling-tube, gradually add 3 cm depth of aqueous sodium hydroxide. Filter the resultant mixture and collect the filtrate. Keep the filtrate for Test 4.	Pale blue/ Blue-green/greenish-blue ppt formed is insoluble in excess NaOH(aq). The residue is dark blue. The filtrate is a colourless solution.
4	To a 1 cm depth of filtrate in a test-tube, add 1 cm depth of nitric acid, followed by $AgNO_3$(aq). To a portion of the resultant mixture, add NH_3(aq).	White ppt formed with $AgNO_3$(aq) is soluble in aqueous ammonia to give a colourless solution.

(b) **(i)** **Cation** Copper(II) ion/ Cu^{2+}

(ii) Either one of the following:

- Colour of the white ppt cannot be seen clearly in dark-coloured **FA 9** solution.
- Addition of NH_3(aq) could precipitate/form dark blue complex with Cu^{2+} if present.

(c) **(i)** $[CuCl_4]^{2-}$

In Test 2, when **FA 10**/ saturated NaY/Y^- is added to **FA 9**, a yellow complex is formed.

The solution is green due to presence of both blue and yellow complexes/ The solution turned green due to more yellow complex formed.

(ii) The conversion of blue to yellow complex is an endothermic process.

Upon warming, the formation of yellow complex is favoured/ POE of Equation 1 shift to the right to absorb some of the added heat.

4 **(a)** **(i)** $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

(ii) Amount of OH^- in 25.0 cm^3 = $2 \times 1.00 \times \frac{25.00}{1000}$

= 0.0500 mol

Volume of HC*l* required for complete neutralisation = $\frac{0.0500}{1.50} \times 1000$

= 33.33 cm^3

Heat evolved = 57 000 × 0.0500 = (25.0 + 33.33) × 4.18 × ΔT

ΔT = 11.7 °C

(b) **<u>Procedure</u>**

1. Fill a 50.00 cm^3 burette to 0.00 cm^3 mark with 1.50 mol dm^{-3} HC*l*(aq).
2. Place a Styrofoam cup in a 250 cm^3 beaker to prevent it from toppling.
3. Pipette 25.0 cm^3 $Ba(OH)_2$ into the Styrofoam cup.
4. Using a thermometer, measure and record the initial temperature of the $Ba(OH)_2$, $\mathbf{T_{initial}}$.
5. Place a lid on the Styrofoam cup to minimise heat loss to the surroundings.
6. Run 5.00 cm^3 of HC*l*(aq) from the burette into the cup though an opening in the lid, stir the solution carefully with the thermometer and record the highest temperature reached, T.
7. Immediately run a further 5.00 cm^3 of HC*l*(aq) from the burette into the cup, stir and record the highest temperature reached, T.
8. Continue the addition of HC*l*(aq), in 5.00 cm^3 portions, until 50.00 cm^3 have been run from the burette.

<u>Recognition of Equivalence Point</u>

After equivalence point, highest temperature reached will be lower for each portion of HC*l*(aq) added.

M47 – Appropriate apparatus for volumes measurement

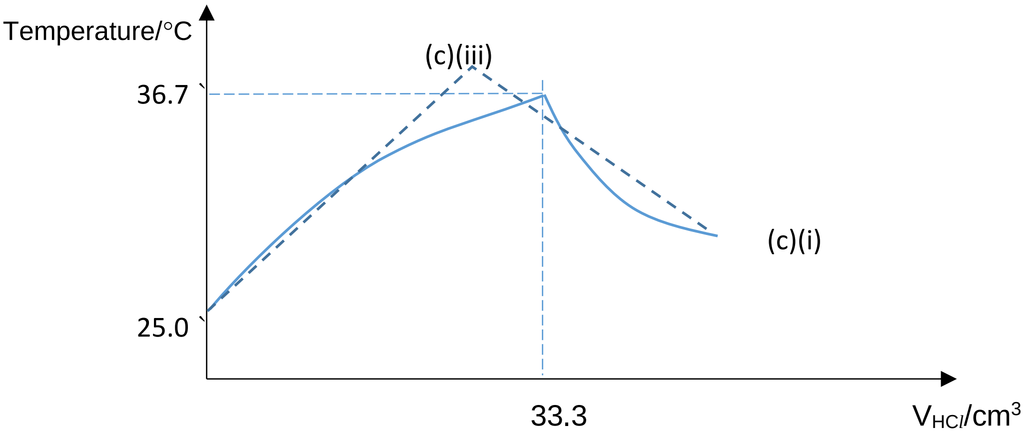
M48 – Details for measurement of volumes using burette

M49 – Measurement of temperatures

M50 – Ensuring accuracy

M51 – Recognition of equivalence point

(c) (i)



(ii) **M53** – Explain T rise due to <u>more</u> reaction occurs.

Initially as V_{HCl} increases, larger amount of $Ba(OH)_2$ is neutralised to produce more heat. Therefore, T increases.

M54 – Explain T fall

(After equivalence point, $Ba(OH)_2$ is completely neutralised.)

Either one of the following:

- Same amount of heat is distributed over an increasingly larger total volume of solution, causing T to fall.
- <u>Excess</u> HC*l* <u>at lower temperature/ room temperature</u> cools down the mixture.