NAME ______________________ CLASS **20S**

JURONG PIONEER JUNIOR COLLEGE
JC2 PRELIMINARY EXAMINATION 2021

CHEMISTRY
9729/04

Higher 2
17 August 2021

Paper 4 Practical
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, class and exam index number 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 a 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 17 and 18.

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	
2	
3	
Total	

This document consists of **17** printed pages and **1** blank page.

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

1 Determination of the solubility product of calcium iodate

Calcium iodate, $Ca(IO_3)_2$, is a sparingly soluble salt and can dissociate in water according to the following equation.

equation 1 $Ca(IO_3)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2IO_3^-(aq)$

In this experiment you will determine the solubility product of calcium iodate, $Ca(IO_3)_2$.

A saturated solution of calcium iodate was prepared, the concentration of iodate ions, IO_3^-, in this saturated solution can be determined by reacting IO_3^- ions with an excess of I^- ions to form iodine, I_2.

equation 2 $IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$

The amount of iodine formed may be titrated against a standard solution of $Na_2S_2O_3$ as shown in equation 3.

equation 3 $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

FA 1 is a filtered saturated solution of calcium iodate, $Ca(IO_3)_2$.

FA 2 is an aqueous strong acid

FA 3 is 0.500 mol dm^{-3} potassium iodide, KI.

FA 4 is 0.150 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.

(a) (i) Titration of FA 1 against FA 4

1. Fill the burette with **FA 4**.
2. Use a pipette to transfer 25.0 cm^3 of **FA 1** into a 250 cm^3 conical flask.
3. Use a measuring cylinder to add 25 cm^3 of **FA 2** to the conical flask.
4. Use another measuring cylinder to add 10 cm^3 of **FA 3** to the conical flask. The solution will turn brown as iodine is produced.
5. Run **FA 4** from the burette into this flask. Near the end-point, when the yellow solution becomes pale, add about 1 cm^3 of starch indicator.
6. Continue adding **FA 4** slowly. The end-point is reached with the solution first becomes colourless.
7. Record your titration results, to an appropriate level of precision, in the space provided on page 3.
8. Repeat points 2 to 7 as necessary until consistent results are obtained.

Titration results

[3]

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

V_{FA4} = ... [3]

(b) **(i)** Calculate the amount of IO_3^- ions present in 25.0 cm^3 of **FA 1**.

amount of IO_3^- ions present in 25.0 cm^3 of **FA 1**= .... [2]

(ii) Calculate the concentration of IO_3^- ions in **FA 1**.

concentration of IO_3^- ions in **FA 1** = ... [1]

(iii) Hence, determine the solubility product of calcium iodate, $Ca(IO_3)_2$.

solubility product of calcium iodate, $Ca(IO_3)_2$ = ... [1]

(c) **(i)** The literature value of the solubility product of $Ca(IO_3)_2$ is 6.47×10^{-6} mol^3 dm^{-9}.

Calculate the experimental error in your experiment by expressing the difference between your calculated solubility product in **(b)(iii)** and the literature value as a percentage of the literature value.

experimental error = ... [1]

(ii) Calculate the maximum percentage error in the volume of **FA 4** used in **1(a)(ii)**.

maximum percentage error in the volume of **FA 4** used = ... [1]

(iii) A student was provided with an **FA 1** solution which was not filtered properly and contained traces of undissolved $Ca(IO_3)_2$ suspended in the solution.

Predict and explain how this would affect the titre obtained in her experiment.

....

....

.... [1]

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(d) Planning

A student suggested that the dissolution of calcium iodate, $Ca(IO_3)_2$, shown in equation **1**, is endothermic.

equation 1 $Ca(IO_3)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2IO_3^-(aq)$

The actual value of the enthalpy change, ΔH, and the entropy change, ΔS, of the reaction can be determined from the following equation.

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

T is the reaction temperature in Kelvin.

R is the molar gas constant.

The procedure you followed in **1(a)(i)** can be modified to investigate the effect of temperature, T, on the solubility product of $Ca(IO_3)_2$. Plotting $\ln K_{sp}$ against $\frac{1}{T}$ gives a straight line of best fit. The enthalpy change, ΔH, and entropy change, ΔS, of the reaction can then be graphically determined.

(i) Plan an investigation to determine the effect of temperature, T, on the solubility product, K_{sp}, of $Ca(IO_3)_2$.

You may assume that you are provided with

- solid calcium iodate, $Ca(IO_3)_2$.
- aqueous strong acid, **FA 2**.
- 0.500 mol dm^{-3} potassium iodide, KI, **FA 3**.
- 0.150 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$, **FA 4**.
- starch indicator
- the equipment normally found in a school or college laboratory.

