NATIONAL JUNIOR COLLEGE
SH 2 Year – End Practical Examination
Higher 2

CANDIDATE NAME

SUBJECT CLASS

REGISTRATION NUMBER

CHEMISTRY **9729/04**

Paper 4 Practical **Tuesday 15 August 2023**

Candidates answer on the Question paper. **2 hours 30 minutes**

READ THESE INSTRUCTIONS FIRST

Write your identification number and name.
Give details of the practical shift and laboratory where appropriate, in the boxes provided.
Write in blue or black pen.
You may use an HB pencil for any diagrams or 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 19 and 20.

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	/ 17
3	/ 13
4	/ 12
Total	/ 55

This document consists of **20** printed pages and **0** blank page.

1 Determine the reacting mole ratio of reagent X with Fe^{2+}

Iron(II) ions, Fe^{2+}, is readily oxidized.

$$Fe^{2+} \longrightarrow Fe^{3+} + e^-$$

A common oxidising agent for Fe^{2+} ion is manganate(VII) ions in an acidic medium.

$$MnO_4^- + 8H^+ + 5e^- \longrightarrow Mn^{2+} + 4H_2O$$

Another oxidising agent that is also able to oxidise Fe^{2+} to Fe^{3+} is reagent **X**.

In this experiment, a limited amount of **X** is added to a solution of Fe^{2+}. The resultant mixture containing unreacted Fe(II) ions will be titrated with acidified MnO_4^-.

FA 1 is a solution containing 0.075 mol dm^{-3} iron(II) sulfate, $FeSO_4$.

FA 2 is 0.010 mol dm^{-3} potassium manganate(VII), $KMnO_4$.

FA 3 is 0.025 mol dm^{-3} of a reagent **X**.

FA 4 is 1.0 mol dm^{-3} sulfuric acid, H_2SO_4.

(a) Procedure

1. Fill a burette with **FA 3** and another burette with **FA 2**.
2. Pipette 25.0 cm^3 of **FA 1** into a conical flask and add 10 cm^3 of **FA 4** using a 10.0 cm^3 measuring cylinder into the same conical flask.
3. Run 12.00 cm^3 of **FA 3** from the burette into the conical flask.
4. Titrate the mixture against **FA 2** until the first permanent pink colour remains in the solution.

You only need to do the titration **ONCE**. Record your titration results in the space below.

Final burette reading / cm^3	
Initial burette reading / cm^3	
Volume of **FA 2** added / cm^3	**24.50 (shift 1)** **24.20 (shift 2)** **24.00 (shift 3)**

Tables with correct headings and units(Do no award if did not put "burette") (Do not award if any final and initial burette readings are inverted / if 50.00 is used as initial burette reading / burette reading is > 50.00)	1
All burette readings recorded to the nearest 0.05 cm^3 + correct computation of titres	1
Accuracy marks ± 0.5 cm^3 from teacher's reading	1

[3]

(b) **(i)** Calculate the amount of Fe^{2+} ions in **FA 1** that reacted with MnO_4^- in **FA 2.**

Amount of MnO_4^- used = $\frac{24.50}{1000} \times 0.01 = 2.45 \times 10^{-4}$ mol

Amount of Fe^{2+} reacted with MnO_4^- = $2.45 \times 10^{-4} \times \mathbf{5} = 1.23 \times 10^{-3}$

amount of Fe^{2+} ions in **FA 1** reacted with MnO_4^- in **FA 2** =…… 1.23×10^{-3} …… mol [1]

(ii) Calculate the total amount of Fe^{2+} ions in **FA 1** that was pipetted into the conical flask initially.

Hence, calculate the amount of Fe^{2+} ions in **FA 1** that reacted with **X** in **FA 3** added.

Total amount of Fe^{2+} = $\frac{25}{1000} \times 0.075 = 1.875 \times 10^{-3}$
= 1.88×10^{-3} mol

Total amount of Fe^{2+} ions in **FA 1** =…………………… mol

Total amount of Fe^{2+} = Fe^{2+} reacted with **X** + Fe^{2+} reacted with MnO_4^-

Amount of Fe^{2+} reacted with **X** = Total amount of Fe^{2+} − Fe^{2+} reacted with MnO_4^-

= $1.875 \times 10^{-3} - \mathbf{1.23 \times 10^{-3}}$

= 6.45×10^{-4}

amount of Fe^{2+} ions in **FA 1** reacted with **X** in **FA 3** =…… 6.45×10^{-4} …… mol [2]

(iii) Calculate the amount of **X** in the 12.00 cm^3 of **FA 3** added.

