VICTORIA JUNIOR COLLEGE
JC 2 PRELIMINARY EXAMINATION
Higher 2

CANDIDATE NAME ……………………………………………………………..

CT GROUP ……………………………..

CHEMISTRY 9729/04

Paper 4 Practical **26 August 2024**

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

Additional Materials: As listed in the instructions below

READ THESE INSTRUCTIONS FIRST

Write your name and CT group on all the work you hand in.
Write in dark blue or black pen.
You may use a 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**.

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	**/ 12**
2	**/ 11**
3	**/ 19**
4	**/ 13**
Total	**/ 55**

This document consists of **18** printed pages and **2** blank pages.

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

1 Determination of the percentage by mass of copper in an alloy

Copper forms compounds containing Cu^{2+} and Cu^{+} ions. Those compounds containing Cu^{2+} ions tend to be relatively stable.

The addition of an excess of potassium iodide, KI, to a solution of Cu^{2+} ions produces iodine, I_2, and a stable precipitate of CuI as shown in equation 1. For a titration to be accurate, it is necessary that all the Cu^{2+} ions are reduced to Cu^{+} ions. The I_2 turns the solution brown.

equation 1 $2Cu^{2+} + 4I^{-} \rightarrow 2CuI + I_2$

I_2 has a relatively low solubility in water. However, the presence of an excess of I^{-} ions in the reaction mixture allows the soluble tri–iodide ion, I_3^{-}, to form as shown in equation 2. This ensures that the I_2 formed as shown in equation 1 is fully dissolved.

equation 2 $I_2 + I^{-} \rightarrow I_3^{-}$

The I_3^{-} ions formed may be titrated against a standard solution of $Na_2S_2O_3$ as shown in equation 3.

equation 3 $I_3^{-} + 2S_2O_3^{2-} \rightarrow S_4O_6^{2-} + 3I^{-}$

In **1(a)**, you will perform titrations to determine the percentage by mass of copper present in an alloy.

FA 2 is an aqueous solution containing 11.54 g dm^{-3} of the alloy in acid.

You are also provided with
- **FA 1**, 0.100 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$,
- 0.600 mol dm^{-3} aqueous potassium iodide, KI,
- aqueous sodium carbonate, Na_2CO_3,
- aqueous ethanoic acid, CH_3CO_2H,
- starch indicator.

The **FA 2** solution has been prepared using 11.54 g of the alloy containing copper and one other metal, in 1 dm^3 of the solution. **FA 2** contains residual traces of acid.
The presence of acid in the titration mixture will affect the accuracy of the results. The procedure described is designed to reduce these errors.
In this experiment, you will determine the percentage by mass of copper in the alloy used to prepare **FA 2**. You will titrate **FA 2** against **FA 1**.

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

1. Fill a burette with **FA 1**. This will be used for both **Question 1** and **Question 2**.
2. Use a pipette to transfer 25.0 cm^3 of **FA 2** into a 250 cm^3 conical flask.
3. Use a teat pipette to add $Na_2CO_3(aq)$, slowly, with shaking, to **FA 2** in the conical flask, until a slight permanent precipitate forms.
4. Use another teat pipette to add $CH_3CO_2H(aq)$ slowly, with shaking, until this precipitate **just** dissolves.
5. Use a measuring cylinder to add about 20 cm^3 of KI(aq) to this flask. A white precipitate forms in a brown solution.
6. Run **FA 1** from the burette into this flask. Near the end–point, when the brown solution becomes pale, add about 1 cm^3 of starch indicator.
7. Continue adding **FA 1** slowly. The end–point is reached when the **solution** first becomes colourless. The white precipitate remains.
8. Record your titration results, to an appropriate level of precision, in the space provided.
9. Repeat points 2 to 8 until consistent results are obtained.

