



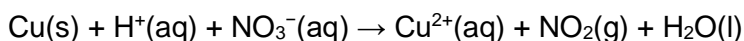
1. Which redox reaction contains the largest change in oxidation state for sulfur?

- A. $\text{H}_2\text{SO}_4 + 8\text{HI} \rightarrow \text{H}_2\text{S} + 4\text{I}_2 + 4\text{H}_2\text{O}$
- B. $\text{S} + \text{O}_2 \rightarrow \text{SO}_2$
- C. $\text{S}_2\text{O}_3^{2-} + 2\text{H}^+ \rightarrow \text{SO}_2 + \text{S} + \text{H}_2\text{O}$
- D. $\text{S} + 6\text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + 6\text{NO}_2 + 2\text{H}_2\text{O}$

Your answer

[1]

2. An unbalanced equation for a redox reaction is shown below.



What is the balancing number for H^+ when the equation is balanced using the smallest whole numbers?

- A. 2
- B. 4
- C. 6
- D. 8

Your answer

[1]

3. Which property is **not** correct for calcium?

- A. It acts as an oxidising agent
- B. It forms a basic oxide
- C. It reacts with water to form hydrogen gas
- D. Its hydroxide is more alkaline than magnesium hydroxide

Your answer

[1]



4. A standard cell is set up from $\text{H}^+(\text{aq})|\text{H}_2(\text{g})$ and $\text{Ag}^+(\text{aq})|\text{Ag}(\text{s})$ half cells.

The standard electrode potentials of the redox systems are shown below.



What is the reaction at the negative electrode of the cell?

- A. $\text{H}^+(\text{aq}) + \text{e}^- \rightarrow \frac{1}{2}\text{H}_2(\text{g})$
- B. $\frac{1}{2}\text{H}_2(\text{g}) \rightarrow \text{H}^+(\text{aq}) + \text{e}^-$
- C. $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$
- D. $\text{Ag}(\text{s}) \rightarrow \text{Ag}^+(\text{aq}) + \text{e}^-$

Your answer

[1]

5. Concentrated nitric acid, HNO_3 , is an oxidising agent. For example, concentrated HNO_3 reacts with sulfur to form sulfuric acid, nitrogen dioxide and one other product.

- Using oxidation numbers, show the element that is oxidised and the element that is reduced in this reaction. Ensure that the oxidation numbers have signs.
- Construct the balanced equation for this reaction.

element oxidised oxidation number change: from to

element reduced oxidation number change: from to

equation

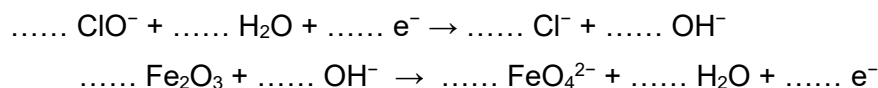
[4]



6. Fe_2O_3 can be oxidised by ClO^- ions under alkaline conditions in a redox reaction.

Unbalanced half-equations for this reaction are shown below.

Balance the half-equations and construct an overall equation for the reaction.



overall equation:

[3]

7. Hydroxylamine, NH_2OH , is a strong reducing agent.

When heated in aqueous solution, NH_2OH reduces Fe^{3+} ions to Fe^{2+} ions.

A student suggests the three possible equations for the reaction, shown below.



The student plans to carry out an investigation to determine which equation is correct.

The method is outlined below.

Stage 1 Using a pipette, add 25.0 cm^3 of $4.32 \times 10^{-2} \text{ mol dm}^{-3}$ NH_2OH to a conical flask.

Add 10 cm^3 of 1 mol dm^{-3} H_2SO_4 to the conical flask followed by an excess of a solution containing $0.0400 \text{ mol dm}^{-3}$ $\text{Fe}^{3+}(\text{aq})$.

Stage 2 Boil the mixture for 5 minutes and allow to cool.

Stage 3 Titrate the cooled mixture with $2.00 \times 10^{-2} \text{ mol dm}^{-3}$ $\text{KMnO}_4(\text{aq})$.



- a. Determine the minimum volume of $0.0400 \text{ mol dm}^{-3} \text{ Fe}^{3+}(\text{aq})$ that the student should plan to use in **Stage 1**.

Explain your reasoning.

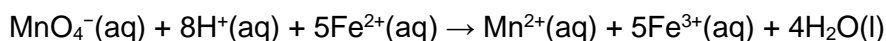
volume = cm^3

explanation:

[4]

- b. In the student's titration, 21.6 cm^3 of $\text{KMnO}_4(\text{aq})$ is required to reach the end point.

The equation that takes place during the titration is shown below.



Analyse the student's results to determine which of the three equations is correct.

Show **all** your working.

[3]



- c. The student intends to repeat the procedure to check their results.

There is insufficient time for the student to repeat all three stages and the student decides to omit **Stage 2**, the boiling stage. Unfortunately the resulting titre is much less than the original titre.

The student rejects the results from the repeated procedure.

- i. Suggest the purpose of the boiling in **Stage 2** and reasons for the second titre being much less than the original titre.

[2]

- ii. The main reason for insufficient time is the need to boil and cool the mixture for each titration.

Suggest how the procedure could be modified so that **Stage 2** does not need to be carried out repeatedly.

Give your reasoning.

[1]

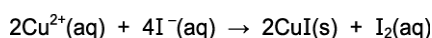
8(a). This question is about redox reactions of transition elements.

Hydrated copper(II) sulfate can crystallise with different numbers of water of crystallisation, x . Compound **A** has the formula $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$.

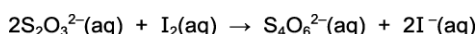
A student carries out an iodine–thiosulfate titration to find the value of x and the formula of Compound **A**. The method is outlined below.

Step 1 A weighed sample of **A** is dissolved in water and made up to 250.0 cm^3 .

Step 2 A 25.00 cm^3 sample of this solution is pipetted into a conical flask, followed by an excess of $\text{KI}(\text{aq})$.



Step 3 The resulting mixture is titrated with $0.120 \text{ mol dm}^{-3} \text{ Na}_2\text{S}_2\text{O}_3(\text{aq})$.





Mass readings

Mass of weighing bottle + **A** = 17.95 g

Mass of weighing bottle = 12.35 g

Titration readings

Titration	Trial	1	2	3
Final burette reading / cm ³	22.50	44.30	22.15	43.85
Initial burette reading / cm ³	0.00	22.50	0.00	22.15
Titre / cm ³				

Table 2.1

- i. Complete **Table 2.1** and calculate the mean titre that the student should use for analysing their results.

mean titre = cm³ [1]

- ii. The uncertainty in each burette reading is ± 0.05 cm³.

Calculate the percentage uncertainty in the titre obtained from **titration 3**.
Give your answer to **two** decimal places.

percentage uncertainty = % [1]

- iii. In **step 2**, why is an excess of KI (aq) added?

[1]

- iv. As the end point of the titration is approached, the student adds a solution to accurately detect the end point.

State the solution and explain the colour change observed at the end point.

[2]



v. Determine the formula of compound **A**.

Show your working.

formula of compound **A** = [4]

(b). A student carries out two experiments based on redox reactions of iron and chromium.

Use the standard electrode potentials below to help you answer the questions that follow.

$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$\text{Fe}(\text{s})$	$E^{\ominus} = -0.44 \text{ V}$
$2\text{H}^{+}(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$\text{H}_2(\text{g})$	$E^{\ominus} = 0.00 \text{ V}$
$\text{Fe}^{3+}(\text{aq}) + \text{e}^{-}$	\rightleftharpoons	$\text{Fe}^{2+}(\text{aq})$	$E^{\ominus} = +0.77 \text{ V}$
$\text{O}_2(\text{g}) + 4\text{H}^{+}(\text{aq}) + 4\text{e}^{-}$	\rightleftharpoons	$2\text{H}_2\text{O}(\text{l})$	$E^{\ominus} = +1.23 \text{ V}$
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^{+}(\text{aq}) + 6\text{e}^{-}$	\rightleftharpoons	$2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	$E^{\ominus} = +1.33 \text{ V}$
$\text{Cl}_2(\text{g}) + 2\text{e}^{-}$	\rightleftharpoons	$2\text{Cl}^{-}(\text{aq})$	$E^{\ominus} = +1.36 \text{ V}$
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^{+}(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$2\text{H}_2\text{O}(\text{l})$	$E^{\ominus} = +1.78 \text{ V}$



For each experiment, identify the species causing the observations shown in bold text and write overall equations for any reactions taking place.

State symbols are **not** required in the equations.

i. **Experiment 1**

1. The student adds iron filings to dilute hydrochloric acid. A **green solution** forms and **gas bubbles** are seen.
2. The student bubbles air through the green solution. The solution turns an **orange-brown colour**.

1:

2:

[3]

ii. **Experiment 2**

The student heats a **green solution** of chromium(III) sulfate with dilute acid and hydrogen peroxide, H_2O_2 .
The solution turns an **orange colour**.

[3]



9(a). Table 21.1 below gives the standard electrode potentials for seven redox systems.

You need to use this information to answer the questions below.

Redox system	Equation	E^\ominus/V
1	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51
2	$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightleftharpoons 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	+1.33
3	$\text{Br}_2(\text{aq}) + 2\text{e}^- \rightleftharpoons 2\text{Br}^-(\text{aq})$	+1.09
4	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80
5	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
6	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
7	$\text{Ce}^{3+}(\text{aq}) + 3\text{e}^- \rightleftharpoons \text{Ce}(\text{s})$	-2.33

Table 21.1

- i. Outline an experimental setup that could be used in the laboratory to measure the standard cell potential of an electrochemical cell based on redox systems 4 and 5.

In your answer you should include details of the apparatus, solutions and the standard conditions required to measure this standard cell potential.

[4]

- ii. An electrochemical cell can be made based on redox systems 2 and 4. The standard cell potential is +0.53 V.

State and explain the effect on the cell potential of this cell if the concentration of silver ions is increased.

[2]



(b). From **Table 21.1**, predict the oxidising agent(s) that **will not** oxidise $\text{Fe}^{2+}(\text{aq})$ to $\text{Fe}^{3+}(\text{aq})$.

[1]

(c). An aqueous solution of iron(II) bromide is mixed with an excess of acidified solution containing manganate(VII) ions.

Using **Table 21.1**, give the formulae of the products of any reactions that take place.

[2]

10(a). Electrochemical cells contain two redox systems, one providing electrons and the other accepting electrons. The tendency to lose or gain electrons is measured using values called standard electrode potentials.

Define the term *standard electrode potential*.

Include all standard conditions in your answer.

[2]

(b). The table below shows two redox systems and their standard electrode potentials, E^\ominus .

Redox system	E^\ominus / V
$\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Cu}(\text{s})$	+0.34
$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80

A standard $\text{Cu}^{2+}(\text{aq}) / \text{Cu}(\text{s})$ half-cell is connected to a standard $\text{Ag}^+(\text{aq}) / \text{Ag}(\text{s})$ half-cell. The potential of the cell is measured.

Water is then added to the $\text{Cu}^{2+}(\text{aq}) / \text{Cu}(\text{s})$ half-cell. This changes the position of equilibrium in the half-cell. The cell potential increases.



- i. Write down the equation for the overall cell reaction.

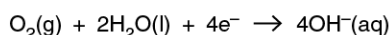
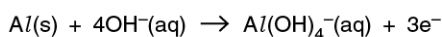
[1]

- ii. Explain, in terms of equilibrium, why the cell potential increases.

[3]

- (c). Aluminium—oxygen cells are being investigated for powering vehicles.

The **reactions** at each electrode are shown below.



- i. The standard electrode potential for the $\text{O}_2 / \text{OH}^{\text{-}}$ redox system is +0.40 V.
The standard cell potential of an aluminium—oxygen cell is 2.71 V.

What is the standard electrode potential of the aluminium redox system in this cell?

standard electrode potential = V [1]

- ii. Construct the overall cell equation for an aluminium—oxygen cell.

[2]



11(a). A student carries out an investigation to prepare and analyse a sample of barium ferrate(VI), BaFeO_4 . The steps in the investigation are shown below.

Step 1

The student adds solid iron(III) oxide to a hot aqueous solution containing an excess of hydroxide ions. The student bubbles chlorine gas through the mixture.

A solution forms containing aqueous ferrate(VI) ions, $\text{FeO}_4^{2-}(\text{aq})$, and aqueous chloride ions.

Step 2

The student adds aqueous barium chloride to the resulting solution.

A precipitate of impure barium ferrate(VI) forms.

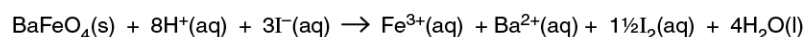
The precipitate is filtered, washed with distilled water and dried.

The student obtains 0.437 g of impure solid barium ferrate(VI).

Step 3

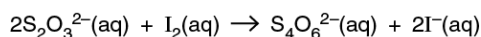
An excess of acidified aqueous potassium iodide is added to the solid from **step 2**.

The BaFeO_4 reacts as shown below, and the impurity does not react. A solution forms containing aqueous iodine, $\text{I}_2(\text{aq})$.



Step 4

The student determines the amount of I_2 formed by carrying out a titration with aqueous sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$.



