

Surname <i>StemUp Model Solution</i>	Centre Number	Candidate Number
First name(s)		2



GCE A LEVEL

1410U30-1



S24-1410U30-1

MONDAY, 10 JUNE 2024 – MORNING

CHEMISTRY – A2 unit 3

Physical and Inorganic Chemistry

1 hour 45 minutes

Section A

Section B

For Examiner's use only		
Question	Maximum Mark	Mark Awarded
1. to 5.	10	
6.	9	
7.	10	
8.	18	
9.	13	
10.	20	
Total	80	

ADDITIONAL MATERIALS

- A calculator, pencil and ruler
- **Data Booklet** supplied by WJEC

INSTRUCTIONS TO CANDIDATES

Use black ink or black ball-point pen. Do not use gel pen or correction fluid.

You may use a pencil for graphs and diagrams only.

Write your name, centre number and candidate number in the spaces at the top of this page.

Section A Answer **all** questions.

Section B Answer **all** questions.

Write your answers in the spaces provided in this booklet. If you run out of space, use the additional page(s) at the back of the booklet, taking care to number the question(s) correctly.

INFORMATION FOR CANDIDATES

The number of marks is given in brackets at the end of each question or part-question.

The maximum mark for this paper is 80.

Your answers must be relevant and must make full use of the information given to be awarded full marks for a question.

The assessment of the quality of extended response (QER) will take place in **Q10(c)(ii)**.



JUN241410U30101

SECTION A

Answer **all** questions.

1. (a) Write an equation, including state symbols, that represents the enthalpy change of lattice formation of solid magnesium chloride. [1]



- (b) Complete the following equation: [1]

standard enthalpy
change of solution

=

*standard enthalpy change
of hydration (of the ions)*

– standard enthalpy
change of lattice formation

2. When chlorine is bubbled into cold sodium hydroxide solution, the following reaction occurs.



- (a) Show that this is a disproportionation reaction. [2]

Oxidation state of Cl in Cl₂ is 0; in NaCl it is -1 (reduced) and in NaOCl it is +1 (oxidised). The same element, chlorine, is both oxidised and reduced, so it is disproportionation.

- (b) Give **one** common use for sodium chlorate(I). [1]

bleach



3. Calcium hydroxide, $\text{Ca}(\text{OH})_2$, is a strong base that is sparingly soluble in water. A saturated solution of calcium hydroxide has a concentration of $2.34 \times 10^{-2} \text{ mol dm}^{-3}$ at 298 K.

Find the pH of this solution.

[2]

$$[\text{OH}^-] = 2 \times 2.34 \times 10^{-2} = 4.68 \times 10^{-2} \text{ mol dm}^{-3}$$

$$[\text{H}^+] = K_w \div [\text{OH}^-] = 1.00 \times 10^{-14} \div 4.68 \times 10^{-2}$$

$$[\text{H}^+] = 2.14 \times 10^{-13} \text{ mol dm}^{-3}$$

$$\text{pH} = -\log(2.14 \times 10^{-13}) = 12.7$$

pH = 12.7

4. One proposal to reduce the carbon dioxide emissions from public transport is to replace diesel trains with trains powered by hydrogen fuel cells. Give **one other** advantage of using hydrogen fuel cells. [1]

Fuel cells release the energy from the fuel more efficiently (a higher efficiency than burning diesel in an engine).

5. Bismuth is the heaviest naturally occurring element in Group 5. Suggest which oxidation state of bismuth would be most stable. Give a reason for your answer. [2]

+3. The inert pair effect increases down the group: the $6s^2$ electrons are not used in bonding, so the oxidation state of (group number - 2), +3, is more stable than +5.



SECTION B

Answer **all** questions.

6. Ammonia is a versatile compound produced on a large scale. It is used in fertilisers and cleaning products and to make a range of other nitrogen-containing compounds.

(a) Ammonia is an example of a weak base.

- (i) State what is meant by a base and explain why the ammonia molecule is able to act as a base. [2]

A base is a proton (H^+) acceptor. The N atom in NH_3 has a lone pair of electrons, which can form a coordinate (dative) bond with an H^+ ion.

