1 **To standardise a 0.03 mol dm^{-3} of iron(III) chloride solution and determine the concentration of sodium thiosulfate solution.**

(a) Planning

FA1 is a standard solution containing 0.03 mol dm^{-3} of FeCl_3 solution.

You are to design an experiment to show how **FA1** is prepared using solid FeCl_3.

You are provided with :

a solid sample of approximately 1.50 g of FeCl_3 (M_r = 162.3),

250 cm^3 volumetric flask and

the usual laboratory apparatus

In your plan, you should include details on:

- the mass of FeCl_3(s) you use with justification,
- the apparatus you would use, and
- the procedures which you would follow.

Do not carry out this plan as FA1 is provided for you to continue with 1(b).

To make standard solution, a 250 cm^3 volumetric flask is used.

Amount of FeCl_3 in 250 cm^3 = 0.03 x 250 / 1000 = 7.50 × 10^{-3} mol

Mass of FeCl_3 = 7.50 × 10^{-3} x 162.3 = 1.22 g

Procedure:

1. Weigh accurately about 1.22 g (2 d.p) of FeCl_3 into a dry clean weighing bottle.
2. Dissolve the solid using distilled water (~ 50 cm^3) in a 100 cm^3 beaker.
3. Transfer the solution after ensuring all the solid has dissolved. Rinse the weighing bottle and beaker with water and transfer all washings into a 250 cm^3 volumetric flask to ensure quantitative transfer
4. Top up the solution with distilled water to the 250 cm^3 mark.
5. Stopper and shake well to obtain a homogenous solution. Label this solution **FA 1.**

procedures 1-5.

[3]

(b) In addition to **FA1**, you are also provided with the following:

FA2 is 0.060 mol dm^{-3} potassium iodide, KI.

FA3 is aqueous sodium thiosulfate, $Na_2S_2O_3$.

When **FA1** is combined with **FA3**, a complex, $[Fe(S_2O_3)_2]^-$, is formed.

Reaction 1: $Fe^{3+}(aq) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Fe(S_2O_3)_2]^-(aq)$

To 1 cm^3 of **FA3** in a test–tube, add 2–3 drops of **FA1** and shake them thoroughly. State the colour of the complex formed.

(Pale) purple/ lilac /Pink/ violet. [1]

(c) A student aimed to determine the concentration of **FA3** using volumetric analysis involving the use of **FA1** and **FA2.**

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

Reaction 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

Using the above reactions, the student planned and carried out the following procedure.

1. Fill a 50.00 cm^3 burette with **FA3** solution.
2. Using a 10 cm^3 measuring cylinder, measure 5 cm^3 of **FA1** and place it in a 100 cm^3 conical flask.
3. Using another 10 cm^3 measuring cylinder, measure 5 cm^3 of **FA2** into the conical flask.
4. Run **FA3** from the burette into the conical flask. Near the endpoint, when the brown solution becomes pale, add 2 drops of starch.
5. Continue adding **FA3** slowly, the endpoint is reached when the solution **first becomes colourless**.
6. Record the titration results, to an appropriate level of precision, in an appropriate table and repeat the experiment once more.

(i) Carry out the student's procedure **twice only** to obtain two titration readings. **Your titre values need not be consistent**.

Note: After the end point is reached, you may observe the appearance of the coloured mixture again. There is no need to titrate any further.

Record your titration in an appropriate table in the space below.

	1	2
Final burette reading/ cm^3	14.50	28.60
Initial burette reading/ cm^3	0.00	14.50
Volume of **FA3**/ cm^3	14.50	14.10

[2]

(ii) Planning

A teacher commented that it is appropriate for the student to use 5 cm^3 instead of 25 cm^3 of **FA1** for this titration. Justify the teacher's comment using suitable calculations.

Assume the concentration of **FA3** is 0.006 mol dm^{-3}.