In your plan you should include brief details of

- the reactants and conditions that you would use,
- how you would prepare a saturated solution,
- the procedure that you would follow and the measurements that you would take,
- how you would determine the K_{sp} of $Ca(IO_3)_2$ for each experiment.

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....

[7]

(ii) Sketch, on Fig. 2.1, the graph you would expect to obtain if the reaction is endothermic as suggested by the student.

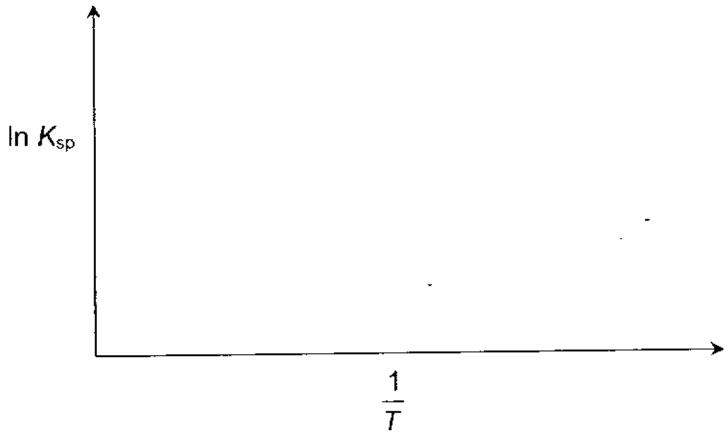

Fig. 1.1

[1]

(iii) Describe how you would use your graph to determine values for ΔH and ΔS, showing your working on the graph clearly.

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.... [3]

[Total: 24]

2 Determination of the enthalpy change of neutralisation, ΔH_{neu}, of an unknown strong acid by a strong base

When an aqueous strong acid is mixed with an aqueous strong base, the neutralisation reaction releases heat causing an increase in the temperature of the solution.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

FA 2 is 0.500 mol dm^{-3} unknown aqueous strong acid.

FA 5 is 1.00 mol dm^{-3} sodium hydroxide, NaOH.

In this experiment you will mix different volumes of **FA 2** and **FA 5**, keeping the total volume constant. For each mixture, the change in temperature, ΔT, will be determined and used to plot a graph of ΔT against volume of **FA 5** used.

You will then use data from the graph to determine a value for the enthalpy change of neutralisation, ΔH_{neu}, and the basicity of the strong acid used.

(a) Determine the change in temperature for a series of reactions between FA 2 and FA 5

1. Fill the burette with **FA 5**.
2. Place the styrofoam cup in a 250 cm^3 beaker to prevent it from tipping over. Transfer 10.00 cm^3 of **FA 5** from the burette into the cup. Measure the temperature of the **FA 5** solution using the thermometer. Record the initial temperature of **FA 5** as $T_{FA\ 5}$.
3. Wash and dry the thermometer.
4. Use a measuring cylinder to measure 40.0 cm^3 of **FA 2**. Measure the temperature of the **FA 2** solution and record the initial temperature of **FA 2** as $T_{FA\ 2}$.
5. Add **FA 2** to **FA 5** in the styrofoam cup. Stir the mixture using the thermometer and record the maximum temperature, T_{max}, reached.
6. Wash and shake dry the styrofoam cup.
7. Repeat steps 2 to 6 using 15.00 cm^3, 20.00 cm^3, 30.00 cm^3, 35.00 cm^3 and 40.00 cm^3 of **FA 5** and appropriate volumes of **FA 2** each time, such that the total volume of the reacting mixture is 50 cm^3.

In an appropriate format in the space provided, record:

- All measurements of volumes used,
- All temperatures measured,
- $T_{weighted\ initial}$ and the change in temperature, ΔT to 1 decimal place.

ΔT should be calculated using the following formula:

$\Delta T = T_{max} - T_{weighted\ initial}$

$$T_{weighted\ initial} = \frac{(\text{volume of } \mathbf{FA\ 5} \times T_{FA\ 5}) + (\text{Volume of } \mathbf{FA\ 2} \times T_{FA\ 2})}{\text{Volume of } \mathbf{FA\ 5} + \text{Volume of } \mathbf{FA\ 2}}$$

Results

[4]

(b) On the grid provided, plot a graph of ΔT (y-axis) against volume of **FA 5** (x-axis) using the data you obtained in **(a)**.

The scale for ΔT should extend at least 2°C above your greatest temperature rise.

(i) Draw two lines of best fit.