Hence, calculate the number of moles of Fe^{2+} ions that react with 1 mol of **X.**

Amount of **X** used = $\frac{12}{1000} \times 0.025 = 3.00 \times 10^{-4}$ mol

amount of **X** =…… 3.00×10^{-4} …… mol

$$\frac{\text{amount of } Fe^{2+}}{\text{amount of } \mathbf{X}} = \frac{\mathbf{6.45 \times 10^{-4}}}{3.00 \times 10^{-4}} = 2.15$$

number of moles of Fe^{2+} ions that react with 1 mol of **X** =…… 2.15 ……[2]

(iv) The redox half equations for the possible identity of **X** are given below.

$$VO^{2+} + 2H^+ + e^- \rightleftharpoons V^{3+} + H_2O$$

$$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$$

$$NO_3^- + 4H^+ + 3e^- \rightleftharpoons NO + 2H_2O$$

$$IO_3^- + 5H^+ + 4e^- \rightleftharpoons HIO + 2H_2O$$

$$NO_3^- + 10H^+ + 8e^- \rightleftharpoons NH_4^+ + 3H_2O$$

Using your answer from **(b)(iii)**, deduce with reasoning, a possible identity for **X**. [You may use number of moles of Fe^{2+} ions that react with 1 mol of **X** = 8.12 for this question if you did not get an answer for **(b)(iii)**. Note that this value is **not** the correct answer for **(b)(iii)**.]

H_2O_2

Fe^{2+} is oxidized to form Fe^{3+} and thus **X must be reduced.**

1 mol of Fe^{2+} gives out 1 mol of electrons to form 1 mol of Fe^{3+}.

The **reacting mole ratio for Fe^{2+} to X is approximately 2:1, indicating that every mol of X gains 2 mol of electrons**.

If using the value given (i.e. 8.12), NO_3^-

Fe^{2+} is oxidized to form Fe^{3+} and thus **X must be reduced.**

1 mol of Fe^{2+} gives out 1 mol of electrons to form 1 mol of Fe^{3+}.

The **reacting mole ratio for Fe^{2+} to X is 8.12 :1, indicating that every mol of X gains 8 mol of electrons**.

....

....

....

....

....[3]

(v) Use the information below to calculate the percentage error for each volume measurement in Table 1.1.

A 25 cm^3 measuring cylinder is graduated to 1 cm^3.

The maximum error for a 25 cm^3 pipette is ±0.06 cm^3.

Table 1.1

Solution	Apparatus used	Volume measured /cm^3	% Error
FA 1	Pipette	25.0	$\frac{\pm 0.06}{25} \times 100\%$ = ±0.240%
FA 2	Burette	23.50	$\frac{2 \times \pm 0.05}{23.50} \times 100\%$ = ±0.426%
FA 4	25 cm^3 Measuring cylinder	10.0	$\frac{\pm 0.5}{10} \times 100\%$ = ±5.00%

[2]

- Correctly calculate all 3 values	Award 1
- must have ± signs and - must have unit of % - all calculated values to 3 sf	Award 1

For burette and measuring cylinder, the uncertainty from measure is half of the smallest division.
Uncertainty for measuring cylinder = ±0.5 cm^3
Uncertainty for burette = ±0.05 cm^3
To measure a volume using a burette, two readings are required, hence total uncertainty for measurement = 2 × (±0.05 cm^3)

[Total: 13]

2 Determine the order of reaction with respect to iodide

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

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

The iodine, I_2, produced can be reacted immediately with thiosulfate ions, $S_2O_3^{2-}$.

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

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

FA 5 is 0.0500 mol dm^{-3} potassium iodide, KI.

FA 6 is 0.0500 mol dm^{-3} acidified iron(III) chloride, $FeCl_3$.

FA 7 is 0.0100 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$.

FA 8 is starch.

(a) Procedure

Experiment 1

1. Using a 25.0 cm^3 measuring cylinder, transfer into a 100 cm^3 beaker
   - 20.0 cm^3 of **FA 5**
   - 10.0 cm^3 of **FA 7**
   - 10.0 cm^3 of **FA 8**
2. Swirl the beaker to mix the solutions thoroughly. Place the beaker on a white tile.
3. Using a 10.0 cm^3 measuring cylinder, transfer 10.0 cm^3 of **FA 6** into the same 100 cm^3 beaker and start timing immediately.
4. Swirl the beaker once and place the beaker on the white tile.
5. Stop timing as soon as the solution turns blue-black.
6. Record this reaction time, *t*, to the nearest second, in Table 2.1.
7. Wash out the beaker and dry it with a paper towel.