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

Reaction 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

No. of moles of 5 cm^3 of **FA1** = 5 / 1000 x 0.03 = 1.5×10^{-4} mol
No. of moles of I_2 = 1.5×10^{-4} / 2 = 7.5×10^{-5} mol
No. of moles of $S_2O_3^{2-}$ = 7.5×10^{-5} x 2 = 1.5×10^{-4} mol

If the concentration of **FA3** is 0.006 mol dm^{-3},
Approximate titre volume = 1.5×10^{-4} / 0.006 x 1000 = 25 cm^3
OR
If 25 cm^3 of **FA1** = 25/ 1000 x 0.03 = 7.5×10^{-4} mol
No. of moles of $S_2O_3^{2-}$ = 7.5×10^{-4} mol
Approximate titre volume = 7.5×10^{-4} / 0.006 x 1000 = 125 cm^3

If 25 cm^3 of **FA1** is used, the titre value will be 5 times (125 cm^3) and exceeds the capacity of the burette. It results in high inaccuracy as refilling of burette is required.

[2]

(iii) Planning

Justify whether the amount of **FA2** that the student used in step **3** of the procedure in **1(c)** is sufficient to determine the concentration of **FA3**.

No. of moles of KI required = 1.5×10^{-4} mol
Minimum volume of **FA2** = $1.5 \times 10^{-4} / 0.06 \times 1000 = 2.5$ cm^3
To ensure excess **FA2** (KI), add 5 cm^3 of KI (accept 3 – 5 cm^3).

[1]

(iv) The student did the titration several times but could not get consistent titre values of **FA3**. Based on your observation during titration, suggest a reason why.

Reaction 1: $Fe^{3+}(aq) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Fe(S_2O_3)_2]^-(aq)$
Reaction 2: $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$
Reaction 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$
After the dark blue colouration disappear to get the colourless solution, the purple colour reappeared because the position of equilibrium of reaction 1 shift left due to the decrease in $[S_2O_3^{2-}]$ during the titration. This result in the increase in the $[Fe^{3+}]$ which reacts with excess KI to produce more I_2. The presence of I_2 in starch is revealed as the reappearance of blue black colouration. Hence, it is difficult to judge the correct end point and titration results is consistent.

[1]

(v) Suggest another reason why the results obtained are not reliable.

The limiting reagent $FeCl_3$ was measured using a 10 cm^3 measuring cylinder which has a **low precision**. Hence, the amount measured in each titration may vary slightly.
OR
The volume of **FA1** and **FA2** used are small, this would results in high percentage error.

[1]

(vi) The student performed the titration and had a titration value of 18.50 cm^3. Given the errors (uncertainties) associated with each reading using a measuring cylinder and burette are ±0.1 cm^3 and ±0.05 cm^3 respectively, calculate the maximum total percentage error (uncertainty) from the apparatus in the titration.

Percentage uncertainty (error) in using the **measuring cylinder**
= (± 0.1 / 5.0) x 100 % = ± **2.00 %**

Each measuring cylinder reading has an uncertainty (error) of ± 2.00 %.
Since two readings using measuring cylinder are made, total percentage error from measuring cylinder is **2.00 x 2 = ± 4.00 %**

Percentage uncertainty (error) in using the **burette**
$= \frac{(\pm 0.05 \times 2)}{18.50} \times 100\ \% = \pm$ **0.541 %**

Hence **maximum percentage error (uncertainty)** from the apparatus in the experiment
= (± 4.00) + (± 0.541) = ± **4.54 %** (3 significant figures) [3]

[Total: 14 marks]

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

In this experiment, you are investigating the rate of reaction between iron(III) ions and iodide ions.

You are provided with:

FA1 is 0.03 mol dm^{-3} of $FeCl_3$ solution

FA2 is 0.060 mol dm^{-3} potassium iodide, KI

FA3 is sodium thiosulfate, $Na_2S_2O_3$

The reaction is started by mixing a solution of iron(III) chloride with sodium thiosulfate, potassium iodide, and starch. The iodine, I_2, produced in **reaction 2** reacts immediately with thiosulfate ions, $S_2O_3^{2-}$ in **reaction 3**.

Reaction 1: $Fe^{3+}(aq) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Fe(S_2O_3)_2]^-(aq)$

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

Reaction 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

When all the thiosulfate have been used, the iodine produced will turn starch indicator blue–black. The rate of the reaction can therefore be determined by finding the time for the blue–black colour to appear.

You are advised to read the instructions before starting any practical work and draw a table, in an appropriate format, for your results in the space on page 10.