Titration results

Final burette reading / cm^3	**25.20**	**25.10**
Initial burette reading / cm^3	**0.00**	**0.00**
Volume of FA 1 used / cm^3	**25.20**	**25.10**

- **Correct headings and units for all recordings**
- **Burette readings recorded to 0.05 cm^3 precision**
- **At least 2 consistent readings within ±0.10 cm^3**

[3]

(ii) From your titrations, obtain a suitable volume of **FA 1**, $V_{FA\ 1}$, to be used in your calculations. Show clearly how you obtained this volume.

- $V_{FA1} = \dfrac{25.20+25.10}{2}$ **(only consistent results should be used.)**

  $= 25.15\ cm^3$ **(2 d.p.; working is required)**

2m for $V_{FA\ 1}$ within ±0.20 cm^3 of supervisor's results
1m for $V_{FA\ 1}$ within ±0.50 cm^3 of supervisor's results

$V_{FA\ 1}$ = …**25.15 cm^3**…

[3]

(b) Calculate the mass of copper in the 11.54 g of the alloy used to prepare 1.00 dm^3 of **FA 2**, using your answer in **1(a)(ii).**

Hence, determine the percentage by mass of copper in the alloy.
[A_r: Cu, 63.5]

Amount of $S_2O_3^{2-} = \dfrac{25.15}{1000} \times 0.100 = 2.52 \times 10^{-3}$ mol

Amount of I_3^- = amount of I_2 = $(2.52 \times 10^{-3}) \div 2 = 1.26 \times 10^{-3}$ mol
Amount of Cu^{2+} in 25.0 cm^3 of FA 2 = $2 \times 1.26 \times 10^{-3} = 2.52 \times 10^{-3}$ mol
Amount of Cu^{2+} in 1.00 dm^3 of FA 2 = $2.52 \times 10^{-3} \div \frac{25.0}{1000}$ = 0.101 mol

- **Mass of Cu^{2+} = 0.101 × 63.5 = 6.39 g (3 s.f.)**

mass of Cu in 11.54 g of the alloy = ...**6.39 g**..........

- **Percentage by mass of Cu in the alloy = $\frac{6.39}{11.54} \times 100$ %**
  = 55.4% (3 s.f.)

percentage by mass of copper in the alloy = ...**55.4%**..........
[2]

(c) A large excess of potassium iodide is used in each titration.

(i) Calculate the amount of iodide ions added in each titration.

- **Amount of iodide used in each titration = $\frac{20}{1000} \times 0.600$**
  = 0.0120 mol

amount of iodide ions = ...**0.0120 mol**..........
[1]

(ii) Hence, determine the ratio of the amount of iodide ions added in each titration to the minimum amount of iodide ions required to reduce the copper(II) ions used.

Amount of iodide needed to reduce Cu^{2+} = amount of Cu^{2+} × 2
= $2.52 \times 10^{-3} \times 2$
= 5.03×10^{-3} mol

- **Ratio = 0.0120 ÷ 0.00503**
  = 2.39 (3 s.f.) (e.c.f)

ratio = ...**2.39**..........
[1]

(iii) Identify **two** different chemical processes that use iodide ions in the experiment.

- **Redox where Cu^{2+} is reduced to Cu^+ while I^- is oxidised to I_2 in equation 1.**
- **Precipitation where a CuI ppt is formed between Cu^+ and I^- ions in equation 1.**
- **A soluble complex ion is formed between I_2 and I^- in equation 2.**

[Any 2 of the 3 points]

[2]
[Total: 12]

2 Investigation of rate between peroxydisulfate ions and iodide ions

When oxidising agent peroxydisulfate ions, $S_2O_8^{2-}(aq)$, reacts with aqueous iodide ions, $I^-(aq)$, a brown solution $I_2(aq)$ is formed as shown in equation 4.

equation 4 $S_2O_8^{2-}(aq) + 2I^-(aq) \rightarrow I_2(aq) + 2SO_4^{2-}$

The iodine, $I_2(aq)$, produced can be reacted immediately with thiosulfate ions, $S_2O_3^{2-}(aq)$, as shown in equation 5.

equation 5 $2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$

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.

You will investigate the rate of reaction between peroxydisulfate ions and iodide ions.