26.4 cm^3 of 0.100 mol dm^{-3} $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$ are required to reach the end point.

Construct an equation for the oxidation of iron(III) oxide (**step 1**).



(b). Write an **ionic** equation for the formation of barium ferrate(VI) (**step 2**).

Include state symbols.

[1]

(c). In **step 3**, what is the reducing agent?

Explain your answer in terms of electrons.

reducing agent

explanation

[2]

(d). The solid sample of barium ferrate(VI) obtained in step 2 is impure.

Determine the percentage, by mass, of barium ferrate(VI) in the 0.437 g of solid formed in **step 2**.

Give your answer to **one** decimal place.

percentage of barium ferrate(VI) = % [4]



(e). When the solution is not alkaline, ferrate(VI) ions react with water.

The reaction forms a gas with a density of $1.333 \times 10^{-3} \text{ g cm}^{-3}$, measured at room temperature and pressure, and an orange-brown precipitate.

- Determine the formulae of the gas and the precipitate.
- Write an equation for the reaction that takes place.

gas

.....

precipitate

.....

equation

[3]

12. Some electrode potentials for ions are shown below.

$\text{Fe}^{2+}(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$\text{Fe}(\text{s})$	$E^{\circ} = -0.44\text{V}$
$\text{Fe}^{3+}(\text{aq}) + \text{e}^{-}$	\rightleftharpoons	$\text{Fe}^{2+}(\text{aq})$	$E^{\circ} = +0.77\text{V}$
$\frac{1}{2}\text{I}_2(\text{aq}) + \text{e}^{-}$	\rightleftharpoons	$\text{I}^{-}(\text{aq})$	$E^{\circ} = +0.54\text{V}$
$\frac{1}{2}\text{Br}_2(\text{aq}) + \text{e}^{-}$	\rightleftharpoons	$\text{Br}^{-}(\text{aq})$	$E^{\circ} = +1.09\text{V}$
$\frac{1}{2}\text{Cl}_2(\text{aq}) + \text{e}^{-}$	\rightleftharpoons	$\text{Cl}^{-}(\text{aq})$	$E^{\circ} = +1.36\text{V}$

- i. Complete the electron configurations for Fe^{2+} and Br^{-} .

Fe^{2+} : $1s^2$

Br^{-} : $1s^2$

.....
.....

[2]



- ii. Predict the products of reacting Fe(s) separately with I₂(aq), Br₂(aq) and Cl₂(aq).

Explain your predictions using the electrode potential data above.

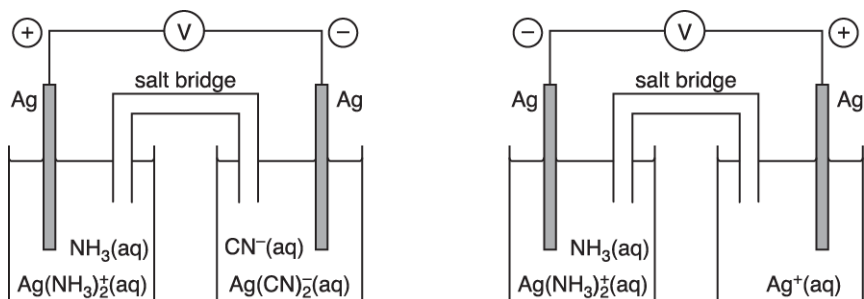
[3]

13(a). Three redox systems, **C**, **D** and **E** are shown in Table 6.1.

C	$\text{Ag}(\text{NH}_3)_2^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + 2\text{NH}_3(\text{aq})$
D	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$
E	$\text{Ag}(\text{CN})_2^-(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s}) + 2\text{CN}^-(\text{aq})$

Table 6.1

The two cells below were set up in an experiment to compare the standard electrode potentials of redox systems **C**, **D** and **E**. The signs on each electrode are shown.





List the three redox systems in order by adding the labels **C**, **D** and **E** to the table below.

E^\ominus	redox system
Most negative	
↑	
Least negative	

[1]

(b). A standard cell is set up between redox system **D** in **Table 6.1** and a standard hydrogen half-cell. The standard cell potential of redox system **D** is +0.34 V.

The cell delivers a current for a length of time.

The pH of the solution in the standard hydrogen half-cell decreases.

- i. What is the pH of the solution in a standard hydrogen half-cell?

pH = [1]

- ii. Explain, in terms of electrode potentials and equilibrium, why the pH of the solution in the hydrogen half-cell decreases as this cell delivers current.

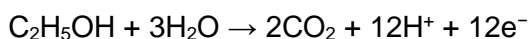
[2]

- iii. Write the equation for the overall cell reaction that takes place in this cell.

[1]

(c). Direct-ethanol fuel cells (DEFCs) are being developed in which the fuel is ethanol rather than hydrogen.

The half-equation for the reaction at the ethanol electrode of the DEFC is shown below:





i. State **one** important difference between a fuel cell and a modern storage cell.

[1]

ii. Suggest **one** advantage of using ethanol, rather than hydrogen, in a fuel cell for vehicles.

[1]

iii. The overall reaction in a DEFC is the same as for the complete combustion of ethanol.

Write the equation for the overall reaction in a DEFC.

[1]

iv. Deduce the half-equation for the reaction at the oxygen electrode in a DEFC.

[1]

v. Using oxidation numbers, show that oxidation and reduction take place in a DEFC.

Oxidation:

.....

Reduction:

.....

[2]



14(a). Chromite is the main ore of chromium. The chromium-containing compound in chromite is $\text{Fe}(\text{CrO}_2)_2$. The percentage of chromium in a sample of chromite can be determined using the method below.

Step 1

A 5.25 g sample of chromite ore is heated with sodium peroxide, Na_2O_2 .



Water is added to the resulting mixture.

Na_2CrO_4 dissolves in the water forming a solution containing CrO_4^{2-} ions.

Step 2

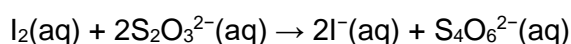
The mixture from **Step 1** is filtered and the filtrate is made up to 1.00 dm^3 in a volumetric flask.

A 25.0 cm^3 sample of this alkaline solution is pipetted into a conical flask and an excess of aqueous potassium iodide is added.

- A redox reaction takes place between I^- ions, CrO_4^{2-} ions and H_2O .
- In this reaction 1 mol CrO_4^{2-} forms 1.5 mol I_2 .

Step 3

The resulting mixture is titrated with $0.100 \text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$ to estimate the I_2 present:



The average titre of $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$ is 25.5 cm^3 .

In **Step 1** Na_2O and NaFeO_2 react with water forming an alkaline solution containing a brown precipitate. This is **not** a redox reaction.

Write equations for:

- the reaction of Na_2O with water
- the reaction of NaFeO_2 with water.



(b). Determine the percentage, by mass, of chromium in the ore.

Give your answer to **one** decimal place.

[6]

(c). This part refers to **Step 2** of the method.

In the redox reaction between I^- ions, CrO_4^{2-} ions and H_2O :

- CrO_4^{2-} ions, are reduced to chromium(III) ions, Cr^{3+}
- I^- ions are oxidised to iodine, I_2
- Construct an overall equation for the redox reaction and write half equations for the oxidation and reduction.

Overall equation:

Half equations:

[3]



15. Two redox reactions of iodine are described below.

Reaction 1: Iodine is reacted with oxygen to form a compound with a molar mass of 333.8 g mol⁻¹.

Reaction 2: In alkaline conditions, iodine disproportionates to form iodide ions, iodate(V) ions and water.

Construct equations for these **two** reactions.

State symbols are **not** required.

Reaction 1:

Reaction 2:

[3]

16(a). Redox reactions can be used to generate electrical energy from electrochemical cells.

A student investigates the redox systems shown in **Table 6.1** below.

	Redox system	E^\ominus/V
1	$A^{\beta+}(aq) + 3e^- \rightleftharpoons A(l)$	-1.66
2	$Cr^{3+}(aq) + e^- \rightleftharpoons Cr^{2+}(aq)$	-0.41
3	$V^{3+}(aq) + e^- \rightleftharpoons V^{2+}(aq)$	-0.26
4	$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
5	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
6	$FeO_4^{2-}(aq) + 8H^+(aq) + 3e^- \rightleftharpoons Fe^{3+}(aq) + 4H_2O(l)$	+2.20

Table 6.1



Define the term *standard electrode potential*.

Include all standard conditions in your answer.

[2]

(b). The student sets up a standard cell based on redox systems **3** and **4**.

- i. Draw a labelled diagram to show how the student could have set up this cell to measure the standard cell potential.

On your diagram,

- show the charge carriers in the circuit joining the two half cells
- label the signs of the electrodes.

[5]

- ii. What is the standard cell potential of this cell?

standard cell potential = V [1]



17. In some gold mines, cyanide is used to extract gold from its ore.

Gold metal in the ore reacts with cyanide ions, water and oxygen to form a water-soluble complex ion, $[\text{Au}(\text{CN})_2]^-$, with a bond angle of 180° . Hydroxide ions are also formed.

i. Name the shape of $[\text{Au}(\text{CN})_2]^-$.

[1]

ii. Using oxidation numbers, show that a redox reaction takes place.

[2]

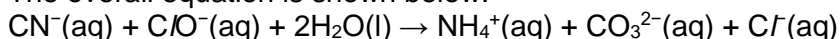
iii. Construct the overall equation for this reaction.

State symbols are **not** required.

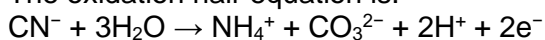
[2]

iv. Some owners of gold mines remove cyanide ions from waste by adding oxidising agents, such as chlorate(I) ions, before discharge into watercourses.

The overall equation is shown below.



The oxidation half-equation is:



Construct the reduction half-equation.

State symbols are **not** required.

[1]



18(a). Hydrated copper(II) methanoate, $\text{Cu}(\text{HCOO})_2 \cdot x\text{H}_2\text{O}$, is a copper salt.

A student carries out the procedure below to prepare $\text{Cu}(\text{HCOO})_2 \cdot x\text{H}_2\text{O}$ and to determine the value of x in its formula.

Step 1

The student prepares $\text{Cu}(\text{HCOO})_2 \cdot x\text{H}_2\text{O}$ by reacting a copper compound with aqueous methanoic acid to form $\text{Cu}(\text{HCOO})_2(\text{aq})$ and allowing the solvent to evaporate.

Step 2

The student dissolves 2.226 g of $\text{Cu}(\text{HCOO})_2 \cdot x\text{H}_2\text{O}$ in water and makes up the solution to 250.0 cm^3 .

Step 3

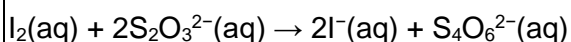
Using a pipette, the student adds 25.0 cm^3 of this solution to a conical flask followed by an excess of $\text{KI}(\text{aq})$.

The $\text{Cu}^{2+}(\text{aq})$ ions react to form a precipitate of copper(I) iodide and $\text{I}_2(\text{aq})$.

In this reaction, 2 mol Cu^{2+} form 1 mol I_2 .

Step 4

The student titrates the iodine in the resulting mixture with 0.0420 mol dm^{-3} $\text{Na}_2\text{S}_2\text{O}_3$ (aq).



23.5 cm^3 0.0420 mol dm^{-3} $\text{Na}_2\text{S}_2\text{O}_3$ (aq) is required to reach the end point.

Complete the electron configuration of copper in

$\text{Cu}(\text{HCOO})_2 \cdot x\text{H}_2\text{O}$: $1s^2$

copper(I) iodide: $1s^2$ [2]



(b). Choose a suitable copper compound for **step 1**, and write the full equation for the reaction that would take place to form $\text{Cu}(\text{HCOO})_2(\text{aq})$.

State symbols are **not** required.

[1]

(c). Write an ionic equation, including state symbols, for the reaction in **step 3**.

[1]

(d). In **step 4**, the student adds a solution to observe the end point accurately.

Name the solution and state the colour change at the end point.

Solution added:.....

Colour change: [2]

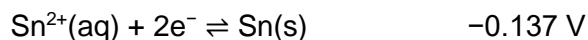
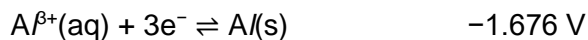
(e). Determine the value of x in $\text{Cu}(\text{HCOO})_2 \cdot x\text{H}_2\text{O}$.

Show your working.

[5]



19. Electrode potentials are given below.



A standard cell is constructed from $\text{Al}^{3+}(\text{aq})|\text{Al}(\text{s})$ and $\text{Sn}^{2+}(\text{aq})|\text{Sn}(\text{s})$ half cells.

Which statement is correct for the standard cell?

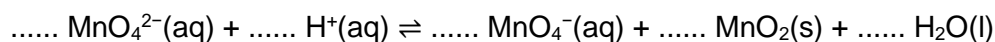
- A Al is oxidised and the cell potential is 1.539 V.
- B Sn is oxidised and the cell potential is 1.539 V.
- C Al is oxidised and the cell potential is 1.813 V.
- D Sn is oxidised and the cell potential is 1.813 V.

Your answer

[1]

20(a). When heated with dilute acid, $\text{MnO}_4^{2-}(\text{aq})$ ions disproportionate into MnO_4^{-} and MnO_2 .