For each of the five experiments, you will need to include the

- volume of **FA2**, $\mathbf{V_{FA2}}$,
- volume of water, $\mathbf{V_{water}}$,
- calculated squared volume of **FA2**, $(\mathbf{V_{FA2}})^2$,
- reaction time, ***t*** to nearest second, and
- calculated rate in mol dm^{-3} s^{-1}

Record all calculated values to 3 significant figures.

(a) Method

Experiment 1

1. Use the measuring cylinders to measure the following:
    - 20 cm^3 of **FA1**
    - 20 cm^3 of **FA3**
    - 1 cm^3 of starch indicator
2. Using another measuring cylinder, measure 10 cm^3 of **FA2**.
3. Pour the measured **FA1** and starch into a dry 100 cm^3 beaker.
4. Pour in the **FA3**, followed by **FA2** immediately into the same beaker. Start the stopwatch on adding **FA2**.
5. Stir the mixture and place the beaker on the printed page on page 2 of the insert.
6. The mixture turns brown and then yellow before turning a blue–black colour. Stop the stopwatch when this blue–black colour appears and obscures the wordings.
7. Record the time to the nearest second in the space on page 10.
8. Wash the beaker thoroughly with water and carefully dry the beaker with paper towel.

Experiment 2

9. Repeat step **1** in **Experiment 1**.
10. Using a measuring cylinder, add 2 cm^3 of **FA2.** Make up the volume to 10 cm^3 using deionised water using another measuring cylinder.
11. Pour the measured **FA1**, starch and deionised water into a dry 100 cm^3 beaker.
12. Pour in the **FA3**, followed by **FA2** immediately into the same beaker. Start the stopwatch on adding **FA2**.
13. Stir the mixture and place the beaker on the printed page on page 2 of the insert.
14. Stop the stopwatch when this blue–black colour appears and obscures the wordings.
15. Record the time to the nearest second in the space on page 10. The timing should not exceed 6 minutes.
16. Wash the beaker thoroughly with water and carefully dry the beaker with paper towel.

Experiments 3−5

Carry out three further experiments to investigate the effect of changing the concentration of **FA2** by altering the volume of aqueous potassium iodide, **FA2**, used.

You should use a volume of **FA2** that is at least 2 cm^3 and the total volume of the reaction mixture must always be 51 cm^3.

2 **(a)** With reference to **1(b)**, consider the colour change observed when **FA2** was added to the solution prepared in step **6** of **Experiment 1** before the blue–black colour appeared. Suggest an explanation for the sequence of colour change in terms of the chemistry involved.

Reaction 1: $Fe^{3+}(aq) + 2S_2O_3^{2-}(aq) \rightleftharpoons [Fe(S_2O_3)_2]^-(aq)$

yellow purple

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

Reaction 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

The brown solution is due to a mixture of purple complex $[Fe(S_2O_3)_2]^-$ and yellow $Fe^{3+}(aq)$.

Since the I_2 formed reacted immediately with thiosulfate, $[S_2O_3^{2-}(aq)]$ decreased, and the position of equilibrium in Reaction 1 shifts left.

Thus, concentration of $[Fe(S_2O_3)_2]^-$ decreased. Hence, the colour of the solution gradually fades until the solution became pale yellow due to Fe^{3+}.

[2]

2 **(b)** Show, by means of calculation, that the change in the concentration of I^-, $\Delta[I^-]$, which occurred when the blue–black colour appeared was 2.35×10^{-3} mol dm^{-3}.

Assume the concentration of $Na_2S_2O_3$ to be 0.006 mol dm^{-3}.

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

Reaction 3: $I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

$\text{Amount of } S_2O_3^{2-} = \frac{20}{1000} \times 0.006 = 1.20 \times 10^{-4}$ mol

$\Delta[I^-] = \frac{1.20 \times 10^{-4}}{\frac{51}{1000}} = 2.35 \times 10^{-3}$ mol dm^{-3}

[1]

(c) Results

The rate of the reaction can be calculated as shown.

$$\text{rate} = \frac{\Delta[I^-]}{\text{reaction time}}$$

Calculate the rate of reaction for each experiment and complete your table.

