

Mark scheme – Redox and Electrode Potentials

Question	Answer/Indicative content	Marks	Guidance
1 a	<p style="text-align: center;">Comparison of E values</p> <p><i>E</i> of redox system 4 (MnO₄⁻/Mn²⁺) is more positive/less negative than <i>E</i> of redox systems 2 (HCOOH/HCHO) OR 1 (CO₂/HCOOH) ✓</p> <p style="text-align: center;">Equilibrium shift related to E values</p> <p>More negative/less positive/system 2 (HCOOH/HCHO) OR system 1 (CO₂/HCOOH) shifts left OR Less negative/more positive/system 4 (MnO₄⁻/Mn²⁺) shifts right ✓</p> <p>• 2 and 4 2MnO₄⁻ + 5HCHO + 6H⁺ → 2Mn²⁺ + 5HCOOH + 3H₂O ✓</p> <p>• 1 and 4 2MnO₄⁻ + 5HCOOH + 6H⁺ → 2Mn²⁺ + 5CO₂ + 8H₂O ✓</p>	4 (AO 3.1× 2)	<p>IGNORE higher/lower</p> <p>ALLOW Overall $E_{\text{reaction}} = (+)1.54\text{V}$ OR $(+)1.62\text{V}$</p> <p>For 'shifts left', ALLOW 'is oxidised' OR 'electrons are lost' OR 'reducing agent'</p> <p>For 'shifts right', ALLOW 'is reduced' OR 'electrons are gained' OR 'oxidising agent'</p> <p>IGNORE state symbols ALLOW multiples DO NOT ALLOW un-cancelled species, e.g. H⁺, on both sides ALLOW for 1 mark two balanced equations with uncancelled species. ALLOW combined equation for 2 marks: 4MnO₄⁻ + 5HCHO + 12H⁺ → 4Mn²⁺ + 5CO₂ + 11H₂O</p> <p>Examiner's Comments</p> <p>Higher-attaining candidates described two oxidations starting from HCHO to end up with CO₂. Many candidates used the data correctly but stopped at the first oxidation to form HCOOH.</p> <p>Lower-attaining candidates did not state that the direction of reaction of redox equilibria is dependent on the relative negativity/positivity of the standard electrode potentials. Some candidates are still using higher/lower to compare the <i>E</i> cell values, and should be encouraged to instead use the phrasing 'more positive' or 'more negative'. Many candidates wrote a correct first equation, although some did not cancel down the H⁺ and/or water.</p>
b	$2\text{H}^+ + \frac{1}{2}\text{O}_2 + 2\text{e}^- \rightarrow \text{H}_2\text{O}$ ✓ $1.34 + (-0.11) = (+)1.23$ (V) ✓	2 (AO 2.6) (AO 2.2×1)	<p>IGNORE state symbols ALLOW multiples</p>
Total		6	

5.2.3 Redox and Electrode Potentials

2		$n(\text{MnO}_4^-) \text{ in titration}$ $= 0.01 \times \frac{12.6}{1000} = 1.26 \times 10^{-4} \checkmark$ $n(\text{SO}_3^{2-}) \text{ in } 25.0 \text{ cm}^3$ $= 1.26 \times 10^{-4} \times 2.5 = 3.15 \times 10^{-4} \text{ (mol)} \checkmark$ $n(\text{SO}_3^{2-}) \text{ in } 250 \text{ cm}^3$ $= 10 \times 3.15 \times 10^{-4} = 3.15 \times 10^{-3} \text{ (mol)} \checkmark$ <p>mass Na_2SO_3 in 525 g meat</p> $= 3.15 \times 10^{-3} \times 126.1 = 0.397 \text{ (g)} \checkmark$ <p>mass Na_2SO_3 in 1 kg of meat</p> $= 0.397215 \times \frac{1000}{525} = 0.7566 \text{ g OR } 756.6 \text{ mg}$ <p>AND less than the maximum permitted level OR AW \checkmark</p>	<p>5 (AO 1.2×1)</p> <p>(AO 2.8×3)</p> <p>(AO 3.2×1)</p>	<p>ALLOW 3 SF or more throughout ALLOW ECF throughout</p> <p>Calculator = 0.397215 g</p> <p>ALLOW within range: 756 to 757 mg</p> <p>ALLOW 0.397 g < 0.446 g per 525 g meat.</p> <p><u>Examiner's Comments</u></p> <p>This question asked the candidate to determine the amount of sodium sulfite in food. Many candidates gained full marks. Most candidates calculated the number of moles of MnO_4^- and SO_3^{2-}. Some candidates calculated the mass in 525 g of meat, although some used the wrong Mr, e.g. 80, for the sulfite ion. The lower ability candidates did not process the scaling to 1 Kg.</p>
		Total	5	
3		<p>Circuit</p> <p>Complete circuit AND voltmeter AND salt bridge linking two half-cells \checkmark</p> <p>Half cells</p> <p>Ag AND Ag^+ AND 1 mol dm^{-3} solution \checkmark</p> <p>Pt AND H^+ AND MnO_4^- AND Mn^{2+} AND 1 mol dm^{-3} /equimolar solution \checkmark</p>	<p>3 (AO 3.4× 1)</p> <p>(AO 1.2× 1)</p> <p>(AO 1.2×1)</p>	<p>Voltmeter must be shown AND salt bridge must be labelled</p> <p>ALLOW small gaps in circuit</p> <p>If species in BOTH half cells are correct but concentration of 1 mol dm^{-3} omitted, ALLOW 1 mark for BOTH half cells.</p> <p>ALLOW acidified as an alternative for H^+</p> <p>IGNORE stated pressure <i>Not relevant here as no gas</i></p> <p><u>Examiner's Comments</u></p> <p>Most candidates drew a circuit containing a voltmeter and the silver half cell but very few candidates included the H^+ in the $\text{MnO}_4^-/\text{Mn}^{2+}$ cell.</p>
		Total	3	
4	i		<p>3 (AO 2.6×3)</p>	<p>All 3 marks are independent.</p> <p>IGNORE charges/oxidation numbers shown around overall equation. <i>Treat as rough working</i></p> <p>ALLOW overall equation shown with some or all ions that are present</p>

5.2.3 Redox and Electrode Potentials

		<p>Overall equation AND state symbols:</p> $\begin{array}{l} \text{M(s)} + \\ 2\text{HCl(aq)} \\ \rightarrow \\ \text{MCl}_2\text{(aq)} \\ + \text{H}_2\text{(g)} \checkmark \end{array}$ <p>STATE SYMBOLS required in overall equation ONLY</p> <p>Half equations:</p> <p>Oxidation $\text{M} \rightarrow \text{M}^{2+} + 2\text{e}^- \checkmark$</p> <p>Reduction $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$ OR $\text{H}^+ + \text{e}^- \rightarrow \frac{1}{2}\text{H}_2 \checkmark$</p>	<p>e.g. (with state symbols)</p> $\text{M} + 2\text{H}^+ \rightarrow \text{M}^{2+} + \text{H}_2$ $\text{M} + 2\text{HCl} \rightarrow \text{M}^{2+} + 2\text{Cl}^- + \text{H}_2$ $\text{M} + 2\text{H}^+ + 2\text{Cl}^- \rightarrow \text{M}^{2+} + 2\text{Cl}^- + \text{H}_2$ <p>In half equations, IGNORE state symbols even is wrong BUT half equations MUST only have species that change.</p> <p>For charges on half equations, ALLOW M^{+2} for M^{2+} OR H^{+1} for H^+ ALLOW $\text{M} - 2\text{e}^- \rightarrow \text{M}^{2+}$</p> <p>If BOTH half equations are correct but shown with oxidation and reduction the wrong way around, award 1 mark from the 2 marks for half equations</p> <p><u>Examiner's Comments</u></p> <p>This question required candidates to write an overall equation and half equations for oxidation and reduction. Many candidates made errors within one or more equations. The overall equation was often written without state symbols, despite the question instruction 'with state symbols'. The oxidation half equation was more likely to be correct than the reduction half equation, which often used Cl instead of H^+. When H^+ was used, the half equation was often unbalanced or electrons had been omitted.</p> <p>It is recommended that candidates carefully use the chemical information in the question.</p>
	ii	<p>Bubbles/effervescence/fizzing stops \checkmark</p> <p>M/metal/solid has disappeared/dissolved \checkmark</p>	<p>Responses must imply that all fizzing has stopped and that all the solid has dissolved i.e. 'metal disappears' is not quite enough. 'All the metal disappears' is enough</p> <p>IGNORE constant mass IGNORE no increase in temperature</p> <p>2 (AO 3.3×2)</p> <p><u>Examiner's Comments</u></p> <p>Most candidates identified that all the metal would have reacted when it had all disappeared and that gas bubbles from the reaction would have stopped. Some responses did not emphasise that these observations would have stopped and this prevented credit being given.</p>
	iii	<p>$\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \checkmark$</p>	<p>ALLOW multiples e.g. $2\text{H}^+ + 2\text{OH}^- \rightarrow 2\text{H}_2\text{O}$</p> <p>1 (AO 2.5)</p> <p>IGNORE state symbols, even if wrong</p> <p><u>Examiner's Comments</u></p>

5.2.3 Redox and Electrode Potentials

		<p>The ionic equation for neutralisation of an acid with an alkali was well known and this question was answered correctly by most candidates.</p>
<p>i v</p>	<p>Mean titre 1 mark $= \frac{(27.30 + 27.20)}{2} = 27.25 \text{ (cm}^3\text{)} \checkmark$</p> <p>Analysis of results 5 marks $n(\text{NaOH}) = 27.25 \times \frac{0.320}{1000} = 8.72 \times 10^{-3} \text{ (mol)} \checkmark$ $n(\text{HCl}) \text{ in } 25.0 \text{ cm}^3 = n(\text{NaOH})$ $n(\text{HCl}) \text{ in } 250 \text{ cm}^3$ $= 8.72 \times 10^{-3} \times 10 = 8.72 \times 10^{-2} \text{ (mol)} \checkmark$</p> <p>$n(\text{HCl}) \text{ that reacted with M}$ $= 0.210 - 8.72 \times 10^{-2} = 0.1228 \text{ (mol)} \checkmark$</p> <p>$n(\text{M}) \text{ that reacted} = \frac{0.1228}{2} = 0.0614 \text{ (mol)} \checkmark$</p> <p>$A_r \text{ of M} = \frac{6.90}{0.0614} = 112.4 \text{ AND M} = \text{cadmium/Cd} \checkmark$</p> <p>COMMON ERRORS: Mean of 27.35 (use of all 3 titres) $\rightarrow 8.752 \times 10^{-3} \rightarrow 8.752 \times 10^{-2} \rightarrow 0.12248$ $\rightarrow 0.06124 \rightarrow 112.7 \text{ AND Cd: } 5 \text{ marks}$</p> <p>No +2 to obtain $n(\text{M})$ $\rightarrow 56.2 \text{ AND Fe (from } 27.25) \quad 5 \text{ marks}$ $\rightarrow 56.3 \text{ AND Fe (from } 27.35) \quad 4 \text{ marks}$</p> <p>No subtraction from 0.210 $A_r \text{ of M} = \frac{6.90}{0.0614} = 112.4 \text{ AND M} = \text{cadmium/Cd} \checkmark$ $\rightarrow 158.2 \text{ to } 158.3 \text{ AND Tb } 5 \text{ marks}$</p> <p>No $\times 10$ to obtain $n(\text{HCl})$ in 250 cm^3 5 marks $0.210 - 8.72 \times 10^{-3} = 0.20128 \text{ OR } 0.201$ $n(\text{M}) = 0.20128/2 = 0.10064$ $A_r = 6.90/0.10064 = 68.56 \rightarrow \text{Zn}$</p> <p>No $\times 10$ and no $\div 2$ 4 marks $0.210 - 8.72 \times 10^{-3} = 0.20128$ $A_r = 6.9/0.20128 = 34.28 \rightarrow \text{Ca}$</p> <p>Omitting initial titration calculation Zero marks $0.210/2 = 0.105 \rightarrow 6.9/0.105 = 65.71 \rightarrow \text{Zn}$</p>	<p>FULL ANNOTATIONS MUST BE USED <hr style="border-top: 1px dashed blue;"/></p> <p>Common error: Incorrect mean from all 3 titres = 27.35 cm^3</p> <p>Use ECF throughout Intermediate values for working to at least 3 SF.</p> <p>TAKE CARE: Value written down may be truncated calculator value. Depending on rounding, either can be credited.</p> <p>ALLOW 0.123 (mol) i.e. 3SF</p> <p>ALLOW 0.0615 (mol) IF 0.1228 rounded to 0.123</p> <p>6 ALLOW 112.2 from 0.0615 AND Cd</p> <p>(AO 2.8×5) ALLOW A_r to nearest whole number ALLOW ECF for metal closest to calculated A_r</p> <p>DO NOT ALLOW Ga OR Sc (Form 3+ ions only)</p> <p>Examiner's Comments</p> <p>Candidates were presented with information about a back titration, a technique that they would be unlikely to have encountered during their course. The question stem to (iv) suggested a three-step strategy. Many candidates followed this guidance and were credited with many of the available marks. Marks were given for a correct method (by error carried forward) even if there was an error or omission in the multi-step calculation. This emphasises the importance of clear working.</p> <p>(AO 3.2)</p> <p>Most candidates determined the correct mean titre of 27.25 cm^3. A few candidates did take the mean of all three titres rather than the closest. Most calculated that $8.72 \times 10^{-3} \text{ mol}$ of NaOH reacted with the same number of moles of HCl in the titration and then scaled up the HCl by a factor of 10 to $8.72 \times 10^{-2} \text{ mol}$ in the 250 cm^3 volumetric flask. These steps are standard for many titration calculations and gave a route to three of the six available marks. The more difficult back titration steps then followed and the higher-attaining candidates recognised the need to subtract this amount of HCl from the original amount of HCl</p>

