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Thursday 23 June 2022 – Morning

A Level Chemistry B (Salters)

H433/03 Practical skills in chemistry

Time allowed: 1 hour 30 minutes



You must have:

- the Practical Insert (inside this document)
- the Data Sheet for Chemistry B

You can use:

- a scientific or graphical calculator
- an HB pencil



Please write clearly in black ink. **Do not write in the barcodes.**

Centre number

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Candidate number

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First name(s)

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Last name

StemUp Model Solution

INSTRUCTIONS

- Use black ink. You can use an HB pencil, but only for graphs and diagrams.
- Write your answer to each question in the space provided. If you need extra space use the lined pages at the end of this booklet. The question numbers must be clearly shown.
- Answer **all** the questions.
- Where appropriate, your answer should be supported with working. Marks might be given for using a correct method, even if your answer is wrong.

INFORMATION

- The total mark for this paper is **60**.
- The marks for each question are shown in brackets [].
- Quality of extended response will be assessed in questions marked with an asterisk (*).
- This document has **16** pages.

ADVICE

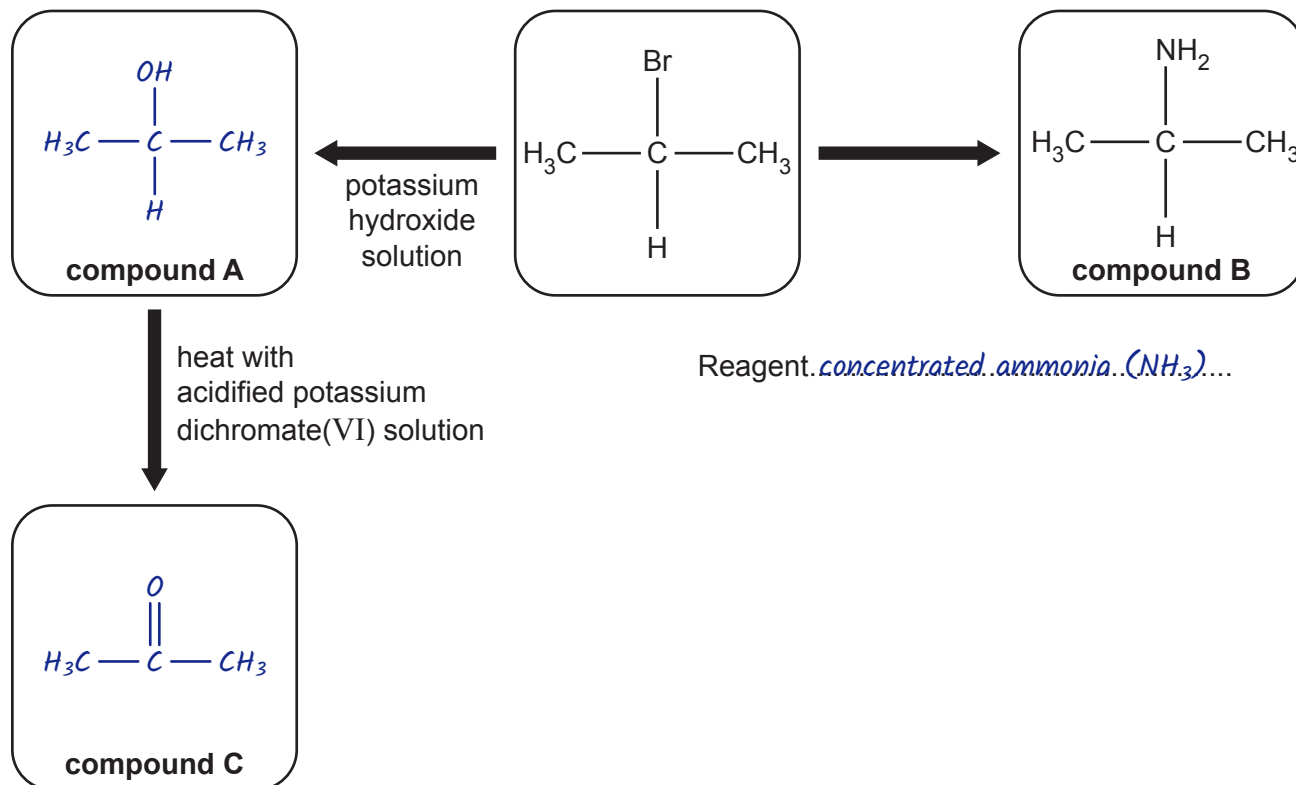
- Read each question carefully before you start your answer.