You are provided with

- **FA 1**, 0.100 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3$,
- **FA 3,** aqueous ammonium peroxydisulfate, $(NH_4)_2S_2O_8$,
- aqueous potassium iodide, KI,
- starch indicator.

You are advised to read the instructions before starting any practical work.

(a) Procedure

You will attempt **four** experiments. In each experiment, **solution 1** will be prepared as described below and a standard volume of **FA 3** will then be added.

For each experiment, you will note the time taken for the solution to turn dark blue.

In the space provided on page 7, prepare a table in which to record results for **four** experiments:

- volumes of **FA 3** and volumes of water to an appropriate level of precision,
- all values of time recorded to the nearest second,
- all values of $\frac{1000}{\text{time}}$ to 3 significant figures.

For each experiment, prepare **solution 1** in a dry 150 cm^3 beaker by adding the following:

- 5.0 cm^3 of **FA 1** using the burette in Question 1
- 20.0 cm^3 of KI(aq) using an appropriate measuring cylinder
- and 2.0 cm^3 starch using an appropriate measuring cylinder

Then stir the contents of the beaker using a glass rod.

(i) Experiment 1

The end–point of the reaction is the first appearance of a blue colour.
Note: Insufficient swirling of the reaction mixture may lead to a blue colour appearing before the true end–point is reached.

1. Place the beaker containing **solution 1** on a white tile.
2. Using an appropriate measuring cylinder, transfer 50.0 cm^3 of **FA 3** to the beaker. Start the stopwatch during this addition.
3. Stir the mixture gently using a glass rod.
4. Stop the stopwatch when the solution **first** turns blue.
5. Record the time taken.
6. Wash the beaker thoroughly with water and dry them.

(ii) Experiment 2

1. Prepare **solution 1** as described earlier.
2. Using a 50 cm^3 measuring cylinder, measure out 10.0 cm^3 of deionised water into the same beaker containing **solution 1**.
3. Place the beaker containing **solution 1** and deionisied water on a white tile.
4. Using an appropriate measuring cylinder, transfer 40.0 cm^3 of **FA 3** to the beaker. Start the stopwatch during this addition.
5. Stir the mixture gently using a glass rod.
6. Stop the stopwatch when the solution **first** turns blue.
7. Record the time taken.
8. Wash the beaker thoroughly with water and dry them.

(iii) Experiments 3 and 4

Carry out **two** further experiments to investigate the effect of changing the concentration of aqueous ammonium peroxydisulfate by altering the volume of **FA 3** used.

You should **not** use a volume of **FA 3** that is more than 40.0 cm^3 and less than 10.0 cm^3 and the total volume of the reaction mixture must always be 77.0 cm^3.

Results

Experiment	Volume of FA 3 / cm^3	Volume of water / cm^3	Time / s	$\left(\frac{1000}{time}\right)$ / s^{-1}
1	50.0	0.0	18	55.6
2	40.0	10.0	23	43.5
3	30.0	20.0	33	30.3
4	20.0	30.0	55	18.2

- **Table has correct headers and units**
- **$V_{FA3} + V_{H2O}$ = 50.0 cm^3 in each experiment & V_{FA3} must be between 10 and 40 cm^3 for expts 3 and 4**
- **All volume readings to 1 d.p, time recording to nearest second, rate calculated in 3sf**