5.2.3 Redox and Electrode Potentials

				<p>used to react with metal M. These candidates then divide this value by two to find the moles of M that reacted (from the 1 : 2 stoichiometry of M : HCl). The correct calculation then gave a relative atomic mass of M as 112 and its identity as cadmium. It was common for candidates to omit the division by two and to arrive at a relative atomic mass of 56 for iron. The mark scheme shows the variety of metals that candidates identified from their calculations, the errors made, and the error carried forward marks that resulted.</p> <p>Many lower-attaining candidates did not follow the 3 steps in the stem, using only the original amount of HCl and ignoring the titration. This approach was not credited with marks.</p> <p>A large range of marks was seen, and the question discriminated extremely well.</p>
		Total	12	
5		<p>TAKE CARE: Correct final answer of -52.3 OR -52.25 can be obtained from two cancelling errors:</p> <ul style="list-style-type: none"> • Use of 50 for energy released (no $\times 2$ of 50 for two solutions mixed) • No $\div 2$ in final step <p>-52.3 OR -52.25 would then be awarded 2 marks out of 4</p> <hr/> <p>Correctly calculates $n(\text{succinic acid})$</p> $= 0.400 \times \frac{50.0}{1000} = 0.02(00) \text{ (mol)} \checkmark$ <p>Energy released in J OR kJ</p> $= 100.00 \times 4.18 \times 5.0 = 2090 \text{ (J)}$ <p>OR 2.090 (kJ) \checkmark</p> <p>Energy released, in kJ or J, for formation of 2 mol H₂O</p> $\pm \frac{2090}{0.0200} = \pm 104500 \text{ (J)}$ <p>OR</p> $\pm \frac{2.090}{0.0200} = \pm 104.5 \text{ OR } \pm 105 \text{ (kJ)} \checkmark$ <p>$\Delta_{\text{neut}}H$ to 3 or more SF AND correct – sign</p> $= -\frac{104.5}{2} = -52.3 \text{ OR } -52.25 \text{ kJ mol}^{-1} \checkmark$	4	<p>ALLOW ECF throughout</p> <p>DO NOT ALLOW less than 3 SF IGNORE units</p> <hr/> <p>ALTERNATIVE METHOD</p> <p>$n(\text{succinic acid}) = 0.02(00) \text{ (mol)} \checkmark$</p> <p>Energy released = 2090 (J) OR 2.090 (kJ) \checkmark</p> <p>$n(\text{H}_2\text{O}) \text{ formed} = 2 \times 0.02(00) = 0.04(00) \text{ (mol)} \checkmark$</p> $\Delta_{\text{neut}}H = -\frac{2.090}{0.0400} = -52.3 \text{ OR } -52.25 \text{ kJ mol}^{-1} \checkmark$ <p><u>Examiner's Comments</u></p> <p>The direct determination of an enthalpy change is a common practical procedure.</p> <p>Most candidates were able to calculate the energy</p>

5.2.3 Redox and Electrode Potentials

change in the reaction as 2.09 kJ, with comparatively few not adding the two 50 cm³ volumes in their calculation. Common errors were using the mass or moles of succinic acid in the $mc\Delta T$ calculation or to use 278 K, rather than 5°C for ΔT .

Most candidates calculated that 0.02 mol of succinic acid had reacted.

The best responses recognised that the enthalpy change of neutralisation (i.e. formation of 1 mol of H₂O) was required and that 0.02 mol succinic acid produced $2 \times 0.02 = 0.04$ mol H₂O. 2.09 kJ was then scaled up to give the correct ΔH value of -52.25 kJ mol⁻¹. Example 1 follows this approach and each stage in the working can be followed easily.

Most candidates scaled up using 0.02 mol to obtain -104.6 kJ mol⁻¹ and then thought that this was the final answer.

Lower ability candidates often calculated the moles of succinic acid as 0.02 but also calculated the moles of NaOH as 0.05 mol. They often then went on to scale up their energy change by 0.05 or 0.07.

Many responses were disorganised and poorly presented with unsubstantiated numbers across the answer space. Imperfect calculations may be partially credited by applying error carried forward but this is only possible if sense can be made of the response.

Compare the responses in Exemplar 13 and Exemplar 14. In Exemplar 14, there is no labelling to show what each stage refers to. It looks as if the candidate has calculated the moles of NaOH rather than succinic acid and has used this in the subsequent stage. Error carried forward has been applied generously on this assumption.

Answer = -52.25 kJ mol⁻¹.

Exemplar 13

5.2.3 Redox and Electrode Potentials

			<p>mixture are the same as for water. $Q_t = 26.5 - 21.5 = 5$</p> $Q = mc\Delta T$ $= (50 + 50) \times 4.18 \times 5 = 2090 \text{ J} \checkmark$ $= 2.09 \text{ kJ}$ <p>moles of acid used = $\frac{50}{1000} \times 0.4 = 0.02 \text{ mol} \checkmark$</p> <p>moles of NaOH used = $\frac{50}{1000} \times 1 = 0.05 \text{ mol}$</p> <p>acid is the limiting reagent.</p> <p>moles of H_2O formed = $0.02 \times 2 = 0.04 \text{ mol} \checkmark$</p> $\Delta_{\text{comb}} H = \frac{-2.09}{0.04} = -52.25$ $\Delta_{\text{comb}} H = \dots -52.25 \dots \text{ kJ mol}^{-1} [4]$ <p>Exemplar 14</p> $q = mc\Delta T \quad 26.5 - 21.5$ $100 \times 4.18 \times 5 = 2090 \checkmark$ $\div 1000 = 2.09$ <p>moles of acid used</p> $\frac{50}{1000} \times 1 = 0.05 \quad \frac{2.09}{0.05} = 41.8 \checkmark$ $\Delta_{\text{comb}} H = \dots -41.8 \dots \text{ kJ mol}^{-1} [4]$
		Total	4
6		<p>$2\text{MnO}_4^- + 5\text{H}_2\text{O}_2 + 6\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{O}_2$</p> <p>Correctly balanced equation for $\text{MnO}_4^-/\text{H}_2\text{O}_2$ reaction but no cancelling of H^+ and/or $\text{e}^- \checkmark$</p> <p>Overall equation correct with all species cancelled \checkmark</p>	<p>2</p> <p>ALLOW multiples ALLOW \rightleftharpoons instead of \rightarrow sign</p> <p>ALLOW 1 mark for final equation with correct balancing numbers AND ONE small slip in a formula OR charge</p> <p>IGNORE annotations around equations, i.e. treat as rough working</p> <p>ALLOW 1 mark for: $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ <i>(H_2O_2 is acting as both reducing and oxidising agent)</i></p> <p>Examiner's Comments</p> <p>This part discriminated extremely well with many candidates obtaining either both marks or zero marks.</p> <p>Candidates needed to select the correct redox pairs, combine them and cancel H^+ and H_2O.</p> <p>Many otherwise correct equations did not receive credit as they contained basic balancing errors. As a final step in constructing an equation, candidates are recommended to check the formulae of their species and that the equation is balanced.</p>
		Total	2

5.2.3 Redox and Electrode Potentials

7	i	Colourless to yellow ✓	<p>IGNORE clear for colourless</p> <p>Examiner's Comments</p> <p>Candidates were expected to apply their knowledge of the colour change in a manganate(VII) titration to this novel situation. Most candidates incorrectly showed the inverse colour change of yellow → colourless.</p>
		<p>Mean titre</p> $= \frac{(23.15 + 23.25)}{2} = 23.2(0) \text{ (cm}^3\text{)} \checkmark$ <p>Analysis of results 5 marks</p> $n(\text{Ce}^{4+}) = 23.20 \times \frac{0.0500}{1000} = 1.16 \times 10^{-3} \text{ (mol)} \checkmark$ <p>ii $n(\text{(COOH)}_2) \text{ in } 25.0 \text{ cm}^3 = \frac{1.16 \times 10^{-3}}{2} = 5.8(0) \times 10^{-4} \text{ (mol)} \checkmark$</p> $n(\text{(COOH)}_2) \text{ in } 250 \text{ cm}^3$ $= 5.8(0) \times 10^{-4} \times 10 = 5.8(0) \times 10^{-3} \text{ (mol)} \checkmark$ <p>Mass $(\text{COOH})_2 = 5.8(0) \times 10^{-3} \times 90.0 = 0.522 \text{ g} \checkmark$</p> $\% \text{ oxalic acid} = \frac{0.522 \times 100}{82.68} = 0.631\% \checkmark$ <p>Percentage MUST be expressed to 3 SF</p>	<p>Common error: Incorrect mean from all 3 titres = 23.30 cm³</p> <p>Use ECF throughout Intermediate values for working to at least 3 SF. TAKE CARE as value written down may be truncated value stored in calculator. Depending on rounding, either can be credited.</p> <hr/> <p>COMMON ERRORS: Mean of 23.30 (use of all 3 titres) → 0.634%: 5 marks</p> <p>TAKE CARE for final answer of 0.63 seen.</p> <ul style="list-style-type: none"> No final mark as only 2 SF 0.63 may have been rounded from 0.631 (from correct mean) OR from 0.634 (using mean from all 3 titres) <p>Check back to mean titre.</p> <p>No +2 to <i>obtain</i> $n(\text{(COOH)}_2)$</p> <p>→ 1.26%: 5 marks from 23.20 → 1.27% 4 marks from 23.30</p> <p>Examiner's Comments</p> <p>Most candidates answered this unstructured titration calculation well, with almost all responses gaining some credit. Most coped well with the 1:2 reaction stoichiometry and the need to scale up to 250 cm³.</p> <p>Common errors included the following.</p> <ul style="list-style-type: none"> Taking the mean of all three titres (23.30 cm³) instead of the closest titres (23.20 cm³). Not considering the 1:2 stoichiometry. Not scaling up to 250 cm³. Giving the final answer to two rather than three significant figures; this was the most common error.

5.2.3 Redox and Electrode Potentials

				<p>The question required the final answer to be given to an appropriate number of significant figures. Many candidates seemed to be unaware that this reflects the least significant figures provided in the data, in this case three significant figures. Candidates are also advised to only round at the end of a multi-step calculation. Rounding of intermediate values introduces rounding errors in the final answer.</p> <p>The example shows a perfect response with each step clearly described, leading to the correct concentration of ethanedioic acid. Candidates should present their calculations clearly. If there is a mistake at any stage, this can be easily identified, allowing for error carried forward to be applied for any subsequent good method.</p> <p>Exemplar 6</p> <p>mean titre value = $\frac{23.29 + 23.15}{2} = 23.20$ ✓</p> <p>volume of $\text{Ce}(\text{SO}_4)_2 = 0.0232 \text{ dm}^3$ conc of $\text{Ce}(\text{SO}_4)_2 = 0.05 \text{ mol dm}^{-3}$ moles of $\text{Ce}(\text{SO}_4)_2 = 1.16 \times 10^{-3} \text{ mol}$</p> <p>25 cm³ of ethanedioic acid reacts with $1.16 \times 10^{-3} \text{ mol of Ce(SO}_4)_2$</p> <p>moles of ethanedioic acid in 25 cm³ = 5.8×10^{-4} moles of ethanedioic acid in 250 cm³ = 5.8×10^{-3}</p> <p>mass of ethanedioic acid extracted: mass = moles \times Mr = $5.8 \times 10^{-3} \times 90 = 0.522 \text{ g}$</p> <p>$\frac{0.522}{8.268} \times 100 = 0.631\%$</p> <p>percentage of ethanedioic acid = 0.631% ✓✓✓✓</p>					
		Total	7						
8	i	<p>Titres correct and ALL recorded to 2 decimal places</p> <table border="1" data-bbox="236 1693 823 1776"> <tr> <td>Titre: 24.00</td> <td>23.40</td> <td>23.75</td> <td>23.85 ✓</td> </tr> </table> <table border="1" data-bbox="236 1816 823 1868"> <tr> <td>mean titre = 23.80 (cm³) ✓</td> </tr> </table>	Titre: 24.00	23.40	23.75	23.85 ✓	mean titre = 23.80 (cm ³) ✓	2	<p>ALLOW 23.8 cm³</p> <p>Examiner's Comments</p> <p>It is clear candidates are not as experienced at filling in titration tables as might be expected. Every value in a titration table should be recorded to a second decimal place to an accuracy of $\pm 0.05 \text{ cm}^3$.</p> <p>The average titre should be calculated by</p>
Titre: 24.00	23.40	23.75	23.85 ✓						
mean titre = 23.80 (cm ³) ✓									