Answer **all** the questions.

1 This question is about haloalkanes and their reactivity.

(a) Haloalkanes are useful intermediates for preparing a range of organic compounds.

A flowchart showing the synthesis of three organic compounds is shown below.



(i) Complete the flowchart showing the structures of compounds **A** and **C**.

On the dotted line show the reagent needed to form compound **B**.

[3]

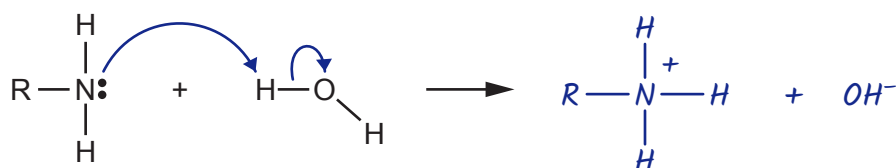
(ii) Which homologous series does compound **B** belong to?

(primary) amines..... [1]

(b) Compound **B** behaves as a base.

Complete the diagram of a mechanism that shows how a molecule of compound **B** reacts with water to give a basic solution.

Show 'curly arrows'.



[2]

- (c) Some haloalkanes act as greenhouse gases that absorb infrared radiation in the troposphere.

Give the source of this infrared radiation.

infrared radiation emitted (re-radiated) by the Earth's surface [1]

- (d)* A student wishes to compare the relative reactivities of haloalkanes.

The student investigates the rates of hydrolysis of 1-chlorobutane, 1-bromobutane and 1-iodobutane by following the reaction in the presence of aqueous silver ions.

Design a suitable procedure that the student should follow.

Give the expected results and how they can be interpreted.

Put 2 cm³ of ethanol (solvent) in each of three test tubes and add 5 drops (same volume) of 1-chlorobutane, 1-bromobutane or 1-iodobutane to each.

Put equal volumes (2 cm³) of 0.1 mol dm⁻³ aqueous silver nitrate into three more tubes.

Stand all six tubes in a water bath at about 50 °C until they reach the same temperature.

Quickly add the silver nitrate from one tube to each haloalkane tube, shake to mix and start a stopwatch.

Time how long it takes for a precipitate to appear in each tube.

Results: in hydrolysis the C-Hal bond breaks, releasing a halide ion: $RX + H_2O \rightarrow ROH + H^+ + X^-$. This forms a silver halide precipitate: $Ag^+(aq) + X^-(aq) \rightarrow AgX(s)$.

The precipitate forms first with 1-iodobutane (yellow AgI), then 1-bromobutane (cream AgBr) and last with 1-chlorobutane (white AgCl). [6]

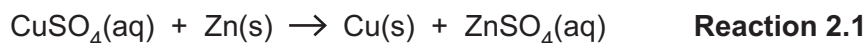
Additional answer space if required.

The quicker the precipitate forms, the faster the rate of hydrolysis, so the rate increases in the order chloro < bromo < iodo.

Reactivity depends on bond strength: the C-Cl bond is the strongest, then C-Br, then C-I, so the weak C-I bond breaks most easily.

2 This question is about the enthalpy change of the reaction of zinc with copper(II) sulfate solution.

(a) A student investigates the temperature change when **Reaction 2.1** occurs.



- (i) The student adds 1.20 g of powdered zinc to 50.0 cm³ of 0.200 mol dm⁻³ copper(II) sulfate solution in a glass beaker.