(ii) Ammonia can be used as part of a mixture that forms a buffer.

- I. Suggest a compound that could be added to ammonia solution to form a buffer. [1]

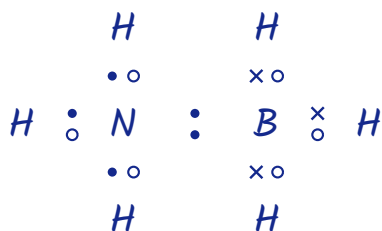
ammonium chloride, NH_4Cl

II. Suggest a use for a buffer. [1]

Using enzymes: keeping the pH constant for an enzyme reaction.

(b) Ammonia can form a compound with borane (BH_3).

- (i) Draw a dot and cross diagram of the compound formed. [1]



$\bullet$ = N electron

$\times$ = B electron

$\circ$ = H electron

*N lone pair forms
the $N \rightarrow B$ bond*



- (ii) Boron-nitrogen bonds are also present in boron nitride. One form of boron nitride (hexagonal boron nitride) is sometimes called white graphite.

Complete the table below giving similarities and differences in the structure and bonding of hexagonal boron nitride and graphite. One similarity has been included for you. [2]

	Similarity between hexagonal boron nitride and graphite	Difference between hexagonal boron nitride and graphite
Structure	<i>layers of hexagonal sheets</i> <i>(rings) present in both</i>	<i>layers in register in BN (B atoms directly above N atoms in the adjacent layer); layers out of register in graphite</i>
Bonding	each atom is bonded by covalent bonds to three others in both graphite and BN	<i>graphite has delocalised electrons, BN does not (localised on N); bonds are polar in BN but non-polar in graphite</i>

- (c) Nitrogen forms a range of compounds such as NH_3 and NF_3 which have three covalent bonds. Phosphorus can form compounds such as PF_3 which has three covalent bonds and PF_5 which has five covalent bonds.

Explain this difference in the chemistry of nitrogen and phosphorus. [2]

Phosphorus (period 3) has available, accessible d-orbitals, so it can expand its octet and use all five outer electrons to form five bonds. Nitrogen (period 2) has no available d-orbitals, so it cannot expand its octet and forms only three covalent bonds.



7. One method of studying the rates of chemical reactions is to use a clock reaction.

One clock reaction involves the reaction between iodate(V) ions, IO_3^- , and hydrogensulfate(IV) ions, HSO_3^- .

A student is given the following three solutions:

- sodium hydrogensulfate(IV) solution containing $\text{HSO}_3^-(\text{aq})$
- potassium iodate(V) solution containing $\text{IO}_3^-(\text{aq})$
- starch solution

He is also provided with deionised water.

The student combines the following volumes of solutions and measures the time taken for the colour to change.

Volume of HSO_3^- / cm^3	Volume of IO_3^- / cm^3	Volume of starch / cm^3	Volume of deionised water / cm^3	Time / s
10	10	5	25	164
10	20	5	15	82
10	30	5	5	55

- (a) Complete the table to show the volume of deionised water that should be used in the final experiment. Give a reason for the value you have chosen. [1]

5 cm^3 ; to keep the total volume constant (50 cm^3) so the concentration of HSO_3^- stays constant.

- (b) Find the order of reaction with respect to iodate(V) ions. Explain how you reached your conclusion. [2]

First order with respect to IO_3^- .

When $[\text{IO}_3^-]$ doubles the time taken halves (164 s to 82 s), so the rate doubles; tripling it gives 55 s, a third of the time.

- (c) The reaction is fourth order overall. Suggest a rate equation for the reaction. [2]

rate = $k[\text{IO}_3^-][\text{HSO}_3^-]^3$

- (d) Suggest a rate determining step for the reaction. [1]

$\text{IO}_3^- + 3\text{HSO}_3^- \rightarrow \text{I}^- + 3\text{HSO}_4^-$



- (e) Addition of lead(II) ions to the mixture after the reaction produces a mixture of products including lead(II) iodide and lead(II) iodate(V).