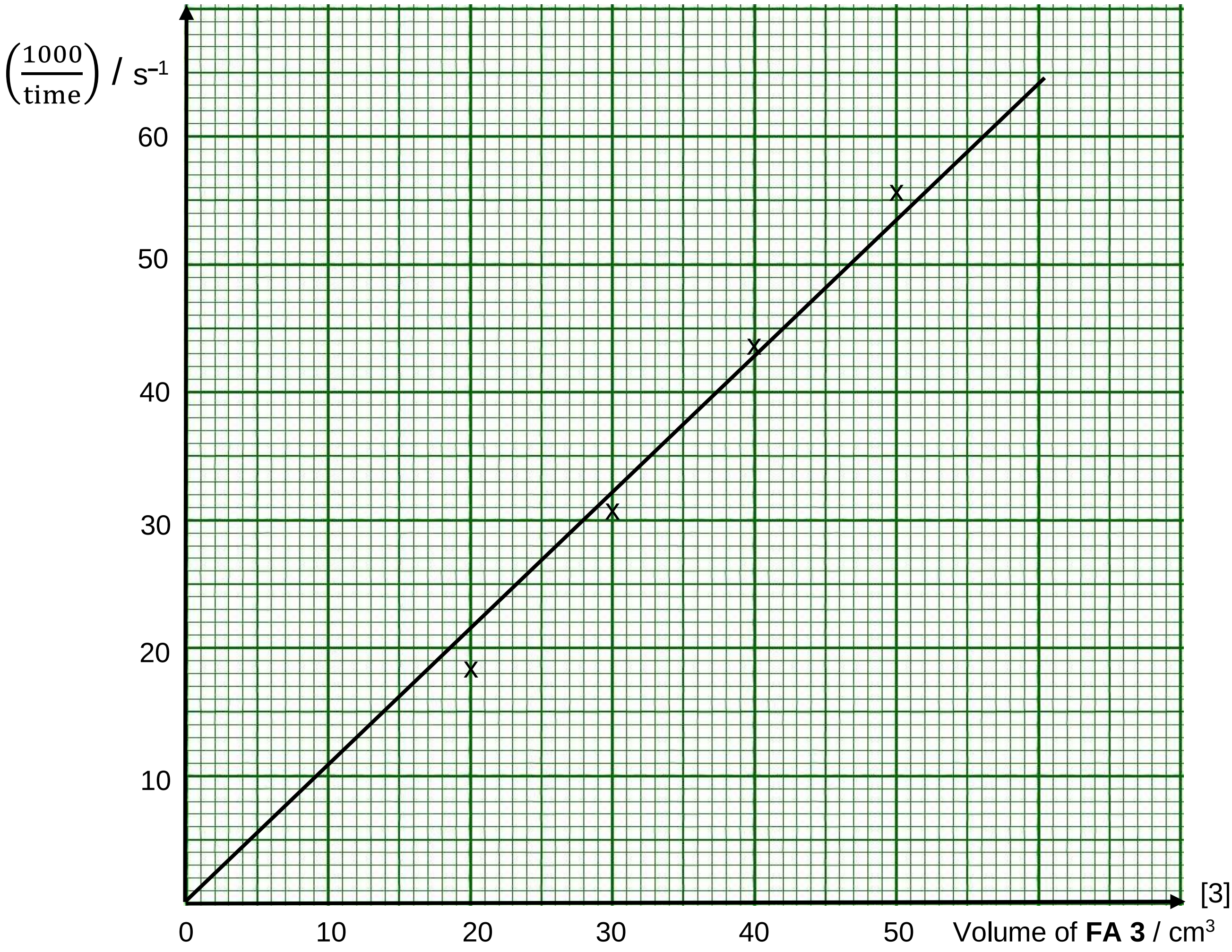
[3]

(b) Explain why $\left(\frac{1000}{\text{time}}\right)$ is a measure of the rate of reaction.

- **Rate of reaction ∝ $\frac{\text{amount of iodine produced}}{\text{time taken}}$. Since a <u>constant amount of $S_2O_3^{2-}$ is used</u> for every experiment, <u>constat amount of iodine</u> is <u>produced (or reacted)</u> for every experiment, <u>rate is inversely proportional to time</u>.**

[1]

(c) **(i)** On the grid, plot a graph of $\frac{1000}{\text{time}}$, on the *y*-axis, against the volume of **FA 3**, on the *x*-axis. Draw a line of best–fit through the points. Include the origin in your graph.



[3]

(ii) Using your graph, deduce the order of reaction with respect to **FA 3**.