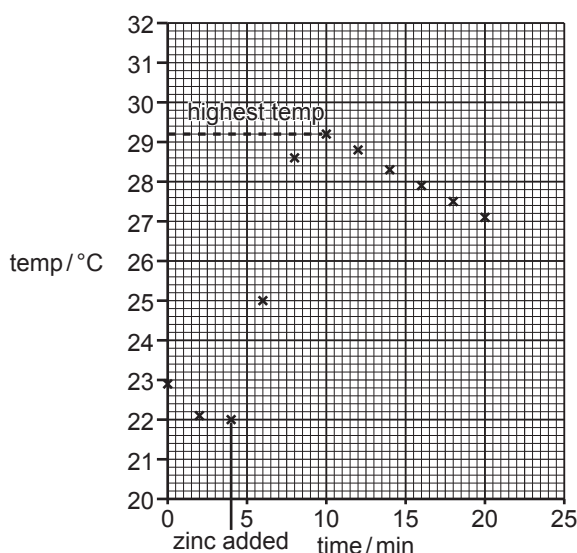
Use calculations to show which is the limiting reagent.

$$n(\text{CuSO}_4) = 0.200 \times 50.0 \div 1000 = 0.0100 \text{ mol}$$

$$n(\text{Zn}) = 1.20 \div 65.4 = 0.0183 \text{ mol, so Zn is in excess (1 : 1)}$$

limiting reagent is ...*copper(II) sulfate*..... [2]

- (ii) The student measures the temperature of the contents of the beaker over a period of 20 minutes and plots the data on the graph below.



The student uses the graph to determine the highest temperature as shown.

The student uses this temperature to calculate the heat energy given to the solution per mole of limiting reagent.

Show the student's calculation, giving the answer in kJ mol⁻¹.

Give your answer to an **appropriate** number of significant figures.

$$\Delta T = 29.2 - 22.0 = 7.2 \text{ }^{\circ}\text{C}$$

$$Q = mc\Delta T = 50.0 \times 4.18 \times 7.2 = 1504.8 \text{ J} = 1.5048 \text{ kJ}$$

$$\text{per mole of CuSO}_4 = 1.5048 \div 0.0100 = 150.48 \text{ kJ mol}^{-1}$$

$$= 150 \text{ kJ mol}^{-1} \text{ (3 s.f.)}$$

heat energy given to the solution = ...*150*..... kJ mol⁻¹ [3]

(b) Calculate $\Delta_r H^\ominus$ for **Reaction 2.1** using the data shown below.

	$\Delta_r H^\ominus / \text{kJ mol}^{-1}$
$\text{Cu}^{2+}(\text{aq})$	+ 64.4
$\text{Zn}^{2+}(\text{aq})$	-152.4

$$\Delta H = -152.4 - (+64.4)$$

$$\Delta_r H^\ominus = \dots\dots\dots -216.8 \dots\dots\dots \text{kJ mol}^{-1} \quad [1]$$

(c)* The value obtained from the student's experiment is considerably less exothermic than the value in part (b).

Evaluate the student's experiment, identifying limitations in both the experimental procedure and the measurements taken. You should also comment on how any of the limitations you have identified will affect the final value.

Suggest possible improvements to the procedure, apparatus and measurements. [6]

Limitations: the reaction was done in an open glass beaker with no lid or insulation, so a lot of heat was lost to the surroundings. This is the most significant error and heat losses give a less exothermic value.

The highest temperature read from the graph (29.2 °C) is too low because the mixture was already cooling while the zinc was still reacting, so ΔT (7.2 °C) is too small. Using this lower temperature gives a less exothermic value.

The calculation used only the mass (50 g) of water, not the whole solution, and ignored the heat absorbed by the metals (zinc and copper) and the beaker. Using the mass of water and ignoring the heat given to the metals gives a less exothermic value.

Improvements: use a lagged polystyrene cup with a lid (or a Dewar flask) to lower heat losses.

Stir continuously and record the temperature every 30 s for a few minutes before adding the zinc and for longer afterwards. Then extrapolate the

You may continue your answer over the page if you need to.

Additional answer space if required for part 2(c).

cooling line back to the time the zinc was added to give the true maximum temperature and a more realistic total heat change.

Use a thermometer with a higher resolution ($\pm 0.1\text{ }^{\circ}\text{C}$) or a temperature probe and data logger, and measure the solution with a pipette or burette.