- The first best-fit line should be drawn using the plotted points before the maximum change in temperature.
- The second best-fit line should be drawn using the plotted points after the maximum change in temperature.

Extrapolate these line until they cross. [3]

(ii) Determine from your graph, the maximum change in temperature, ΔT_{max}, and the volume, V_{max}, of **FA 5** required to obtain this value.

ΔT_{max} = ...

V_{max} = ... [1]

(c) **(i)** Calculate the amount of H^+ that would react completely with the amount of NaOH contained in V_{max} of **FA 5**.

amount of H^+ that reacted with NaOH in V_{max} of **FA 5** = ... [1]

(ii) Using your answer in **(c)(i)**, determine if **FA 2** is monobasic, dibasic or tribasic.

Basicity of **FA 2** = ... [2]

(iii) Calculate the heat change for the neutralisation reaction at ΔT_{max}.

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

Heat change = ... [1]

Using your answers in **(c)(iii)**, calculate a value for the enthalpy change of neutralisation, ΔH_{neu}.

ΔH_{neu} = ... [4]

(d) A student performed a similar experiment by mixing **FA 5** with 0.500 mol dm^{-3} aqueous ethanoic acid.

Predict how the values of ΔT_{max} and V_{max} will change for this experiment by ticking on the appropriate boxes provided.

	decrease	no change	increase
ΔT_{max}			
V_{max}			

[1]

[Total: 17]

3 Investigation of the chemistry of some vanadium ions

FA 6 is an acidic solution containing VO_2^+ ions.

Vanadium, is able to exhibit variable oxidation states in its compounds. When **FA 6** is reacted with **FA 7**, a **series of colour changes occur**.

Table 3.1 gives some standard electrode potential values relevant to this question.

Table 3.1

Electrode reaction	E^0/ V
$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq)$	−0.26
$VO^{2+}(aq) + 2H^+ + e^- \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
$VO_2^+(aq) + 2H^+ + e^- \rightleftharpoons VO^{2+}(aq) + H_2O(l)$	+1.00
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14

Perform the tests described in Tables 3.2, 3.3 and 3.4, and record your observations in the table. Unless otherwise stated, the volumes given below are approximate and should be estimated rather than measured. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 3.2

		tests	observations
(a)	**(i)**	To about 5 cm depth of **FA 6** in a test-tube, add a spatula measure of **FA 7**. Shake the mixture. Leave the mixture for 5-15 minutes, observing the mixture occasionally. The reaction is complete when a violet solution is obtained. Filter the reaction mixture into a test-tube. This is solution **X** which contains V^{2+}. Label this test-tube **X**.	
	(ii)	Wash a teat-pipette with solution **X**. Using this teat-pipette, transfer 20 drops of **X** into a test-tube. Add aqueous potassium manganate(VII) dropwise with shaking, counting the number of drops needed to give the first permanent pink or orange colour.	

[5]

(b) **FA 8** is an aqueous solution that has been made by reacting solid **FA 7** with dilute sulfuric acid.

Table 3.3

	tests	observations
(i)	To about 1 cm depth of **FA 8** in a test-tube, add aqueous sodium hydroxide until no further changes are observed.	
(ii)	To about 1 cm depth of **FA 8** in a test-tube, add aqueous ammonia until no further changes are observed.	

[1]

(iii) Identify **FA 7**.

FA 7 is .... [1]

Table 3.4

		tests	observations
(c)	(i)	To about 3 cm depth of **FA 6** in a boiling tube, add a spatula measure of tin powder. Shake the mixture. Gently warm the mixture. In order to view the colours clearly you may find it necessary, from time to time, to allow the tin to settle. Warm the mixture until no further changes are seen. The solution obtained is solution **Y** which contains another of the reduction products of the vanadium ion present in **FA 6**.	

[1]

(c) **(ii)**

Table 3.1

Electrode reaction	$E^\ominus$/ V
$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq)$	–0.26
$VO^{2+}(aq) + 2H^+ + e^- \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
$VO_2^+(aq) + 2H^+ + e^- \rightleftharpoons VO^{2+}(aq) + H_2O(l)$	+1.00
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	–0.14

Suggest which ion is responsible for the final colour observed in **3(c)(i)**. Explain your answer using data from Table 3.1.

Ion: ....

Explanation:

....

....

....

....

....

....

....

.... [3]

(iii) Write an overall chemical equation to represent the reaction occurring in **(c)(i)**.

.... [1]

(iv) If the experiment in **3(a)(ii)** was repeated using solution **Y**, calculate the number of drops of aqueous potassium manganate(VII) that would be required to give the first permanent pink or orange colour. Explain your calculation in terms of the redox reactions involved.