Experiment 2

1. Using a 25.0 cm^3 measuring cylinder, transfer into a 100 cm^3 beaker
   - 10.0 cm^3 of **FA 5**
   - 10.0 cm^3 of deionised water
   - 10.0 cm^3 of **FA 7**
   - 10.0 cm^3 of **FA 8**

2. Swirl the beaker to mix the solutions thoroughly. Place the beaker on a white tile.

3. Using a 10.0 cm^3 measuring cylinder, transfer 10.0 cm^3 of **FA 6** into the same 100 cm^3 beaker and start timing immediately.

4. Swirl the beaker once and place the beaker on the white tile.

5. Stop timing as soon as the solution turns blue-black.

6. Record this reaction time, *t*, to the nearest second, in Table 2.1.

7. Wash out the beaker and dry it with a paper towel.

Experiments 3–5

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

2. The combined volume of **FA 5** and distilled water must always be 20.0 cm^3.

3. Do not use a volume of **FA 5** that is less than 5.0 cm^3.

4. Record all your values of *t* and volume of **FA 5, FA 6**, **FA 7**, **FA 8** and deionised water used in Table 2.1.

(b) Results

The rate of reaction can be calculated as rate = $\frac{1000}{\text{reaction time}}$

Table 2.1

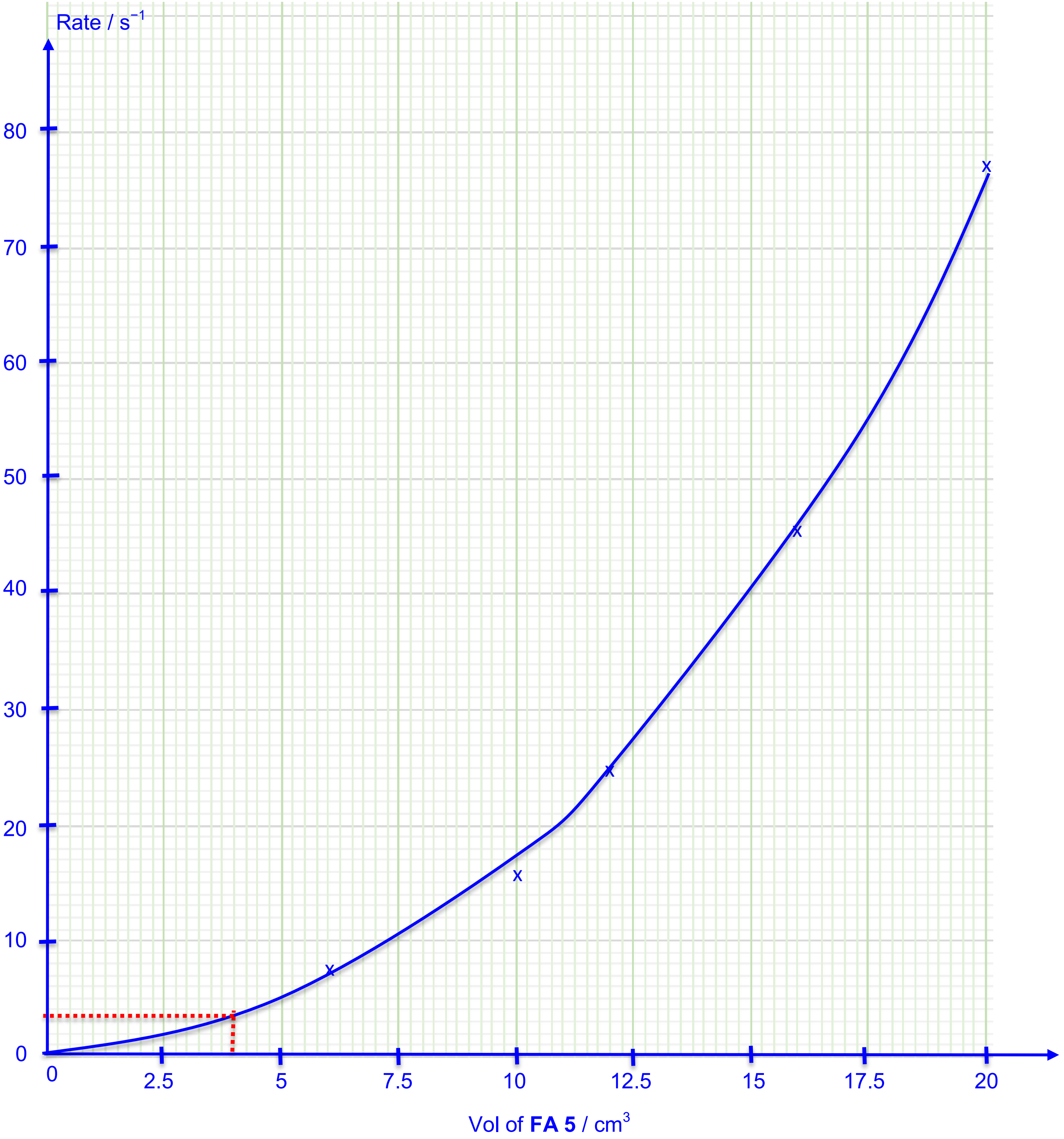
experiment	vol. of **FA 6** / cm^3	vol. of **FA7** / cm^3	vol. of **FA 8** / cm^3	vol. of **FA 5** / cm^3	vol. of deionised water / cm^3	total volume / cm^3	time, *t* /s	rate / s^{-1}
1	10.0	10.0	10.0	20.0	0.0	50.0	13	76.9
2	10.0	10.0	10.0	10.0	10.0	50.0	65	15.4
3	10.0	10.0	10.0	16.0	4.0	50.0	22	45.5
4	10.0	10.0	10.0	12.0	8.0	50.0	24	24.4
5	10.0	10.0	10.0	6.0	14.0	50.0	129	7.75