In the calculation use the mass and specific heat capacity of the solution, and allow for the heat given to the metals.

.....

.....

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Turn over for the next question

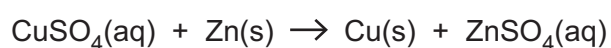
- 3 This question is about redox and electrochemical cells.

Table 3.1 shows standard electrode potentials, some of which will be needed for the rest of this question.

Half reaction	Standard electrode potential, $E^\ominus$ / V
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Fe}(\text{s})$	-0.44
$2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2(\text{g})$	0.00
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.34
$\frac{1}{2}\text{O}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightleftharpoons 2\text{OH}^-(\text{aq})$	+0.40
$\frac{1}{2}\text{I}_2(\text{aq}) + \text{e}^- \rightleftharpoons \text{I}^-(\text{aq})$	+0.54
$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80
$\text{Cl}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-(\text{aq})$	+1.36
$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51

Table 3.1

- (a) The reaction between zinc metal and aqueous Cu^{2+} ions can be arranged in a cell to produce electrical energy.



A diagram of a copper/zinc cell is shown in **Fig. 3.1**.

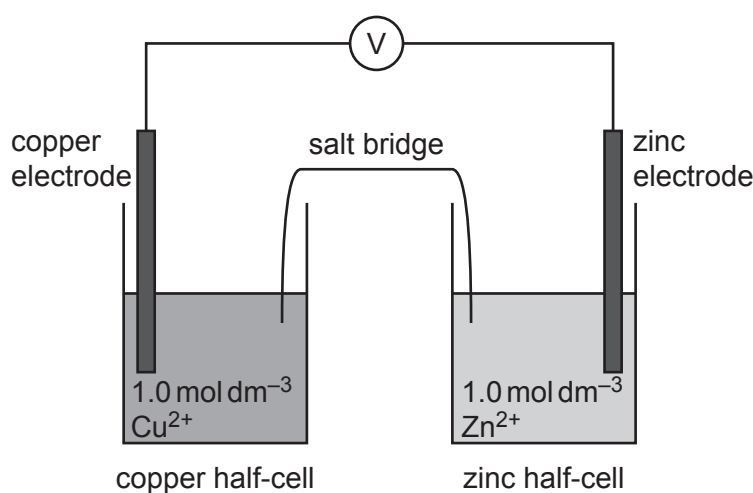
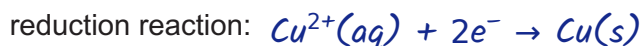


Fig. 3.1

When the cell delivers a current, oxidation takes place in one half-cell and reduction in the other.

- (i) Write half-equations for the reactions that take place in each half-cell.

Show state symbols.



[2]

- (ii) Explain the purpose of the salt bridge.

*It allows ions to flow (move) between the two half-cells,
so it completes the circuit (and keeps the solutions neutral).*

[2]

- (iii) Calculate $E^{\ominus}_{\text{cell}}$ for the copper/zinc cell in **Fig. 3.1**.

$$E^{\ominus}_{\text{cell}} = +0.34 - (-0.76)$$

$$E^{\ominus}_{\text{cell}} = 1.10 \text{ V [1]}$$

- (b) Standard electrode potentials can be used to decide whether a reaction is feasible.

Predict, with reasons, if any reaction could take place in each of the following situations.

Use the data in **Table 3.1**.

- (i) Metallic silver is added to aqueous iron(II) sulfate.

No reaction: $E^{\ominus}$ of Ag^{+}/Ag (+0.80 V) is more positive than Fe^{2+}/Fe (-0.44 V), and silver can only lose electrons to a more positive $E^{\ominus}$.

[1]

- (ii) Aqueous potassium chloride is added to acidified potassium manganate(VII) solution.