- i. Balance the equation for this disproportionation reaction.



[1]

- ii. Although $\text{MnO}_4^{2-}(\text{aq})$ ions disproportionate in acidic conditions, $\text{MnO}_4^{2-}(\text{aq})$ ions are stable under alkaline solutions.

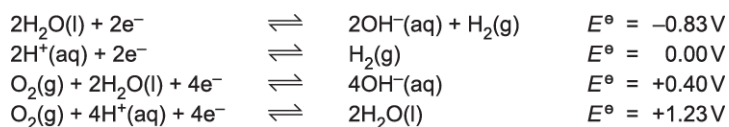
Explain this difference in stability, in terms of equilibrium.

[2]



(b). Hydrogen fuel cells generate energy from redox reactions. Hydrogen fuel cells have been developed to operate in either an acidic or alkaline electrolyte.

Standard electrode potentials for several redox systems are shown below.



Using this information, show how acidic and alkaline hydrogen fuel cells produce the same overall cell equation and cell potential, despite different reactions taking place at each electrode.

[4]

21. A student investigates reactions of cobalt ions, as outlined below.

- A student dissolves cobalt(II) chloride in water. A pink solution forms containing the hexaaqua complex ion **B**.
- The student adds an excess of concentrated ammonia solution to the pink solution until there is no further change.
- A pale brown solution forms which contains the complex ion $[\text{Co}(\text{NH}_3)_6]^{2+}$.

i. Write the equation for the formation of $[\text{Co}(\text{NH}_3)_6]^{2+}$ from complex ion **B**.

State the type of reaction.

Equation

Type of reaction

[2]



- ii. Draw a 3-D diagram of the $[\text{Co}(\text{NH}_3)_6]^{2+}$ ion.

On your diagram, show the value of the bond angles involving Co.

[2]

- iii. A solution containing $[\text{Co}(\text{NH}_3)_6]^{2+}$ is reacted as outlined below.
- The solution is warmed with aqueous hydrogen peroxide, $\text{H}_2\text{O}_2(\text{aq})$. The H_2O_2 oxidises cobalt(II) to cobalt(III), to form a red-brown solution containing a six-coordinate complex ion **C**.
 - Concentrated hydrochloric acid is added to the red-brown solution. Yellow crystals of a complex **D** are formed.

Complex **D** has the percentage composition by mass:

Co, 22.03%; N, 31.41%, H, 6.73%; Cl, 39.83%.

Determine the formulae of **C** and **D**, showing clearly the ligands and any charges.

Show all your working.

[4]



- iv. Write half equations and an overall equation for the oxidation of $[\text{Co}(\text{NH}_3)_6]^{2+}$ to **C** by hydrogen peroxide in (iii).

Half equations

Overall equation

[3]

22. Concentrated sulfuric acid is a powerful oxidising agent. Concentrated sulfuric acid oxidises hydrogen iodide, HI, to form iodine, hydrogen sulfide and one other product.

Construct an equation for this reaction.

[2]

23. A cell is constructed from the two redox systems below.



Which statement(s) is/are correct for the cell?

1. The cell potential is 1.14 V.
2. The reaction at the copper electrode is $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^{-}$.
3. The silver electrode increases in mass.

- A** 1, 2 and 3
B Only 1 and 2
C Only 2 and 3
D Only 1

Your answer

[1]



24(a). This question is about redox, electrode potentials and feasibility.

Table 22.1 shows standard electrode potentials for four redox systems.

You need to use this information to answer the questions below.

Redox system	Equation	E^\ominus / V
1	$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{Zn}(\text{s})$	-0.76
2	$\text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	+0.17
3	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$	+0.77
4	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51

Table 22.1

A standard cell is set up in the laboratory based on redox systems **1** and **3** and the standard cell potential is measured.

- i. Draw a labelled diagram to show how this cell could be set up to measure its standard cell potential.

Include details of the apparatus, solutions and the standard conditions required to measure this standard cell potential.

Standard conditions

[4]



ii. Predict the standard cell potential of this cell.

standard cell potential =

V
[1]

(b). In **Table 22.1**, what is the strongest reducing agent and the strongest oxidising agent?

Strongest reducing agent

Strongest oxidising agent

[2]

(c). Electrode potentials can be used to predict the feasibility of reactions.

Construct an overall equation for the predicted reaction between the species in redox systems **2** and **4**.

[2]



25. Sodium oxide, Na_2O , can be prepared by the redox reaction of NaNO_2 and sodium metal. Nitrogen gas is also formed.

i. What is the systematic name for NaNO_2 ?

[1]

ii. Using oxidation numbers, with signs, show the element that is oxidised and the element that is reduced in this reaction.

Element oxidised

Oxidation number change from

to

Element reduced

Oxidation number change from

to

[2]

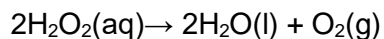
iii. Construct the equation for this reaction.

Equation

[1]



26. Solutions of hydrogen peroxide decompose slowly into water and oxygen:



This reaction is catalysed by manganese dioxide, $\text{MnO}_2(\text{s})$.

Standard electrode potentials are shown below.



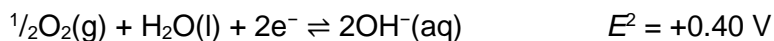
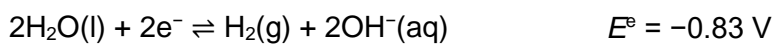
Using the electrode potentials, explain how MnO_2 is able to act as a catalyst for the decomposition of hydrogen peroxide.

Your answer should include relevant equations.

[4]



27. The redox equilibria for a hydrogen–oxygen fuel cell in alkaline solution are shown below.



What is the equation for the overall cell reaction?

- A $\text{H}_2(\text{g}) + 4\text{OH}^-(\text{aq}) \rightarrow 3\text{H}_2\text{O}(\text{l}) + \frac{1}{2}\text{O}_2(\text{g})$
- B $3\text{H}_2\text{O}(\text{l}) + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2(\text{g}) + 4\text{OH}^-(\text{aq})$
- C $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g})$
- D $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$

Your answer

[1]

28. Which enthalpy change(s) is/are endothermic?

- 1 The bond enthalpy of the C–H bond
- 2 The second electron affinity of oxygen
- 3 The standard enthalpy change of formation of magnesium

- A 1, 2 and 3
- B Only 1 and 2
- C Only 2 and 3
- D Only 1

Your answer

[1]



29. Compound **B** is an iodate(V) salt of a Group 1 metal.

The iodate(V) ion has the formula IO_3^- .

A student carries out a titration to find the formula of compound **B**.

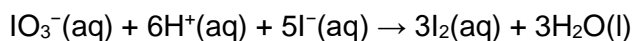
Step 1:

The student dissolves 1.55 g of **B** in water and makes up the solution to 250.0 cm^3 in a volumetric flask.

The student pipettes 25.00 cm^3 of the solution of **B** into a conical flask, followed by 10 cm^3 of dilute sulfuric acid and an excess of $\text{KI}(\text{aq})$.

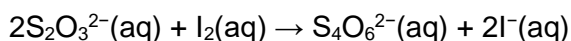
Step 2:

The iodate(V) ions are reduced to iodine, as shown below.



Step 3:

The resulting mixture is titrated with 0.150 mol dm^{-3} $\text{Na}_2\text{S}_2\text{O}_3(\text{aq})$.



The student repeats **step 2** and **step 3** until concordant titres are obtained.

Titration readings

Titration	Trial	1	2	3
Final burette reading / cm^3	24.00	47.40	23.75	47.05
Initial burette reading / cm^3	0.00	24.00	0.00	23.20
Titre / cm^3				

Table 20.1

- i. Complete **Table 20.1** and calculate the mean titre that the student should use for analysing the results.

mean titre = cm^3 [2]



- ii. The uncertainty in each burette reading is $\pm 0.05 \text{ cm}^3$.

Calculate the percentage uncertainty in the titre obtained from **titration 1**.

Give your answer to **two** decimal places.

percentage uncertainty = % [1]

- iii. Describe and explain how the student should determine the end point of this titration accurately.

[2]

- iv. Determine the relative formula mass and formula of the Group 1 iodate(V), **B**.

Show your working.

relative formula mass of **B** =

formula of **B** = [5]



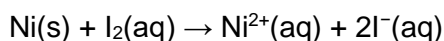
30(a). This question is about some reactions of d block elements and their ions.

Table 21.1 shows standard electrode potentials which will be needed within this question.

$\text{Zn}^{2+}(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$\text{Zn}(\text{s})$	$E^{\ominus} = -0.76\text{V}$
$\text{Cr}^{3+}(\text{aq}) + \text{e}^{-}$	\rightleftharpoons	$\text{Cr}^{2+}(\text{aq})$	$E^{\ominus} = -0.42\text{V}$
$\text{Ni}^{2+}(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$\text{Ni}(\text{s})$	$E^{\ominus} = -0.25\text{V}$
$\text{I}_2(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$2\text{I}^{-}(\text{aq})$	$E^{\ominus} = +0.54\text{V}$
$\text{Fe}^{3+}(\text{aq}) + \text{e}^{-}$	\rightleftharpoons	$\text{Fe}^{2+}(\text{aq})$	$E^{\ominus} = +0.77\text{V}$
$\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^{+}(\text{aq}) + 6\text{e}^{-}$	\rightleftharpoons	$2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$	$E^{\ominus} = +1.33\text{V}$
$\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^{+}(\text{aq}) + 2\text{e}^{-}$	\rightleftharpoons	$2\text{H}_2\text{O}(\text{l})$	$E^{\ominus} = +1.78\text{V}$

Table 21.1

A standard cell is set up in the laboratory with the cell reaction shown below.



- i. Draw a labelled diagram to show how this cell could be set up to measure its standard cell potential.

Include details of apparatus, solutions and the standard conditions required.

Standard conditions

[4]

- ii. Predict the standard cell potential of this cell.

standard cell potential = V **[1]**



(b). Use the information in **Table 21.1** to help you answer both parts of this question.

i. Write the overall equation for the oxidation of Fe^{2+} by acidified H_2O_2 .

[1]

ii. Zinc reacts with acidified $\text{Cr}_2\text{O}_7^{2-}$ ions to form Cr^{2+} ions in two stages.

Explain why this happens in terms of electrode potentials and equilibria.

Include overall equations for the reactions which occur.

[4]

(c). * 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.

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.

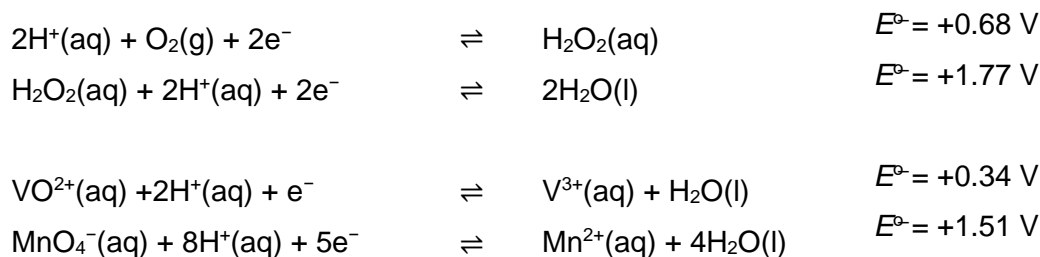
Reaction 3: Copper(II) oxide reacts with warm dilute nitric acid in a neutralisation reaction, to form a blue solution. Unreacted copper(II) oxide is filtered off, and the solution is left overnight in an evaporating basin.

A hydrated salt, **F**, crystallises, with the percentage composition by mass:
Cu, 26.29%; H, 2.48%; N, 11.59%; O, 59.63%.



31. Hydrogen peroxide can act as an oxidising agent or as a reducing agent.

Some standard electrode potentials are shown below.



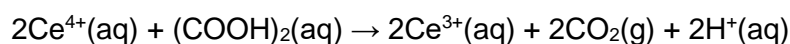
Use this information to write an equation for a reaction in which hydrogen peroxide acts as a reducing agent.

[2]

32. Ethanedioic acid, $(\text{COOH})_2$, is present in rhubarb leaves.

A student carries out a redox titration using aqueous cerium(IV) sulfate, $\text{Ce}(\text{SO}_4)_2(\text{aq})$, to determine the percentage, by mass, of ethanedioic acid in rhubarb leaves.

In the titration, $\text{Ce}^{4+}(\text{aq})$ ions oxidise ethanedioic acid in hot acid conditions:



$\text{Ce}^{4+}(\text{aq})$ ions have a yellow colour. $\text{Ce}^{3+}(\text{aq})$ ions are colourless.

The student weighs 82.68 g of rhubarb leaves and extracts ethanedioic acid from the leaves.

The ethanedioic acid is added to dilute sulfuric acid to form a colourless solution which is made up to 250.0 cm^3 with distilled water.

The student heats 25.00 cm^3 of this solution to 70 °C and titrates this volume with 0.0500 mol dm^{-3} $\text{Ce}(\text{SO}_4)_2$ from the burette.



The student repeats the titration to obtain concordant (consistent) titres.

Titration results

The trial titre has been omitted.