[3]

All volume readings recorded to the nearest 0.1 cm^3 for **all volumes** **All** time readings recorded to the nearest 1 s **All** Rate values to 3 sf	1
3 additional experiments, • volume of **FA 5** must be > 5 cm^3 and < 20 cm^3 • the difference between 2 values cannot be less than 2 cm^3 Volume of water chosen to ensure so that **FA5** + water = 20.0 cm^3	1
Correctly calculate **all** rate values	1

(i) On the grid below, plot a graph of rate of reaction against volume of **FA 5**.

Include the origin, (0,0), in your scales and draw a line of best fit.



Axis labelled **rate** in **s^{-1}** and **vol** of FA 5 in **cm^3** Linear scales chosen to use more than half of each axis Best fit line passes through point of origin	1
All 5 points accurately plotted (within 1⁄2 small square and in the correct square).	1

[2]

(ii) **Use your graph** to calculate the time that the reaction would have taken if 4.00 cm^3

of **FA 5** had been used. **Show on the graph** how you obtained your answer.

From graph

rate = 3.5 s^{-1} **[show construction lines on the graph]**

time = $\frac{1000}{3.5}$

= 286 s

time = 286 s [2]

(iii) Suggest why rate is inversely proportional to time as shown in the relationship.

$$\text{rate} = \frac{1000}{\text{reaction time}}$$

The same amount of iodine /product is produced before the colour change is observed./ **same amount of thiosulfate added** into each expt thus the amount of iodine to be produced before colour change is observed is the same.

The longer it takes for the colour to appear the **slower the rate.** [1]

(c) **(i)** Calculate the concentration of KI in the reaction mixture for **Experiments 1** and **2.** Hence, state the relationship between volume of KI used and its concentration in the reaction mixture.

Expt 1: 20 cm^3 of KI in total volume 50 cm^3

$[KI]_{mixture} = \frac{20}{50} \times 0.05 = 2.00 \times 10^{-2}$ mol dm^{-3}

Expt 2:
10 cm^3 of KI in total volume 50 cm^3

$[KI]_{mixture} = \frac{10}{50} \times 0.05 = 1.00 \times 10^{-2}$ mol dm^{-3}

Concentration of KI in the reaction mixture is directly **proportional** to its volume used.

[2]

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

As concentration / **volume of iodide ions increases,** the **rate increases with an increasing gradient**

[2]

Concentration of iodide doubled and rate quadrupled

Concentration of iodide doubled and rate quadruple with stating values.

Remark not directly proportional as line is a curve / not a straight line Straight line: rate is proportional to concentration (of iodide ions) / proportional as line has a positive gradient Reject directly proportional unless the straight line passes within 5 small squares of (0,0).	1
Reason: plotted points **should** give line of increasing gradient (This mark is not available if a straight line was drawn.)	1

(d) **(i)** A student carries out another experiment, **Experiment 6**, to investigate the effect of iron(III) concentration on the rate of this reaction.

The rate is compared to that of **Experiment 2**. Suggest the volumes that the student could use for **Experiment 6** in Table 2.2 provided below.