- **Since total volume is kept constant in each experiment, volume of FA 3 used ∝ [$S_2O_8^{2-}$] in the mixture. A straight line graph through the origin (or 1000/time ∝ $V_{FA\ 3}$ from the graph) is obtained, i.e. rate ∝ volume of FA 3. Thus, rate ∝ [$S_2O_8^{2-}$]. Reaction is first order with respect to $S_2O_8^{2-}$.**

[1]

(d) Experiment 1 in **2(a)** can be modified to enable the study of the order of reaction with respect to iodide ions. Concentration of $(NH_4)_2S_2O_8$ should also be kept the same as that in experiment 1.

Suggest the volume of KI, H_2O, $Na_2S_2O_3$ and starch that could be used in this experiment.

Experiment	volume of $(NH_4)_2S_2O_8$ / cm^3	volume of KI / cm^3	volume of H_2O / cm^3	volume of $Na_2S_2O_3$ / cm^3	Volume of starch / cm^3
1	50.0	20.0	0.0	5.0	2.0
5	50.0	**10.0**	**10.0**	**5.0**	**2.0**

1m for the correct volumes of KI, H_2O, $Na_2S_2O_3$ and starch for experiment 5

[1]

(e) Another student performed experiment 1 in **2(a)**. She repeated the experiment but, before adding the **FA 3** to **solution 1**, she added a few drops of $FeCl_3$(aq) to **FA 3**. She found that this addition considerably reduced the time required for a dark blue colour to form.

With the aid of relevant equations, explain why the time required for a dark blue colour to form was reduced considerably when $FeCl_3$(aq) is added.

- $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$
  $2Fe^{2+} + S_2O_8^{2-} \rightarrow 2Fe^{3+} + 2SO_4^{2-}$

- **$FeCl_3$(aq) acts as a homogenous catalyst which provides an alternative pathway with lower activation energy so that it speeds up the reaction and the time taken will be reduced.**

[2]
[Total: 11]

3 Determination of the enthalpy change of dehydration of hydrated copper(II) sulfate

Hydrated copper(II) sulfate, $CuSO_4•5H_2O$, can lose its water of crystallisation to form anhydrous copper(II) sulfate.

The enthalpy change for the dehydration of hydrated copper(II) sulfate is shown in equation 6.

equation 6 $CuSO_4•5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$

You will carry out experiments to determine the enthalpy changes for the solution of hydrated and anhydrous copper(II) sulfate and then use Hess's law to determine the enthalpy change of dehydration.

FA 4 is hydrated copper(II) sulfate, $CuSO_4•5H_2O$.
FA 5 is anhydrous copper(II) sulfate, $CuSO_4$.

(a) Determination of the enthalpy change of solution of hydrated copper(II) sulfate

Method

1. Weigh the container with **FA 4**. Record the mass.
2. Support a styrofoam cup in a 250 cm^3 beaker.
3. Use a measuring cylinder to transfer 25.0 cm^3 of deionised water into the cup.
4. Measure and record the temperature of the water.
5. Tip all the **FA 4** into the water and stir until the solid dissolves.
6. Measure and record the lowest temperature reached.
7. Rinse and dry the styrofoam cup ready for the next experiment.
8. Weigh the container with any residual **FA 4**. Record the mass.
9. Calculate and record the change in temperature.
10. Calculate and record the mass of **FA 4** used.

Results

Mass of container and FA 4 / g	**10.656**
Mass of container and residual FA 4 / g	**5.589**
Mass of FA 4 used / g	**5.067**
Initial temperature of water / °C	**30.0**
Lowest temperature of solution / °C	**27.6**
Change in temperature / °C	**2.4**

- **Correct headers and units for 3(a) and 3(c)**
- **3 d.p. for mass readings and 1 d.p. for temperature readings for 3(a) and 3(c)**

[2]

(b) Calculations

(i) Calculate the energy change, in J, during this reaction.

Assume that the specific capacity of the reaction mixture is 4.18 J g^{-1} K^{-1}, and that the density of the reaction mixture is 1.00 g cm^{-3}.

- **Energy change = 25.0 × 4.18 ×2.4**
  = 251 J

energy change = ...**251**.......... J
[1]

(ii) Calculate the amount, in mol, of hydrated copper(II) sulfate, **FA 4**, used in the experiment. Show your working.