Expt	V_{FA2}/ cm^3	V_{water}/ cm^3	***t*** /s	$(V_{FA2})^2$/ cm^6	Rate / 10^{-5} mol dm^{-3} s^{-1}
1	10.0	0.0	26	100	9.04
2	2.0	8.0	304	4.00	0.773
3	4.0	6.0	128	16.0	1.84
4	6.0	4.0	69	36.0	3.41
5	8.0	2.0	34	64.0	6.91

[4]

2 (d) (i) Plot a graph of rate on the y–axis against $(V_{FA2})^2$ on the x–axis on the grid in **Figure 2.1**. Draw the best fit line taking account of all the plotted data points.

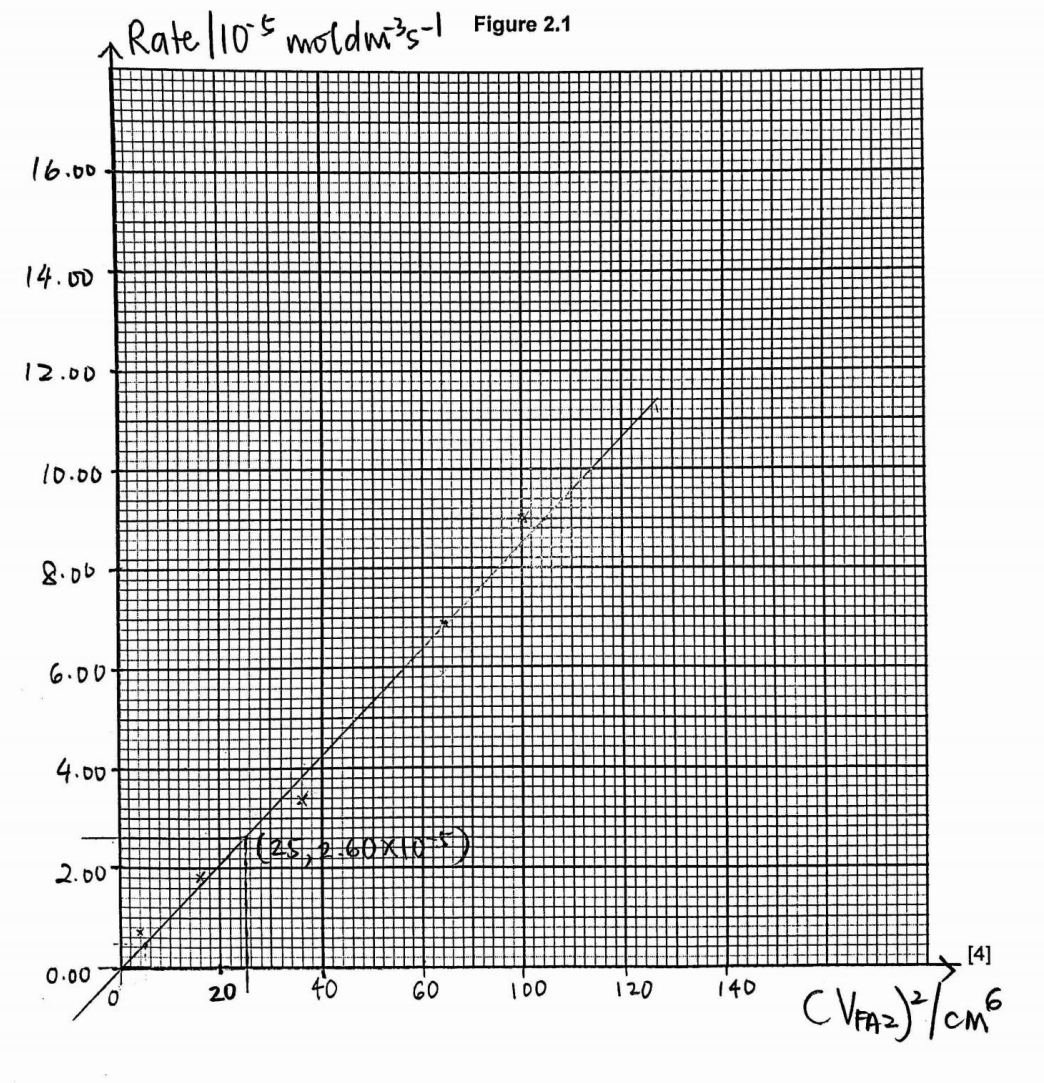


Figure 2.1

[4]

(d) (ii) Explain why the volume of **FA2** in each experiment can be used as the $[I^-]$.