Number of drops: ....

Explanation: ....

....

....

.... [2]

[Total: 14]

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 Aqueous Anions

Anion	*Reaction*
carbonate, CO_3^{2-}	CO_2 liberated by dilute acids
choride, $Cl^-(aq)$	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 → (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 on warming with dilute acids; gives white ppt. with $Ba^{2+}(aq)$ (soluble in excess dilute strong acids)

(c) Tests for Gases

gas	*Test and test results*
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 acidified aqueous 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

NAME ______________________ CLASS **20S**

JURONG PIONEER JUNIOR COLLEGE
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CHEMISTRY (Ans Scheme)

9729/04

Higher 2

Paper 4 Practical

17 August 2021
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Candidates answer on the Question paper.

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Shift
Laboratory

For Examiner's Use	
1	
2	
3	
Total	

This document consists of **17** printed pages and **1** blank page.

[Turn Over

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

1 Determination of the solubility product of calcium iodate

Calcium iodate, $Ca(IO_3)_2$, is a sparingly soluble salt and can dissociate in water according to the following equation.

equation 1 $Ca(IO_3)_2(s) \ll Ca^{2+}(aq) + 2IO_3^-(aq)$

In this experiment you will determine the solubility product of calcium iodate, $Ca(IO_3)_2$.

A saturated solution of calcium iodate was prepared, the concentration of iodate ions, IO_3^-, in this saturated solution can be determined by reacting IO_3^- ions with an excess of I^- ions to form iodine, I_2.

equation 2 $IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$

The amount of iodine formed may be titrated against a standard solution of $Na_2S_2O_3$ as shown in equation 3.

equation 3 $I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

FA 1 is a filtered saturated solution of calcium iodate, $Ca(IO_3)_2$.

FA 2 is an aqueous strong acid

FA 3 is 0.500 mol dm^{-3} potassium iodide, KI.

FA 4 is 0.150 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.

(a) (i) **Titration of FA 1 against FA 4**

1. Fill the burette with **FA 4**.
2. Use a pipette to transfer 25.0 cm^3 of **FA 1** into a 250 cm^3 conical flask.
3. Use a measuring cylinder to add 25 cm^3 of **FA 2** to the conical flask.
4. Use another measuring cylinder to add 10 cm^3 of **FA 3** to the conical flask. The solution will turn brown as iodine is produced.
5. Run **FA 4** from the burette into this flask. Near the end-point, when the yellow solution becomes pale, add about 1 cm^3 of starch indicator.
6. Continue adding **FA 4** slowly. The end-point is reached with the solution first becomes colourless.
7. Record your titration results, to an appropriate level of precision, in the space provided.
8. Repeat points 2 to 7 as necessary until consistent results are obtained.

Titration results

Titration number	1	2
Final burette reading / cm^3	23.30	46.70
Initial burette reading / cm^3	0.00	23.30
Volume of FA4 used / cm^3	23.30	23.40
	✓	✓

[3]

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

$$V_{FA4} = \frac{23.30 + 23.40}{2} = \underline{\mathbf{23.35}}\ \mathbf{cm^3}$$

V_{FA4} = $\underline{\mathbf{23.35}}$ $\mathbf{cm^3}$ [3]

(b) **(i)** Calculate the amount of IO_3^- ions present in 25.0 cm^3 of **FA 1**.

Amount of $S_2O_3^{2-}$ reacted = $\frac{23.35}{1000} \times 0.150 = 0.00350$ mol

$IO_3^- \sim 3I_2 \sim 6S_2O_3^{2-}$

Amount of IO_3^- reacted = $\frac{1}{6} \times 0.00350 = 0.000584$ mol

amount of IO_3^- ions present in 25.0 cm^3 of **FA 1**= $\underline{0.000584}$ mol [2]

(ii) Calculate the concentration of IO_3^- ions in **FA 1**.

$[IO_3^-] = 0.000584 \div \frac{25.0}{1000} = 0.0234$ mol dm^{-3}

concentration of IO_3^- ions in **FA 1** = $\underline{0.0234}$ mol dm^{-3} [1]

(iii) Hence, determine the solubility product of calcium iodate, $Ca(IO_3)_2$.