- (i) Give the colour of lead(II) iodide. [1]

bright (canary) yellow

- (ii) A sample of lead(II) iodate(V) decomposes on heating to give a mixture of solids and a mixture of oxygen gas and iodine vapour.

At a temperature of 200 °C, the volume of gas is 220 cm³, but when cooled to 20 °C the iodine solidifies and the volume of gas is 110 cm³.

Calculate the number of moles of gas at these two temperatures and hence find the percentage of iodine molecules in the original gas mixture. [3]

[All volumes are measured at a pressure of 1.01×10^5 Pa]

$$n = pV \div RT$$

$$200\text{ }^{\circ}\text{C} = 473\text{ K: } n = (1.01 \times 10^5 \times 220 \times 10^{-6}) \div (8.31 \times 473)$$

$$= 22.22 \div 3930.6 = 5.653 \times 10^{-3} \text{ mol } (O_2 + I_2)$$

$$20\text{ }^{\circ}\text{C} = 293\text{ K: } n = (1.01 \times 10^5 \times 110 \times 10^{-6}) \div (8.31 \times 293)$$

$$= 11.11 \div 2434.8 = 4.562 \times 10^{-3} \text{ mol } (O_2 \text{ only})$$

$$\text{moles } I_2 = 5.653 \times 10^{-3} - 4.562 \times 10^{-3} = 1.091 \times 10^{-3} \text{ mol}$$

$$\% I_2 = (1.091 \times 10^{-3} \div 5.653 \times 10^{-3}) \times 100 = 19.3\%$$

Number of moles of gas at 200 °C = 5.653×10^{-3} mol

Number of moles of gas at 20 °C = 4.562×10^{-3} mol

Percentage of iodine molecules = 19.3 %



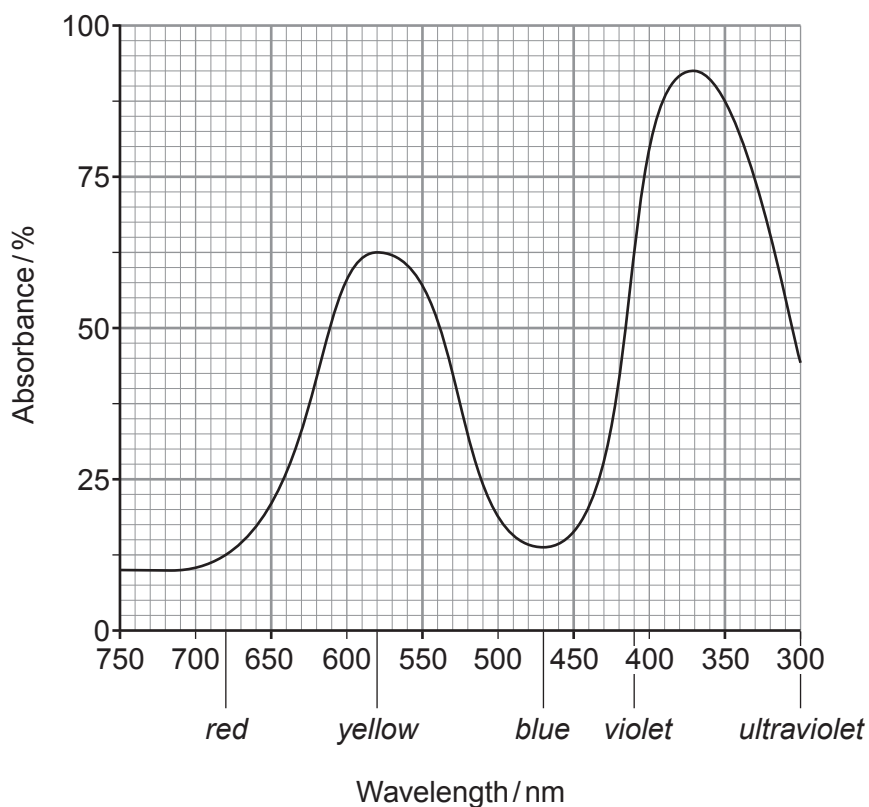
8. Vanadium is a transition element that forms compounds with a wide range of oxidation states. Its compounds are also used in a number of catalysts.