5.2.3 Redox and Electrode Potentials

				averaging concordant titres, i.e. those within 0.10 cm ³ of each other.
	ii	<p>Percentage uncertainty</p> $= \frac{0.05 \times 2}{23.40} \times 100 = 0.43 (\%) \checkmark$	1	<p>ALLOW ECF from incorrect subtraction in (i) or incorrect mean</p> <p>ALLOW 0.42% from titre values 2, 3 or 4 or mean titre or trial titre.</p> <p>2 DP required</p> <p><u>Examiner's Comments</u></p> <p>Candidates are unfamiliar with determination of percentage uncertainty. Marks were credited for any percentage uncertainty calculation correctly determined from any titre value, as many opted to choose the trial value as titre 1 or used an average titre.</p>
	iii	<p>Add starch (near the end point) \checkmark</p> <p>Blue to colourless \checkmark</p>	2	<p>ALLOW blue/black OR black OR purple for colour of mixture</p> <p>ALLOW blue colour disappears (to colourless)</p> <p>IGNORE 'clear'</p> <p>IGNORE 'colorimetry'</p> <p><u>Examiner's Comments</u></p> <p>Only the higher ability candidates realised starch needed to be added close to the end-point and this made the resulting colour change (blue-black to colourless) easier to see.</p> <p>The common error was to assume this was an acid-base titration and indicators such as methyl orange or phenolphthalein should be added.</p>
	i v	<p>FIRST CHECK THE ANSWER ON THE ANSWER LINE</p> <p>IF B = RbIO₃ AND relative formula mass = 260.5 award 5 marks</p> <p>IF relative formula mass = 260.5 award 4 marks</p> <p>-----</p> <p><i>n</i>(S₂O₃²⁻) in titration</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px 0;"> $= \frac{0.150 \times 23.80}{1000} = 3.57 \times 10^{-3} \text{ (mol)} \checkmark$ </div> <p><i>n</i>(IO₃[✓]) in titration</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px 0;"> $= 10 \times 5.95 \times 10^{-4} = 5.95 \times 10^{-3} \text{ (mol)} \checkmark$ </div> <p><i>n</i>(IO₃[✓]) in original 250 cm³</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px 0;"> $= \frac{3.57 \times 10^{-3}}{6} = 5.95 \times 10^{-4} \text{ (mol)} \checkmark$ </div>	5	<p>ALLOW ECF from incorrect mean titre in (a)(i)</p> <p>ECF from <i>n</i>(S₂O₃²⁻) in titration</p> <p>ALLOW a two-step calculation</p> <p>$n(I_2) = n(S_2O_3^{2-}) \div 2$ and $n(IO_3^-) = n(I_2) \div 3$</p> <p>ECF from <i>n</i>(IO₃⁻) in titration</p>

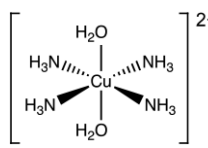
5.2.3 Redox and Electrode Potentials

		<p>Relative formula mass of B $= \frac{1.55}{5.95 \times 10^{-3}} = 260.5 \text{ (g mol}^{-1}\text{)} \checkmark$</p> <p>Formula of B (must be derived from relative formula mass)</p> <div style="border: 1px solid black; padding: 5px; margin: 10px 0;"> <p>Iodate of Group 1 metal that most closely matches calculated molar mass of B</p> <p>Formula from 260.5 = $\text{RbIO}_3 \checkmark$</p> </div>		<p>ECF from $n(\text{IO}_3^-)$ in original 250 cm³</p> <p>IF scaling $\times 10$ is omitted, ALLOW ECF from $n(\text{IO}_3^-)$ in titration</p> <p>ALLOW ECF from incorrect RFM of B provided metal is from Group 1 ALLOW RbIO_3^- DO NOT ALLOW RbIO_3 without relative formula mass value. DO NOT ALLOW 260.4 (without working) and RbIO_3 IF B = RbIO_3 AND relative formula mass = 261 award 5 marks</p> <p><u>Examiner's Comments</u></p> <p>This unstructured calculation was done well by the higher ability candidates. Lower ability candidates struggled to show what they were attempting to calculate and in particular did not appreciate the 1 : 6 ratio of $\text{S}_2\text{O}_3^{2-}(\text{aq})$ to $\text{IO}_3^-(\text{aq})$.</p> <p>Candidates might be advised to start $n(\text{formula}) = \dots$ at the start of each line of calculation</p> <p>eg $n(\text{S}_2\text{O}_3^{2-}) = \dots \text{ mol}$</p> <p>No credit was given to candidates who grasped the identity of the Group 1 iodate from nowhere and calculated the theoretical relative formula mass.</p>
		Total	10	
9	a	i	4	<p>Voltmeter must be shown AND salt bridge must be labelled ALLOW small gaps in circuit</p> <p>ALLOW half cells drawn either way around IGNORE 2 before $\text{I}^-(\text{aq})$ DO NOT ALLOW $\text{I}_2(\text{g})$ OR $\text{I}_2(\text{s})$ OR $\text{I}_2(\text{l})$</p> <p>ALL conditions required BUT ALLOW 1 mol dm⁻³/1M if omitted here but shown for just one solution in diagram Look on diagram in addition to answer lines</p> <p>IGNORE pressure <i>Not relevant for this cell</i></p> <p>DO NOT ALLOW 1 mol for concentration</p>

5.2.3 Redox and Electrode Potentials

				<p>Examiner's Comments</p> <p>A significant number of candidates did not draw a simple circuit. Most candidates scored the mark for the Ni electrode but fewer scored the mark for the Pt electrode in the $I^-(aq)/I_2(aq)$ mixture. The last mark, for the conditions (including concentrations of 1.00 mol dm^{-3}), was almost universally scored even by those who omitted the solutions in their diagram.</p> <p>Common errors were: diagrams with electrodes not entering solutions; salt bridges not labelled or not entering solutions; making solid iodine an electrode; and assuming iodine to be a gas and adapting a standard hydrogen electrode to accommodate this.</p>
	ii	$E = 0.79 \text{ (V)} \checkmark$	1	<p>IGNORE sign</p> <p>Examiner's Comments</p> <p>This was the most successfully answered single mark question on the paper.</p>
	b i	$H_2O_2(aq) + 2H^+(aq) + 2Fe^{2+}(aq) \rightarrow 2Fe^{3+}(aq) + 2H_2O(l) \checkmark$	1	<p>ALLOW multiples IGNORE state symbols, even if wrong</p> <p>Examiner's Comments</p> <p>Most candidates got the equation correct. Occasionally a number was omitted causing incorrect stoichiometry and occasionally the equation was reversed.</p>
	ii	<p>Equations</p> $3Zn(s) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 3Zn^{2+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l) \checkmark$ $Zn(s) + 2Cr^{3+}(aq) \rightarrow Zn^{2+}(aq) + 2Cr^{2+}(aq) \checkmark$ <p>Comparison of E values (seen once)</p> <p>E of Zn is more negative/less positive than E of $Cr_2O_7^{2-}$ OR E of Zn is more negative/less positive than E of $Cr^{3+} \checkmark$</p> <p>Equilibrium shift related to E values</p>	4	<p>ALLOW multiples IGNORE state symbols, even if wrong</p> <p>ALLOW E_{cell} is (+) 2.09V for Zn/$Cr_2O_7^{2-}$ cell OR ALLOW E_{cell} is (+) 0.34V for Zn/Cr^{3+} cell IGNORE 'lower/higher'</p>

5.2.3 Redox and Electrode Potentials

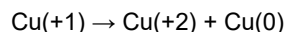
	<p>More negative/less positive OR Zn system shifts left</p> <p>OR</p> <p>Less negative/more positive $\text{Cr}_2\text{O}_7^{2-}$ system shifts right OR Less negative/more positive Cr^{3+} system shifts right ✓</p>	<p>For 'shifts left': ALLOW '(Zn) is oxidised' OR 'electrons are lost (from Zn)'</p> <p>For 'shifts right', ALLOW '(Cr) is reduced' OR 'electrons are gained'</p> <p>Examiner's Comments</p> <p>Only the higher ability candidates achieved full marks.</p> <p>Most candidates were able to come up with both redox equations and some could state that zinc was oxidised or chromium was reduced as a result of zinc having a more negative electrode potential than either chromium species.</p> <p>Common errors were: stating that zinc had a lower (not more negative) electrode potential; stating that zinc was an oxidising agent (as well as being oxidised); confusing the direction of change of equilibria shown; and not balancing the overall redox equations.</p>
c	<p><i>Please refer to the marking instructions on page 5 of this mark scheme for guidance on how to mark this question.</i></p> <p>Level 3 (5–6 marks) All three reactions are covered in detail with C, D, E and F identified with clear explanations.</p> <p><i>There is a well-developed line of reasoning which is clear and logically structured with clear chemical communication and few omissions. The information presented is relevant and substantiated.</i></p> <p>Level 2 (3–4 marks) All three reactions are covered but explanations may be incomplete</p> <p>OR Two reactions are explained in detail.</p> <p><i>There is an attempt at a logical structure with a line of reasoning. The information is relevant e.g. formulae may contain missing brackets or numbers and supported by some evidence.</i></p> <p>Level 1 (1–2 marks) Make two simple explanations from any one reaction.</p> <p>OR Makes one simple explanation from each of two reactions</p>	<p>Indicative scientific points may include:</p> <p>REACTION 1 ($\text{CuSO}_4/\text{NH}_3$) Product</p> <p>C : $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$</p> <p>Equation</p> $[\text{Cu}(\text{H}_2\text{O})_6]^{2+} + 4\text{NH}_3 \rightarrow [\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+} + 4\text{H}_2\text{O}$ <p>Structure of trans stereoisomer</p> <p>6</p>  <p>Correct connectivity</p> <p>REACTION 2 ($\text{Cu}_2\text{O}/\text{H}_2\text{SO}_4$) Products</p> <p>D : CuSO_4 OR $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ E: Cu</p> <p>Equation</p> $\text{Cu}_2\text{O} + \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + \text{Cu} + \text{H}_2\text{O}$

5.2.3 Redox and Electrode Potentials

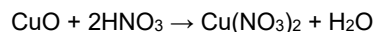
There is an attempt at a logical structure with a line of reasoning. The information is in the most part relevant.

0 marks No response worthy of credit.

Oxidation numbers



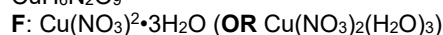
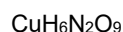
REACTION 3 (CuO/HNO₃) Equation



Molar ratios

$$\begin{array}{cccc} \text{Cu} & : & \text{H} & : & \text{N} & : & \text{O} \\ = & \frac{26.29}{63.5} & : & \frac{2.49}{1.0} & : & \frac{11.59}{14.0} & : & \frac{59.63}{16.0} \end{array}$$

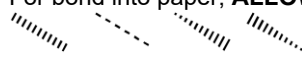
Formula of F



Further guidance on use of wedges

- Must contain 2 'out wedges', 2 'in wedges' and 2 lines in plane of paper **OR** 4 lines, 1 'out wedge' and 1 'in wedge':

For bond into paper, **ALLOW**:

- 

ALLOW following geometry:

- 

Examiner's Comments

Many candidates had a stab at identifying **C-F** but neglected to include equations for the three reactions described or to show relevant working.

Most candidates recognised **C** as the ammoniacal copper(II) ion but the formula was frequently incorrect and correct attempts at a ligand substitution equation from $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ was rarely seen. Diagrams showing the *trans* isomer were attempted but often of poor quality due to incorrect linking.

Candidates recognised **D** as being CuSO_4 but often did not identify **E** as Cu due to a lack of familiarity with this common disproportionation reaction. $\text{Cu}(\text{OH})_2(\text{s})$ was a common incorrect identification of **E**. Only the best responses described the oxidation number changes which made this a disproportionation reaction.

F was identified by a percentage by mass calculation to determine an empirical formula and then by deduction to produce $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$. Having done this, many candidates did not give the relatively simple equation for reaction 3 between copper(II) oxide and dilute nitric acid.

Exemplar 2

(d)* Three different reactions of copper compounds are described below.

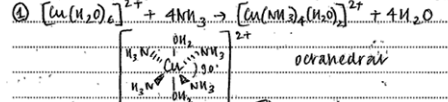
Reaction 1: Aqueous copper(II) sulfate reacts with excess aqueous ammonia in a ligand substitution reaction. A deep blue solution is formed, containing an octahedral complex ion, C, which is a *trans* isomer. $\Rightarrow \text{100}^*$

Reaction 2: Copper(I) oxide reacts with hot dilute sulfuric acid in a disproportionation reaction. A blue solution, D, and a brown solid, E are formed. $\text{Cu}_2\text{O} + \text{H}_2\text{SO}_4$

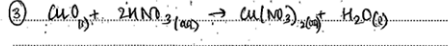
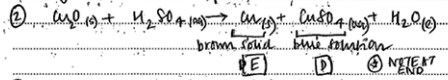
Reaction 3: Copper(II) oxide reacts with warm dilute nitric acid in a neutralisation reaction, to form a blue solution. Unreacted copper(II) oxide is filtered off, and the solution is left overnight in an evaporating basin. A hydrated salt, F, crystallises, with the percentage composition by mass: Cu, 26.29%; H, 2.48%; N, 11.59%; O, 59.63%.

Identify C-F by formulae or structures, as appropriate.

Include equations, any changes in oxidation number, and working. [6]

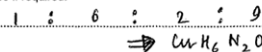


This is the *trans* isomer (C) because the H_2O ligands are 180° apart.