	$Ca(IO_3)_2(s)$	$\ll$	$Ca^{2+}(aq)$	+	$2IO_3^-(aq)$
Eqm conc.			$\frac{1}{2} \times 0.0234$ $= 0.0117$		0.0234

Ksp of $Ca(IO_3)_2 = [Ca^{2+}][IO_3^-]^2$

$= (0.0117)(0.0234)^2$

$= 6.41 \times 10^{-6}$ mol^3dm^{-9}

solubility product of calcium iodate, $Ca(IO_3)_2$ = $\underline{6.41 \times 10^{-6}}$ mol^3dm^{-9} [1]

[Turn Over

(c) **(i)** The literature value of the solubility product of $Ca(IO_3)_2$ is 6.47×10^{-6} $mol^3\ dm^{-9}$.

Calculate the experimental error in your experiment by expressing the difference between your calculated solubility product in **1(b)(iii)** and the literature value as a percentage of the literature value.

$$\text{Experimental error} = \frac{6.47\times10^{-6} - 6.41\times10^{-6}}{6.47\times10^{-6}}\times100 = \underline{0.927}\%$$

[1]

(ii) Calculate the maximum percentage error in the volume of **FA 4** used in **1(a)(ii)**.

$$\text{Max. \% error in volume of } \textbf{FA4} \text{ used} = \frac{2(0.05)}{23.35}\times100 = \pm\ \underline{0.428}\ \%$$

maximum percentage error in the volume of **FA 4** used = ± **<u>0.428</u>** % [1]

(iii) A student was provided with an **FA 1** solution which was not filtered properly and contained traces of undissolved $Ca(IO_3)_2$ suspended in the solution.

Predict and explain how this would affect the titre obtained in her experiment.

<u>Amount of $Ca(IO_3)_2$</u> in the conical flask is <u>larger</u>, which produces more I_2 that required more $S_2O_3^{2-}$ to reach end point. Thus, **<u>titre value will be larger</u>**.

[1]

(d) Planning

A student suggested that the dissolution of calcium iodate, $Ca(IO_3)_2$, shown in equation **1**, is endothermic.

equation 1 $Ca(IO_3)_2(s) \ll Ca^{2+}(aq) + 2IO_3^-(aq)$

The actual value of the enthalpy change, ΔH, and the entropy change, ΔS, of the reaction can be determined from the following equation.

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

T is the reaction temperature in Kelvin.

R is the molar gas constant.

The procedure you followed in **1(a)(i)** can be modified to investigate the effect of temperature, T, on the solubility product of $Ca(IO_3)_2$. Plotting $\ln K_{sp}$ against $\frac{1}{T}$ gives a straight line of best fit. The enthalpy change, ΔH, and entropy change, ΔS, of the reaction can then be graphically determined.

(i) Plan an investigation to determine the effect of temperature, T, on the solubility product, K_{sp}, of $Ca(IO_3)_2$.

You many assume that you are provided with

- solid calcium iodate, $Ca(IO_3)_2$.
- aqueous strong acid, **FA 2**.
- 0.500 mol dm^{-3} potassium iodide, KI, **FA 3**.
- 0.150 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$, **FA 4**.
- starch indicator
- the equipment normally found in a school or college laboratory.

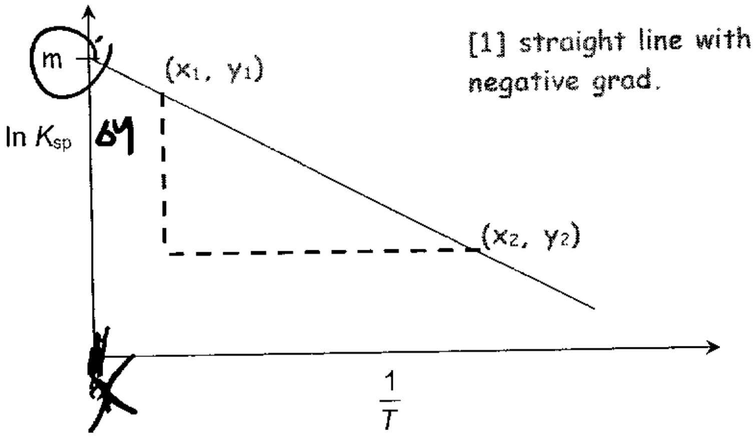
In your plan you should include brief details of

- the reactants and conditions that you would use,
- how you would prepare a saturated solution,
- the procedure that you would follow and the measurements that you would take,
- how you would determine the K_{sp} of $Ca(IO_3)_2$ for each experiment.