	1	2	3
Final reading/cm ³	24.30	47.80	23.65
Initial reading/cm ³	1.05	24.30	0.50

- i. This titration is self-indicating and the student does not need to add an indicator.

What colour change would the student observe at the end point?

[1]

Colour change from _____ to _____

- ii. Calculate the percentage, by mass, of ethanedioic acid in the rhubarb leaves.

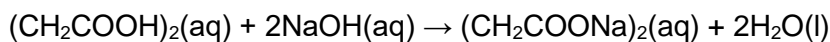
Give your answer to an **appropriate** number of significant figures.

percentage of ethanedioic acid = _____ % [6]



33. A student carries out two experiments in the laboratory based on succinic acid (butanedioic acid), $(\text{CH}_2\text{COOH})_2$.

Aqueous succinic acid can be neutralised by aqueous sodium hydroxide, $\text{NaOH}(\text{aq})$:



This reaction can be used to determine a value for the enthalpy change of neutralisation, $\Delta_{\text{neut}}H$.

The student follows this method:

- Add 50.0 cm^3 of $0.400 \text{ mol dm}^{-3}$ succinic acid to a polystyrene cup.
- Measure out 50.0 cm^3 of 1.00 mol dm^{-3} $\text{NaOH}(\text{aq})$, which is in excess.
- Measure the temperature of both solutions.
- Add the $\text{NaOH}(\text{aq})$ to the aqueous succinic acid in the polystyrene cup, stir the mixture, and record the maximum temperature.

Temperature readings

Maximum temperature of mixture/ $^{\circ}\text{C}$	26.5
Initial temperature of both solutions/ $^{\circ}\text{C}$	21.5

Calculate a value for the enthalpy change of neutralisation, $\Delta_{\text{neut}}H$, in kJ mol^{-1} .

Assume that the density of all solutions and the specific heat capacity, c , of the reaction mixture are the same as for water.

$\Delta_{\text{neut}}H =$

kJ mol^{-1}
[4]



34. Four redox systems relevant to hydrogen–oxygen fuel cells are shown below.

	E° / V
$\text{H}_2\text{O}(\text{l}) + \text{e}^- \rightleftharpoons \text{OH}^-(\text{aq}) + \frac{1}{2}\text{H}_2(\text{g})$	-0.83
$\text{H}^+(\text{aq}) + \text{e}^- \rightleftharpoons \frac{1}{2}\text{H}_2(\text{g})$	0.00
$\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) + 2\text{e}^- \rightleftharpoons 2\text{OH}^-(\text{aq})$	+0.40
$\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}(\text{l})$	+1.23

Which statement(s) is/are correct for an alkaline hydrogen–oxygen fuel cell?

- 1 The reaction at the positive electrode is: $\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2\text{O}(\text{l})$
- 2 The overall cell reaction is: $\text{H}_2(\text{g}) + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$.
- 3 The cell potential is 1.23 V.

- A** 1, 2 and 3
B Only 1 and 2
C Only 2 and 3
D Only 1

Your answer

[1]



35. A student is provided with a sample of a metal **M**.

The student analyses metal **M** using a 'back-titration' technique:

- The metal is reacted with excess acid.
- The resulting solution is titrated to determine the amount of acid remaining after the reaction.

Stage 1

The student adds 100 cm³ of 2.10 mol dm⁻³ HCl (aq) to 6.90 g of **M**.

An excess of HCl (aq) has been used to ensure that all of metal **M** reacts.

A redox reaction occurs, forming a solution containing **M** in the +2 oxidation state.

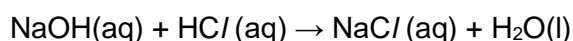
Stage 2

The resulting solution from **Stage 1** is made up to 250.0 cm³ with distilled water.

Stage 3

A 25.00 cm³ sample of the diluted solution from **Stage 2** is titrated with 0.320 mol dm⁻³ NaOH(aq).

The NaOH(aq) reacts with excess HCl (aq) that remains in **Stage 1**:



The student repeats the titration to obtain concordant titres.

Titration results (The trial titre has been omitted.)

The burette readings have been recorded to the nearest 0.05 cm³.

	1	2	3
Final reading / cm³	27.80	37.55	32.20
Initial reading / cm³	0.50	10.00	5.00

- i. In **Stage 1**, a redox reaction takes place between **M** and HCl (aq), forming hydrogen and a solution containing **M** in the +2 oxidation state.

Write an overall equation, with state symbols, for this reaction. Write half-equations for the oxidation and reduction processes.



Overall
equation

Oxidation half-
equation

Reduction half-
equation

[3]

- ii. In **Stage 1**, suggest **two** observations that would confirm that all of metal **M** has reacted.

1

2

[2]

- iii. In **Stage 3**, write the ionic equation for the reaction taking place in the titration.

[1]

- iv. Metal **M** can be identified following the steps below.

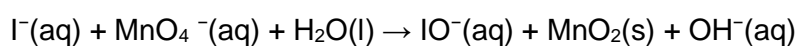
1. The amount, in mol, of excess $\text{HCl}(\text{aq})$ that remains after the reaction of **M** with $\text{HCl}(\text{aq})$.
2. The amount, in mol, of $\text{HCl}(\text{aq})$ that reacted with **M**.
3. The identity of metal **M**.

Analyse the results to identify metal **M**.



Metal **M** = [6]

36. Iodide ions, $\text{I}^{-}(\text{aq})$, react with $\text{MnO}_4^{-}(\text{aq})$. The unbalanced equation is shown below.



What is the ratio of $\text{MnO}_2(\text{s})$ to $\text{OH}^{-}(\text{aq})$ in the balanced equation?

- A** 1 : 3
- B** 1 : 2
- C** 1 : 1
- D** 3 : 2

Your answer

[1]



37(a). Standard electrode potentials for four redox systems are shown in **Table 19.1**.

Redox system	Half-equation	E^\ominus / V
1	$\text{CO}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{HCOOH}(\text{aq})$	-0.11
2	$\text{HCOOH}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{HCHO}(\text{aq}) + \text{H}_2\text{O}(\text{l})$	-0.03
3	$\text{Ag}^+(\text{aq}) + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80
4	$\text{MnO}_4^-(\text{aq}) + 8\text{H}^+(\text{aq}) + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+1.51

Table 19.1

A student sets up a standard cell in the laboratory based on redox systems **3** and **4**.

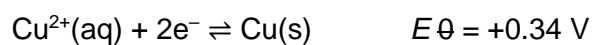
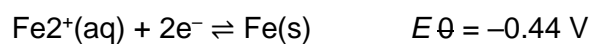
Draw a labelled diagram to show how this cell could be set up to measure its standard cell potential at 298 K.

[3]



[5]

39. Two redox systems are shown below.



Which species in the two redox systems is the strongest oxidising agent?

- A $\text{Fe}^{2+}(\text{aq})$
- B $\text{Fe}(\text{s})$
- C $\text{Cu}^{2+}(\text{aq})$
- D $\text{Cu}(\text{s})$

Your answer

[1]



40(a). Storage cells and fuels cells are types of electrochemical cell.

The electrode potentials for five redox systems are shown in **Table 19.1**.

Redox system	Half-equation	E^{\ominus} / V
1	$\text{Cr}^{3+}(\text{aq}) + 3\text{e}^{-} \rightleftharpoons \text{Cr}(\text{s})$	-0.74
2	$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^{-} \rightleftharpoons 4\text{OH}^{-}(\text{aq})$	+0.40
3	$\text{MnO}_4^{-}(\text{aq}) + \text{e}^{-} \rightleftharpoons \text{MnO}_4^{2-}(\text{aq})$	+0.56
4	$\text{MnO}_4^{-}(\text{aq}) + 8\text{H}^{+}(\text{aq}) + 5\text{e}^{-} \rightleftharpoons \text{Mn}^{2+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l})$	+0.51
5	$\text{MnO}_4^{2-}(\text{aq}) + 4\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightleftharpoons \text{MnO}_2(\text{s}) + 2\text{H}_2\text{O}(\text{l})$	+1.70

Table 19.1

A student sets up an electrochemical cell based on redox systems **1** and **4** in **Table 19.1**.

- i. Draw a labelled diagram to show how this cell could be set up in the laboratory.

[3]

- ii. Construct the equation for the overall cell reaction.

[1]

(b). In acid conditions, $\text{MnO}_4^{2-}(\text{aq})$ disproportionates to form $\text{MnO}_2(\text{s})$ and $\text{MnO}_4^{-}(\text{aq})$.

- i. Explain, in terms of oxidation numbers, why disproportionation has taken place.

[2]

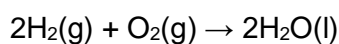


- ii. Explain, in terms of electrode potentials and equilibrium shifts why $\text{MnO}_4^{2-}(\text{aq})$ disproportionates in acid conditions. Use the information in **Table 19.1**.

[2]

- (c). An **alkaline** hydrogen-oxygen fuel cell is set up.

The overall equation for the cell reaction is shown below.



Redox system **2** in **Table 19.1** is the positive electrode of this cell.

- i. Write the half-equation at the negative electrode.

[1]

- ii. The cell potential is 1.23 V.

Calculate the electrode potential of the negative electrode.

electrode potential = V [1]

- iii. State **one** important feature of a fuel cell that is different from a conventional storage cell.

[1]



41(a). A student carries out an experiment to determine the percentage by mass of copper in an ore containing copper in its +2 oxidation state.

The student is provided with a sample of the copper ore, 1 mol dm^{-3} potassium iodide, KI(aq) , and $0.0200 \text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$.

The student's method is outlined below.

- Step 1** Add an excess of warm nitric acid to 2.50 g of the ore.
The copper(II) compounds in the ore react, forming aqueous copper(II) nitrate.
- Step 2** Filter the mixture to remove the unreacted rock. Neutralise the filtrate.
- Step 3** Add an excess of aqueous potassium iodide, KI(aq) .
A precipitate of copper(I) iodide and a solution of iodine, $\text{I}_2(\text{aq})$, forms.
- Step 4** Titrate the mixture from **Step 3** using $0.0200 \text{ mol dm}^{-3}$ sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$ in the burette.
- $$\text{I}_2(\text{aq}) + 2\text{S}_2\text{O}_3^{2-}(\text{aq}) \rightarrow 2\text{I}^{-}(\text{aq}) + \text{S}_4\text{O}_6^{2-}(\text{aq})$$
- 26.55 cm^3 of $0.0200 \text{ mol dm}^{-3}$ $\text{Na}_2\text{S}_2\text{O}_3$ are required to reach the end point.

In **Step 1**, the student observed that bubbles of gas were produced.

Suggest the formula of the copper(II) compound which reacted with HNO_3 to form the gas, and write a full equation for the reaction.

Formula:

Equation: [2]

(b). Write an **ionic** equation, including state symbols, for the reaction in **Step 3**.

[1]

(c). Suggest a suitable indicator for this titration and state the colour change at the end point in **Step 4**.

Indicator:

Colour from _____ to _____ [1]



(d). Determine the percentage, by mass, of copper in the copper ore.

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

percentage = % [4]

(e). Explain whether the calculated percentage by mass of copper would be higher, lower or the same if the following changes were made to the method.

i. The potassium iodide was not in excess, in **Step 3**.

[1]

ii. The burette readings were read from the top of the meniscus, in **Step 4**.

[1]

(f). The student then modifies the method in order to obtain a more accurate value for the percentage by mass of copper in the ore. The student decides to use 25.00 g of the copper ore in **Step 1**.

What further modifications should the student make to produce a more accurate value for the percentage by mass of copper in the ore?

[2]



Construct the equation for this reaction. State symbols are not required.

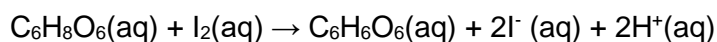
[2]

43. A student carries out an investigation on vitamin C, $C_6H_8O_6$.

The label on a carton of orange juice lists the mass of vitamin C, in mg, in a typical serving of 150 cm^3 .

The student carries out an investigation to check the vitamin C content in the orange juice.

Vitamin C can be oxidised by iodine:



The student dilutes 150 cm^3 of the orange juice with water to 250.0 cm^3 in a volumetric flask.

The student then titrates 25.0 cm^3 volume of this solution with $9.60 \times 10^{-4}\text{ mol dm}^{-3}$ iodine solution, $I_2(aq)$.

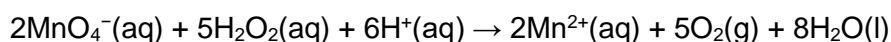
The mean titre of $I_2(aq)$ is 22.50 cm^3 .

Determine the mass, in mg, of vitamin C in a 150 cm^3 serving of the orange juice.

mass of vitamin C in the 150 cm^3 serving of orange juice = mg [4]



44. Hydrogen peroxide, H_2O_2 , can be oxidised by manganate(VII) ions under acid conditions as shown below.



In a titration, 25.00 cm^3 of a disinfectant containing hydrogen peroxide reacts with 22.00 cm^3 of $0.125 \text{ mol dm}^{-3} \text{ KMnO}_4(\text{aq})$.

What is the concentration of H_2O_2 , in mol dm^{-3} , in the disinfectant?