	Cu	H	N	O
mass	26.29	2.48	11.59	59.63
rfm	63.5	1	14	16
mol	0.414	2.48	0.828	3.73

Additional answer space if required.



A hydrated salt is made up of an anhydrous salt with water of crystallisation.
 $\text{Cu}(\text{NO}_3)_2 \cdot 3\text{H}_2\text{O}$ (F) and molecular

\Rightarrow This fits the empirical formula.

NOTE: ② The oxidation number of Cu goes from +1 to 0 in Cu_2O , and from +1 to +2 in Cu_2SO_4 . Cu^+ is reduced to form Cu and oxidised to form Cu^{2+} in CuSO_4 .

This exemplifies how considered structuring of candidate responses can enhance their clarity. The clear labelling of C-F in the response, along with numbering which corresponds to the reactions in the question stem, make the candidate's line of reasoning easy to follow. The underlining and annotations in the question stem show good

5.2.3 Redox and Electrode Potentials

					practice in picking out and interpreting key information.
			Total	16	
1 0		i	sodium nitrate(III)	1	<p>ALLOW sodium nitrite OR sodium nitrite(III)</p> <p>Examiner's Comment: This part was very poorly answered, the most common answer being sodium nitrate. The examiners were expecting sodium nitrate(III) but the mark scheme was extended to also allow sodium nitrite.</p> <p>Sodium(III) nitrate was sometimes seen, indicating that candidates are not fully conversant with rules for showing oxidation states in names.</p>
		ii	<p>Sodium / Na oxidised from 0 to +1 ✓</p> <p>Nitrogen / N reduced from +3 to 0 ✓</p>	2	<p>ALLOW 1+ for +1 and 3+ for +3</p> <p>ALLOW N₂ for nitrogen</p> <p>ALLOW 1 mark for elements AND all oxidation numbers correct, but N on oxidised line and Na on reduced line</p> <p>'+' is required in +3 and +1 oxidation numbers</p> <p>Examiner's Comment: This part was generally answered well although a significant number of candidates managed to get one of the oxidation numbers wrong, usually for N. It was rare to see the sign for an oxidation number omitted.</p>
		iii	<p>$2\text{NaNO}_2 + 6\text{Na} \rightarrow 4\text{Na}_2\text{O} + \text{N}_2$ ✓</p> <p>IGNORE state symbols</p>	1	<p>ALLOW multiples, e.g. $\text{NaNO}_2 + 3\text{Na} \rightarrow 2\text{Na}_2\text{O} + \frac{1}{2}\text{N}_2$ ✓</p> <p>Examiner's Comment: The examiners were impressed with the responses for this part with just over half the candidates producing a correct balanced equation for this unfamiliar reaction. Most used whole numbers for balancing but it was common to also see the half-multiple version including $\frac{1}{2}\text{N}_2$.</p>
			Total	4	
1 1			ALLOW equilibrium sign in equations provided reactants on left	4	<p>ALLOW correct multiples IGNORE state symbols</p> <p>..... .</p> <p>ALLOW uncancelled H₂O and H⁺</p> <p>$\text{H}_2\text{O}_2 + \text{MnO}_2 + 4\text{H}^+ \rightarrow \text{O}_2 + \text{Mn}^{2+} + 2\text{H}_2\text{O} + 2\text{H}^+$</p>

5.2.3 Redox and Electrode Potentials

		<p>Reaction of H₂O₂ with MnO₂: $\text{H}_2\text{O}_2 + \text{MnO}_2 + 2\text{H}^+ \rightarrow \text{O}_2 + \text{Mn}^{2+} + 2\text{H}_2\text{O} \checkmark$</p> <p>Reaction of H₂O₂ with Mn²⁺: $\text{H}_2\text{O}_2 + \text{Mn}^{2+} \rightarrow \text{MnO}_2 + 2\text{H}^+ \checkmark$</p> <p>Use of E data</p> <p>Use of E data to support equation(s) above or half direction of provided half equations (one including MnO₂) \checkmark <i>Also look for evidence around half equations</i></p> <p>MnO₂ regenerated / reformed \checkmark</p> <p><i>Must be linked to an equation showing MnO₂ as reactant and an equation showing MnO₂ as product</i></p>		<p>$\text{H}_2\text{O}_2 + \text{Mn}^{2+} + 2\text{H}_2\text{O} + 2\text{H}^+ \rightarrow \text{MnO}_2 + 4\text{H}^+ + 2\text{H}_2\text{O}$</p> <p>Examples</p> <ul style="list-style-type: none"> • More negative E moves to left ORA • Reduction half equation to the right ORA • Most positive E is reduced ORA • Calculated E cell = +0.81 V (from top 2) OR +0.27 V (from bottom 2) <p>ALLOW combining of equations above to show that MnO₂ is used and reformed</p> <p>Examiner's Comment: Many candidates found this part challenging and there was a wide variety of answers and marks awarded. There were two equations to construct showing how MnO₂, and Mn²⁺ react with H₂O₂. Many combined the two equations involving H₂O₂ to obtain the overall equation for the decomposition of H₂O₂ which was given at the top of the paper. Of the equations seen, many had species uncancelled. Many candidates only tackled one of the equations.</p> <p>Candidates were expected to provide evidence for their equations based on the electrode potentials provided. Use of this data was patchy and only the best candidates linked the relative E values to the direction of movement or redox details. A significant number gave cell potentials.</p> <p>Regeneration of MnO₂ was well understood but often just stated with no reference to the equations. This part discriminated very well.</p>
		Total	4	
1 2	a	<p>AWARD 2 marks for correct balancing AND all species cancelled on both sides of equation: $2\text{MnO}_4^- + 6\text{H}^+ + 5\text{SO}_3^{2-} \rightarrow 2\text{Mn}^{2+} + 3\text{H}_2\text{O} + 5\text{SO}_4^{2-} \checkmark \checkmark$</p> <p>AWARD 1 mark for correct balancing but not all species (H₂O, H⁺) cancelled on both sides of equation \checkmark e.g. $2\text{MnO}_4^- + 16\text{H}^+ + 5\text{SO}_3^{2-} + 5\text{H}_2\text{O}$</p>	2	<p>ALLOW correct multiples e.g. $\text{MnO}_4^- + 3\text{H}^+ + 2\frac{1}{2}\text{SO}_3^{2-} \rightarrow \text{Mn}^{2+} + 1\frac{1}{2}\text{H}_2\text{O} + 2\frac{1}{2}\text{SO}_4^{2-}$</p> <p>IGNORE state symbols</p> <p>e.g. $\text{MnO}_4^- + 8\text{H}^+ + 2\frac{1}{2}\text{SO}_3^{2-} + 2\frac{1}{2}\text{H}_2\text{O}$</p>

5.2.3 Redox and Electrode Potentials

			$\rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{SO}_4^{2-} + 10\text{H}^+$		$\rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 2\frac{1}{2}\text{SO}_4^{2-} + 5\text{H}^+$
					<p>Examiner's Comments This was successfully answered by the majority of candidates.</p>
	b	i	<p><i>Circuit:</i> complete circuit AND voltmeter AND labelled salt bridge linking two half-cells ✓</p> <p><i>Half cells:</i> Pt AND Fe²⁺ AND Fe³⁺ ✓</p> <p>Zn AND Zn²⁺ ✓</p> <p><i>Standard conditions:</i></p> <p>1 mol dm⁻³ (solution(s)) AND 298 K / 25°C ✓</p>	4	<p>Electrodes / salt bridge must at least touch the surface</p> <p>ALLOW small gaps in circuit wires</p> <p>ALLOW half cells drawn either way around</p> <p>ALLOW 1 mol / dm³ OR 1 M ALLOW 1 mol dm⁻³/1M if omitted here but shown for just one solution in diagram IGNORE pressure DO NOT ALLOW 1 mol(e) for concentration</p> <p>Examiner's Comments A significant number of candidates handicapped themselves by failing to draw a simple circuit. Diagrams with electrodes not entering solutions, salt bridges not labelled or not entering solutions were frequently seen. Most candidates scored the mark for the Zn electrode but less scored the mark for the Pt electrode in the Fe²⁺(aq)/Fe³⁺(aq) mixture. The last mark, for the conditions (including concentrations of 1.00 mol dm⁻³), was almost universally scored even by those who omitted the solutions in their diagram.</p>
		ii	1.53 (V) ✓	1	<p>IGNORE sign</p> <p>Examiner's Comments This was the most successfully answered one mark question on the paper.</p>
	c		<p><i>strongest reducing agent:</i> Zn ✓</p> <p><i>strongest oxidizing agent:</i> MnO₄⁻ ✓</p>	2	<p>NOTE: H⁺ has been ignored</p> <p>Examiner's Comments Most candidates were able to use the standard electrode potentials given in the question to work out that zinc (Zn(s)) was the strongest reducing agent and that manganate(VII) ions were the strongest oxidising agent.</p>
			Total	9	
1 3			<p>H₂SO₄ + 8HI → 4I₂ + H₂S + 4H₂O</p> <p>All species correct</p>	2	

5.2.3 Redox and Electrode Potentials

			<p>OR $\text{H}_2\text{SO}_4 : \text{HI}$ ratio = 1 : 8 ✓</p> <p>Equation complete and balanced ✓</p>		
			Total	2	
1 4	a	i	$3 \text{MnO}_4^{2-} + 4 \text{H}^+ \rightarrow 2 \text{MnO}_4^- + \text{MnO}_2 + 2 \text{H}_2\text{O}$ ✓	1	ALLOW 1 in front of MnO_2
		ii	<p>In acidic conditions (Concentration of) H^+ increases AND equilibrium (position) shifts to the right to reduce concentration of H^+/remove H^+ ✓</p> <p>In alkaline conditions OH^- reacts with H^+ AND equilibrium (position) shifts to the left to increase concentration of H^+/add H^+ ✓</p>	2	ALLOW $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$
		b	<p>In acid: $\text{H}_2(\text{g}) \rightarrow 2\text{H}^+(\text{aq}) + 2\text{e}^-$ AND $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$ ✓</p> <p>$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ AND Cell potential = $1.23 - 0.00 = 1.23$ (V) ✓</p> <p>In alkali: $2\text{OH}^-(\text{aq}) + \text{H}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + 2\text{e}^-$ AND $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq})$ ✓ $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$</p> <p>AND Cell potential = $0.40 - (-0.83) = 1.23$ (V) ✓</p>	4	ALLOW $\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$
			Total	7	
1 5	a		<p>Cu^{2+}: $(1\text{s}^2) 2\text{s}^2 2\text{p}^6 3\text{s}^2 3\text{p}^6 3\text{d}^9$ ✓</p> <p>Cu^+: $(1\text{s}^2) 2\text{s}^2 2\text{p}^6 3\text{s}^2 3\text{p}^6 3\text{d}^{10}$ ✓</p>	2	<p>IGNORE repeated 1s^2 after 1s^2 prompt on answer line</p> <p>ALLOW 4s^0, either before or after 3d</p> <p>ALLOW upper case D, etc and subscripts, e.g.....3S₂3P⁶</p> <p>DO NOT ALLOW [Ar] as shorthand for $1\text{s}^2 2\text{s}^2 2\text{p}^6 3\text{s}^2 3\text{p}^6$</p> <p>Examiner's Comments</p>

5.2.3 Redox and Electrode Potentials

				<p>The responses seen were very mixed. Able candidates scored the two marks easily but many errors were seen, particularly by removal of 3d electrons rather than 4s electrons from copper atoms to give the electron configurations of the ions (especially for Cu⁺ in CuI).</p>
	b	<p>IGNORE any charges shown within formulae (treat as rough working)</p> <p>$\text{CuCO}_3 + 2\text{HCOOH} \rightarrow \text{Cu}(\text{HCOO})_2 + \text{H}_2\text{O} + \text{CO}_2$</p> <p>OR $\text{CuO} + 2\text{HCOOH} \rightarrow \text{Cu}(\text{HCOO})_2 + \text{H}_2\text{O}$</p> <p>OR $\text{Cu}(\text{OH})_2 + 2\text{HCOOH} \rightarrow \text{Cu}(\text{HCOO})_2 + 2\text{H}_2\text{O} \checkmark$</p>	1	<p>IGNORE state symbols</p> <p>In formula of HCOOH / HCOO, ALLOW H, C and O in ANY order</p> <p>ALLOW H₂CO₃ for H₂O and CO₂ in carbonate equation</p> <p>ALLOW (HCOO)₂Cu for Cu(HCOO)₂</p> <p>DO NOT ALLOW equation with CuSO₄</p> <p>Examiner's Comments</p> <p>Most candidates attempted an equation using CuO, Cu(OH)₂ or CuCO₃. Marks were then sometimes lost by not balancing the equation. It was not uncommon to see equations using CuSO₄ or CuCl₂ as reactant and consequently this mark was often not awarded.</p>
	c	<p>$2\text{Cu}^{2+} + 4\text{I}^- \rightarrow 2\text{CuI}(\text{s}) + \text{I}_2 \checkmark$</p> <p>State symbol for CuI(s) ONLY required</p>	1	<p>ALLOW multiples, e.g. $\text{Cu}^{2+} + 2\text{I}^- \rightarrow \text{CuI}(\text{s}) + \frac{1}{2}\text{I}_2$</p> <p>IGNORE other state symbols, even if incorrect</p> <p>Examiner's Comments</p> <p>This equation proved to be much more difficult than in 8(b), with only the best candidates producing a correctly balanced equation. As with 4(c) and 7(b)(iii), equations were often unbalanced in terms of charge and oxidation number.</p>
	d	<p>Starch \checkmark</p> <p>Blue / black to colourless / white \checkmark</p> <p>MARK INDEPENDENTLY</p>	2	<p>IGNORE 'brown' in composite colour with blue or black, i.e.</p> <p>ALLOW blue / brown to colourless</p> <p>ALLOW black / brown to colourless</p> <p>DO NOT ALLOW just 'it turns colourless / is decoloured'</p> <p><i>Initial colour required</i></p> <p>IGNORE clear for colourless</p> <p>Examiner's Comments</p> <p>Most candidates seemed unaware that starch is used to identify the end point in iodine–thiosulfate</p>