Since total volume of the mixture is kept constant by adding deionised water, volume of **FA2** used is proportional to $[I^-]$. [1]

2 **(d)** **(iii)** Deduce the order of reaction with respect to $[I^-]$. Use evidence from your graph to support your deduction.

Observe from the graph, any one of the following:

- the graph is linear/gradient and pass through origin is a constant, showing that rate $\propto$ squared volume/ concentration of **FA2/** state that it is a graph of rate against $[I^-]^2$.

OR

- the rate increase by xxx for every additional xxx cm^3 of **FA2** used
- when squared volume of **FA2** double, rate double (or other multiples)

✓ Therefore, the reaction is second order with respect to $[I^-]$

[1]

(e) **(i)** Use your graph and the formula in **2(c)** to calculate the time that the reaction would have taken if 5.0 cm^3 of **FA2** had been used. Show your working clearly.

If 5.0 cm^3 of **FA2** is used, V_{FA2}^2 = 25 cm^6, rate = 2.6 x 10^{-5} mol dm^{-3} s^{-1} (read from graph) [1]

Time = $\Delta[I^-]$/ rate = $\frac{2.35x\ 10^{-3}}{2.6\ x\ 10^{-5}}$ = 90.4 s (to 3 sf or nearest seconds).

time = ... [2]

(d) **(ii)** Calculate the initial concentration of iron(III) ions and the initial concentration of iodide ions in the reaction mixture if 5.0 cm^3 of **FA2** had been used.

$[Fe^{3+}] = \frac{0.030\times20.0}{51.0}$ = 0.0118 mol dm^{-3}

$[I^-] = \frac{0.060\times5.0}{51.0}$ = 5.88 x 10^{-3} mol dm^{-3}

initial $[Fe^{3+}]$ = ...

initial $[I^-]$ = ... [1]

2 (d) (iii) Given that the reaction is first order with respect to $[Fe^{3+}]$, calculate the rate constant.

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

Ecf order of reaction from d(iii)

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

$= \frac{2.6 \times 10^{-5}}{(0.00118) \times (5.88 \times 10^{-3})^2}$

$= 637$ [1] $mol^{-2}\ dm^6\ s^{-1}$

rate constant = ... [2]

(f) At the end of Experiment **1**, a student washed the beaker and used it immediately for Experiment **2**. State and explain how the student's action would affect the time ***t*** for Experiment **2**.

Dilution of the reacting mixture will occur due to **residual water** left after washing. Time, **t**, for experiment 2 will be **longer/increase**.

[2]

(g) List 1 possible source of experimental errors and suggest the corresponding improvement to reduce the experimental error.

	experimental errors	improvements
1	The temperature of the set-up may not be constant throughout the experiment. A change of temperature can affect the rate and the time recorded.	Performed the experiment in a water bath with controlled temperature
2	It was difficult to start the stopwatch while reagents was added. The time for the cross to be obscured would thus be inaccurately determined.	Allow the stopwatch to run and add reagents at a convenient time.

[2]

[Total: 22 marks]

3 Qualitative Analysis

You are provided the following reagents and the *Qualitative Analysis Notes* on page 21 and 22.

FA1 is 0.03 mol dm^{-3} of $FeCl_3$ solution

FA4 is 6% volume H_2O_2

FA5 is 0.00215 mol dm^{-3} of $\mathbf{X^-}$ solution

(a) You are to perform the tests described in **Table 3.1** and record your observations in the same table.

Your answers should include

- details of colour changes throughout the tests and precipitates formed,
- the identity of any gases evolved, and details of the test used to identify each one.

You should indicate clearly at what stage in a test a change occurs.
Marks are not given for chemical equations.
No additional or confirmatory tests for ions present should be attempted.

