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.

[Turn over

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

[3]

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

$V_{\textbf{FA 1}}$ = ………………...

[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]

mass of Cu in 11.54 g of the alloy = ……………………………………

percentage by mass of copper in the alloy = …………………………………
[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 ions = …………………………………
[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.

ratio = ……………………………
[1]

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

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[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

[3]

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

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[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**.

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[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				

[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.

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[2]

[Total: 11]

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

Hydrated copper(II) sulfate, $CuSO_4 \bullet 5H_2O$, 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 $CuSO_4 \bullet 5H_2O(s) \rightarrow CuSO_4(s) + 5H_2O(l)$

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, $CuSO_4 \bullet 5H_2O$.
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

[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 = ... 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 $CuSO_4$•$5H_2O$ = ... 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.

enthalpy change of solution = ... 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

[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]

enthalpy change = ... 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.

enthalpy change = ... 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.

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(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, $CuSO_4 \bullet xH_2O$, gravimetrically.

You may assume that you are provided with:

- approximately 10 g of hydrated copper(II) sulfate crystals, $CuSO_4 \bullet xH_2O$
- 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

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[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]

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[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.

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[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.	
Add a 1 cm depth of dilute nitric acid followed by a 1 cm depth of aqueous silver nitrate.	
Add a 1 cm depth of dilute nitric acid followed by a 1 cm depth of aqueous barium nitrate.	

[3]

(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.	

[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 …………………………

anion present in the filtrate …………………………

[1]

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

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[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)**.

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[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.

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[2]

[Total: 13]

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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 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