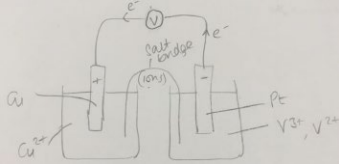
5.2.3 Redox and Electrode Potentials

				<p>titrations. Even when starch was given, the colour change was often incorrect. Random responses were seen to this part, e.g. methyl orange, phenolphthalein, potassium manganate and sodium thiosulfate.</p>
	e	<p>WORKING REQUIRED Correct answer: $x = 4$ required evidence of working $n(\text{S}_2\text{O}_3^{2-})$ OR $n(\text{Cu}^{2+}) = \frac{0.0420 \times 23.5}{1000} = 9.87 \times 10^{-4}$ (mol) ✓ In 250.0 cm³ solution, $n(\text{Cu}^{2+}) = 9.87 \times 10^{-3}$ (mol) ✓ $M(\text{Cu}(\text{HCOO})_2 \cdot 4\text{H}_2\text{O}) = \frac{2.226}{9.87 \times 10^{-3}} = 225.5$ (g mol⁻¹) ✓ $x(\text{H}_2\text{O})$ has mass of $225.5 - M(\text{Cu}(\text{HCOO})_2)$ $= 225.5 - 153.5$ $= 72(.0)$ ✓ $x = \frac{72(.0)}{18(.0)} = 4$ WHOLE NUMBER needed AND evidence of working ✓</p>	<p>FULL ANNOTATIONS MUST BE USED At least 3 SF required throughout <i>Alternative approach for final 3 marks based on mass:</i> $\text{mass Cu}(\text{HCOO})_2 = 9.87 \times 10^{-3} \times 153.5 = 1.515$ g ✓ $n(\text{H}_2\text{O}) = \frac{2.226 - 1.515}{18(.0)} = \frac{0.711}{18(.0)} = 0.0395$ (mol) ✓ $x = \frac{0.0395}{9.87 \times 10^{-3}} = 4$ ✓ ALLOW $\text{Cu}(\text{HCOO})_2 \cdot 4\text{H}_2\text{O}$ COMMON ERRORS for 4 marks $x = 117$ (calc 116.78) <i>Use of 9.87×10^{-4} (no scaling $\times 10$) $\rightarrow M = 2255.319$</i> $x = 17$ (calc 16.53) 4 marks <i>Use of 4.935×10^{-4} (Use of $0.5 \times 9.87 \times 10^{-3}$)</i> Check $n(\text{Cu}^{2+})$ for other ECFs Check for ECFs from incorrect $M(\text{anhydr salt})$ Actual = 153.5 Examiner's Comments Many candidates were on firm territory with a redox titration problem. The majority went through a well-rehearsed sequence of steps to obtain all five marks for showing that x was 4. Where '4' had not been obtained, marks could still be awarded for intermediate working if correct. Answer: $x = 4$</p>	5
		Total	11	
1 6	i	linear ✓	1	<p>IGNORE planar Examiner's Comments Most candidates identified the shape as linear.</p>

5.2.3 Redox and Electrode Potentials

	ii	<p>Au / Gold has been oxidised from 0 to +1 ✓</p> <p>O / Oxygen / O₂ has been reduced from 0 to -2 ✓</p>	2	<p>IF Ag referred to, rather than Au, treat as a slip and apply BOD</p> <p>ALLOW 0 to 1 (i.e. no + sign for +1)</p> <p>ALLOW 1 mark for ALL oxidation numbers correct with no oxidised or reduced OR oxidation and reduction wrong way round, e.g. Au goes from 0 to +1 and O goes from 0 to -2 ✓ Au is reduced from 0 to +1 and O is oxidised from 0 to -2 ✓</p> <p>Examiner's Comments</p> <p>Although most identified correct oxidation numbers, a significant number thought that the CN⁻ ion was reduced. Overall though, candidates demonstrated a good understanding of basic redox.</p>
	iii	<p>IGNORE any charges shown within complexes (treat as rough working)</p> <p>$4\text{Au} + 8\text{CN}^- + 2\text{H}_2\text{O} + \text{O}_2 \rightarrow 4[\text{Au}(\text{CN})_2]^- + 4\text{OH}^-$ ✓ ✓</p> <p>First mark for all 6 species</p> <p>Second mark for balancing</p>	2	<p>IF Ag referred to, rather than Au, treat as a slip and apply BOD</p> <p>IGNORE state symbols</p> <p>CARE: In $[\text{Au}(\text{CN})_2]^-$, - sign is OUTSIDE square brackets</p> <p>For 1st mark, IGNORE e⁻ present</p> <p>ALLOW 1 mark for balanced equation with CN⁻ missing, i.e. $4\text{Au} + 2\text{H}_2\text{O} + \text{O}_2 \rightarrow 4\text{Au}^+ + 4\text{OH}^-$</p> <p>ALLOW 1 mark rogue e⁻ on either side</p> <p>ALLOW multiples, e.g. $2\text{Au} + 4\text{CN}^- + \text{H}_2\text{O} + \frac{1}{2}\text{O}_2 \rightarrow 2[\text{Au}(\text{CN})_2]^- + 2\text{OH}^-$ $\text{Au} + 2\text{CN}^- + \frac{1}{2}\text{H}_2\text{O} + \frac{1}{4}\text{O}_2 \rightarrow [\text{Au}(\text{CN})_2]^- + \text{OH}^-$</p> <p>Examiner's Comments</p> <p>Almost all candidates wrote the correct species but balancing the equation was much more demanding. It was often balanced by species but not by charge and oxidation number.</p>
	i v	<p>$\text{ClO}^- + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{Cl}^- + \text{H}_2\text{O}$ ✓</p>	1	<p>IGNORE state symbols</p> <p>ALLOW e for electron</p> <p>ALLOW multiples</p> <p>Examiner's Comments</p> <p>Candidates found it easy to derive the correct half equation.</p>
		Total	6	

5.2.3 Redox and Electrode Potentials

1 7	a	<p>Definition The e.m.f. (of a half-cell) compared with / connected to a (standard) hydrogen half-cell / (standard) hydrogen electrode ✓</p> <p>Standard conditions <i>Units essential</i> Temperature of 298 K / 25°C AND (solution) concentrations of 1 mol dm⁻³ AND pressure of 100kPa OR 10⁵ Pa OR 1 bar ✓</p>	<p>For e.m.f., ALLOW voltage OR potential difference / p.d. OR electrode / reduction / redox potential ALLOW e.m.f. of a cell.... ALLOW / (standard) hydrogen cell IGNORE S.H.E. (as abbreviation for standard hydrogen electrode) DO NOT ALLOW hydrogen fuel cell</p> <p>2 ALLOW 1M OR 1 mol/dm³ DO NOT ALLOW 1 mol OR 1 mole ALLOW 1 atmosphere / 1 atm OR 101 kPa OR 101325 Pa</p> <p>Examiner's Comments</p> <p>This definition was well known with most candidates achieving both marks. An error was more likely for the units of the standard conditions, usually 1 mol and 1 mol dm³ for concentration. A less common error was comparison with a hydrogen fuel cell.</p>
	b i	<p>Complete circuit with voltmeter AND labelled salt bridge linking two half-cells ✓</p>  <p>Cu electrode in Cu²⁺ ✓</p> <p>Pt electrode in V²⁺ AND V³⁺ ✓</p> <p>Cu shown as + AND Pt shown as - ✓</p> <p>electrons in wire AND ions in salt bridge ✓ <i>On diagram or stated</i></p>	<p>Half cells can be drawn in either order Half cells must show electrodes dipping into solutions ALLOW small gaps in circuit DO NOT ALLOW half-cell with H₂ added</p> <p>IGNORE any stated concentrations IGNORE 'anode' and 'cathode'</p> <p>In salt bridge, ALLOW any stated ion that may be present, e.g. K⁺, NH₄⁺, NO₃⁻, Cu²⁺, V²⁺, V³⁺</p> <p>5 IGNORE direction of travel of ions and electrons.</p> <p>ALLOW Cu half cell as + AND V half cell as-</p> <p>Examiner's Comments</p> <p>Most candidates drew a full circuit with half cells containing the correct species in a suitably labelled diagram. The signs of the electrodes were usually correct but the charge carriers were omitted much more often than they were included. The overall quality of the diagrams was poor, usually resembling drawings or even rough sketches.</p>
	ii	0.60 OR 0.6 (V) ✓	1 IGNORE any sign

5.2.3 Redox and Electrode Potentials

			<p>Examiner's Comments</p> <p>This part was answered correctly by almost all candidates. Answer: standard cell potential = 0.60 V</p>
c		<p>Definitions: 1 mark Oxidising agent removes / accepts / gains electrons OR increases oxidation number (of another species) AND Reducing agent adds / donates / loses electrons OR decreases oxidation number (of another species) ✓</p> <p>Oxidising agent: 2 marks Cr³⁺ oxidises Al OR Cr³⁺ acts as oxidising agent AND 3Cr³⁺ + Al → 3Cr²⁺ + Al³⁺ ✓ Explanation (dependent on Cr³⁺ oxidising Al above) E of redox system 2 (Cr³⁺/Cr²⁺) is more positive /less negative (than E of system 1 (Al³⁺/Al)) ORA, i.e. in terms of 1 being more negative (than 2) ✓</p> <p>Reducing agent: 3 marks Cr³⁺ reduces FeO₄²⁻/(H⁺) ✓ 2Cr³⁺ + 2FeO₄²⁻ + 2H⁺ → Cr₂O₇²⁻ + 2Fe³⁺ + H₂O ✓ Explanation (dependent on Cr³⁺ reducing FeO₄²⁻ above) E of redox system 5 (Cr₂O₇²⁻/Cr³⁺) is less positive / more negative (than E of system 6 (FeO₄²⁻/Fe³⁺)) ORA, i.e. in terms of 6 being more positive (than 5) ✓</p>	<p>FULL ANNOTATIONS MUST BE USED ALLOW oxidising agent decreases its oxidation number AND reducing agent increases its oxidation number</p> <p>IGNORE oxidising agent oxidises / is reduced OR reducing agent reduces / is oxidised In equations,</p> <ul style="list-style-type: none"> • IGNORE state symbols (even if incorrect) • ALLOW ⇌ in equation <p>IF more than one equation shown for Cr³⁺ as oxidising agent, CON and zero marks for 2 oxidising agent marks IGNORE equations with Cr²⁺ as reactant</p> <p>Explanations MUST be in terms of positive / negative: IGNORE 'higher' E OR 'greater'</p> <p>ALLOW E_{cell} = +1.25 V (+ sign required)</p> <p>IF more than one equation shown for Cr³⁺ as a reducing agent, CON and zero marks for 3 reducing agent marks IGNORE equations with Cr²⁺ as reactant</p> <p>Explanations MUST be in terms of positive / negative: IGNORE 'higher' E OR 'greater'</p> <p>ALLOW E_{cell} = +0.87 V (+ sign required)</p> <p>Examiner's Comments</p> <p>Although most candidates produced lengthy answers to this part, there was often little that could be rewarded. The terms oxidising agent and reducing agent were usually described but the expected response needed to be in terms of electrons or oxidation number. An alarming number of candidates claimed that an oxidising agent is oxidised and a reducing agent reduced.</p> <p>Many candidates went on to discuss the role of Cr³⁺ ions in oxidising aluminium, with a correct</p>