Table 2.2

experiment	vol. of **FA 6** / cm^3	vol. of **FA7** / cm^3	vol. of **FA 8** / cm^3	vol. of **FA 5** / cm^3	vol. of deionised water / cm^3	total volume / cm^3
2	10.0	10.0	10.0	10.0	10.0	50.0
6	**5.0**	10.0	10.0	10.0	**15.0**	50.0

[1]

volume of FA 5, FA 7 and FA 8 same as Experiment 2

Total volume of FA 6 + water = 20 cm^3

(ii) This student records a time of 215 s for **Experiment 2**.

The rate of the reaction is directly proportional to the concentration of iron(III) ions.

Suggest how long it would take the reaction mixture proposed for **Experiment 6** in **(d)(i)** to turn blue-black. Assume that **Experiment 6** is carried out at the same temperature as **Experiment 2**.

Do not carry out Experiment 6.

FA 6 + water = 20 cm^3

Constant total vol, thus $[Fe^{3+}]$ is directly proportional to vol of FA 6 used.

Comparing expt 2 and 6, with concentration FA 5, FA 7 and FA 8 kept constant.

When $[Fe^{3+}]$ ÷ 2, rate should also ÷ 2.

rate 6 = ½ × rate 2

$$\frac{1000}{t_6} = \frac{1}{2} \times \frac{1000}{215}$$

t_6 = 430 s

Alternatively,

$$\frac{Rate\ 2}{Rate\ 6} = \frac{\frac{1000}{215}}{\frac{1000}{t}} = \frac{[Fe^{3+}][I^-]^2}{[Fe^{3+}][I^-]^2} = \frac{(10)(10)^2}{(5)(10)^2}$$

t = 430s

to be awarded the [1], some explanation is needed to how the proportionality relationship

time = 430 s [1]

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

(i) Write an equation for the reaction between thiosulfate ions and hydrogen ions in aqueous solution.

$S_2O_3^{2-} + 2H^+ \rightarrow S + SO_2 + H_2O$

....[1]

(ii) A student carries out the same investigation as **Experiment 2** in **(d)(ii)** but the solutions are mixed in a different order.

The student places **FA 5** and distilled water in one beaker and all the other reactants in a second beaker. The student then transfers the mixture **after 5 minutes** from the second beaker to the first and starts timing.

Suggest if the time taken for this experiment would be greater than, less than or no change from 215 s.

Time taken will be **less than 215s**

Some of the **$S_2O_3^{2-}$ in FA 7 would have reacted with the acid present in the acidified FA 6 in the 2nd beaker**. There will be **less $S_2O_3^{2-}$ present** when the 2 solutions are mixed. With less $S_2O_3^{2-}$(aq) there will be **less iodine that needs to be formed** before the blue black colour appears.

[2]

[Total: 17]

3. Identification of ions in inorganic compounds

FA 9 and FA 10 are aqueous solutions each containing a cation and chloride.

FA 11 is an aqueous solution containing a cation and an anion.

FA 12 is a solution of potassium iodide, KI.

FA 13 is a solution of sodium thiosulfate, $Na_2S_2O_3$.

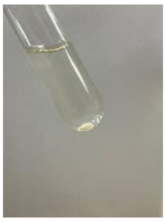
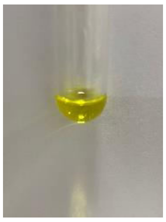
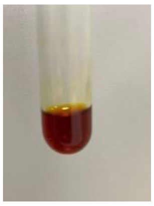
(a) Carry out the following tests. Carefully record your observations in Table 3.1.

Unless stated otherwise, 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**.

(i) **Table 3.1**

tests		observations	
		FA 9	**FA 10**
1	To a 1 cm depth of the respective **FA** in a test tube, add an equal volume **FA 12**.	**White ppt** formed in a **brown solution**. ***don't accept grey ppt**	Yellow solution turn **red/ orange / red brown / reddish brown / brown solution** formed. ***don't accept yellow**
	To the resultant mixture, add excess **FA 13**. 	**Brown solution turns colourless.** White ppt present. (Note: white ppt dissolves in excess FA13)  	**Red brown solution decolourised** or turns pale green. (Note: after some time, pale yellow ppt of sulfur could be observed)  
2	To a 1 cm depth of the respective **FA** in a test tube, add aqueous ammonia, slowly with shaking, until no further change is observed.	**Blue ppt formed**, **soluble** in **excess** aqueous ammonia **to give a dark blue solution**.   	**Red-brown ppt** formed, **insoluble in excess** aqueous ammonia.  