1. Use a 100 cm^3 measuring cylinder to add 100 cm^3 of deionised water into a 250 cm^3 conical flask.
2. Leave the conical flask containing the solution to stand in a thermostatically controlled water bath, with temperature set at 30 °C.
3. Using a spatula, add a few spatulas of solid $Ca(IO_3)_2$ into the flask. Stir to dissolve all the solids using a glass rod. Keep adding more solids, with stirring after each addition, until some solids are left undissolved.
4. To ensure that the solution is saturated, leave the conical flask containing undissolved solids to stand in a thermostatically controlled water bath for at least 5 mins.
5. Filter the solution through a dry filter paper and dry filter funnel into a dry 100 cm^3 conical flask placed in the water bath set at 30 °C.
6. Record the temperature of the filtrate.
7. Fill the burette with FA4.
8. Use a pipette to transfer 25.0 cm^3 of the filtrate into conical flask placed in the water bath set at 30 °C.
9. Use a measuring cylinder to add 25 cm^3 of FA 2 and another measuring cylinder to add 10 cm^3 of FA 3 to the conical flask. Iodine will be produced.
10. Quickly titrate the iodine in the conical flask with FA 4 until the colour of the solution turns pale yellow. Add a few drops of starch indicator and continue adding FA 4 dropwise until the colour of the solution changes from blue black to colourless.
11. Repeat points 7 and 10 until consistent results are obtained for the saturated solution at 30 °C.
12. Carry out the calculations from 1(b)(i) to 1(b)(iii) to determine the Ksp value at each temperature.
13. Repeat points 1 to 11 for temperatures set at 35, 40, 45 and 50°C

[7]

(d) **(ii)** Sketch, on Fig. 2.1, the graph you would expect to obtain if the reaction is endothermic as suggested by the student.



[1] straight line with negative grad.

Fig. 2.1

[1]

(iii) Describe how you would use your graph to determine values for ΔH and ΔS, showing your working on the graph clearly.

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

$\text{Gradient} = -\frac{\Delta H}{R}$

$\Delta H = -\text{gradient} \times R$

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

$\Delta S = m \times R$

[3]

[Total: 11]

2 **Determination of the enthalpy change of neutralisation, ΔH_{neu}, of an unknown strong acid by a strong base**

When an aqueous strong acid is mixed with an aqueous strong base, the neutralisation reaction releases heat causing an increase in the temperature of the solution.

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

FA 2 is 0.500 mol dm^{-3} unknown aqueous strong acid.

FA 5 is 1.00 mol dm^{-3} sodium hydroxide, NaOH.

In this experiment you will mix different volumes of **FA 2** and **FA 5**, keeping the total volume constant. For each mixture, the change in temperature, ΔT, will be determined and used to plot a graph of ΔT against volume of **FA 5** used.

You will then use data from the graph to determine a value for the enthalpy change of neutralisation, ΔH_{neut}, and the basicity of the strong acid used.

(a) Determine the change in temperature for a series of reactions between FA 2 and FA 5

1. Fill the burette with **FA 5**.
2. Place the styrofoam cup in a 250 cm^3 beaker to prevent it from tipping over. Transfer 10.00 cm^3 of **FA 5** from the burette into the cup. Measure the temperature of the **FA 5** solution using the thermometer. Record the initial temperature of **FA 5** as $T_{FA\ 5}$.
3. Wash and dry the thermometer.
4. Use a measuring cylinder to measure 40.0 cm^3 of **FA 2**. Measure the temperature of the **FA 2** solution and record the initial temperature of **FA 2** as $T_{FA\ 2}$.
5. Add **FA 2** to **FA 5** in the styrofoam cup. Stir the mixture using the thermometer and record the maximum temperature, T_{max}, reached.
6. Wash and shake dry the styrofoam cup.
7. Repeat steps 2 to 6 using 15.00 cm^3, 20.00 cm^3, 30.00 cm^3, 35.00 cm^3 and 40.00 cm^3 of **FA 5** and appropriate volumes of **FA 2** each time, such that the total volume of the reacting mixture is 50 cm^3.

In an appropriate format in the space provided, record:

- All measurements of volumes used,
- All temperatures measured,
- $T_{weighted\ initial}$ and the change in temperature, ΔT to 1 decimal place.

ΔT should be calculated using the following formula:

$\Delta T = T_{max} - T_{weighted\ initial}$

$$T_{weighted\ initial} = \frac{(\text{volume of } \textbf{FA 5} \times T_{FA\ 5}) + (\text{Volume of } \textbf{FA 2} \times T_{FA\ 2})}{\text{Volume of } \textbf{FA 5} + \text{Volume of } \textbf{FA 2}}$$

Results

Expt	**Volume of FA 5 / cm^3**	**Volume of FA 2 / cm^3**	**$T_{FA\ 5}$ / °C**	**$T_{FA\ 2}$ / °C**	**$T_{weighted\ initial}$ / °C**	**T_{max} / °C**	**ΔT / °C**
1	10.00	40.0	31.6	32.0	31.9	34.3	2.4
2	15.00	35.0					
3	20.00	30.0	31.7	31.6	31.6	37.5	5.9
4	30.00	20.0	31.7	31.6	31.7	37.5	5.8
5	35.00	15.0					
6	40.00	10.0	31.8	31.6	31.6	34.4	2.8

[4]

(b) On the grid provided, plot a graph of *Δ**T*** (y-axis) against volume of **FA 5** (x-axis) using the data you obtained in **(a)**.