5.2.3 Redox and Electrode Potentials

				<p>equation. The expected explanation in terms of electrode potentials was seen more rarely.</p> <p>The role of Cr^{3+} in reducing FeO_4^{2-} was seen much more rarely. The commonest responses showed a list of equations, all showing reactions of Cr^{2+} rather than Cr^{3+}.</p> <p>The best candidates produced superb answers demonstrating excellent understanding of electrode potentials.</p> <p>When comparing electrode potentials, candidates are advised to use terms such 'more negative' or 'more positive, rather than 'greater' or 'higher', which are ambiguous when comparing negative numbers. An alternative approach in terms of positive cell potentials were seen and credited, although the positive sign was required.</p>
		Total	14	
1 8		<p>in all equations ALLOW equilibrium signs IGNORE state symbols</p> <p>.....</p> <p>Reaction 1: 1 mark $2\text{I}_2 + 5\text{O}_2 \rightarrow 2\text{I}_2\text{O}_5 \checkmark$</p> <p>.....</p> <p>Reaction 2: 2 marks 1st mark: ALL CORRECT species</p> <p>e.g.: $\text{I}_2 + \text{OH}^- \rightarrow \text{I}^- + \text{IO}_3^- + \text{H}_2\text{O}$</p> <p>2nd mark for CORRECT balanced equation $3\text{I}_2 + 6\text{OH}^- \rightarrow 5\text{I}^- + \text{IO}_3^- + 3\text{H}_2\text{O}$ $\checkmark\checkmark$</p>	3	<p>FULL ANNOTATIONS MUST BE USED</p> <p>ALLOW correct multiples throughout, e.g. $\text{I}_2 + 2\frac{1}{2}\text{O}_2 \rightarrow \text{I}_2\text{O}_5$</p> <p>.....</p> <p>For 1st mark, IGNORE e^- present</p> <p>ALLOW species / equation with NaOH or KOH, e.g. $3\text{I}_2 + 6\text{NaOH} \rightarrow 5\text{I}^- + \text{IO}_3^- + 3\text{H}_2\text{O} + 6\text{Na}^+$ $3\text{I}_2 + 6\text{NaOH} \rightarrow 5\text{NaI} + \text{NaIO}_3 + 3\text{H}_2\text{O}$</p> <p>ALLOW Species: $\text{I}_2 + \text{OH}^- \rightarrow \text{I}^- + \text{IO}_2^+ + \text{H}_2\text{O} \checkmark$ OR Equation: $3\text{I}_2 + 4\text{OH}^- \rightarrow 5\text{I}^- + \text{IO}_2^+ + 2\text{H}_2\text{O} \checkmark\checkmark$</p> <p>Species: $\text{I}_2 + \text{OH}^- \rightarrow \text{I}^- + \text{IO}_3^+ + \text{H}_2\text{O} \checkmark$ OR Equation: $3\text{I}_2 + 2\text{OH}^- \rightarrow 5\text{I}^- + \text{IO}_3^+ + \text{H}_2\text{O} \checkmark\checkmark$</p> <p>Examiner's Comments</p> <p>This question was about equilibrium, set in the context of the solubility of iodine.</p> <p>This part required candidates to interpret unfamiliar information to construct reactions for redox reactions of iodine. Candidates were far more successful with the first equation than the more demanding second equation.</p> <p>For the second equation, the key difficulty was the formula of the iodate(V) ion, IO_3^-, with I^{5+} or the formula of I_2O_5 often being shown.</p> <p>The most difficult skill was balancing the second redox equation, suggesting that many candidates need more practice in this area.</p>

5.2.3 Redox and Electrode Potentials

		Total	3	
1 9	a	<p>Equations can be in either order</p> <p>$\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH} \checkmark$</p> <p>$\text{NaFeO}_2 + 2\text{H}_2\text{O} \rightarrow \text{Fe}(\text{OH})_3 + \text{NaOH} \checkmark$</p>	<p>ALLOW multiples throughout IGNORE state symbols ALLOW $\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{Na}^+ + 2\text{OH}^-$</p> <p>DO NOT ALLOW equations with uncanceled species. e.g. $\text{Na}_2\text{O} + 2\text{H}_2\text{O} \rightarrow 2\text{NaOH} + \text{H}_2\text{O}$</p> <p>ALLOW $2\text{NaFeO}_2 + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2\text{NaOH}$ OR $2\text{NaFeO}_2 + \text{H}_2\text{O} \rightarrow \text{Fe}_2\text{O}_3 + 2\text{Na}^+ + 2\text{OH}^-$</p> <p>Examiner's Comments</p> <p>The information needed to write the equation was within the information provided for step 1 and the stem. Candidates were much more successful with the first than the second equation. The clue that an alkaline solution had been formed should have helped with the identification of NaOH as a product of both reactions. The brown precipitate provided a clue that Fe(OH)₃ had been formed although the examiners also credited an equation producing Fe₂O₃.</p>	2
	b	<p>FIRST, CHECK THE ANSWER ON ANSWER LINE IF answer = 33.7%, award 6 marks. IF there is an alternative answer, check to see if there is any ECF credit possible using working below</p> <p>..... amount $\text{S}_2\text{O}_3^{2-}$ used = $0.1000 \times \frac{25.50}{1000}$ = 2.550×10^{-3} (mol) ? amount I_2 = $2.550 \times 10^{-3} \div 2$ = 1.275×10^{-3} (mol) \checkmark</p> <p>amount CrO_4^{2-} $\frac{2}{3} \times 1.275 \times 10^{-3}$ OR $1.275 \times 10^{-3} \div 1.5$ = $8.5(00) \times 10^{-4}$ (mol) \checkmark</p> <p>amount CrO_4^{2-} in original 1000 cm^3 = $40 \times 8.5(00) \times 10^{-4}$ = $3.4(00) \times 10^{-2}$ mol \checkmark</p> <p>Mass of Cr / Cr^{3+} in ore = $52.0 \times 3.4(00) \times 10^{-2}$ g = 1.768 g \checkmark</p> <p>Percentage Cr in ore = $\frac{1.768}{5.25} \times 100$ = 33.7% \checkmark</p> <p>MUST be to one decimal place (in the question)</p>	<p>FULL ANNOTATIONS MUST BE USED</p> <p>IF a step is omitted but subsequent step subsumes previous, then award mark for any missed step Working: at least 3 SF throughout until final % mark BUT ignore trailing zeroes, ie for 0.490 allow 0.49</p> <p>.....</p> <p>ECF answer above $\div 2$</p> <p>ECF answer above $\div 1.5$</p> <p>ECF answer above $\times 40$</p> <p>ECF answer above $\times 52.0$ IMPORTANT: The last two marks are ONLY available by using 52.0 for Cr</p> <p>..... Common ECFs:</p> <p>0.8% $\times 40$ missing 5 marks (scaling error)</p>	6

5.2.3 Redox and Electrode Potentials

			<p>0.84% × 40 missing 4 marks (scaling error and 2 DP)</p> <p>33.68% 5 marks (2 DP)</p> <p>16.8% 5 marks (divide Cr somewhere by 2)</p> <p>144.9%; 72.5% 4 marks (Final 2 marks unavailable) Use of $M(\text{Fe}(\text{CrO}_2)_2) = 223.8$ instead of $M(\text{Cr})$.</p> <p>Examiner's Comments</p> <p>Many candidates were on firm territory with a redox titration problem. The majority went through a well-rehearsed sequence of steps to obtain four marks for reaching the amount of CrO_4^{2-} ions in the original solution. Sometimes, candidates used the 1:1.5 molar ratio for $\text{CrO}_4^{2-} : \text{I}_2$ the wrong way around to obtain 1.9125×10^{-3} rather than 8.50×10^{-4} mol CrO_4^{2-}. More candidates had problems in scaling up by 40 to obtain the original amount of CrO_4^{2-} as 3.40×10^{-2} mol. Strangely many used a factor of 4 instead.</p> <p>The last two marks proved to be more elusive, with many candidates calculating the percentage of $\text{Fe}(\text{CrO}_2)_2$ rather than Cr in the sample of chromite.</p> <p>The responses seen show just how far candidates have travelled since early structured titration calculations for AS to complex unstructured calculations at the end of the A-level course.</p> <p>Answer: 33.7%</p>
c		<p><i>Overall:</i></p> $4^{2-} + 3\text{I}^- + 4\text{H}_2\text{O} \rightarrow \text{Cr}^{3+} + 1\frac{1}{2} \text{I}_2 + 8\text{OH}^- \checkmark$ <p>CrO</p> <p><i>Half equations:</i></p> $4^{2-} + 4\text{H}_2\text{O} + 3\text{e}^- \rightarrow \text{Cr}^{3+} + 8\text{OH}^- \checkmark$ <p>CrO</p> $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^- \checkmark$	<p>ALLOW multiples and equilibrium signs throughout</p> <p>IGNORE state symbols throughout</p> <p>e.g. $2\text{CrO}_4^{2-} + 6\text{I}^- + 8\text{H}_2\text{O} \rightarrow 2\text{Cr}^{3+} + 3\text{I}_2 + 16\text{OH}^-$</p> <p>ALLOW equation using H^+. i.e.</p> $\text{CrO}_4^{2-} + 3\text{I}^- + 8\text{H}^+ \rightarrow \text{Cr}^{3+} + 1\frac{1}{2} \text{I}_2 + 4\text{H}_2\text{O}$ <p>OR</p> $2\text{CrO}_4^{2-} + 6\text{I}^- + 16\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 3\text{I}_2 + 8\text{H}_2\text{O}$ <p>ALLOW CrO_4^{2-} half equation using H^+. i.e.</p> $\text{CrO}_4^{2-} + 8\text{H}^+ + 3\text{e}^- \rightarrow \text{Cr}^{3+} + 4\text{H}_2\text{O}$ <p>Examiner's Comments</p> <p>This part required candidates to construct three equations for an unfamiliar reaction. The examiners allowed equations using H^+ rather than OH^-. It was then possible to credit many candidates with the full three marks with many excellent responses seen. Predictably the</p>

5.2.3 Redox and Electrode Potentials

				equations involving CrO_4^{2-} were more difficult but even some weaker candidates were able to construct an equation for the oxidation of iodide ions. The very best candidates did manage to construct the equations in alkaline conditions.
		Total	11	
20	i	Fe^{2+} : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$ ✓		<p>ALLOW 4s before 3d, ie $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^10 4p^6$</p> <p>ALLOW $1s^2$ written after answer prompt (ie $1s^2$ twice)</p> <p>ALLOW upper case D, etc and subscripts, e.g.4S23D1</p> <p>ALLOW for Fe^{2+}4s⁰</p> <p>DO NOT ALLOW [Ar] as shorthand for $1s^2 2s^2 2p^6 3s^2 3p^6$</p> <p>Look carefully at $1s^2 2s^2 2p^6 3s^2 3p^6$ – there may be a mistake</p> <p>Examiner's Comments</p>
	i	Br^- : $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$ ✓	2	<p>Few candidates produced two incorrect electron configurations but there were many mistakes seen for either species. For Fe^{2+}, the commonest error was for loss of electrons from the 3d rather than 4s sub-shell of an Fe atom. For a Br^- ion, it was common to see the electron configuration of a Br atom. Surprisingly a common error was to see $4p^4$ rather than $4p^6$ from loss rather than gain of an electron. Only just over half the candidates showed two correct configurations so clearly more care is needed when answering.</p>
	ii	<p>With Cl_2 AND Br_2 AND I_2 products are Fe^{2+} (AND halide ion) FeCl_2 AND FeBr_2 AND FeI_2 ✓</p> <p>OR</p> <p>Evidence that two electrode potentials have been compared for at least ONE reaction, ✓ e.g. $\text{Fe} -0.44$ AND $\text{Cl}_2 +1.36$ e.g. Iron has more / most negative electrode potential</p> <p>With Cl_2 AND Br_2, products are Fe^{3+} (AND halide ion) FeCl_3 AND FeBr_3 ?</p>	3	<p>FULL ANNOTATIONS NEEDED</p> <p>ALLOW products within equations (even if equations are not balanced)</p> <p>IF stated, IGNORE reactants</p> <p>ALLOW response in terms of positive 'cell reactions', e.g. $\text{Fe} + \text{Cl}_2 \rightarrow \text{Fe}^{2+} + 2\text{Cl}^-$ $E = (+)1.80 \text{ V}$</p> <p>IGNORE comments about reducing and oxidising agents and electrons</p> <p>Examiner's Comments</p> <p>The majority of candidates predicted that Fe would react with all three halogens to form Fe^{2+} ions, supported by equations and electrode potential data. Many simply stated that Fe has the more negative E value (or the halogens the more</p>

5.2.3 Redox and Electrode Potentials

					<p>positive value). It was also common to see cell voltages used, such as +0.98 V for a reaction between iron and iodine. Both approaches were credited.</p> <p>The most able candidates correctly predicted that Fe²⁺ ions, initially formed from the reaction of iron with bromine and chlorine, would then be oxidised to Fe³⁺. The best answers showed exceptional understanding. Candidates are advised to consider all the information supplied in a question as the majority had ignored completely the Fe³⁺/Fe²⁺ data.</p>
			Total	5	
2 1	a	i	pH = 0 ✓	1	<p>Guidance</p> <p>Examiner's Comments</p> <p>Surprisingly, this part caused many candidates a problem with less than half the candidates obtaining the correct answer of 0 V. The commonest incorrect responses seen were pH values of 1 or even 7.</p> <p>Answer: pH = 0 V</p>
		ii	<p>H redox system is more negative (e.g. has a more -ve <i>E</i> OR less +ve <i>E</i> OR is -ve electrode)</p> <p>OR H redox system releases electrons (May be in equation, e.g. H₂ → 2H⁺ + 2e⁻) ✓</p> <p>Equilibrium shifts to increase [H⁺] OR H⁺ OR standard hydrogen equation shifts to increase [H⁺] OR H⁺ ✓</p>	2	<p>ALLOW ORA, ie Ag redox system (D) has more positive <i>E</i> / less negative <i>E</i></p> <p>ALLOW equilibrium sign</p> <p>IGNORE H is more reactive ORA</p> <p>IGNORE direction of equilibrium shift</p> <p>Examiner's Comments</p> <p>This part discriminated well with many candidates unable to interpret the information. Although the question asked for an explanation in terms of electrode potentials and equilibrium, these were often absent in the responses. Surprisingly, many identified the incorrect polarity for the hydrogen half-cell and then concluded that a decrease in pH is a result of a decrease, rather than an increase in [H⁺(aq)]. To score well required a very good understanding of equilibrium in the context of electrode potentials.</p>
		iii	H ₂ + 2Ag ⁺ → 2Ag + 2H ⁺ ✓	1	<p>ALLOW multiples e.g. ½H₂ + Ag⁺ → Ag + H⁺ State symbols NOT required ALLOW equilibrium sign</p>