[3]

(ii) Identify the cation in **FA 9** and **FA 10** and state the evidence for each cation by completing Table 3.2.

Table 3.2

	cation	evidence
FA 9	Cu^{2+}	Test 2 : **Blue ppt** $Cu(OH)_2$ **formed** **soluble in excess aqueous ammonia to give a dark blue** $[Cu(NH_3)_4]^{2+}$ **solution** **OR** Test 1 **: Brown solution,** I_2 , **formed** in redox reaction between I^- and Cu^{2+} **and a white ppt of** CuI **is formed** $^{+}$
FA 10	Fe^{3+}	Test 2 : **red brown ppt** $Fe(OH)_3$ **formed** **insoluble in excess aqueous ammonia**

[2]

(iii) When a spatula of magnesium powder is added to a 1 cm depth of **FA 10** in a test tube, effervescence was observed.

Write equations to explain the observation.

$[Fe(H_2O)_6]^{3+} \rightleftharpoons [Fe(H_2O)_5(OH)]^{2+} + H^+$ **must be reversible arrow**

$Mg(s) + 2H^+(aq) \longrightarrow H_2(g) + Mg^{2+}(aq)$

……………………………………………………………………………………………………

……………………………………………………………………………………………………[1]

(b) **FA 11** is an aqueous solution containing one anion and one cation. Only one of these ions contains nitrogen. **FA 11** does not contain sulfite, SO_3^{2-} ion.

(i) Carry out the following tests to identify the **cation** in **FA 11** and record the observations in Table 3.3.

Table 3.3

tests		observations
1	To 1 cm depth of **FA 11** in a boiling tube, add aqueous sodium hydroxide slowly with shaking, until no further change is observed. Warm the resultant solution.	**No ppt formed with aqueous sodium hydroxide** **Upon warming, gas evolved turns moist red litmus paper blue**
2	To 1 cm depth of **FA 11** in a test tube, add aqueous ammonia slowly with shaking, until no further change is observed.	**No ppt formed with aqueous ammonia**

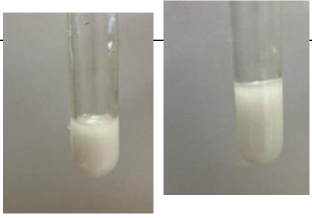
The cation in **FA 11** NH_4^+ . [2]

(ii) Carry out further tests to identify the anion in **FA 11,** using only any bench reagent provided, and hence write the formula of **FA 11**.

Record, in a table in the space below, the reagents, conditions and observations for the tests.

You **must** use a boiling tube if any liquid is heated.

Test	Observations
To a 1 cm depth of **FA 11** in a test tube, add dilute nitric acid.	**No effervescence.**
To a 1 cm depth of **FA 11** in a test tube, add a few drops of aqueous silver nitrate Add aqueous ammonia to the resultant solution.	**No ppt formed.** **No ppt formed.**
To a 1 cm depth of **FA 11** in a test tube, add equal volume of barium nitrate. Add dilute nitric acid until no further change to the resultant mixture	**White ppt formed** **White ppt insoluble in dilute nitric acid**



The formula of **FA 11** $(NH_4)_2SO_4$ [5]

[Total: 13]

4. Planning

A more reactive metal will displace a less reactive metal from a solution of its salt. This reaction is exothermic. If the same reaction is set up in an electrochemical cell, instead of an enthalpy change, electrical energy is produced, and a cell voltage can be measured. This relationship can be shown in the equation below.

$$\Delta H - T\Delta S = -nFE_{cell}$$

You are to plan an investigation of the reaction of three different metals (zinc, magnesium and iron) with aqueous copper(II) sulfate. You will plan to investigate whether there is a relationship between their cell potential values, and their enthalpy changes of reaction, ΔH_r.

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

$$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$$

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

(a) The first part of the investigation is to determine the enthalpy change, ΔH_r, for the reaction of zinc with 1.00 mol dm^{-3} copper(II) sulfate.

You are provided with a sample of powdered zinc and aqueous copper(II) sulfate.

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s) \qquad \Delta H_r$$

The maximum temperature change occurring during this reaction can be determined graphically. The maximum temperature change, ΔT_{max}, obtained from the graph can be used to calculate the enthalpy change of reaction, ΔH_r.

In this question, you are to plan a procedure that would provide sufficient data to allow you to determine a reliable value for the enthalpy change of reaction, ΔH_r.

(i) Calculate the minimum mass of zinc that needs to be added to 100 cm^3 of 1.00 mol dm^{-1} copper(II) sulfate to ensure that zinc is in excess.