Table 3.1

Test	Observations	Explanation (teacher's notes)
i. To 5 cm^3 of **FA1**, add aqueous silver nitrate in excess. Filter the mixture.	White ppt formed.The solution remained pale yellow/ colourless.	Cl^- is present.
Separate the filtrate into three test tubes for **(ii)**, **(iii)** and **(iv)**.	The filtrate was a pale yellow/ colourless solution. The residue was a white ppt.	The filtrate contains Fe^{3+} ions. The white ppt is $AgCl$.
Place the filter funnel containing the residue on a new test tube, add aqueous ammonia over the residue.	The white ppt was soluble aqueous ammonia to form a pale yellow/colourless solution.	Cl^- is present. $[Ag(NH_3)_2]^+$ is formed. It should be a colourless solution but the ppt is covered with traces of Fe^{3+}, hence the solution appears pale yellow.
	With HNO_3, white ppt formed and	

	Add HNO_3 dropwise to the filtrate until no more change.	white ppt insoluble in excess nitric acid in a colourless solution.	Reversal of equilibrium. HNO_3 neutralise the NH_3 and hence AgCl is formed again.
ii.	To a new test tube containing 1 cm^3 of **FA4**, add a few drops of filtrate from the first test tube from **(i)**. Wait and observe the effervescence. Test for the identity of the gas.	The pale yellow solution turned light brown/orange. Effervescence of colourless, odourless gas observed. Gas relighted a glowing splinter. Oxygen evolved.	**FA4** acts an reducing agent as oxygen is evolved. Hence, hydrogen peroxide FA4 undergoes oxidation.
iii.	To the second test tube containing 1 cm^3 of filtrate from **(i)**, add aqueous sodium hydroxide till excess. Add 1 cm^3 of **FA4**.	Brown/ orange ppt, insoluble in excess aqueous NaOH. The brown ppt insoluble in **FA4**. Vigorous effervescence of colourless, odourless gas is seen. The gas relighted a glowing splinter. Oxygen evolved.	Fe^{3+} is present. $Fe(OH)_3$ cannot be reduced by FA4. However, more O_2 released as H_2O_2 exists as HO_2^- and have a greater reducing ability (gets oxidised more easily by remaining Fe^{3+}).
iv.	To the last test tube containing 1 cm^3 of filtrate from **(i)**, add 1 cm^3 of **FA5**.	(Blood) red solution formed.	$Fe^{3+}(aq)$ undergoes a ligand exchange.
v.	To a new test tube containing a fresh sample of **FA1**, test the pH of the solution with a Universal Indicator paper. Add aqueous ammonia dropwise. Add **FA1** dropwise to dissolve the precipitate formed. Stop adding **FA1** once the precipitate disappears.	The UI paper turned red/brown or. The pH was 1 or 2. Brown/ orange ppt formed on adding NH_3. A yellow/ orange/ light brown solution was formed on adding FA1.	Fe^{3+} solution is acidic because of appreciable hydrolysis. Ammonia is neutralised by the acidic Fe^{3+}. There are no excess Fe^{3+}. Ideally, the pH drops to 7. This is how neutral $FeCl_3$ is made.

[6]

3 (b) With reference to the observation made in test **(i),** identify the ppt and explain its solubility in aqueous ammonia.

The white ppt is AgC*l*.[1] The ppt is soluble in aqueous ammonia because of the formation of complex[1] , $[Ag(NH_3)_2]^+$.

[2]

(c) In the presence of strong base,

$H_2O_2 + OH^- \rightarrow HO_2^- + H_2O$

Electrode Reaction	$E^\ominus$/V
$Fe^{3+} + e \rightleftharpoons Fe^{2+}$	+0.77
$Fe(OH)_3 + e \rightleftharpoons Fe(OH)_2$	–0.56
$O_2 + 2H_2O + 2e \rightleftharpoons H_2O_2$	+0.68
$H_2O_2 + 2H^+ + 2e \rightleftharpoons 2H_2O$	+1.23
$O_2 + H_2O + 2e \rightleftharpoons HO_2^- + OH^-$	–0.08

With reference to the electrode potential given, explain the difference in the observation for effervescence made when **FA4** is added in test **(ii)** and **(iii)**.