Assume that KMnO_4 only reacts with H_2O_2 in the disinfectant.

- A 0.0440
- B 0.110
- C 0.275
- D 0.550

Your answer

[1]

45. This question is about the reactions of Group 2 metals and their compounds.

A student adds magnesium to dilute hydrochloric acid in one test tube.

The student adds calcium to dilute hydrochloric acid in a second test tube.

A redox reaction takes place in each test tube.

- i. Suggest **two** observations from the student's experiment that would show that calcium is more reactive than magnesium.

1

2

[1]



- ii. Write half-equations for the reaction of magnesium with hydrochloric acid.

Oxidation half-equation:

Reduction half-equation:

[2]

46. This question is about reactions of transition metal compounds.

Aqueous sodium hydroxide is added to an aqueous solution of iron(II) sulfate.

A pale green precipitate forms which turns brown when left to stand in air.

- i. Write an ionic equation for the formation of the pale green precipitate.
- ii. Use the information below to explain why the pale green precipitate turns brown when left to stand in air and construct an equation for the reaction which occurs.

[1]

Redox System	Equation	E°/V
1	$\text{Fe}(\text{OH})_3(\text{s}) + \text{e}^- \rightleftharpoons \text{Fe}(\text{OH})_2(\text{s}) + \text{OH}^-(\text{aq})$	-0.56V
2	$\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightleftharpoons 4\text{OH}^-(\text{aq})$	+0.40V

[4]



47. The equations show the electrode potentials of the half-cells used in a lithium-ion cell.

	E^\ominus / V
$\text{Li}^+ + \text{e}^- \rightleftharpoons \text{Li}$	-3.04
$\text{Li}^+ + \text{CoO}_2 + \text{e}^- \rightleftharpoons \text{LiCoO}_2$	+1.16

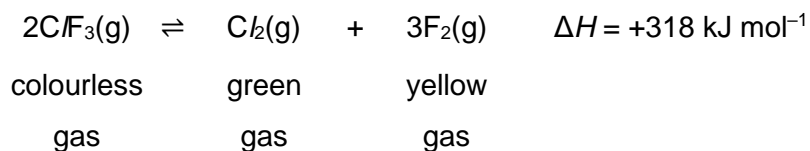
Which statement is correct in a lithium-ion cell?

- A The cell potential is 2.88 V.
- B The reaction at the positive electrode is: $\text{LiCoO}_2 \rightarrow \text{Li}^+ + \text{CoO}_2 + \text{e}^-$
- C The overall cell reaction is: $\text{Li} + \text{CoO}_2 \rightarrow \text{LiCoO}_2$
- D The oxidation number of Co changes from +2 to +1.

Your answer

[1]

48. Chlorine trifluoride can be decomposed into its elements forming the equilibrium mixture below.



Which statement(s) is/are correct?

- 1 The decomposition is a redox reaction.
 - 2 When the equilibrium mixture is cooled, the colour fades.
 - 3 The decomposition has a negative entropy change.
- A 1, 2 and 3
 - B Only 1 and 2
 - C Only 2 and 3
 - D Only 1

Your answer

[1]

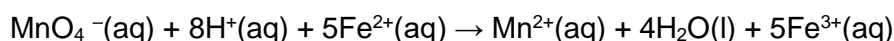


49(a). Some grass fertilisers contain compounds of iron.

During heavy rain, a fertiliser is washed into a nearby river causing the water to be polluted with a mixture of iron(II) and iron(III) ions.

A student determines the concentration of iron(II) ions in a sample of river water by titration with potassium manganate(VII).

25.0 cm³ portions of river water are acidified with dilute sulfuric acid. Each portion is titrated with 0.00250 mol dm⁻³ potassium manganate(VII) until a colour change is seen.



i. State the colour change seen at the end point of the titration.

from to

[1]

ii. The student's titration results are shown in the table below.
The trial titre has been omitted.

	1	2	3
Final volume / cm³	12.65	25.60	38.35
Initial volume / cm³	0.00	12.65	25.60
Titre volume / cm³

Complete the table above and calculate the mean titre that the student should use to determine the concentration of iron(II) ions in the river water.

mean titre = cm³ **[2]**



iii. Determine the concentration, in mol dm⁻³, of iron(II) ions in the river water.

concentration = mol dm⁻³ [3]

(b). The student modifies the experiment in (a) to determine the combined concentration of iron(II) and iron(III) ions in the river water.

The student's method is shown below.

Step 1 Add excess zinc to a 250.0 cm³ sample of river water and warm gently.

Step 2 Cool the solution and remove excess zinc by filtration.

Step 3 Acidify 25.0 cm³ portions of the filtrate from **Step 2**. Then titrate each portion with 0.00250 mol dm⁻³ potassium manganate(VII) until a colour change is seen.

The table below shows information about three redox systems.

Redox system	Half-equation	E° / V
1	$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	-0.76
2	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
3	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51



Use the information in the table above to explain the reasons for **Step 1** and **Step 2**.

Reason(s) for **Step 1**

Reason(s) for **Step 2**

[4]

50. An acidified solution containing $\text{Cr}_2\text{O}_7^{2-}$ ions reacts with vanadium(III) ions in a redox reaction to form a solution containing Cr^{3+} ions and VO_2^+ ions.

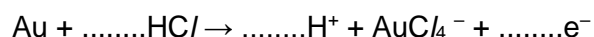
Construct the overall equation for this reaction.

[2]

51(a). A mixture of concentrated nitric and hydrochloric acid is called 'aqua regia'. Aqua regia can dissolve gold.

The reaction of aqua regia with gold is a redox reaction which forms chlorauric acid, HAuCl_4 .

- i. Balance the half-equation for the oxidation process in this reaction.



[1]



- ii. In the reduction process in this reaction, HNO_3 and H^+ react together to form 2 oxides: **X** ($M_r = 30$) and **Z** ($M_r = 18$).

Determine the formulae of **X** and **Z** and write the half-equation for this reduction.

X =

Z =

half-equation

[3]

(b). Three reactions involving sulfuric acid are shown below.

Reaction 1

Dilute sulfuric acid is reacted with nickel(II) hydroxide to form a green solution.

The solvent is allowed to evaporate leaving hydrated crystals of compound **D**, with the percentage composition by mass: Ni, 22.33%; S, 12.20%; O, 60.87%; H, 4.60%.

Reaction 2

Concentrated sulfuric acid is reacted with hydrogen bromide, HBr , to form three products:

- an element which exists as diatomic molecules
- a gaseous compound **E**
- a liquid.

At RTP, 1.00 dm^3 of compound **E** has a mass of 2.67 g.

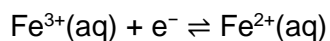
Reaction 3

Concentrated sulfuric acid acts as a catalyst when 2-hydroxypropanoic acid reacts to form compound **F** ($M_r = 144$).



52(a). Standard electrode potentials are measured by comparison with a reference half-cell.

Draw a labelled diagram to show how the standard electrode potential could be measured for the redox system below.



Include details of the apparatus, solutions and the standard conditions needed when measuring this standard electrode potential.

Standard
conditions

[4]



(b). Many electric vehicles are powered by lithium-ion cells.

Hydrogen-oxygen fuel cells can also be used to power vehicles.

Six redox systems are shown in the table. State symbols have been omitted.

Redox system	Half-equation	E°/V
1	$\text{Li}^+ + \text{e}^- \rightleftharpoons \text{Li}$	-3.04
2	$2\text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{H}_2 + 2\text{OH}^-$	-0.83
3	$2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2$	0.00
4	$\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightleftharpoons 4\text{OH}^-$	+0.40
5	$\text{Li}^+ + \text{CoO}_2 + \text{e}^- \rightleftharpoons \text{LiCoO}_2$	+1.16
6	$\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$	+1.23

i. A lithium-ion cell involves **redox systems 1 and 5**.

Construct the overall cell equation for a lithium-ion cell.

[1]

ii. Hydrogen-oxygen fuel cells can operate in acidic or in alkaline conditions.

Show that for acidic and alkaline hydrogen-oxygen fuel cells, the standard cell potentials, and the overall cell equations, are the same.

Acidic



Alkaline

[3]

53. This question is about compounds and ions of iron(II) and iron(III) that contain ethanedioate ions, $\text{C}_2\text{O}_4^{2-}$.

A student plans an investigation to find the number of waters of crystallisation, x , in a sample of hydrated iron(II) ethanedioate, $\text{FeC}_2\text{O}_4 \cdot x\text{H}_2\text{O}$.

The student decides to carry out a redox titration between solutions of iron(II) ethanedioate and potassium manganate(VII) in acidic conditions.

- i. In the titration, both iron(II) ions and ethanedioate, $\text{C}_2\text{O}_4^{2-}$, ions are oxidised.

Construct half-equations for the oxidation of iron(II) and ethanedioate ions.

Oxidation of iron(II) ions

Oxidation of ethanedioate ions

[2]

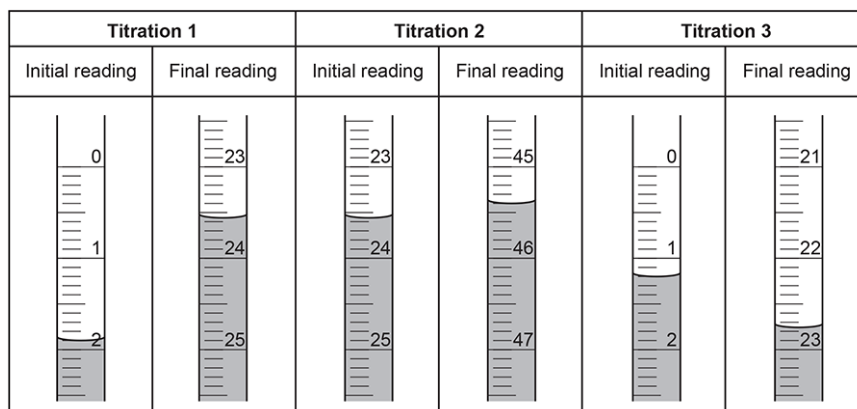
Redox and Electrode Potentials OCR Questions



- ii. The student prepares a 250.0 cm³ solution of iron(II) ethanedioate by dissolving 1.295 g of FeC₂O₄·xH₂O, in dilute sulfuric acid.

The student titrates 25.0 cm³ samples of this solution with 0.0200 mol dm⁻³ KMnO₄ in the burette. The student carries out a trial, followed by three further titrations.

The diagrams show the initial burette readings and the final burette readings for the student's three further titrations.



All burette readings are measured to the nearest 0.05 cm³.

Complete the titration table.

	1	2	3
Final reading/ cm³			
Initial reading/ cm³			
Titre / cm³			

[3]

- iii. The uncertainty in each burette reading is ±0.05 cm³.

Calculate the percentage uncertainty for the titre in **Titration 1**.

percentage uncertainty = % [1]



54. This question is about the chemistry of compounds containing phosphorus.

Phosphine, PH_3 , is a poisonous gas.

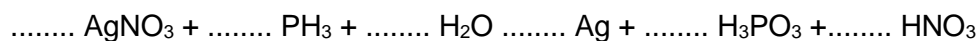
- i. Phosphine reacts with oxygen gas to form phosphorus(V) oxide and water.

Write the equation for this reaction.

[1]

- ii. Aqueous silver nitrate, AgNO_3 , is reduced by PH_3 .
The unbalanced equation is shown below.

Balance the equation and use oxidation numbers to explain why this is a redox reaction.