[A_r : Zn, 65.4 ; Cu, 63.5 ; S, 32.1 ; O, 16.0]

If 100 cm^3 of $CuSO_4$ is used, amount of $CuSO_4$ = $\frac{100}{1000} \times 1.00 = 0.100$ mol

= Amount of Zn required.

Minimum mass of Zn required = 0.100 × 65.4 = 6.54 g

[1]

(ii) Plan an investigation to determine the maximum temperature change, ΔT_{max}, graphically for the reaction between aqueous copper(II) sulfate and zinc powder.

Measurements should be taken

- before reaction starts
- during the reaction
- for some time after the reaction is complete

You may assume that you are provided with

- 100 cm^3 of 1.00 $mol dm^{-3}$ copper(II) sulfate
- 10 g powdered zinc
- The equipment normally found in a school or college laboratory

In your plan you should include brief details of

- the apparatus you would use
- the quantities you would use
- the procedures you would follow
- the measurements you would make to allow a suitable temperature-time graph to be drawn
- how you would ensure that an **accurate** value for ΔT_{max} is obtained.

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………[5]

Procedures:

1 **Using a weighing balance**, record the mass of a **clean and dry weighing bottle**.

2 Weigh accurately about **7.0 g of Zn (or more)** into the weighing bottle. Record the mass in **Table 1**. (Zn is in excess)

3 **Use 100.0 cm^3 measuring cylinder** to transfer **100.0 cm^3** of 1.00 mol dm^{-3} $CuSO_4$ into a **polystyrene cup** with lid, supported in a 250 cm^3 beaker. *cannot use burette for 100 cm^3.

4 **Start the stopwatch,** gently stir the solution in the cup with a **0.2 °C division thermometer** and measure and **record the temperature of $CuSO_4$ solution every 0.5 min for 2.5 min** in **Table 2**.

5 At **exactly 3.0 min**, add Zn solid into the polystyrene cup. Cover the cup with a lid to minimize heat exchange with environment. Continue to stir the mixture, **measure** and **record the temperature at 3.5 min**, then **every 0.5 min after mixing** until the temperature reaches a maximum.

6 **Continue to measure temperature every 0.5 min for a further 2 - 3 min**. Record in data in **Table 2**.

7 Weigh and record the **mass of the weighing bottle containing residual Zn powder**.

8 Plot a temperature vs time graph using the data collected in **Table 2**.

9 Draw a best-fit straight line **for the points before t = 3.0 min**

10 Draw another best-fit straight line taking into account of all the points **after the temperature of the mixture has started to fall steadily**.

11 **Extrapolate both lines to t = 3.0 min**. From the graph, read the temperatures T_{max} and T_{min} at t = 3.0 min and calculate the temperature change, ΔT_{max}.

Table 1

Mass of weighing bottle / g	
Mass of Zn and weighing bottle / g	
Mass of residual Zn and weighing bottle / g	
Mass of Zn added / g	

Table 2

Time/ min	**Temp/°C**
0.00	
0.50	
1.00	
1.50	
2.00	
2.50	

Time/ min	**Temp/°C**
3.00	No reading
3.50	
4.00	
4.50	
5.00	

Correct choice of apparatus - Clean, **dry** weighing bottle - 100 cm^3 measuring cylinder - styrofoam cup - 0.2 °C division thermometer	1
Record the mass of empty weighing bottle Record mass of weighing bottle + solid Record the mass of weighing bottle after experiment *can infer if table drawn	1
Measure the initial temperature of the solution Measure the temperature of the solution for every 0.5 min time interval Measure beyond the final temperature of the solution an additional 4-6 more reading (about 2-3 more mins) *can infer if table drawn	1
Logical flow of the procedure	1
Indicate in the procedure graph to be drawn is a temperature against time graph Indicate in the procedure, the need to do extrapolation to **estimate the maximum, ΔT_{max}** and minimum temperatures of the solution at the time of mixing.	1

(iii) Sketch, on Fig. 4.1, the graph you would expect to obtain using the measurements you planned to make in **4(a)(ii).**