(a) State why transition elements can form a range of oxidation states in their compounds.

[1]

The 3d and 4s electrons have similar energies, so successive ionisation energies are similar and different numbers of electrons can be lost.

- (b) The visible spectrum of an aqueous solution of a compound containing V^{3+} ions is shown below.



- (i) Suggest the colour expected for aqueous solutions of V^{3+} compounds.

Give a reason for your answer.

[2]

Blue. Red/yellow (and violet) light is absorbed; the light that is not absorbed (transmitted), around 450 to 500 nm, is in the blue region of the spectrum, so blue is seen.

- (ii) In aqueous solutions, most vanadium ions form complexes with water acting as a ligand.

I. Give the meaning of the term ligand.

[1]

A molecule or ion with a lone pair of electrons that can form a coordinate bond to a transition metal ion.

II. Describe the bonding that occurs between the ligand and the transition metal ion.

[2]

A lone pair of electrons on the ligand (on O in H_2O) is donated into an empty orbital of the metal ion; both electrons of the shared pair come from the ligand; a coordinate bond.

- (c) VO_2^+ ions can be reduced to VO^{2+} ions by iodide ions as shown below.



- (i) Write the half-equation for the oxidation of iodide ions.

[1]



- (ii) Write the half-equation for the reduction of VO_2^+ ions in acid solution.

[1]



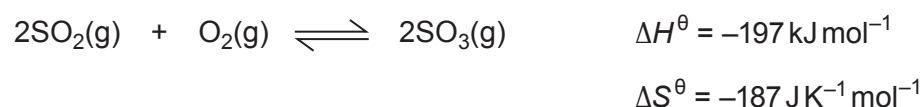
(d) Vanadium(V) oxide, V_2O_5 , is used as a heterogeneous catalyst in the contact process.

(i) State what is meant by the term heterogeneous in this context. [1]

The catalyst is in a different physical state (phase) from the reactants:

solid V_2O_5 , gaseous reactants.

(ii) This catalyst is used to catalyse the oxidation of sulfur dioxide.



I. The standard enthalpy change of formation of sulfur dioxide is -297 kJ mol^{-1} . Calculate the standard enthalpy change of formation of sulfur trioxide, SO_3 . [2]

$$\Delta H = \Sigma \Delta_f H(\text{products}) - \Sigma \Delta_f H(\text{reactants})$$

$$-197 = 2\Delta_f H(SO_3) - 2(-297)$$

$$2\Delta_f H(SO_3) = -197 - 594 = -791$$

$$\Delta_f H^\theta = -395.5 \text{ kJ mol}^{-1}$$

II. A student calculates the value at which $\Delta G = 0$ and states that this is the minimum temperature needed for the reaction to occur.

Find the temperature at which $\Delta G = 0$ and state, giving a reason, whether this is the **minimum** temperature needed for the reaction to occur. [3]

$$\Delta G = 0 \text{ when } T = \Delta H \div \Delta S$$

$$T = -197\,000 \div -187 = 1053 \text{ K}$$

$$\text{Temperature} = 1053 \text{ K}$$

The student is incorrect. ΔH and ΔS are both negative, so $\Delta G =$

$\Delta H - T\Delta S$ is negative below 1053 K; the reaction is feasible at

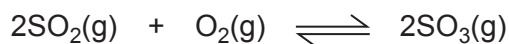
temperatures lower than this, so 1053 K is the maximum

temperature for feasibility, not the minimum.



- III. Samples of 0.040 mol of SO_2 and 0.040 mol of O_2 are placed in a sealed vessel of volume 2.00 dm^3 . The reaction is allowed to come to equilibrium, giving a concentration of $0.014 \text{ mol dm}^{-3}$ of SO_3 .