5.2.3 Redox and Electrode Potentials

				<p>Examiner's Comments</p> <p>Perhaps influenced by incorrect polarity, many equations were shown the wrong way round. Many equations were not balanced by charge with Ag^+ and Ag often having no balancing number.</p>
	b	i	Fuel reacts with oxygen / oxidant to give electrical energy / voltage ✓	<p>1</p> <p>ALLOW named fuel. e.g. hydrogen / H_2; ethanol; methanol, etc ALLOW fuel cell requires constant supply of fuel AND oxygen / an oxidant OR fuel cell operates continuously as long as a fuel AND oxygen / an oxidant are added IGNORE 'reactants' 'products' and comments about pollution and efficiency</p> <p>Examiner's Comments</p> <p>Many responses seen here were far too vague. It was common to see answers about a fuel reacting but with no reference to oxygen. Responses in terms of less pollution or greenhouse gases gained no credit.</p>
		ii	ethanol is a liquid OR is less volatile	Assume that 'it' refers to ethanol
		ii	OR ethanol is easier to store / transport / stored more safely	ALLOW ORA throughout
		ii	OR hydrogen is explosive / more flammable	IGNORE ethanol has a higher boiling point
		ii	OR ethanol has more public / political acceptance ✓	<p>1</p> <p>IGNORE H_2 is a gas IGNORE 'produces no CO_2' OR less pollution IGNORE comments about efficiency IGNORE comments about biomass and renewable</p> <p>Examiner's Comments</p> <p>Most candidates were credited here, recognising that ethanol is a liquid and can therefore be more easily stored or transported.</p>
		iii	$\text{C}_2\text{H}_5\text{OH} + 3\text{O}_2 \rightarrow 2\text{CO}_2 + 3\text{H}_2\text{O} \rightarrow$	<p>1</p> <p>Correct species AND balancing needed ALLOW multiples ALLOW $\text{C}_2\text{H}_6\text{O}$ for formula of ethanol IGNORE state symbols</p> <p>Examiner's Comments</p> <p>Almost all candidates wrote correct species but the equation was not always balanced correctly with $3\frac{1}{2}\text{O}_2$ or 2O_2 being the commonest errors. Candidates need to take great care when balancing an equation containing an alcohol to account for the O atom with the alcohol formula.</p>
		i	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$ ✓	<p>1</p> <p>Correct species AND balancing needed ALLOW multiples, e.g. $3\text{O}_2 + 12\text{H}^+ + 12\text{e}^- \rightarrow$</p>

5.2.3 Redox and Electrode Potentials

				<p>6H₂O $\frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}$ ALLOW e (ie no - sign) ALLOW $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightarrow 4\text{OH}^-$ OR $3\text{O}_2 + 6\text{H}_2\text{O} + 12\text{e}^- \rightarrow 12\text{OH}^-$ IGNORE state symbols</p> <p>Examiner's Comments</p> <p>The equations seen were impressive with nearly half the candidates providing an equation that could be credited. The examiners credited many different balancing numbers of $\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}$. The commonest seen were multiples using $\frac{1}{2}\text{O}_2$ and especially 3O_2. As correct answers often showed no working, perhaps the oxygen half equation had been memorised by many candidates.</p>
		v	<p>oxidation: C from -2 to +4 '+' sign not required ✓ reduction: O from 0 to -2 ✓</p>	<p>ALLOW 2- and 4+ ALLOW $\text{C}^{2-} \rightarrow \text{C}^{4+}$</p> <p>ALLOW 0 and 2- ALLOW $\text{O}^0 \rightarrow \text{O}^{2-}$ ALLOW 1 mark if correct oxidation numbers shown for BOTH C and O but wrong way around (ie C on reduction line and O on oxidation line)</p> <p>2 IGNORE O₂ reduced IGNORE any reference to electron transfer (<i>not in question</i>)</p> <p>Examiner's Comments</p> <p>This part required identification of oxidation and reduction in an unfamiliar context. The reduction of O was far easier than the oxidation of C and this was reflected in 1/2 being the commonest mark awarded. Discrimination was very good with able candidates collecting both marks.</p>
			Total	10
2 2	a		<p>$\text{Fe}_2\text{O}_3 + 3\text{Cl}_2 + 10\text{OH}^- \rightarrow 2\text{FeO}_4^{2-} + 6\text{Cl}^- + 5\text{H}_2\text{O}$ ✓✓ First mark for all 6 species Second mark for balancing</p>	<p>2 ALLOW multiples ALLOW oxidation half equation for two marks $\text{Fe}_2\text{O}_3 + 10\text{OH}^- \rightarrow 2\text{FeO}_4^{2-} + 5\text{H}_2\text{O} + 6\text{e}^-$ Correct species would obtain 1 mark – <i>question: equation for oxidation</i></p> <p>ALLOW variants forming H⁺ for 1 mark, e.g: $\text{Fe}_2\text{O}_3 + 3\text{Cl}_2 + 5\text{OH}^- \rightarrow 2\text{FeO}_4^{2-} + 6\text{Cl}^- + 5\text{H}^+$ $\text{Fe}_2\text{O}_3 + 3\text{Cl}_2 + 5\text{OH}^- \rightarrow 2\text{FeO}_4^{2-} + 5\text{HCl} + \text{Cl}^-$</p> <p>Examiner's Comments</p>

5.2.3 Redox and Electrode Potentials

			<p>The information needed to write the equation was largely within the information provided for step 1. In step 1, candidates were provided with three reactants and two of the products. They were also told that the reaction was carried out using an excess of hydroxide ions, so any potential H⁺ ions produced would be neutralised to water. Only the very best candidates were able to interpret this information to score both marks for the correct equation. Many attempts seen did not start with iron(III) oxide. When arriving at a complete equation, candidates are recommended to check the overall charge on either side. This must balance, a feature not seen in the majority of responses. One mark was available for an equation with all species correct, including water as the third product, or a 'correct' equation but with H⁺ produced.</p>
b	$\text{Ba}^{2+}(\text{aq}) + \text{FeO}_4^{2-}(\text{aq}) \rightarrow \text{BaFeO}_4(\text{s}) \checkmark$	1	<p>Balanced ionic equation AND state symbols required DO NOT ALLOW +2 or -2 for ionic charges</p> <p>Examiner's Comments</p> <p>As with 8(a), the relevant information was mostly included within the referenced part: step 2. The responses were very disappointing as the required equation is very similar to a simple precipitation reaction between silver and halide ions. The requirement for state symbols was clearly stated but often omitted from otherwise correct equations.</p>
c	<p>Reason can ONLY be correct from correct reducing agent </p> <p>reducing agent: I⁻ OR KI ✓</p> <p>I⁻ adds / donates / loses electrons AND to FeO₄²⁻ OR to BaFeO₄ OR to Fe(VI) or to Fe(+6) ✓ ALLOW Fe(6+) OR Fe⁶⁺</p>	2	<p>IGNORE H⁺ OR acidified ALLOW iodide / potassium iodide but DO NOT ALLOW iodine</p> <p>ALLOW I⁻ loses electrons AND to form I₂</p> <p>ALLOW Fe(6+) OR Fe⁶⁺</p> <p>Examiner's Comments</p> <p>The majority of candidates identified iodide ions or potassium iodide as the oxidising agent. Iodine was often recognised as the product but the explanation was usually in terms of oxidation number despite the question asking for electrons – very much a case of reading the question. Precise language was also required as iodine and iodide are rather different, especially as iodine is the product. The best responses discussed the species being reduced, BaFeO₄ or Fe(VI).</p>

5.2.3 Redox and Electrode Potentials

d		<p>FIRST, CHECK THE ANSWER ON ANSWER LINE IF answer = 51.8%, award 4 marks.</p> <p>..... $n(\text{S}_2\text{O}_3^{2-}) \text{ used} = 0.1000 \times \frac{26.4}{1000} = 2.64 \times 10^{-3} \text{ (mol)} \checkmark$</p> <p>$n(\text{FeO}_4^{2-}) = \frac{1}{2} \times \frac{2}{3} \times 2.64 \times 10^{-3} = 8.8(0) \times 10^{-4} \text{ (mol)}$ \checkmark</p> <p>Mass BaFeO_4 in sample $= 8.8 \times 10^{-4} \times 257.1 \text{ g} = 0.226248 \text{ g} \checkmark$</p> <p>$\% \text{ purity} = \frac{0.226248}{0.437} \times 100 = 51.8\% \checkmark$</p> <p>MUST be to one decimal place (in the question) </p> <p>As an alternative for the final two marks, ALLOW: $\frac{0.437}{257.1} = 0.00170 \text{ (mol)}$ \checkmark</p> <p>$\% \text{ purity} = \frac{8.8 \times 10^{-4}}{1.70 \times 10^{-3}} \times 100 = 51.8\% \checkmark$</p>	<p>FULL ANNOTATIONS MUST BE USED </p> <p>For alternative answers, look first at common ECFs below. Then check for ECF credit possible using working below IF a step is omitted but subsequent step subsumes previous, then award mark for any missed step </p> <p>Working must be to at least 3 SF throughout until final % mark BUT ignore trailing zeroes, ie for 0.880 allow 0.88</p> <p>ECF answer above $\times \frac{1}{2} \times \frac{2}{3}$ This mark may be seen in 2 steps via l_2 but the mark is for both steps combined</p> <p>ECF 257.1 \times answer above $\frac{\text{answer above}}{0.437} \times 100$ ECF</p> <p>ALLOW 51.7% FROM 0.226 g BaFeO_4 (earlier rounding) </p> <p>Common ECFs: No $\times \frac{2}{3}$ for $n(\text{FeO}_4^{2-})$: $\% \text{ purity} = 77.7\%/77.6\%$ 3 marks No $\div 2$ for $n(\text{FeO}_4^{2-})$: $\% \text{ purity} = 25.9\%$ 3 marks 24.6 used instead of 26.4: $\% \text{ purity} = 48.2\%$ 3 marks</p> <p>Examiner's Comments</p> <p>After the information-finding demands of parts (a)–(c), candidates were on much firmer territory here with a stock redox titration problem. Many candidates secured all 4 marks and most were able to obtain some marks along the way. The hardest mark was the step from the initial amount of $\text{Na}_2\text{S}_2\text{O}_3$ to the amount of BaFeO_4.</p> <p>Answer: 51.8%</p>
e		<p>gas: $\text{O}_2 \checkmark$</p> <p>precipitate: $\text{Fe}(\text{OH})_3 \checkmark$</p> <p>equation: $2\text{FeO}_4^{2-} + 5\text{H}_2\text{O} \rightarrow 1\frac{1}{2}\text{O}_2 + 2\text{Fe}(\text{OH})_3 + 4\text{OH}^-$ OR $2\text{FeO}_4^{2-} + \text{H}_2\text{O} + 4\text{H}^+ \rightarrow 1\frac{1}{2}\text{O}_2 + 2\text{Fe}(\text{OH})_3 \checkmark$</p>	<p>DO NOT ALLOW names IGNORE a balancing number shown before a formula</p> <p>ALLOW $\text{Fe}(\text{OH})_3(\text{H}_2\text{O})_3$</p> <p>ALLOW multiples ALLOW $2\text{FeO}_4^{2-} + 11\text{H}_2\text{O} \rightarrow 1\frac{1}{2}\text{O}_2 + 2\text{Fe}(\text{OH})_3(\text{H}_2\text{O})_3 + 4\text{OH}^-$</p>

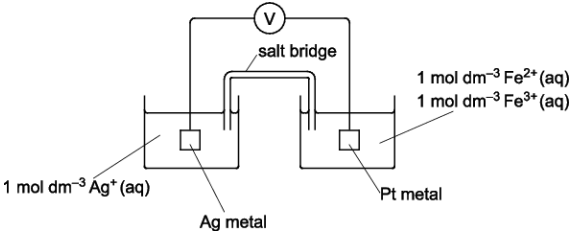
5.2.3 Redox and Electrode Potentials

				<p>Examiner's Comments</p> <p>This part required candidates to construct an equation for an unfamiliar reaction. Candidates were reasonably competent in identifying the gas as O₂ and precipitate as Fe(OH)₃. Unfortunately, some responded with 'oxygen' despite the formulae being asked for in the question. The correct equation proved to be the hardest mark on the paper, being seen extremely rarely. As with the equation in 8(a), often the overall charge didn't balance on either side of the equation, a consideration that would have led to many more correct responses.</p>	
			Total	12	
2 3	a	<p>Definition The e.m.f. (of a half-cell) compared with / connected to a (standard) hydrogen half-cell / (standard) hydrogen electrode ✓</p> <p>Standard conditions <i>Units essential</i> Temperature of 298 K / 25°C</p> <p>AND (solution) concentrations of 1 mol dm⁻³</p> <p>AND pressure of 100 kPa OR 10⁵ Pa OR 1 bar ✓</p>	<p>As alternative for e.m.f., ALLOW voltage OR potential difference OR p.d. OR electrode potential OR reduction potential OR redox potential ALLOW / (standard) hydrogen cell IGNORE S.H.E. (as abbreviation for standard hydrogen electrode)</p> <p>ALLOW 1M</p> <p>DO NOT ALLOW 1 mol ALLOW 1 atmosphere / 1 atm OR 101 kPa OR 101325 Pa</p> <p>Examiner's Comments</p> <p>This definition was well known with most candidates achieving both marks. An error was more likely for the units of the standard conditions, usually 1 mol and 1 mol dm³ for concentration. A less common error was comparison with a hydrogen fuel cell.</p>	2	
	b	i	$2\text{Ag}^+(\text{aq}) + \text{Cu}(\text{s}) \rightarrow 2\text{Ag}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \checkmark$	<p>State symbols not required ALLOW \rightleftharpoons provided that reactants on LHS</p> <p>Examiner's Comments</p> <p>This proved to be a relatively easy mark although some showed the equation the wrong way round or did not balance the silver species on one side of the equation.</p>	1
		ii	<p>Assume Cu²⁺ Cu OR Cu half cell unless otherwise stated.</p> <p>[Cu²⁺] decreases OR < 1 mol dm⁻³</p> <p>AND</p>	<p>FULL ANNOTATIONS MUST BE USED</p> <p>ALLOW [Cu²⁺] less than standard concentration / 1 mol dm⁻³</p> <p>DO NOT ALLOW water reacts with Cu²⁺ OR Cu</p>	3