[A_r: H, 1.0 O, 16.0 S, 32.1 Cu, 63.5]

- **Amount of FA 4 = 5.067 ÷ (63.5 + 32.1 + 64.0 + 5(18.0))**
  = 0.0203 mol

amount of $CuSO_4 \cdot 5H_2O$ = ...**0.0203**.......... mol
[1]

(iii) Hence, calculate the enthalpy change, in kJ mol^{-1}, when 1.00 mol of hydrated copper(II) sulfate dissolves in water. This is the enthalpy of solution.

- **ΔH = +251 ÷ 0.0203**
  = +12365 J mol^{-1}
  = +12.4 kJ mol^{-1}

Must include sign

enthalpy change of solution = ...**+12.4**.......... kJ mol^{-1}
[1]

(c) Determination of the enthalpy change of solution of anhydrous copper(II) sulfate

Method

1. Weigh the container with **FA 5**. Record the mass.
2. Support a styrofoam cup in a 250 cm^3 beaker.
3. Use the measuring cylinder to transfer 25.0 cm^3 of deionised water into the cup.
4. Measure and record the temperature of the water.
5. Tip all the **FA 5** into the water and stir until the solid dissolves.
6. Measure and record the highest temperature reached.
7. Weigh the container with any residual **FA 5**. Record the mass.
8. Calculate and record the change in temperature.
9. Calculate and record the mass of **FA 5** used.

Results

Mass of container and FA 5 / g	**6.314**
Mass of container and residual FA 5 / g	**4.785**
Mass of FA5 used / g	**1.529**
Initial temperature of water / °C	**30.0**
Highest temperature of solution / °C	**35.0**
Change in temperature / °C	**5.0**

Accuracy:
ΔT/m = 3.27 ±0.5 (2 marks)
ΔT/m = 3.27 ±1.0 (1 mark)

[2]

(d) Calculations

(i) Calculate the enthalpy change, in kJ mol^{-1}, for the enthalpy change of solution of anhydrous copper(II) sulfate.

[A_r: H, 1.0 O, 16.0 S, 32.1 Cu, 63.5]

- **Energy change = 25.0 × 4.18 × 5.0**
  = 523 J
  Amount of FA 5 = 1.529 ÷ (63.5 + 32.1 + 64.0)
  = 0.00959 mol
  Δ*H* = –523 ÷ 0.00959
  = –54536 J mol^{-1}
  = –54.5 kJ mol^{-1}

enthalpy change = ...**–54.5**...... kJ mol^{-1}

[1]

(ii) Use the enthalpy changes calculated in **3(b)(iii)** and **3(d)(i)** to calculate the enthalpy change of dehydration of hydrated copper(II) sulfate, in kJ mol^{-1}.

$$CuSO_4{\bullet}5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$$

Show clearly how you obtained your answer.

$$\begin{tikzcd}
\mathbf{CuSO_4{\bullet}5H_2O(s)} \arrow[rr, "\Delta H"] \arrow[dr, "+12.4"'] & & \mathbf{CuSO_4(s) + 5H_2O(l)} \arrow[dl, "-54.5"] \\
& \mathbf{Cu^{2+}(aq) + SO_4^{2-}(aq) + 5H_2O(l)} &
\end{tikzcd}$$

$\Delta H = +12.4 - (-54.5)$
$= +66.9$ kJ mol^{-1}

- **1 mark for using relevant values from (b)(iii) and (d)(i)**
- **Apply Hess's Law with or without cycle**
- **Correct statements, units and 3sf for final answers in 3(b) and 3(d)**

enthalpy change = ...**+66.9**.......... kJ mol^{-1}
[3]

(e) In the experiments in **3(a)** and **3(c),** you used the same method to determine the enthalpy change of solution of two solids. Compare the relative percentage error in **3(b)(iii)** and **3(d)(i)**. Explain your answer.

Ignore the effect of differences in mass used.