Yes: $\text{Cl}_2/\text{Cl}^{-}$ (+1.36 V) is less positive than $\text{MnO}_4^{-}/\text{Mn}^{2+}$ (+1.51 V), so Cl^{-} loses electrons to MnO_4^{-} and is oxidised to chlorine:



[2]

- (c) The relationship between the ion concentration and the electrode potential for a metal/ion electrode is given (at 298 K) by **Equation 3.1**, where n is the number of electrons transferred in the half reaction.

$$E = E^{\ominus} + \frac{0.059 \times \log_{10}[\text{ion}]}{n} \quad \text{Equation 3.1}$$

- (i) Explain how **Equation 3.1** shows that $E = E^{\ominus}$ under standard conditions.

Under standard conditions $[\text{ion}] = [\text{Zn}^{2+}] = 1 \text{ mol dm}^{-3}$

$\log_{10}(1) = 0$, so the term $0.059 \times \log_{10}[\text{ion}] \div n = 0$ and $E = E^{\ominus}$.

..... [2]

- (ii) Calculate the electrode potential of a $\text{Zn}^{2+}(\text{aq})/\text{Zn}(\text{s})$ half-cell when $[\text{Zn}^{2+}(\text{aq})] = 0.20 \text{ mol dm}^{-3}$.

$$\begin{aligned} E &= -0.76 + (0.059 \times \log_{10} 0.20) \div 2 \\ &= -0.76 + (0.059 \times -0.699) \div 2 = -0.76 - 0.0206 = -0.781 \end{aligned}$$

$$E = \text{.....} -0.78 \text{..... V [2]}$$

Half reaction	Standard electrode potential, $E^{\ominus} / \text{V}$
$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Zn}(\text{s})$	-0.76
$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Fe}(\text{s})$	-0.44
$2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{H}_2(\text{g})$	0.00
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{Cu}(\text{s})$	+0.34
$\frac{1}{2}\text{O}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^{-} \rightleftharpoons 2\text{OH}^{-}(\text{aq})$	+0.40
$\frac{1}{2}\text{I}_2(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{I}^{-}(\text{aq})$	+0.54
$\text{Fe}^{3+}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
$\text{Ag}^{+}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{Ag}(\text{s})$	+0.80
$\text{Cl}_2(\text{g}) + 2\text{e}^{-} \rightleftharpoons 2\text{Cl}^{-}(\text{aq})$	+1.36
$\text{MnO}_4^{-}(\text{aq}) + 8\text{H}^{+}(\text{aq}) + 5\text{e}^{-} \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51

Table 3.1

- (d) The rusting of iron is an electrochemical process.

A solution of 'ferroxyl' indicator can be used to investigate the reactions taking place in rusting. This indicator turns blue in the presence of Fe^{2+} ions and pink in alkaline solution.

In an experiment to investigate rusting, two iron nails are placed in a solution of 'ferroxyl' indicator (containing sodium chloride), in two separate petri dishes. One of the nails has copper wire wrapped around it, the other zinc wire.

The results are shown in **Fig. 3.2** below.

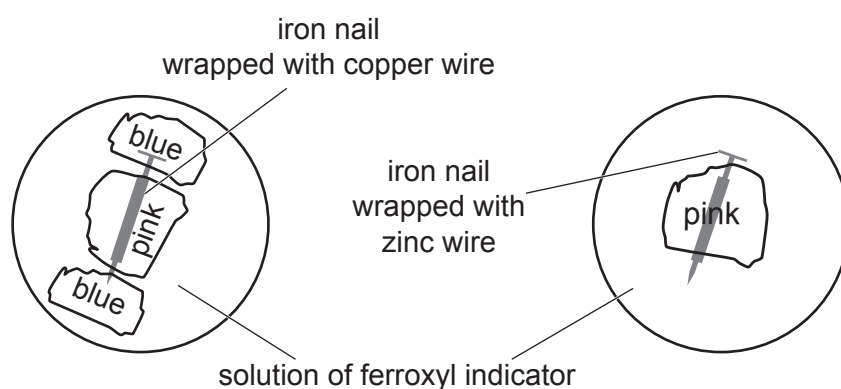


Fig. 3.2

A student suggests that the results in **Fig. 3.2** show that the attached zinc wire prevents the iron nail from rusting, but the attached copper wire does not.