The scale for *Δ**T*** should extend at least 2°C above your greatest temperature rise.

(i) Draw two lines of best fit.

- The first best-fit line should be drawn using the plotted points before the maximum change in temperature.
- The second best-fit line should be drawn using the plotted points after the maximum change in temperature.

Extrapolate these line until they cross. [3]

(ii) Determine from your graph, the maximum change in temperature, ΔT_{max}, and the volume, $\mathbf{V}_{\mathbf{max}}$, of **FA 5** required to obtain this value.

ΔT_{max} = **7.45** °C

V_{max} = **24.25** cm^3 [1]

(c) **(i)** Calculate the amount of H^+ that would react completely with the amount of NaOH contained in V_{max}, of **FA 5**.

Amount of NaOH reacted = amount of H^+ in 24.25 cm^3 of **FA 1**

= (24.25 / 1000) x 1 = 0.0243 mol

amount of H^+ that reacted with NaOH in V_{max}, of **FA 5** = 0.0243 mol [1]

(ii) Using your answer in **(c)(i)**, determine if **FA 2** is monobasic, dibasic or tribasic.

Volume of FA 2 required = 50 – 24.25 = **<u>25.75</u>** cm^3

Amount of acid in FA 2 = (25.75 / 1000) x 0.500 = **<u>0.0129</u>** mol

Ratio of amount of H^+/amount of FA 2 = $\frac{0.0243}{0.0129} = \underline{\mathbf{1.88}} \approx 2$

The acid is **<u>dibasic</u>**. [2]

(iii) Calculate the heat change for the neutralisation reaction at ΔT_{max}.

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

Heat change = (50.0 × 1.00)(4.18)(7.45) = **<u>1560</u>** J = **<u>1.56</u>** kJ

[1]

(iv) Using your answers in **2(c)(iii)**, calculate a value for the enthalpy change of neutralisation, ΔH_{neut}.

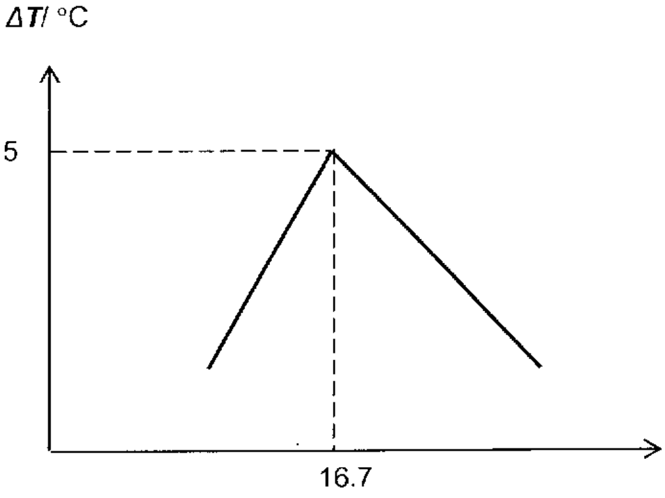
Amount of H_2O formed = amount of H^+ = 0.0243 mol

ΔH_{neut} = -(1.57 / 0.0243) = **<u>– 64.2</u>** kJ mol^{-1} [4]

(d) A student performed a similar experiment by mixing **FA 5** with 0.500 mol dm^{-3} aqueous ethanoic acid.

Predict how the values of ΔT_{max} and V_{max} will change for this experiment by ticking on the appropriate boxes provided.

	decrease	No change	increase
ΔT_{max}	✓ **[1]**		
V_{max}	✓		



Volume of **FA 5** used / cm^3

Fig. 2.1

[1]

[Total: 14]

3 Investigation of the chemistry of some vanadium ions

FA 6 is an acidic solution containing VO_2^+ ions.

Vanadium, is able to exhibit variable oxidation states in its compounds. When **FA 6** is reacted with **FA 7**, a **series of colour changes occur**.

Table 3.1 gives some standard electrode potential values relevant to this question.