The effervescence is more vigorous for step **(iii)** as the H_2O_2 is in alkaline medium and exists as HO_2^-. HO_2^- is a stronger reducing agent as the $E^\ominus(O_2/ HO_2^-)$ is more negative than that of H_2O_2. Hence the $E^\ominus$(cell) is more positive when hydrogen peroxide reacts with Fe^{3+} in alkaline medium.

[2]

(d) Write an equation to explain the colour of the Universal Indicator paper on addition of **FA1** before adding aqueous ammonia.

$[Fe(H_2O)_6]^{3+} \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+} + H^+$

$[Fe(H_2O)_6]^{3+} + H_2O \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+} + H_3O^+$

The presence of H^+ result in a pH of 1 or 2.

[1]

3 (e) The pH of the resultant solution in test **(v)** is 7. Explain the chemistry behind the changes of pH in test **(v)** when aqueous ammonia is added dropwise followed by **FA1**.

When ammonia is added, the reddish brown ppt of $Fe(OH)_3$ was formed, and the pH was >7.

When more FA1 was added, the H^+ from FA1 neutralised the ammonia, hence the pH dropped to 7 and the ppt disappeared.

(neutralised by the H^+ from the **FA1**.

OR

The ppt disappeared as the OH^- in $Fe(OH)_3$ is removed by H^+.

There are no excess Fe^{3+} to make the solution acidic . Hence, the pH drops to 7.) [2]

(f) Predict what happen when the resultant solution in test **(v)** is added to phenol.

A purple/violet complex is formed. [1]

(g) Planning

Thermometry is sometimes used to determine the number of ligands attached to the central metal atom or ion. One of such monodentate ligand is X^- present in **FA5**.

A student performed an experiment to determine the structural formula of the hexa–coordinated complex.

$$[Fe(H_2O)_6]^{3+} + nX^- \rightarrow [Fe(H_2O)_{6-n}X_n]^{3-n} + nH_2O \qquad \Delta H < 0$$

He used varying volumes of **FA1** and **FA5** only, while keeping the total volume of the mixture at 80 cm^3. He recorded the change in temperature.

(i) Outline how you would:

- determine the effect of changing the volumes of **FA1** and **FA5** on the temperature change of the experiment and hence,
- vary the volumes of **FA1** and **FA5** to form the needed hexa–coordinated complex, $[Fe(H_2O)_{6-n}X_n]^{3-n}$.

You are provided with the same solutions, which were used in the test **(iv)** described. **There is no need to perform the experiment.**

No details regarding the use of specific apparatus are required. [3]

For experiment **1**

A 10 cm^3 portion of **FA1** was added to 70 cm^3 of **FA5**. The mixture was mixed thoroughly and the highest temperature reached was noted down.

The procedure was repeated for experiments **2** to **7** (minimum 5), with the following changes in the volume of **FA1** and **FA5** as stated in the table below.

Experiment	Volume of **FA1** / cm^3	Volume of **FA5** / cm^3	ΔT / °C
1	10	70	
2	20	60	
3	30	50	
4	40	40	
5	50	30	
6	60	20	
7	70	10	

(ii) The student plotted a graph and concluded that n=1 in the structural formula of $[Fe(H_2O)_{6-n}X_n]^{3-n}$. On the axes of **Figure 3.1**, sketch a graph that the student would obtain from your experiment in **3(g)(i)** to draw this conclusion. Indicate the volume of **FA1** that the student would use to reach this conclusion on the x-axis of the graph.

3 (h)