Comment on the student's suggestion and use appropriate half-equations from **Table 3.1** to explain the colours in **Fig. 3.2**.

The student is correct: there is no blue (no Fe^{2+}) around the nail with zinc, so it does not rust, but the blue around the nail with copper shows Fe^{2+} forms, so it rusts.

Zinc (-0.76 V) is more negative than iron (-0.44 V), so zinc is oxidised instead; with copper ($+0.34 \text{ V}$) the iron is oxidised: $\text{Fe(s)} \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{e}^-$ (blue).

The pink in both dishes shows hydroxide ions (OH^-) form, as oxygen gains the electrons: $\frac{1}{2}\text{O}_2(\text{aq}) + \text{H}_2\text{O(l)} + 2\text{e}^- \rightarrow 2\text{OH}^-(\text{aq})$ [4]

4 This question refers to the Practical Insert that is provided as an insert to this paper.

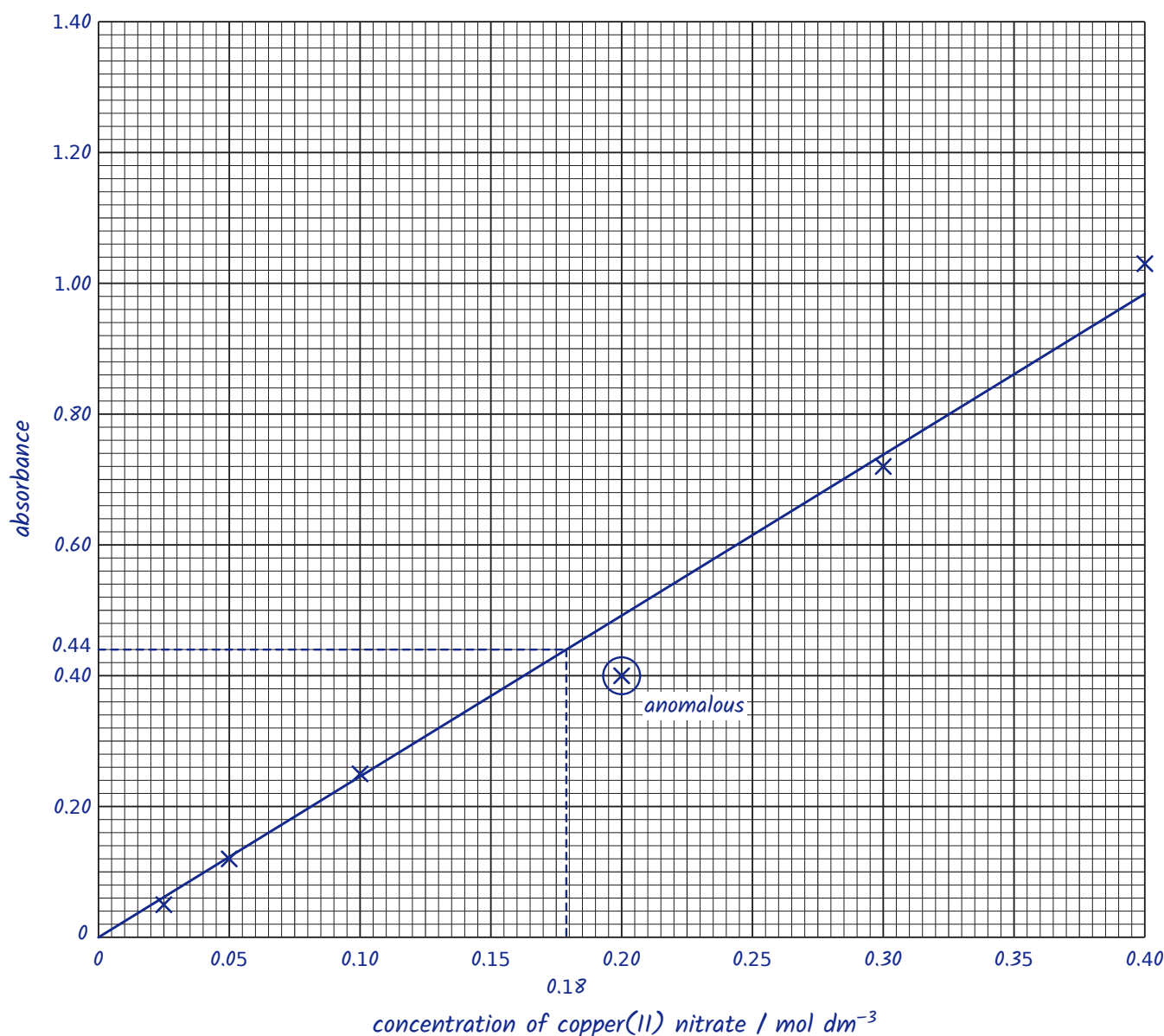
- (a) Suggest how the students accurately made up a solution of $0.300 \text{ mol dm}^{-3}$ copper(II) nitrate from their standard $0.400 \text{ mol dm}^{-3}$ solution.