Explanation

[3]

END OF QUESTION PAPER



Mark scheme

Question	Answer/Indicative content	Marks	Guidance
1	A	1	
	Total	1	
2	B	1	
	Total	1	
3	A	1	
	Total	1	
4	B	1	
	Total	1	
5	<p>Element oxidised: sulfur / S 0 to +6</p> <p>Element reduced: nitrogen / N +5 to +4</p> <p>$6\text{HNO}_3 + \text{S} \rightarrow 6\text{NO}_2 + \text{H}_2\text{SO}_4 + 2\text{H}_2\text{O}$</p> <p>Correct species</p> <p>Balance</p>	4	<p>ALLOW 5+, 4+ and 6+</p> <p>Signs required</p> <p>ALLOW $4\text{H}^+ + 6\text{NO}_3^- + \text{S} \rightarrow 6\text{NO}_2 + \text{SO}_4^{2-} + 2\text{H}_2\text{O}$</p>
	Total	4	
6	<p>$\text{..ClO}^- + \text{..H}_2\text{O} + 2\text{..e}^- \rightarrow \text{..Cl}^- + 2\text{..OH}^-$</p> <p>$\text{Fe}_2\text{O}_3 + 10\text{..OH}^- \rightarrow 2\text{..FeO}_4^{2-} + 5\text{..H}_2\text{O} + 6\text{..e}^-$</p> <p>$\text{Fe}_2\text{O}_3 + 3\text{..ClO}^- + 4\text{..OH}^- \rightarrow 2\text{..FeO}_4^{2-} + 3\text{..Cl}^- + 2\text{..H}_2\text{O}$</p>	3	ALLOW multiples throughout
	Total	3	
7	a	4	<p>Factor 3 must be included in second mark for ECF on third mark.</p> <p>ALLOW 2 sig figs</p>



		$\text{volume} = \frac{3.24 \times 10^{-2} \times 1000}{0.0400} = 81.0 \text{ cm}^3$ <p>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>)</p>		
	b	$n(\text{MnO}_4^-) = 2.00 \times 10^{-2} \times \frac{21.6}{1000} = 4.32 \times 10^{-4} \text{ (mol)}$ $n(\text{Fe}^{2+}) = 4.32 \times 10^{-4} \times 5 = 2.16 \times 10^{-3} \text{ (mol)}$ <p>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</p>	3	<p>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</p> <p>ECF answer above $\times 5$</p> <p>This mark is only possible from correct answers above, i.e. no ECF</p>
	c	i	2	
		ii	1	ALLOW increase scale of Stage 1
		Total	10	
8	a	i	1	Working not required for mean titre
		ii	1	ALLOW ECF from incorrect subtraction in (i)
		iii	1	ALLOW All Cu^{2+} converted to $\text{CuI}/\text{Cu(I)}$
		iv	2	



		Blue to colourless AND all I ₂ has reacted ✓		
	v	$n(\text{S}_2\text{O}_3^{2-})$ in titration = $\frac{0.120 \times 21.75}{1000} = 2.61 \times 10^{-3}$ (mol) ✓ $n(\text{I}_2) = 1.305 \times 10^{-3}$ (mol) OR $n(\text{Cu}^{2+}) = 2.61 \times 10^{-3}$ (mol) AND $n(\text{Cu}^{2+})$ in original 250 cm ³ = $10 \times 2.61 \times 10^{-3} = 2.61 \times 10^{-2}$ (mol) Mass of A (CuSO ₄ ·xH ₂ O) used = 17.95 – 12.35 = 5.60 (g) $M(\text{CuSO}_4 \cdot x\text{H}_2\text{O}) = \frac{5.60}{2.61 \times 10^{-2}} = 214.6$ (g mol ⁻¹) ✓ $x = \frac{214.6 - 159.6}{18.0} = 3.05 = 3$ Formula = CuSO ₄ ·3H ₂ O	4	ALLOW ECF from incorrect mean titre in (i) ECF from $n(\text{S}_2\text{O}_3^{2-})$ ECF from $n(\text{Cu}^{2+})$ BUT DO NOT ALLOW incorrect mass ECF from $M(\text{CuSO}_4 \cdot x\text{H}_2\text{O})$ AND $x =$ nearest whole number
	b i	green solution: Fe ²⁺ (aq) OR [Fe(H ₂ O) ₆] ²⁺ AND gas bubbles: H ₂ (g) AND orange-brown solution: Fe ³⁺ (aq) OR [Fe(H ₂ O) ₆] ³⁺ ✓ Fe(s) + 2H ⁺ (aq) → Fe ²⁺ (aq) + H ₂ (g) ✓ 4Fe ²⁺ (aq) + O ₂ (g) + 4H ⁺ (aq) → 4Fe ³⁺ (aq) + 2H ₂ O(l) ✓	3	State symbols are not required in this part IGNORE , even if incorrect ALLOW full equation: Fe(s) + 2HCl(aq) → FeCl ₂ (aq) + H ₂ (g)
	ii	orange solution: Cr ₂ O ₇ ²⁻ AND green solution (anywhere) Cr ³⁺ OR [Cr(H ₂ O) ₆] ³⁺ ✓ 2Cr ³⁺ (aq) + H ₂ O(l) + 3H ₂ O ₂ (aq) → Cr ₂ O ₇ ²⁻ (aq) + 8H ⁺ (aq) H ⁺ , H ₂ O and e ⁻ all cancelled ✓✓	3	State symbols are not required in this part IGNORE , even if incorrect IGNORE Cr(VI) <i>The question asks for species</i> ALLOW 1 mark for H ⁺ /H ₂ O/e ⁻ not cancelled, e.g. 2Cr ³⁺ (aq) + 7H ₂ O(l) + 3H ₂ O ₂ (aq) + 6H ⁺ (aq) → Cr ₂ O ₇ ²⁻ (aq) + 14H ⁺ (aq) + 6H ₂ O(l) ✓
		Total	15	



9	a	i	<p>Half-cells (2 marks) Ag(s) and $1 \text{ mol dm}^{-3} \text{ Ag}^+(\text{aq})$ $1 \text{ mol dm}^{-3} \text{ Fe}^{2+}(\text{aq})$ 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 / $25 \text{ }^\circ\text{C}$ 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	
10	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 10^5 Pa OR 1 bar ✓</p>	2	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) ALLOW 1M DO NOT ALLOW 1 mol ALLOW 1 atmosphere / 1 atm OR 101 kPa OR 101325 Pa Examiner's Comments This definition was well known with most



				<p>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	$2\text{Ag}^+(\text{aq}) + \text{Cu}(\text{s}) \rightarrow 2\text{Ag}(\text{s}) + \text{Cu}^{2+}(\text{aq}) \checkmark$	1	<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>
	ii	<p>Assume Cu²⁺ Cu OR Cu half cell unless otherwise stated.</p> <p>[Cu²⁺] decreases OR < 1 mol dm⁻³ AND</p> <p>Equilibrium (shown in table) shifts to left \checkmark</p> <p>more electrons are released by Cu \checkmark</p> <p>The cell has a bigger difference in $E \checkmark$</p>	3	<p>FULL ANNOTATIONS MUST BE USED </p> <p>ALLOW [Cu²⁺] less than standard concentration / 1 mol dm⁻³ DO NOT ALLOW water reacts with Cu²⁺ OR Cu</p> <p>ALLOW E (for Cu²⁺ 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</p>



				<p>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	$E = -2.31 \text{ (V)} \checkmark$	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>$4\text{Al(s)} + 4\text{OH}^{\text{-}}(\text{aq}) + 3\text{O}_2(\text{g}) + 6\text{H}_2\text{O(l)} \rightarrow 4\text{Al(OH)}_4^{\text{-}}(\text{aq})$</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. $4\text{Al(s)} + 16\text{OH}^{\text{-}}(\text{aq}) + 3\text{O}_2(\text{g}) + 6\text{H}_2\text{O(l)} \rightarrow 4\text{Al(OH)}_4^{\text{-}}(\text{aq}) + 12\text{OH}^{\text{-}}(\text{aq})$</p> <p>ALLOW 1 mark if charge on Al(OH)₄ is omitted, i.e. $4\text{Al(s)} + 4\text{OH}^{\text{-}}(\text{aq}) + 3\text{O}_2(\text{g}) + 6\text{H}_2\text{O(l)} \rightarrow 4\text{Al(OH)}_4(\text{aq})$</p> <p>ALLOW 1 mark for an 'correct equation' reversed, i.e. $4\text{Al(OH)}_4^{\text{-}}(\text{aq}) \rightarrow 4\text{Al(s)} + 4\text{OH}^{\text{-}}(\text{aq}) + 3\text{O}_2(\text{g}) + 6\text{H}_2\text{O(l)}$</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</p>



				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 $\text{Al}(\text{OH})_4^-$ ion.
			Total	9
1 1	a	$\text{Fe}_2\text{O}_3 + 3\text{Cl}_2 + 10\text{OH}^- \rightarrow 2\text{FeO}_4^{2-} + 6\text{Cl}^- + 5\text{H}_2\text{O} \checkmark\checkmark$ <p>First mark for all 6 species Second mark for balancing</p>	2	<p>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 – question: equation for oxidation</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> <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	Balanced ionic equation AND state symbols required DO NOT ALLOW +2 or -2 for ionic charges



				<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>.....</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) ✓</p> <p>ALLOW Fe(6+) OR Fe⁶⁺</p>	2	<p>IGNORE H⁺ OR acidified</p> <p>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>
d		<p>FIRST, CHECK THE ANSWER ON ANSWER LINE</p> <p>IF answer = 51.8%, award 4 marks.</p> <p>.....</p>	4	<p>FULL ANNOTATIONS MUST BE USED</p> <p>.....</p> <p>For alternative answers, look first at common ECFs below.</p> <p>Then check for ECF credit possible using working below</p> <p>IF a step is omitted but subsequent step subsumes previous, then award mark for any missed step</p> <p>.....</p> <p>Working must be to at least 3 SF</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: </p> <p> $\text{Theoretical amount of } \text{BaFeO}_4 = \frac{0.437}{257.1} = 0.00170 \text{ (mol)} \checkmark$ </p> <p> $\% \text{ purity} = \frac{8.8 \times 10^{-4}}{1.70 \times 10^{-3}} \times 100 = 51.8\% \checkmark$ </p>		<p> 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 I_2 but the mark is for both steps combined </p> <p> ECF $257.1 \times$ answer above </p> <p> ECF $\frac{\text{answer above}}{0.437} \times 100$ </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>	3	<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> <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	
1 2	i	Fe ²⁺ : 1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ⁶ ✓		<p>ALLOW 4s before 3d, ie 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁶ ALLOW 1s² written after answer prompt (ie 1s² twice)</p> <p>ALLOW upper case D, etc and subscripts, e.g.4S₂3D₁ ALLOW for Fe²⁺4s⁰ DO NOT ALLOW [Ar] as shorthand for 1s²2s²2p⁶3s²3p⁶</p> <p>Look carefully at 1s²2s²2p⁶3s²3p⁶ – there may be a mistake</p> <p>Examiner's Comments</p> <p>Few candidates produced two incorrect electron configurations but there were many mistakes seen for either species. For Fe²⁺, 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 4p⁴ rather than 4p⁶ 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>
	i	Br ⁻ : 1s ² 2s ² 2p ⁶ 3s ² 3p ⁶ 3d ¹⁰ 4s ² 4p ⁶ ✓	2	
	ii	With Cl ₂ AND Br ₂ AND I ₂ products are Fe ²⁺ (AND halide ion)	3	<p>FULL ANNOTATIONS NEEDED</p> <p>ALLOW products within equations (even if equations are not balanced) IF stated, IGNORE reactants</p>



		<p>FeCl₂ AND FeBr₂ AND FeI₂ ✓</p> <p>OR Evidence that two electrode potentials have been compared for at least ONE reaction, ✓ e.g. Fe -0.44 AND Cl₂ +1.36 e.g. Iron has more / most negative electrode potential</p> <p>With Cl₂ AND Br₂, products are Fe³⁺ (AND halide ion) FeCl₃ AND FeBr₃ ?</p>		<p>ALLOW response in terms of positive 'cell reactions', e.g Fe + Cl₂ → Fe²⁺ + 2Cl⁻ E = (+)1.80 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 Fe²⁺ ions, supported by equations and electrode potential data. Many simply stated that Fe has the more negative E value (or the halogens the more 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									
1 3	a	<table border="1"> <thead> <tr> <th>E°</th> <th>redox system</th> </tr> </thead> <tbody> <tr> <td>Most negative</td> <td>E</td> </tr> <tr> <td></td> <td>C</td> </tr> <tr> <td>Least negative</td> <td>D</td> </tr> </tbody> </table> <p style="text-align: center;">✓</p>	E°	redox system	Most negative	E		C	Least negative	D	1	<p>ALL 3 correct for 1 mark</p> <p>Examiner's Comments</p> <p>This part required candidates to apply their knowledge and understanding to three cells. Just over half the candidates were able to list the redox systems by E value. There was no clear pattern with incorrect responses.</p>
E°	redox system											
Most negative	E											
	C											
Least negative	D											
	b 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</p>								



				<p>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>
		<p>ii</p> <p>H redox system is more negative (e.g. has a more -ve E OR less +ve E OR is -ve electrode) OR H redox system releases electrons (May be in equation, e.g. $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$) ✓</p> <p>Equilibrium shifts to increase $[\text{H}^+]$ OR H^+ OR standard hydrogen equation shifts to increase $[\text{H}^+]$ OR H^+ ✓</p>	2	<p>ALLOW ORA, <i>ie</i> Ag redox system (D) has more positive E / less negative E</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 $[\text{H}^+(\text{aq})]$. To score well required a very good understanding of equilibrium in the context of electrode potentials.</p>
		<p>iii</p> <p>$\text{H}_2 + 2\text{Ag}^+ \rightarrow 2\text{Ag} + 2\text{H}^+$ ✓</p>	1	<p>ALLOW multiples e.g. $\frac{1}{2}\text{H}_2 + \text{Ag}^+ \rightarrow \text{Ag} + \text{H}^+$ State symbols NOT required ALLOW equilibrium sign</p> <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>
c	i	<p>Fuel reacts with oxygen / oxidant to give electrical energy / voltage ✓</p>	1	<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</p>



				<p>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	1	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>IGNORE H₂ is a gas IGNORE 'produces no CO₂' 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	$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O \rightarrow$	1	<p>Correct species AND balancing needed ALLOW multiples ALLOW C₂H₆O 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 3½O₂ or 2O₂ 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>
	iv	$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O \checkmark$	1	<p>Correct species AND balancing needed ALLOW multiples, e.g. 3O₂ + 12H⁺ + 12e⁻ → 6H₂O</p>