- **As magnitude of ΔT is smaller for FA 4, the percentage error in 3(b)(iii) is larger than that in 3(d)(i).**

[1]

(f) **Planning**

Another sample of hydrated copper(II) sulfate crystals is taken for analysis. The amount of water of crystallisation in the hydrated salt can be determined gravimetrically (by weighing).

The water of crystallisation in hydrated copper(II) sulfate crystals can be removed by heating. Anhydrous copper(II) sulfate does not decompose at temperatures achieved by a Bunsen flame.

(i) Plan an investigation to determine the value of x in hydrated copper(II) sulfate crystals, $CuSO_4 \cdot xH_2O$, gravimetrically.

You may assume that you are provided with:

- approximately 10 g of hydrated copper(II) sulfate crystals, $CuSO_4 \cdot xH_2O$
- all other 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 procedure you would follow
- the measurements you would make (you may find it useful to label measurements in your plan as M1, M2, M3 etc.)
- how you would ensure that an **accurate** and **reliable** value of x is obtained

1. **Using a weighing balance, weigh a dry and empty boiling tube. Record this reading as M1.**

2. **Using a weighing balance, add about 5.0 g of $CuSO_4 \cdot xH_2O$ crystals. Record mass of solid and boiling tube as M2.**

3. **Heat the boiling tube gently at first (~ 1 min), then strongly (~5 min) to decompose the sample.**

4. **Allow the boiling tube and the solid residue to cool for some time on a heat–proof mat.**

5. **Using a weighing balance, weigh the boiling tube and solid residue. Record this reading.**

6. **Repeat steps 3 to 5 until a constant mass is achieved to ensure complete decomposition has occurred. Record the constant mass as M3.**

- **1 m for correct choice of apparatus (weighing balance, dry boiling tube)**
- **1 m for correct sequence of steps (heating, cooling, weighing and reheating)**
- **1 m for correct measurements made (mass of empty boiling tube, mass of boiling tube & contents before and after heating, M1, M2, M3 etc.)**
- **1 m for repeat heating until constant mass**

Reject TARE for empty boiling tube before weighing solid.
Accept use of crucible in place of boiling tube for heating solid.

[4]

(ii) Describe how you would use your measurements in **3(f)(i)** to calculate a value of x.

[A_r: H,1.0 O,16.0 S, 32.1, Cu, 63.5]

Before heating

Mass of boiling tube and $CuSO_4 \cdot xH_2O$ / g	**M2**

Mass of empty boiling tube / g	M1
Mass of $CuSO_4 \bullet xH_2O$ used / g	M2 – M1

After heating

Mass of residue after 3rd heating / g	M3 – M1

- Amount of $CuSO_4 = \frac{M3 - M1}{159.6}$ mol

(mass of residue + tube – mass of tube)

- Amount of H_2O lost $= \frac{M2 - M3}{18.0}$ mol

(mass of hyd solid + tube – mass of residue + tube)

- x = Amount of H_2O lost ÷ Amount of $CuSO_4$

$$= \frac{M2 - M3}{18.0} \div \frac{M3 - M1}{159.6}$$

Alternative answer 1:

- Amount of $CuSO_4 = \frac{M3 - M1}{159.6}$ mol

(mass of residue + tube – mass of tube)

- Amount of $CuSO_4 \bullet xH_2O = \frac{M2 - M1}{159.6 + 18.0x}$ mol

(mass of hyd solid + tube – mass of tube)

- Amount of $CuSO_4$ = Amount of $CuSO_4 \bullet xH_2O$

$$\frac{M3 - M1}{159.6} = \frac{M2 - M1}{159.6 + 18.0x}$$

$$x = \left[\frac{159.6(M2 - M1)}{M3 - M1} - 159.6\right] \div 18.0$$

Alternative Answer2:

Accept $x = \frac{\text{Amount of } H_2O}{\text{Amount of } CuSO_4 \cdot xH_2O}$

$$x = \frac{M2 - M3}{18.0} \div \frac{M2 - M1}{159.6 + 18.0x}$$

$$x = \frac{M2 - M3}{18.0} \times \frac{159.6 + 18.0x}{M2 - M1}$$

$$x = \frac{(159.6 + 18.0x)(M2 - M3)}{18.0(M2 - M1)}$$

$$x = \frac{159.6(M2 - M3)}{18.0(M3 - M1)}$$

[3]
[Total: 19]

4 Qualitative Analysis

FA 6 is a solid containing one cation and one anion.
FA 7 is a solution containing one cation and one anion.