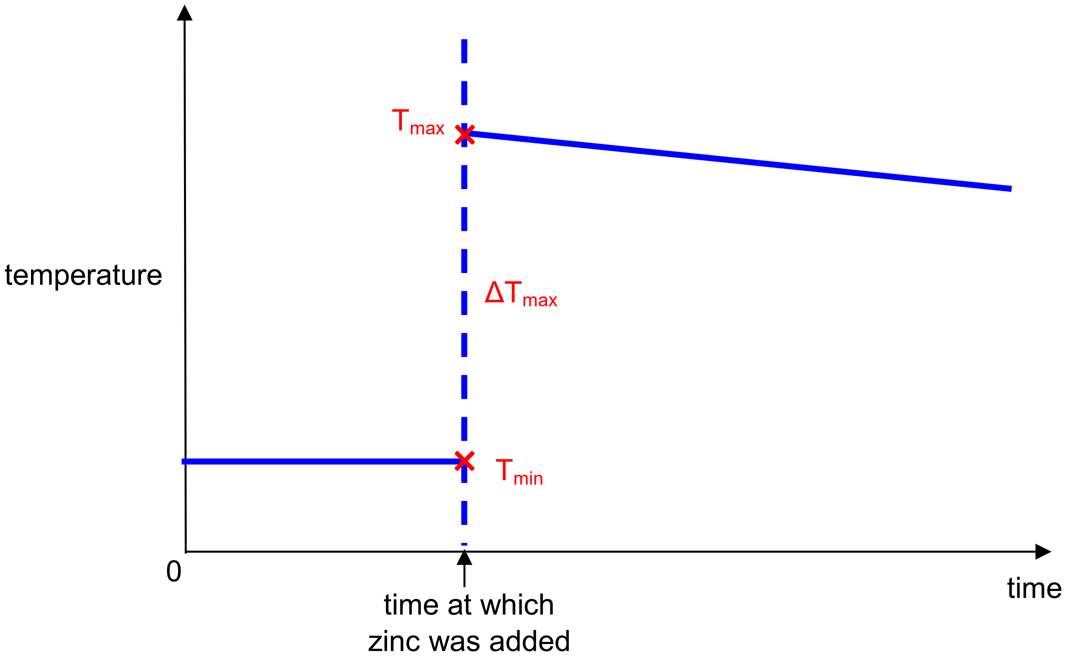

Fig 4.1 [1]

(b) In one experiment, the increase in temperature when the excess zinc powder is added to 50.0 cm^3 of 0.500 mol dm^{-3} aqueous copper(II) sulfate is 25.9 °C.

Calculate the enthalpy change for this reaction, ΔH_r, in kJ mol^{-1}. Include a sign in your answer.

$$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$$

Assume the specific heat capacity, *c*, of the reaction mixture is 4.18 J g^{-1} K^{-1}.
Assume 1.0 cm^3 of 0.500 mol dm^{-3} aqueous copper(II) sulfate has a mass of 1.0 g.

No of mol of Cu^{2+} = $\frac{50}{1000} \times 0.5 = 0.025$ mol (L.R.) zinc is in excess

$$\Delta H_r = -\frac{mc\Delta T}{\text{amount of limiting reagent}}$$

$$= -\frac{(50)(4.18)(+25.9)}{0.025}$$

$$= -216500 \text{ J mol}^{-1}$$

$$\approx -217 \text{ kJ mol}^{-1}$$

ΔH_r = ...-217.... kJ mol^{-1} [2]

(c) Predict how ΔH_r may change as E_{cell} increases. Give a reason for your prediction.

$\Delta G = \Delta H - T\Delta S = -nFE_{cell}$

When **E_{cell} increases, ΔG becomes more negative**, ΔS is similar as no change in states, thus =**ΔH would be more negative**. [1]

OR

ΔH_r become more negative as increases E_{cell}.

Larger E_{cell} value suggests that more electrical energy is converted from heat energy released from the reaction.

(d) Explain why the enthalpy change determination and cell potential determination should not be carried out at different temperature from each other.

$\Delta H - T\Delta S = -nFE_{cell}$

So that ΔH_r and E_{cell} can be compared, the **states of the products and reactants may be different than at standard state**, thus the values of **ΔH_r will be different from the standard ΔH_r**. [1]

or

When the temperature is different, the **E° of each half equation will be different** thus the E_{cell} calculated will be different.

(e) E_{cell} values are shown for the cell reactions.

Use your answer to **(b),** prediction in **(c)** and data from the table below to predict ΔH_r values for reactions 2 and 3.

Complete the table with these values.

	cell reaction	E_{cell} / V	ΔH_r / kJ mol^{-1}
1	$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$	+1.10	−217 (ecf from (b)
2	$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$	+2.72	−489 (any value that is more exo than −217)
3	$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$	+0.78	−150 (any value that is less exo than −217)

*Cannot accept positive values[1]

[Total: 12]

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 → (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 aq. 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