5.2.3 Redox and Electrode Potentials

		<p>Equilibrium (shown in table) shifts to left ✓</p> <p>more electrons are released by Cu ✓</p> <p>The cell has a bigger difference in E ✓</p>		<p>ALLOW E (for $\text{Cu}^{2+} \text{Cu}$) is less positive / more negative / decreases IGNORE standard electrode potential (<i>Cell no longer standard</i>) IGNORE E^\ominus decreases CARE DO NOT ALLOW statements about silver E changing (CON)</p> <p>IGNORE just 'cell potential increases' (in the question) <i>The final mark is more subtle and is a consequence of the less positive E value of the copper half cell</i></p> <p>Examiner's Comments</p> <p>This part was designed to be very demanding and this proved to be the case with weaker candidates made little if any headway. The key requirements were a decreased concentration of copper ions resulting in an equilibrium shift to the left, increasing the electron flow from the copper electrode and decreasing the electrode potential of the copper half cell. This results in a larger difference in electrode potential between the two half cells and an increase in the cell potential. Even able candidates often responded in terms of the standard electrode potential changing which can never be the case. Many candidates suggested that the added water would react with copper ions forming a copper hydroxide precipitate with an equilibrium shift to the right. To score well required a very good understanding of equilibrium in the context of electrode potentials.</p>
c	i	<p>$E = -2.31 \text{ (V)}$ ✓</p>	1	<p>± sign AND 2.31 required for the mark</p> <p>Examiner's Comments</p> <p>In past exams, candidates have found it easy to calculate a standard cell potential. This slightly different slant on the question produced far more incorrect responses with +2.31 V (wrong sign) and 3.11 V (wrong subtraction) being common. Answer: -2.31 V</p>
	ii	<p>$4\text{Al(s)} + 4\text{OH}^{\ominus}\text{(aq)} + 3\text{O}_2\text{(g)} + 6\text{H}_2\text{O(l)} \rightarrow 4\text{Al(OH)}_4^{\ominus}\text{(aq)}$</p> <p>species ✓ balance ✓</p>	2	<p>IGNORE state symbols ALLOW multiples ALLOW 1 mark for an equation in which OH^- are balanced but have not been cancelled, e.g. $4\text{Al(s)} + 16\text{OH}^{\ominus}\text{(aq)} + 3\text{O}_2\text{(g)} + 6\text{H}_2\text{O(l)} \rightarrow 4\text{Al(OH)}_4^{\ominus}\text{(aq)} + 12\text{OH}^{\ominus}\text{(aq)}$</p> <p>ALLOW 1 mark if charge on Al(OH)_4 is omitted, i.e</p>

5.2.3 Redox and Electrode Potentials

				$4\text{Al(s)} + 4\text{OH}^{\text{(aq)}} + 3\text{O}_2\text{(g)} + 6\text{H}_2\text{O(l)} \rightarrow 4\text{Al(OH)}_4^{\text{(aq)}}$ <p>ALLOW 1 mark for an 'correct equation' reversed, i.e. $4\text{Al(OH)}_4^{\text{(aq)}} \rightarrow 4\text{Al(s)} + 4\text{OH}^{\text{(aq)}} + 3\text{O}_2\text{(g)} + 6\text{H}_2\text{O(l)}$</p> <p>Examiner's Comments</p> <p>The attempts at this equation were impressive with many securing both marks for the correct balanced equation. One mark was available and awarded often for the following: a correctly balanced equation but with hydroxide ions not cancelled; the correct equation but reversed; an otherwise correct equation but with the charge missing from the Al(OH)_4^- ion.</p>	
		Total	9		
2 4	a i	 <p>Half-cells (2 marks) Ag(s) and $1 \text{ mol dm}^{-3} \text{Ag}^{\text{(aq)}}$</p> <p>$1 \text{ mol dm}^{-3} \text{Fe}^{2+}\text{(aq)}$ AND $1 \text{ mol dm}^{-3} \text{Fe}^{3+}\text{(aq)}$ AND Pt metal</p> <p>Complete circuit (1 mark) salt bridge AND voltmeter AND wires</p> <p>Standard conditions (1 mark) 298 K / 25 °C AND 100 kPa / 101 kPa pressure</p>	4	ALLOW 1 atm	
		ii	(Electrode potential of) Ag^+ / Ag becomes more positive ... therefore, E_{cell} becomes smaller OR less positive.	2	ALLOW equilibrium Ag / Ag^+ shifts to right ALLOW more negative 2nd mark only available if deduced from 1st mark ALLOW ECF for 2nd mark
	b	Ce^{3+} and Zn^{2+}	1		
	c	Mn^{2+} , H_2O , Fe^{3+} , Br_2 Three species correct Four species correct	2		
		Total	9		
2 5	a i	Titres correct and ALL recorded to 2 decimal places	1		

5.2.3 Redox and Electrode Potentials

		<p>Titre: 22.50 21.80 22.15 21.70</p> <p>AND mean titre = 21.75 (cm³) ✓</p>		Working not required for mean titre
	ii	<p>Percentage uncertainty = $\frac{0.05 \times 2}{21.70} \times 100 = 0.46 (\%)$ ✓</p> <p>2 DP required</p>	1	ALLOW ECF from incorrect subtraction in (i)
	iii	(Excess KI) ensures that all I ₂ is formed for titration ✓	1	ALLOW All Cu ²⁺ converted to CuI/Cu(I)
	i v	<p>Add starch ✓</p> <p>Blue to colourless AND all I₂ has reacted ✓</p>	2	
	v	<p>$n(\text{S}_2\text{O}_3^{2-})$ in titration = $\frac{0.120 \times 21.75}{1000} = 2.61 \times 10^{-3} (\text{mol})$ ✓</p> <p>$n(\text{I}_2) = 1.305 \times 10^{-3} (\text{mol})$ OR $n(\text{Cu}^{2+}) = 2.61 \times 10^{-3} (\text{mol})$</p> <p>AND $n(\text{Cu}^{2+})$ in original 250 cm³ = $10 \times 2.61 \times 10^{-3} = 2.61 \times 10^{-2} (\text{mol})$</p> <p>Mass of A (CuSO₄·xH₂O) used = 17.95 – 12.35 = 5.60 (g)</p> <p>$M(\text{CuSO}_4 \cdot x\text{H}_2\text{O}) = \frac{5.60}{2.61 \times 10^{-2}} = 214.6 (\text{g mol}^{-1})$ ✓</p> <p>$x = \frac{214.6 - 159.6}{18.0} = 3.05 = 3$</p> <p>Formula = CuSO₄·3H₂O</p>	4	<p>ALLOW ECF from incorrect mean titre in (i)</p> <p>ECF from $n(\text{S}_2\text{O}_3^{2-})$</p> <p>ECF from $n(\text{Cu}^{2+})$</p> <p>BUT DO NOT ALLOW incorrect mass</p> <p>ECF from $M(\text{CuSO}_4 \cdot x\text{H}_2\text{O})$</p> <p>AND x = nearest whole number</p>
	b i	<p>green solution: Fe²⁺(aq) OR [Fe(H₂O)₆]²⁺</p> <p>AND gas bubbles: H₂(g)</p> <p>AND orange-brown solution: Fe³⁺(aq) OR [Fe(H₂O)₆]³⁺ ✓</p> <p>Fe(s) + 2H⁺(aq) → Fe²⁺(aq) + H₂(g) ✓</p> <p>4Fe²⁺(aq) + O₂(g) + 4H⁺(aq) → 4Fe³⁺(aq) + 2H₂O(l) ✓</p>	3	<p>State symbols are not required in this part IGNORE, even if incorrect</p> <p>ALLOW full equation: Fe(s) + 2HCl(aq) → FeCl₂(aq) + H₂(g)</p>
	ii	<p>orange solution: Cr₂O₇²⁻</p> <p>AND green solution (anywhere) Cr³⁺ OR [Cr(H₂O)₆]³⁺ ✓</p>	3	<p>State symbols are not required in this part IGNORE, even if incorrect</p> <p>IGNORE Cr(VI) <i>The question asks for species</i></p>

5.2.3 Redox and Electrode Potentials

		$2\text{Cr}^{3+}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + 3\text{H}_2\text{O}_2(\text{aq}) \rightarrow \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 8\text{H}^+(\text{aq})$ $\text{H}^+, \text{H}_2\text{O} \text{ and } \text{e}^- \text{ all cancelled } \checkmark\checkmark$		ALLOW 1 mark for $\text{H}^+/\text{H}_2\text{O}/\text{e}^-$ not cancelled, e.g. $2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l}) + 3\text{H}_2\text{O}_2(\text{aq}) + 6\text{H}^+(\text{aq}) \rightarrow \text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{H}_2\text{O}(\text{l})$ \checkmark
		Total	15	
2 6	a	$n(\text{NH}_2\text{OH}) = 4.32 \times 10^{-2} \times 0.0250 = 1.08 \times 10^{-3} \text{ mol}$ $n(\text{Fe}^{3+}) = 3 \times 1.08 \times 10^{-3} = 3.24 \times 10^{-3} \text{ mol}$ (assuming Equation 3) $\text{volume} = \frac{3.24 \times 10^{-2} \times 1000}{0.0400} = 81.0 \text{ cm}^3$ Explanation: minimum amount of Fe^{3+} required is maximum amount theoretically required to react with all NH_2OH , i.e. if Equation 3 is correct (greatest amount of Fe^{3+} required) (<i>owtte</i>)	4	Factor 3 must be included in second mark for ECF on third mark. ALLOW 2 sig figs
	b	$n(\text{MnO}_4^-) = 2.00 \times 10^{-2} \times \frac{21.6}{1000} = 4.32 \times 10^{-4} \text{ (mol)}$ $n(\text{Fe}^{2+}) = 4.32 \times 10^{-4} \times 5 = 2.16 \times 10^{-3} \text{ (mol)}$ Ratio $\text{NH}_2\text{OH} : \text{Fe}^{2+}$ OR $\text{NH}_2\text{OH} : \text{Fe}^{2+}$ $= 1.08 \times 10^{-3} : 2.16 \times 10^{-3} = 1 : 2$ AND Equation 2 is correct	3	Working must be to at least 3 sig figs throughout until final numerical answer BUT ignore trailing zeroes, e.g. for 0.490 allow 0.49 ECF answer above $\times 5$ This mark is only possible from correct answers above, i.e. no ECF
	c i	Boiling speeds up the reaction OR Ensures that reaction is complete (Titre is less because) there is less Fe^{2+}	2	
	ii	In Stage 1 , increase quantities so that there is sufficient solution for more than one titration	1	ALLOW increase scale of Stage 1
		Total	10	
2 7		$\dots\text{ClO}^- + \dots\text{H}_2\text{O} + 2\text{e}^- \rightarrow \dots\text{Cl}^- + 2\text{OH}^-$ $\text{Fe}_2\text{O}_3 + 10\text{OH}^- \rightarrow 2\text{FeO}_4^{2-} + 5\text{H}_2\text{O} + 6\text{e}^-$ $\text{Fe}_2\text{O}_3 + 3\text{ClO}^- + 4\text{OH}^- \rightarrow 2\text{FeO}_4^{2-} + 3\text{Cl}^- + 2\text{H}_2\text{O}$	3	ALLOW multiples throughout
		Total	3	
2 8		Element oxidised: sulfur / S 0 to +6 Element reduced: nitrogen / N +5 to +4	4	ALLOW 5+, 4+ and 6+ Signs required

5.2.3 Redox and Electrode Potentials

			$6\text{HNO}_3 + \text{S} \rightarrow 6\text{NO}_2 + \text{H}_2\text{SO}_4 + 2\text{H}_2\text{O}$		ALLOW $4\text{H}^+ + 6\text{NO}_3^- + \text{S} \rightarrow 6\text{NO}_2 + \text{SO}_4^{2-} + 2\text{H}_2\text{O}$
			Correct species		
			Balance		
			Total	4	