Name the apparatus involved.

Using a pipette (and filler), transfer 75.0 cm^3 ($3 \times 25.0 \text{ cm}^3$) of the $0.400 \text{ mol dm}^{-3}$ solution into a 100 cm^3 volumetric flask. Make up to the mark with deionised water and invert to mix. This dilutes it by 4/3: $0.400 \times 75.0 \div 100 = 0.300 \text{ mol dm}^{-3}$.

[3]

- (b) (i) Plot a graph of absorbance against concentration on the graph paper below and draw an appropriate line of best fit. Label the axes. [3]



- (ii) Use your graph on page 12 and data from the insert to calculate the percentage by mass of copper in the brass sample.

From the graph, absorbance 0.44 gives $[Cu^{2+}] = 0.18 \text{ mol dm}^{-3}$

amount of Cu in $250 \text{ cm}^3 = 0.18 \times 250 \div 1000 = 0.045 \text{ mol}$

mass of Cu = $0.045 \times 63.5 = 2.86 \text{ g}$ (in 250 cm^3)

percentage = $2.86 \div 4.04 \times 100 = 70.7\%$

percentage by mass of copper in the brass sample = ...70.7..... % [3]

- (iii) Explain why an orange filter is placed into the colorimeter before taking the readings.

*Orange light is the complementary (opposite) colour to the blue
copper(II) solution, so it is the colour the solution absorbs most.
This gives the maximum absorption values (most sensitive readings).*

..... [2]

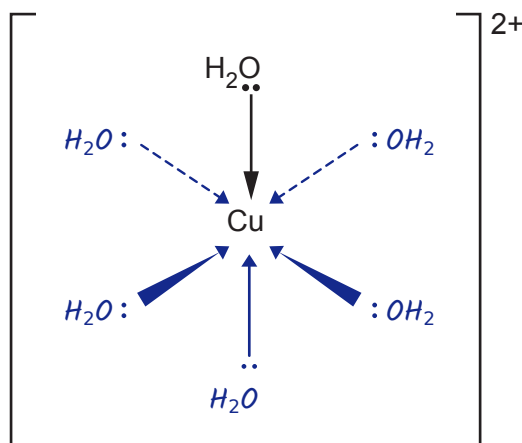
Question 4 continues on page 14

(c) A copper(II) ion is said to form an octahedral complex with water ligands.

(i) Explain the term **ligand**.

A molecule or ion that forms a coordinate (dative) bond to the central metal ion by donating a lone pair of electrons. [1]

(ii) Complete the diagram below to show the **shape** of the octahedral copper(II) complex ion with water ligands.



[3]

(d) d-block elements form a variety of differently coloured complex ions.

Explain, in terms of electronic structure, why these complex ions are coloured and why different complexes of the same cation have different colours.

The ligands cause the d orbital energy levels of the metal ion to split, giving an energy gap ΔE . Electrons absorb visible light and are excited to the higher d level; the light absorbed has energy equal to the gap, $\Delta E = h\nu$. The colour seen is the wavelengths (frequencies) not absorbed, the complementary colour. Different ligands give different gaps, so different frequencies are absorbed and colours differ. [4]

END OF QUESTION PAPER

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