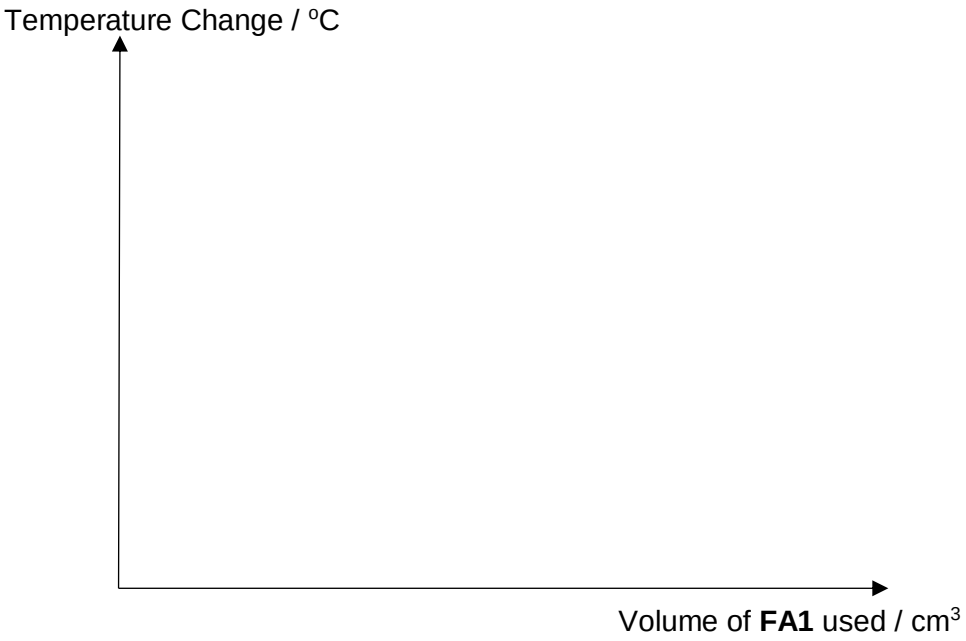
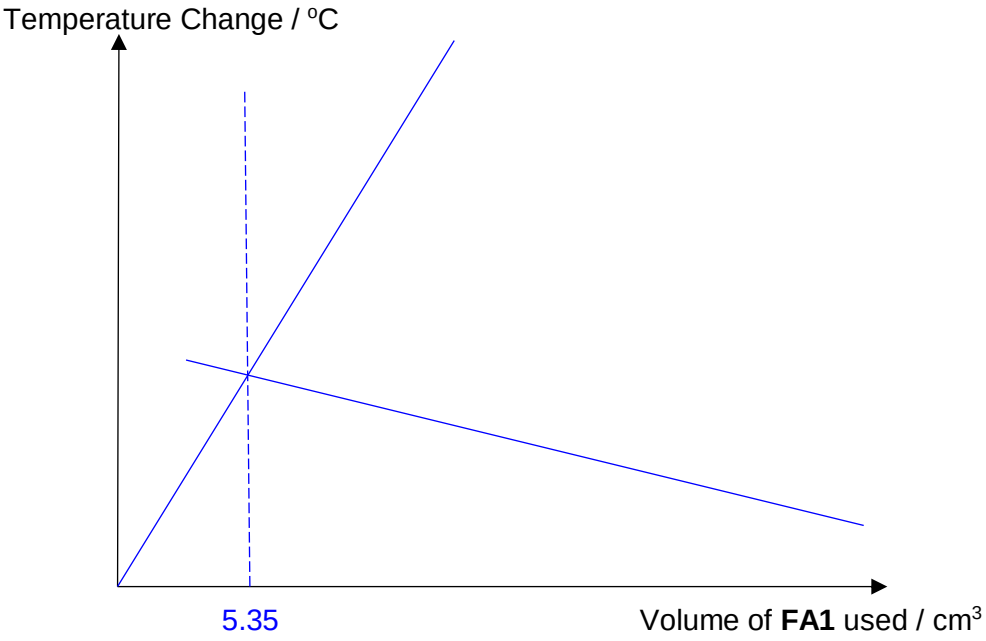


Figure 3.1



Two straight lines drawn

Concentration of **FA1** = 0.03 mol dm^{-3}

Concentration of X^- = 0.00215 mol dm^{-3}

Let the intersection occur at y cm^3 of **FA1**

Amount of **FA1** used $= \frac{y}{1000} \times 0.03 = 3 \times 10^{-5}$ y mol

Amount of X^- used = $\frac{80 - y}{1000} \times 0.00215 = (1.72 \times 10^{-4} - 2.15 \times 10^{-6}$ y) mol

Since ratio of Fe^{3+} : X^{-} is 1 : 1;

$3 \times 10^{-5}\ y = 1.72 \times 10^{-4} - 2.15 \times 10^{-6}\ y$

Intersection occurred at 5.35 cm^3 of FA1

[2]

[Total: 19 marks]