Calculate the value of the equilibrium constant, K_c , under these conditions, giving its unit. [4]



$$\text{initial } [\text{SO}_2] = [\text{O}_2] = 0.040 \div 2.00 = 0.020 \text{ mol dm}^{-3}$$

$$\text{SO}_3 \text{ formed} = 0.014, \text{ so } \text{SO}_2 \text{ used} = 0.014 \text{ and } \text{O}_2 \text{ used} = 0.007$$

$$\text{equilibrium: } [\text{SO}_2] = 0.020 - 0.014 = 0.006 \text{ mol dm}^{-3}$$

$$[\text{O}_2] = 0.020 - 0.007 = 0.013 \text{ mol dm}^{-3}$$

$$K_c = [\text{SO}_3]^2 \div ([\text{SO}_2]^2 [\text{O}_2])$$

$$= (0.014)^2 \div ((0.006)^2 \times 0.013)$$

$$= 1.96 \times 10^{-4} \div 4.68 \times 10^{-7} = 419$$

$$K_c = 419$$

$$\text{Unit } \text{mol}^{-1} \text{ dm}^3$$



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9. (a) Addition of a solution containing $\text{Ag}^+(\text{aq})$ to solutions containing halide ions such as chloride, bromide or iodide causes precipitates to form.

- (i) Give the colours of the precipitates formed with chloride, bromide and iodide ions. [1]

chloride ions white

bromide ions cream

iodide ions yellow

- (ii) The standard enthalpy change of formation of $\text{AgCl}(\text{s})$ is $-127.1 \text{ kJ mol}^{-1}$.

Use the data below to calculate the standard enthalpy change of formation of $\text{AgBr}(\text{s})$ and hence show which of the two silver halides is more stable with respect to its elements. You **must** show your working. [4]

Reaction	Standard enthalpy change, $\Delta H^\theta / \text{kJ mol}^{-1}$
$\text{Ag}(\text{s}) \longrightarrow \text{Ag}^+(\text{aq}) + \text{e}^-$	+105.6
$\text{Br}_2(\text{l}) + 2\text{e}^- \longrightarrow 2\text{Br}^-(\text{aq})$	-243.1
$\text{AgBr}(\text{s}) \longrightarrow \text{Ag}^+(\text{aq}) + \text{Br}^-(\text{aq})$	+84.4



$$\Delta_f H(\text{AgBr}) = 105.6 + \frac{1}{2}(-243.1) + (-84.4)$$

$$= 105.6 - 121.55 - 84.4 = -100.35 \text{ kJ mol}^{-1}$$

($\text{AgBr} \rightarrow \text{Ag}^+ + \text{Br}^-$ is +84.4, so the reverse is -84.4)

$\Delta_f H(\text{AgCl}) = -127.1 \text{ kJ mol}^{-1}$ is more negative than $\Delta_f H(\text{AgBr}) = -100.35 \text{ kJ mol}^{-1}$, so AgCl is more stable with respect to its elements than AgBr .



- (b) The standard electrode potential for the half-equation below is +0.80 V.



- (i) When a piece of copper metal is placed in a solution of silver nitrate, a displacement reaction occurs.

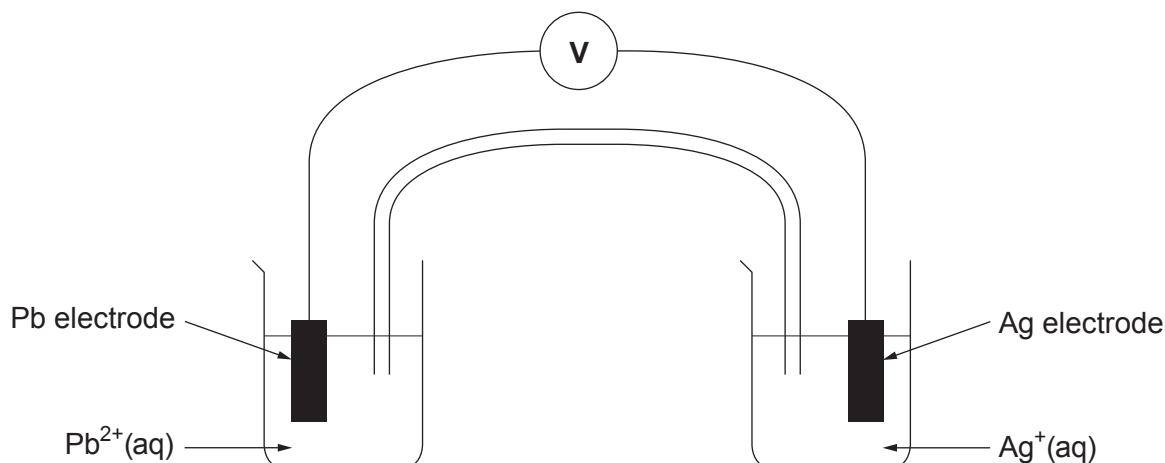
I. Write the **ionic** equation for this displacement reaction. [1]