(a) Carry out the following tests and record your observations.
The volumes given are approximate and should be estimated rather than measured.
You do **not** need to test and identify any gases evolved.
If any solution is warmed, a **boiling tube** must be used.

(i) Warm (do **not** boil) a 5 cm depth of **FA 7** in a boiling tube. Stop warming the **FA 7**, add all of the **FA 6** and shake the boiling tube. Filter the mixture into a second boiling tube. The filtrate will be used in the tests in **4(a)(ii)**. Write your observations in the space below.

- **Brick-red/Red/Red-brown solid dissolved to give a pale blue solution.**
- **Pale blue filtrate (or pale blue solution) was obtained together with a red–brown solid as residue.**

[2]

(ii) Use a 1 cm depth of the filtrate from **4(a)(i)** in separate test–tubes for each of the following tests.

test	observations
Add aqueous ammonia.	• **Blue ppt. soluble in excess NH_3(aq) to form a dark blue solution.**
Add a 1 cm depth of dilute nitric acid followed by a 1 cm depth of aqueous silver nitrate.	• ***No effervescence. No ppt. formed / solution remains blue.**
Add a 1 cm depth of dilute nitric acid followed by a 1 cm depth of aqueous barium nitrate.	• ***No effervescence. White ppt formed.**

[3]

*** No effervescence for test 2 and test 3. Do not penalise as long as it has been mentioned once**

(iii) **FA 8** is a dry sample of the residue obtained by filtration in **4(a)(i)**.

test	observations
Add a 1 cm depth of dilute nitric acid to all of the **FA 8** in its test–tube. Allow the mixture to stand for about 1 minute. Then add aqueous sodium hydroxide.	• **Effervescence. Brown gas evolved.** **Red-brown solid dissolved to give a pale blue solution.** **Blue ppt insoluble in excess NaOH(aq).**

[2]

(b) **(i)** From your observations in **4(a)(ii)**, suggest the identity of the cation and the anion present in the filtrate produced in **4(a)(i)**.

cation present in the filtrate ……**Cu^{2+}**……

anion present in the filtrate ……………SO_4^{2-}……………

- **Correct cation and anion**

[1]

(ii) Write ionic equations for reactions in **4(a)(ii)** where a precipitate was formed. Include state symbols.

- **$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$**
- **$Cu^{2+}(aq) + 2OH^{-}(aq) \rightarrow Cu(OH)_2(s)$**

Equations must be balanced with state symbols

[2]

(iii) State the type of reaction that occurred when dilute nitric acid is added to **FA 8**. Explain your answer with reference to your observations in **(a)(iii)**.

- **Redox reaction. Red-brown Cu is oxidised as the oxidation number of Cu increased from 0 in Cu to +2 in Cu^{2+} which gives a blue solution.**

[1]

(c) A student suggested that **FA 7** is an acid.

Apart from using an indicator, suggest and carry out a chemical test to determine whether the student was correct. **You may use any of the chemicals that you are provided with in this paper.**

Describe the test by including the name of the reagent, your observations and your conclusion.

- **Test: Add aqueous sodium carbonate into a test–tube containing 2 cm^3 of FA 7.**
- **Observation: Effervescence of CO_2 gas is produced which formed a white ppt with limewater. Student is correct (or FA 7 is an acid).**

[Alternative]

- **Test: Add FA1 into a test-tube containing 2 cm^3 of FA 7.**
- **Observation: Off-white / Yellow ppt is formed. Student is correct (or FA 7 is an acid).**

[2]

[Total: 13]

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 A*l* foil
nitrite, $NO_2^-(aq)$	NH_3 liberated on heating with $OH^-(aq)$ and A*l* 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 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