				$\frac{1}{2}\text{O}_2 + 2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2\text{O}$ 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 Examiner's Comments 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.
		v	oxidation: C from -2 to +4 '+' sign not required ✓ reduction: O from 0 to -2 ✓	2 ALLOW 2- and 4+ ALLOW $\text{C}^{2-} \rightarrow \text{C}^{4+}$ 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) IGNORE O_2 reduced IGNORE any reference to electron transfer (<i>not in question</i>) Examiner's Comments 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.
			Total	11
1 4	a		Equations can be in either order $\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{NaOH}$ ✓	2 ALLOW multiples throughout IGNORE state symbols ALLOW $\text{Na}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{Na}^+ + 2\text{OH}^-$



	$\text{NaFeO}_2 + 2\text{H}_2\text{O} \rightarrow \text{Fe}(\text{OH})_3 + \text{NaOH} \checkmark$		<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 $\text{Fe}(\text{OH})_3$ had been formed although the examiners also credited an equation producing Fe_2O_3.</p>
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) ✓</p> <p>amount CrO_4^{2-} $2/3 \times 1.275 \times 10^{-3}$ OR $1.275 \times 10^{-3} \div 1.5$ = $8.5(00) \times 10^{-4}$ (mol) ✓</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 ✓</p> <p>Mass of Cr / Cr^{3+} in ore = 52.0 $\times 3.4(00)$ $\times 10^{-2}$ g</p> <p>1.768 g ✓</p> <p>Percentage Cr in ore = $\frac{1.768}{5.25} \times 100$</p>	6	<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>= 33.7% ✓</p> <p>MUST be to one decimal place (in the question)</p>		<p>5 marks (scaling error)</p> <p>0.8% × 40 missing</p> <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		Overall:	3	<p>ALLOW multiples and equilibrium signs throughout</p> <p>IGNORE state symbols throughout</p> <p>e.g. $2\text{CrO}_4^{2-} + 6\text{I}^- + 8\text{H}_2\text{O} \rightarrow 2\text{Cr}^{3+} + 3\text{I}_2$</p>

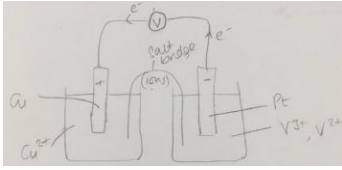


		$4^{2-} + 3I^- + 4H_2O \rightarrow Cr^{3+} + 1\frac{1}{2} I_2 + 8OH^-$ ✓ CrO <i>Half equations:</i> $4^{2-} + 4H_2O + 3e^- \rightarrow Cr^{3+} + 8OH^-$ ✓ CrO $2I^- \rightarrow I_2 + 2e^-$ ✓		+ 16OH ⁻ ALLOW equation using H ⁺ . i.e. $CrO_4^{2-} + 3I^- + 8H^+ \rightarrow Cr^{3+} + 1\frac{1}{2} I_2 + 4H_2O$ OR $2CrO_4^{2-} + 6I^- + 16H^+ \rightarrow 2Cr^{3+} + 3I_2 + 8H_2O$ ALLOW CrO ₄ ²⁻ half equation using H ⁺ . i.e. $CrO_4^{2-} + 8H^+ + 3e^- \rightarrow Cr^{3+} + 4H_2O$
		Total	11	
1 5		<p>in all equations ALLOW equilibrium signs IGNORE state symbols</p> <p>.....</p> <p>Reaction 1: 1 mark $2I_2 + 5O_2 \rightarrow 2I_2O_5$ ✓</p> <p>.....</p> <p>Reaction 2: 2 marks 1st mark: ALL CORRECT species</p> <p>e.g.: $I_2 + OH^- \rightarrow I^- + IO_3^- + H_2O$</p> <p>2nd mark for CORRECT balanced equation $3I_2 + 6OH^- \rightarrow 5I^- + IO_3^- + 3H_2O$ ✓✓</p>	3	<p>FULL ANNOTATIONS MUST BE USED </p> <p>ALLOW correct multiples throughout, e.g. $I_2 + 2\frac{1}{2}O_2 \rightarrow I_2O_5$</p> <p>.....</p> <p>For 1st mark, IGNORE e⁻ present</p> <p>ALLOW species / equation with NaOH or KOH, e.g. $3I_2 + 6NaOH \rightarrow 5I^- + IO_3^- + 3H_2O + 6Na^+$ $3I_2 + 6NaOH \rightarrow 5NaI + NaIO_3 + 3H_2O$</p> <p>ALLOW Species: $I_2 + OH^- \rightarrow I^- + IO_2^+ + H_2O$ ✓ OR Equation: $3I_2 + 4OH^- \rightarrow 5I^- + IO_2^+ + 2H_2O$ ✓✓</p> <p>Species: $I_2 + OH^- \rightarrow I^- + IO_3^+ + H_2O$ ✓ OR Equation: $3I_2 + 2OH^- \rightarrow 5I^- + IO_3^+ +$</p>



				<p>H₂O ✓ ✓</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. For the second equation, the key difficulty was the formula of the iodate(V) ion, IO₃⁻, with I⁵⁺ or the formula of I₂O₅ often being shown. The most difficult skill was balancing the second redox equation, suggesting that many candidates need more practice in this area.</p>
		Total	3	
1 6	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>	2	<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>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	Complete circuit with voltmeter	5	



	<p>AND labelled salt bridge linking two half-cells ✓</p>  <p>Cu electrode in Cu^{2+} ✓</p> <p>Pt electrode in V^{2+} AND V^{3+} ✓</p> <p>Cu shown as + AND Pt shown as - ✓</p> <p>electrons in wire AND ions in salt bridge ✓</p> <p><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 H_2 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^+, NH_4^+, NO_3^-, Cu^{2+}, V^{2+}, V^{3+}</p> <p>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	<p>IGNORE any sign</p> <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 Cr^{3+} oxidises Al OR Cr^{3+} acts as oxidising agent</p>	6	<p>FULL ANNOTATIONS MUST BE USED</p> <p>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)



	<p>AND $3\text{Cr}^{3+} + \text{Al} \rightarrow 3\text{Cr}^{2+} + \text{Al}^{3+}$ ✓ Explanation (dependent on Cr³⁺ oxidising Al above) <i>E</i> of redox system 2 (Cr³⁺/Cr²⁺) is more positive /less negative (than <i>E</i> 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⁺) ✓ $2\text{Cr}^{3+} + 2\text{FeO}_4^{2-} + 2\text{H}^+ \rightarrow \text{Cr}_2\text{O}_7^{2-} + 2\text{Fe}^{3+} + \text{H}_2\text{O}$ ✓ Explanation (dependent on Cr³⁺ reducing FeO₄²⁻ above) <i>E</i> of redox system 5 (Cr₂O₇²⁻/Cr³⁺) is less positive / more negative (than <i>E</i> of system 6 (FeO₄²⁻/Fe³⁺)) ORA, i.e. in terms of 6 being more positive (than 5) ✓</p>	<ul style="list-style-type: none"> 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' <i>E</i> OR 'greater'</p> <p>ALLOW <i>E</i>_{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' <i>E</i> OR 'greater'</p> <p>ALLOW <i>E</i>_{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 equation. The expected explanation in terms of electrode potentials was seen more rarely.</p> <p>The role of Cr³⁺ in reducing FeO₄²⁻ was seen much more rarely. The commonest responses showed a list of equations, all showing reactions of Cr²⁺ rather than Cr³⁺.</p> <p>The best candidates produced superb</p>
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				<p>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 7	i	linear ✓	1	<p>IGNORE planar</p> <p>Examiner's Comments</p> <p>Most candidates identified the shape as linear.</p>
	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 ✓</p> <p>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>
		iv	$\text{ClO}^- + 2\text{H}^+ + 2e^- \rightarrow \text{Cl}^- + \text{H}_2\text{O} \checkmark$	<p>1</p> <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
1 8	a		<p>$\text{Cu}^{2+}: (1s^2) 2s^2 2p^6 3s^2 3p^6 3d^9 \checkmark$</p> <p>$\text{Cu}^+: (1s^2) 2s^2 2p^6 3s^2 3p^6 3d^{10} \checkmark$</p>	<p>2</p> <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.....$3S_2 3P^6$</p> <p>DO NOT ALLOW [Ar] as shorthand for $1s^2 2s^2 2p^6 3s^2 3p^6$</p> <p>Examiner's Comments</p> <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 In formula of HCOOH / HCOO, ALLOW H, C and O in ANY order ALLOW H_2CO_3 for H_2O and CO_2 in carbonate equation</p> <p>ALLOW $(\text{HCOO})_2\text{Cu}$ for $\text{Cu}(\text{HCOO})_2$</p> <p>DO NOT ALLOW equation with CuSO_4</p> <p>Examiner's Comments</p> <p>Most candidates attempted an equation using CuO, $\text{Cu}(\text{OH})_2$ or CuCO_3. Marks were then sometimes lost by not balancing the equation. It was not uncommon to see equations using CuSO_4 or CuCl_2 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 $\text{CuI}(\text{s})$ 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 ALLOW black / brown to colourless</p> <p>DO NOT ALLOW just 'it turns colourless / is decoloured' <i>Initial colour required</i></p> <p>IGNORE clear for colourless</p> <p>Examiner's Comments</p>

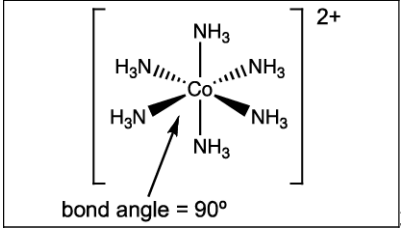
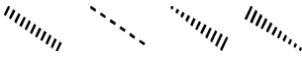
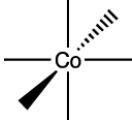


				<p>Most candidates seemed unaware that starch is used to identify the end point in iodine–thiosulfate titrations. Even when starch was given, the colour change was often incorrect. Random responses were seen to this part, e.g. methyl orange, phenolphthalein, potassium manganate and sodium thiosulfate.</p>
e		<p>WORKING REQUIRED Correct answer: x = 4 required evidence of working</p> <p>..... $n(\text{S}_2\text{O}_3^{2-}) \text{ OR } n(\text{Cu}^{2+}) = \frac{0.0420 \times 23.5}{1000} = 9.87 \times 10^{-4} \text{ (mol)}$</p> <p>In 250.0 cm³ solution, $n(\text{Cu}^{2+}) = 9.87 \times 10^{-3} \text{ (mol)} \checkmark$</p> <p>$M(\text{Cu}(\text{HCOO})_2 \cdot 4\text{H}_2\text{O}) = \frac{2.226}{9.87 \times 10^{-3}} = 225.5 \text{ (g mol}^{-1}\text{)} \checkmark$</p> <p>$x(\text{H}_2\text{O})$ has mass of 225.5 – $M(\text{Cu}(\text{HCOO})_2)$ = 225.5 – 153.5 = 72(.0) \checkmark</p> <p>$x = \frac{72(.0)}{18(.0)} = 4$</p> <p style="text-align: center;">WHOLE NUMBER needed</p> <p>AND evidence of working \checkmark</p>	5	<p>FULL ANNOTATIONS MUST BE USED </p> <p>At least 3 SF required throughout</p> <p><i>Alternative approach for final 3 marks based on mass:</i></p> <p>mass $\text{Cu}(\text{HCOO})_2 = 9.87 \times 10^{-3} \times 153.5 = 1.515 \text{ g } \checkmark$</p> <p>$n(\text{H}_2\text{O}) = \frac{2.226 - 1.515}{18(.0)} = \frac{0.711}{18(.0)} = 0.0395 \text{ (mol)} \checkmark$</p> <p>$x = \frac{0.0395}{9.87 \times 10^{-3}} = 4 \checkmark$</p> <p>ALLOW $\text{Cu}(\text{HCOO})_2 \cdot 4\text{H}_2\text{O}$</p> <p>.....</p> <p>COMMON ERRORS for 4 marks $x = 117$ (calc 116.78) Use of 9.87×10^{-4} (no scaling $\times 10$) \rightarrow $M = 2255.319$</p> <p>$x = 17$ (calc 16.53) 4 marks Use of 4.935×10^{-4} (Use of $0.5 \times 9.87 \times 10^{-3}$)</p> <p>Check $n(\text{Cu}^{2+})$ for other ECFs Check for ECFs from incorrect $M(\text{anhydr salt})$ Actual = 153.5</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 all five marks for showing that x was 4.</p> <p>Where '4' had not been obtained, marks could still be awarded for intermediate</p>



				working if correct. Answer: x = 4
		Total	11	
1 9		A	1	
		Total	1	
2 0	a	i	$3 \text{MnO}_4^{2-} + 4 \text{H}^+ \rightarrow 2 \text{MnO}_4^- + \text{MnO}_2 + 2 \text{H}_2\text{O} \checkmark$	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 $\text{H}^+ \checkmark$</p> <p>In alkaline conditions OH^- reacts with H^+ AND equilibrium (position) shifts to the left to increase concentration of H^+/add $\text{H}^+ \checkmark$</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 $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l}) \checkmark$</p> <p>$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ AND Cell potential = $1.23 - 0.00 = 1.23 \text{ (V)} \checkmark$</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 $\text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^- \rightarrow 4\text{OH}^-(\text{aq}) \checkmark$ $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$</p> <p>AND Cell potential = $0.40 - (-0.83) = 1.23 \text{ (V)} \checkmark$</p>	4 ALLOW $\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$ ALLOW $\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$
		Total	7	
2 1		i	<p>$[\text{Co}(\text{H}_2\text{O})_6]^{2+} + 6\text{NH}_3 \rightarrow [\text{Co}(\text{NH}_3)_6]^{2+} + 6\text{H}_2\text{O} \checkmark$</p> <p>ligand substitution \checkmark</p>	2 ALLOW ligand exchange