II. State what information this provides regarding the standard electrode potential for the $\text{Cu}(\text{s})|\text{Cu}^{2+}(\text{aq})$ half-cell. Give a reason for your answer. [2]

$E^\ominus$ for $\text{Cu}^{2+}|\text{Cu}$ must be less positive (more negative) than +0.80 V. The reaction is feasible, so the emf must be positive: the reduction half-cell (Ag) is more positive than the oxidation half-cell (Cu).

- (ii) The apparatus below was assembled with the $\text{Pb}(\text{s})|\text{Pb}^{2+}(\text{aq})$ half-cell under standard conditions connected to the $\text{Ag}(\text{s})|\text{Ag}^+(\text{aq})$ half-cell. The silver is the positive electrode.



The value on the high-resistance voltmeter was recorded with different concentrations of silver ions used in the $\text{Ag(s)}|\text{Ag}^+(\text{aq})$ half-cell. The results are shown below.

Concentration of $\text{Ag}^+(\text{aq})$ / mol dm^{-3}	Value recorded on high-resistance voltmeter / V
1.0	0.93
0.1	0.87
0.01	0.81
0.001	0.75
0.0001	0.69

- I. Calculate the value of the standard electrode potential for the $\text{Pb(s)}|\text{Pb}^{2+}(\text{aq})$ half-cell.

[2]

standard conditions: $[\text{Ag}^+] = 1.0 \text{ mol dm}^{-3}$, $\text{emf} = 0.93 \text{ V}$

$\text{emf} = E^\ominus(\text{Ag}) - E^\ominus(\text{Pb})$: $0.93 = 0.80 - E^\ominus(\text{Pb})$

$E^\ominus(\text{Pb}) = 0.80 - 0.93 = -0.13 \text{ V}$

$E^\ominus = \underline{\underline{-0.13}} \text{ V}$

- II. Explain why the value recorded on the voltmeter becomes less positive as the concentration of silver ions decreases.

[3]

As $[\text{Ag}^+]$ decreases, the equilibrium $\text{Ag}^+ + \text{e}^- \rightleftharpoons \text{Ag}$ shifts to the left, so

E for the Ag half-cell becomes less positive. The emf of the cell = $E_{\text{Ag}} -$

E_{Pb} , and E_{Pb} is unchanged, so the value recorded becomes less positive.



10. Abandoned metal mines in Wales contribute to the pollution of some rivers and streams. The mines on Parys mountain on Anglesey have led the Afon Goch stream to become acidic and carry a range of metal ions.

- (a) The pH of the initial stretch of Afon Goch was found to be 3.24. Calculate the concentration of H^+ ions in this water.

[2]

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-3.24}$$

$$[\text{H}^+] = 5.75 \times 10^{-4} \text{ mol dm}^{-3}$$

- (b) The three main metals contaminating the water of Afon Goch are iron, copper and zinc. A research group wishes to find the concentrations of all three metals in a water sample from the river.

An initial reduction step converts all the iron ions present to Fe^{2+} and all the copper ions present to copper metal. The **amphoteric** zinc ions are not affected. This produces test water **A**.

Two students are asked to find the concentration of Fe^{2+} ions in test water **A**.

- (i) The reduction step produced 4.55 mg of copper metal from a 250 cm^3 sample of river water.

Suggest how this could be separated from test water **A**.