Table 3.1

Electrode reaction	E^θ/ V
$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq)$	−0.26
$VO^{2+}(aq) + 2H^+ + e^- \rightleftharpoons V^{3+}(aq) + H_2O(l)$	+0.34
$VO_2^+(aq) + 2H^+ + e^- \rightleftharpoons VO^{2+}(aq) + H_2O(l)$	+1.00
$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s)$	−0.14

Perform the tests described in Tables 3.2, 3.3 and 3.4, and and record your observations in the table. Unless otherwise stated, the volumes given below are approximate and should be estimated rather than measured. Test and identify any gases evolved. If there is no observable change, write **no observable change**.

Table 3.2

		tests	observations
(a)	**(i)**	To about 5 cm depth of **FA 6** in a test-tube, add a spatula measure of **FA 7**. Shake the mixture. Leave the mixture for 5-15 minutes, observing the mixture occasionally. The reaction is complete when a violet solution is obtained. Filter the reaction mixture into a test-tube. This is solution **X** which contains V^{2+}. Label this test-tube **X**.	Effervescence H_2 gas pops with a lighted splint Expected colour change of solution Yellow → green (intermediate) → blue→ green/ blue-green → violet
	(ii)	Wash a teat-pipette with solution **X**. Using this teat-pipette, transfer 20 drops of **X** into a test-tube. Add aqueous potassium manganate(VII) dropwise with shaking, counting the number of drops needed to give the first permanent pink or orange colour.	Expected colour change of solution violet → green/ blue-green→ blue → green (intermediate) → Yellow → pink/orange 39 drops of $KMnO_4$ to give first permanent pink colour. [5]

(b) **FA 8** is an aqueous solution that has been made by reacting solid **FA 7** with dilute sulfuric acid.

Table 3.3

	tests	observations
(i)	To about 1 cm depth of **FA 8** in a test-tube, add aqueous sodium hydroxide until no further changes are observed.	**White ppt** formed **soluble in excess** NaOH to give a colourless solution
(ii)	To about 1 cm depth of **FA 8** in a test-tube, add aqueous ammonia until no further changes are observed.	**White ppt** formed **soluble in excess** NH_3 to give a colourless solution

[1]

(iii) Identify **FA 7**.

FA 7 is Zn [1]

Table 3.4

		tests	observations
(c)	**(i)**	To about 3 cm depth of **FA 6** in a boiling tube, add a spatula measure of tin powder. Shake the mixture. Gently warm the mixture. In order to view the colours clearly you may find it necessary, from time to time, to allow the tin to settle. Warm the mixture until no further changes are seen. The solution obtained is solution **Y** which contains another of the reduction products of the vanadium ion present in **FA 6**.	Expected colour change of solution **Yellow** → green (intermediate) → **blue** → **green/ blue-green** [1]

(c) **(ii)**

Table 3.1

Electrode reaction	E^θ/ V
$V^{3+}(aq) + e^- \ll V^{2+}(aq)$	−0.26
$VO^{2+}(aq) + 2H^+ + e^- \ll V^{3+}(aq) + H_2O(l)$	+0.34
$VO_2^+(aq) + 2H^+ + e^- \ll VO^{2+}(aq) + H_2O(l)$	+1.00
$Sn^{2+}(aq) + 2e^- \ll Sn(s)$	−0.14

Suggest which ion is responsible for the final colour observed in **3(c)(i)**. Explain your answer using data from Table 3.1.

Ion: $\underline{V^{3+}}$

Explanation:

Sn and VO_2^+: E^θ_{cell} = +1.00 - (- 0.14) = **+1.14** V **(energetically feasible)**

Sn and VO^{2+}: E^θ_{cell} = +0.34 - (- 0.14) = **+0.48** V **(energetically feasible)**

Sn and V^{3+}: E^θ_{cell} = (- 0.26) - (- 0.14) = - 0.12 V (not energetically feasible)

[3]

(iii) Write an overall chemical equation to represent the reaction occurring in **(c)(i)**.

$\underline{Sn + VO_2^+ + 4H^+ \rightarrow V^{3+} + 2H_2O + Sn^{2+}}$

[1]

(iv) If the experiment in **3(a)(ii)** was repeated using solution **Y**, calculate the number of drops of aqueous potassium manganate(VII) that would be required to give the first permanent pink or orange colour. Explain your calculation in terms of the redox reactions involved.

Number of drops: 2/3 × no. of drops required for X.

Explanation: n no. of drops are required to **oxidise V** from an oxidation number of **+2 to +5** (V^{2+} to VO_2^+). **2/3**n will be required to **oxidise V** from an oxidation number of **+3 to +5**. (V^{3+} to VO_2^+).

[2]

[Total: 17]