[1]

filtration



- (ii) One student attempts to find the iron content in 25.0 cm^3 of test water **A** by gravimetric analysis. He adds excess aqueous sodium hydroxide, filters and then heats the precipitate in an oxygen-rich atmosphere.

- I. Give **two** reasons why the aqueous sodium hydroxide that is added must be in excess. [2]

Excess NaOH dissolves the amphoteric zinc hydroxide precipitate (forming $[\text{Zn}(\text{OH})_4]^{2-}$), so zinc does not add to the mass. Excess also ensures all of the iron ions are precipitated.

- II. The experiment produced 14.5 mg of Fe_2O_3 . Use this value to show that the concentration of Fe^{2+} ions in test water **A** is $7.27 \times 10^{-3} \text{ mol dm}^{-3}$. [2]

$$M_r(\text{Fe}_2\text{O}_3) = 159.6$$

$$\text{moles Fe}_2\text{O}_3 = 14.5 \times 10^{-3} \div 159.6 = 9.085 \times 10^{-5} \text{ mol}$$

$$\text{moles Fe}^{2+} = 2 \times 9.085 \times 10^{-5} = 1.817 \times 10^{-4} \text{ mol}$$

$$[\text{Fe}^{2+}] = 1.817 \times 10^{-4} \div 0.0250 = 7.27 \times 10^{-3} \text{ mol dm}^{-3}$$

- III. The total error when measuring the mass of Fe_2O_3 by difference was 0.2 mg .

Find the percentage error in this measurement. [1]

$$(0.2 \div 14.5) \times 100$$

Percentage error = 1.4 %



- (iii) The second student titrated 25.0 cm³ samples of test water **A** against a solution of acidified potassium manganate(VII) of concentration $2.10 \times 10^{-3} \text{ mol dm}^{-3}$.



The mean volume of acidified potassium manganate(VII) solution required for reaction was 18.10 cm³ and the percentage error in the titration was 0.8%.

- I. State what observation would be made at the end-point of the titration. [1]

The solution turns a pale pink colour that stays (permanent pink).

- II. Calculate the concentration of Fe²⁺ ions in the sample of test water **A**. [2]

$$\text{moles MnO}_4^- = 2.10 \times 10^{-3} \times 18.10 \div 1000 = 3.80 \times 10^{-5} \text{ mol}$$

$$\text{moles Fe}^{2+} = 5 \times 3.80 \times 10^{-5} = 1.90 \times 10^{-4} \text{ mol}$$

$$[\text{Fe}^{2+}] = 1.90 \times 10^{-4} \div 0.0250 = 7.60 \times 10^{-3} \text{ mol dm}^{-3}$$

$$\text{Concentration} = 7.60 \times 10^{-3} \text{ mol dm}^{-3}$$

- (iv) The two students compared their results and decided that they are in agreement.

Calculate the percentage difference between the concentrations of Fe²⁺ ions found by the two students in parts (ii) and (iii) and hence show whether the students' decision is valid. [2]

$$\text{difference} = 7.60 \times 10^{-3} - 7.27 \times 10^{-3} = 0.33 \times 10^{-3}$$

$$\% \text{ difference} = (0.33 \div 7.27) \times 100 = 4.5\%$$

$$\text{Percentage difference} = 4.5\%$$

Not valid: the total percentage error is 1.4% + 0.8% = 2.2%, which is smaller than the 4.5% difference, so the results do not agree.



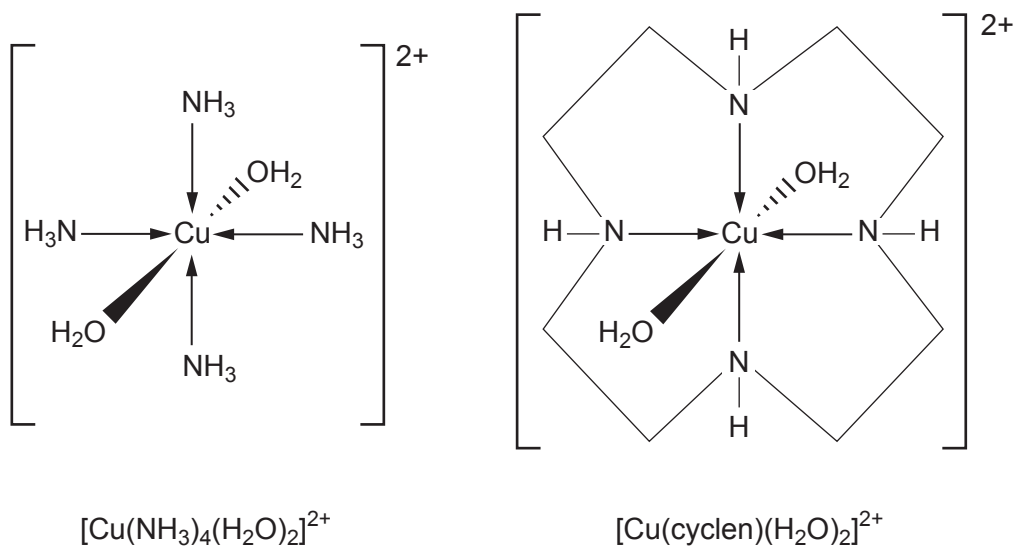
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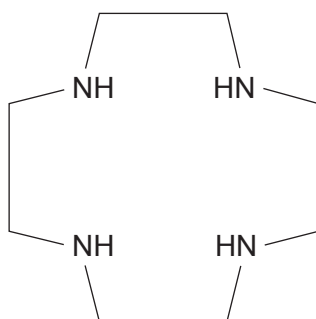
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- (c) One method of determining the concentration of copper ions present in a solution is to form a complex ion with a characteristic colour and measure the concentration of this using colorimetry. Two such copper complexes are $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ and $[\text{Cu}(\text{cyclen})(\text{H}_2\text{O})_2]^{2+}$.



Cyclen is a cyclic tetradentate ligand with four nitrogen atoms that can bond to a transition metal ion. The structure of the cyclen ligand is shown below.



cyclen

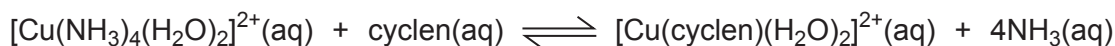
- (i) Give the colour of the $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ complex.

[1]

deep (royal) blue



- (ii) An equilibrium exists between $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ and $[\text{Cu}(\text{cyclen})(\text{H}_2\text{O})_2]^{2+}$ in aqueous solution.



$$\Delta H^\theta = +1.2 \text{ kJ mol}^{-1}$$

Measurements show that the strengths of the Cu—N bonds in $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ and $[\text{Cu}(\text{cyclen})(\text{H}_2\text{O})_2]^{2+}$ are similar and that the Gibbs free energy change for the equilibrium is negative.

Discuss the information given above. Your answer should:

- explain why the enthalpy change for the reaction is very small
 - suggest why the Gibbs free energy change for this equilibrium would be negative
 - state and explain the effect of a negative Gibbs free energy change on the value of K_c for an equilibrium
- [6 QER]

Enthalpy: in the forward reaction four Cu—N bonds to ammonia are broken and four Cu—N bonds to cyclen are formed. As these bonds have similar strengths, the energy needed to break bonds is about the same as the energy released when the new bonds form, so ΔH is very small ($+1.2 \text{ kJ mol}^{-1}$).

Entropy: $\Delta G = \Delta H - T\Delta S$. As ΔH is small and positive, ΔG can only be negative if ΔS is positive so that $T\Delta S$ is greater than ΔH . There are 2 particles on the left (complex + cyclen) but 5 on the right (complex + 4NH_3), because one tetradentate cyclen replaces four monodentate NH_3 ligands. The four free ammonia molecules have more entropy than one cyclen molecule, so ΔS is positive and ΔG is negative.

Effect on K_c : $\Delta G = -RT \ln K_c$, so a negative ΔG makes $\ln K_c$ positive and K_c greater than 1. The equilibrium lies to the right: products are more likely to be present than reactants, and the higher concentration of products gives a large value of K_c .

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