	ii	 <p>Shape ✓ bond angle 90 ° ✓</p> <p>Bonds must be to N of NH₃ ligands</p>	2	<p>IGNORE charges (anywhere) and labels (even if wrong)</p> <p>Square brackets NOT required</p> <p>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':</p> <p>For bond into paper, ALLOW:</p>  <p>ALLOW following geometry throughout:</p> 
	iii	<p>Empirical formula of complex D</p> <p>Co : N : H : Cl $\frac{22.03}{58.9} : \frac{31.41}{14.0} : \frac{6.73}{1.00} : \frac{39.83}{35.5}$</p> <p>OR 0.374 : 2.24 : 6.73 : 1.12 ✓</p> <p>= 1 : 6 : 18 : 3 = CoN₆H₁₈Cl₃ ✓</p> <p>complex ion C [Co(NH₃)₆]³⁺ ✓</p> <p>complex D [Co(NH₃)₆]³⁺[Cl⁻]₃ ✓</p>	4	<p>Correct empirical formula subsumes previous mark</p> <p>ALLOW [Co(NH₃)₆]³⁺ 3Cl⁻</p>
	iv	<p>Half equations [Co(NH₃)₆]²⁺ → [Co(NH₃)₆]³⁺ + e⁻ ✓</p> <p>H₂O₂ + 2e⁻ → 2OH⁻ ✓</p> <p>Overall equation 2[Co(NH₃)₆]²⁺ + H₂O₂ → 2[Co(NH₃)₆]³⁺ + 2OH⁻ ✓</p>	1	<p>ALLOW multiples throughout</p> <p>ALLOW H₂O₂ + 2H⁺ + 2e⁻ → 2H₂O</p> <p>ALLOW</p> <p>2[Co(NH₃)₆]²⁺ + H₂O₂ + 2H⁺ → 2[Co(NH₃)₆]³⁺ + 2H₂O</p>



			Total	11	
2			H ₂ SO ₄ + 8HI → 4I ₂ + H ₂ S + 4H ₂ O		
2			All species correct OR H ₂ SO ₄ : HI ratio = 1 : 8 ✓	2	
			Equation complete and balanced ✓		
			Total	2	
2			C	1	
3			Total	1	
2	a	i	<p>complete circuit AND voltmeter</p> <p><i>Circuit:</i> 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> 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</p> <p>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	IGNORE sign



				<p>Examiner's Comments This was the most successfully answered one mark question on the paper.</p>
	b	<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>
	c	<p>AWARD 2 marks for correct balancing AND all species cancelled on both sides of equation: 2MnO₄⁻ + 6H⁺ + 5SO₃²⁻ → 2Mn²⁺ + 3H₂O + 5SO₄²⁻ ✓ ✓</p> <p>AWARD 1 mark for correct balancing but not all species (H₂O, H⁺) cancelled on both sides of equation ✓ e.g. 2MnO₄⁻ + 16H⁺ + 5SO₃²⁻ + 5H₂O → 2Mn²⁺ + 8H₂O + 5SO₄²⁻ + 10H⁺</p>	2	<p>ALLOW correct multiples e.g. MnO₄⁻ + 3H⁺ + 2½SO₃²⁻ → Mn²⁺ + 1½H₂O + 2½SO₄²⁻</p> <p>IGNORE state symbols</p> <p>e.g. MnO₄⁻ + 8H⁺ + 2½SO₃²⁻ + 2½H₂O → Mn²⁺ + 4H₂O + 2½SO₄²⁻ + 5H⁺</p> <p>Examiner's Comments This was successfully answered by the majority of candidates.</p>
		Total	9	
2 5	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	Sodium / Na oxidised from 0 to +1 ✓	2	<p>ALLOW 1+ for +1 and 3+ for +3</p> <p>ALLOW N₂ for nitrogen</p>



			<p>Nitrogen / N</p> <p>reduced from +3 to 0 ✓</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>	<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
2 6			<p>ALLOW equilibrium sign in equations provided reactants on left</p> <p>Reaction of H_2O_2 with MnO_2: $\text{H}_2\text{O}_2 + \text{MnO}_2 + 2\text{H}^+ \rightarrow \text{O}_2 + \text{Mn}^{2+} + 2\text{H}_2\text{O}$ ✓</p> <p>Reaction of H_2O_2 with Mn^{2+}: $\text{H}_2\text{O}_2 + \text{Mn}^{2+} \rightarrow \text{MnO}_2 + 2\text{H}^+$ ✓</p> <p>Use of E data</p> <p>Use of E data to support equation(s) above or half direction of provided</p>	<p>ALLOW correct multiples IGNORE state symbols</p> <p>..... ALLOW uncanceled H_2O 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> <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



		<p>half equations (one including MnO_2) ✓</p> <p>Also look for evidence around half equations</p> <p>MnO_2 regenerated / reformed ✓</p> <p>Must be linked to an equation showing MnO_2 as reactant and an equation showing MnO_2 as product</p>		<ul style="list-style-type: none"> Calculated $E_{\text{cell}} = +0.81 \text{ V}$ (from top 2) OR $+0.27 \text{ V}$ (from bottom 2) <p>ALLOW combining of equations above to show that MnO_2 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 MnO_2, and Mn^{2+} react with H_2O_2. Many combined the two equations involving H_2O_2 to obtain the overall equation for the decomposition of H_2O_2 which was given at the top of the paper. Of the equations seen, many had species uncanceled. 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 MnO_2 was well understood but often just stated with no reference to the equations. This part discriminated very well.</p>
		Total	4	
27		D	1 (AO 2.5)	<p>Examiner's Comments</p> <p>Most candidates knew the correct equation.</p>
		Total	1	
28		B	1 (AO 1.1)	<p>Examiner's Comments</p> <p>D was the common distractor given as the answer by many candidates, suggesting confusion with the first</p>



				electron affinity and second electron affinity of oxygen.																
			Total	1																
2 9	i	<p>Titres correct and ALL recorded to 2 decimal places</p> <table border="1" style="width: 100%; border-collapse: collapse;"> <tr> <td style="width: 15%;">Titre:</td> <td style="width: 15%;"></td> <td style="width: 15%;"></td> <td style="width: 15%;"></td> <td style="width: 15%; text-align: right;">23.85</td> </tr> <tr> <td></td> <td style="text-align: center;">23.40</td> <td style="text-align: center;">23.75</td> <td style="text-align: center;">✓</td> <td></td> </tr> <tr> <td style="text-align: left;">24.00</td> <td></td> <td></td> <td></td> <td></td> </tr> </table> <table border="1" style="width: 100%; border-collapse: collapse;"> <tr> <td style="padding: 5px;">mean titre = 23.80 (cm³) ✓</td> </tr> </table>	Titre:				23.85		23.40	23.75	✓		24.00					mean titre = 23.80 (cm ³) ✓	2	<p>ALLOW 23.8 cm³</p> <p><u>Examiner's Comments</u></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 ± 0.05 cm³.</p> <p>The average titre should be calculated by averaging concordant titres, i.e. those within 0.10 cm³ of each other.</p>
Titre:				23.85																
	23.40	23.75	✓																	
24.00																				
mean titre = 23.80 (cm ³) ✓																				
	ii	<p>Percentage uncertainty $= \frac{0.05 \times 2}{23.40} \times 100 = 0.43 (\%) \checkmark$</p>	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) ✓</p> <p>Blue to colourless ✓</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>
iv	<p>FIRST CHECK THE ANSWER ON THE ANSWER LINE IF B = RbIO₃ AND relative formula mass = 260.5 award 5 marks IF relative formula mass = 260.5 award 4 marks</p> <p>----- -</p> <p>$n(\text{S}_2\text{O}_3^{2-})$ in titration</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px auto;"> $= \frac{0.150 \times 23.80}{1000} = 3.57 \times 10^{-3} \text{ (mol) } \checkmark$ </div> <p>$n(\text{IO}_3^-)$ in titration</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px auto;"> $= 10 \times 5.95 \times 10^{-4} = 5.95 \times 10^{-3} \text{ (mol) } \checkmark$ </div> <p>$n(\text{IO}_3^-)$ in original 250 cm³</p> <div style="border: 1px solid black; padding: 5px; width: fit-content; margin: 5px auto;"> $= \frac{3.57 \times 10^{-3}}{6} = 5.95 \times 10^{-4} \text{ (mol) } \checkmark$ </div> <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; width: fit-content; margin: 5px auto;"> <p>Iodate of Group 1 metal that most closely matches calculated molar mass of B</p> <p>Formula from 260.5 = RbIO₃ ✓</p> </div>	5	<p>ALLOW ECF from incorrect mean titre in (a)(i)</p> <p>ECF from $n(\text{S}_2\text{O}_3^{2-})$ in titration ALLOW a two-step calculation $n(\text{I}_2) = n(\text{S}_2\text{O}_3^{2-}) \div 2$ and $n(\text{IO}_3^-) = n(\text{I}_2) \div 3$</p> <p>ECF from $n(\text{IO}_3^-)$ in titration</p> <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 RbIO₃ DO NOT ALLOW RbIO₃ without relative formula mass value. DO NOT ALLOW 260.4 (without working) and RbIO₃ IF B = RbIO₃ AND relative formula mass = 261 award 5 marks</p>



				<p>Examiner's Comments</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 $n(\text{S}_2\text{O}_3^{2-}) = \dots$ 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	
3 0	a	i	<p>complete circuit AND voltmeter AND salt bridge linking two half-cells ✓</p> <p>Circuit:</p> <p>Pt AND I^- AND I_2 ✓</p> <p>Half cells:</p> <p>Ni AND Ni^{2+} ✓</p> <p>Standard conditions:</p> <p>1 mol dm^{-3} solutions AND 298 K / 25°C ✓</p>	<p>4</p> <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 $\text{dm}^{-3}/1\text{M}$ 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> <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</p>



				<p>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	$E = 0.79 \text{ (V)}$ ✓	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	$H_2O_2(aq) + 2H^{+}(aq) + 2Fe^{2+}(aq) \rightarrow$ $2Fe^{3+}(aq) + 2H_2O(l)$ ✓	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> $3Zn(s) + Cr_2O_7^{2-}(aq) + 14H^{+}(aq) \rightarrow 3Zn^{2+}(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$ <p>✓</p> $Zn(s) + 2Cr^{3+}(aq) \rightarrow Zn^{2+}(aq) + 2Cr^{2+}(aq)$ <p>✓</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>



	<p>Comparison of E values (seen once)</p> <p>E of Zn is more negative/less positive than E of $\text{Cr}_2\text{O}_7^{2-}$</p> <p>OR</p> <p>E of Zn is more negative/less positive than E of Cr^{3+} ✓</p> <p>Equilibrium shift related to E values</p> <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><u>Examiner's Comments</u></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</p>	6	<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> <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> <p>Structure of trans stereoisomer</p>



Two reactions are explained in detail.

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.

Level 1 (1–2 marks)

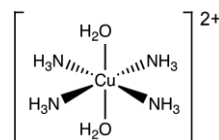
Make two simple explanations from any one reaction.

OR

Makes one simple explanation from each of two reactions

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.



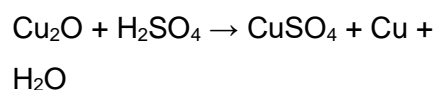
Correct connectivity

REACTION 2 (Cu₂O/H₂SO₄)

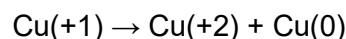
Products



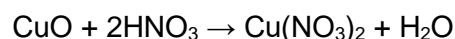
Equation



Oxidation numbers



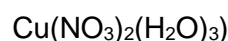
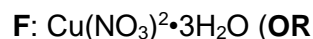
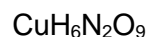
REACTION 3 (CuO/HNO₃) Equation



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



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Further guidance on use of wedges



Exemplar 2

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

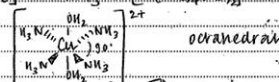
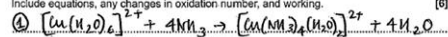
Reaction 1: Aqueous copper(II) sulfate reacts with excess aqueous ammonia in a ligand substitution reaction. A deep-blue solution is formed, containing an octahedral complex ion, C, which is a *trans* isomer. $\Rightarrow 100$

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. $+ H_2O$

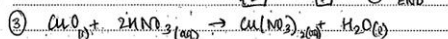
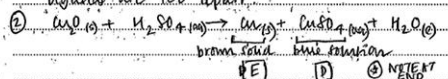
Reaction 3: Copper(II) oxide reacts with warm dilute nitric acid in a neutralisation reaction, to form a blue solution. Unreacted copper(II) oxide is filtered off, and the solution is left overnight in an evaporating basin. A hydrated salt, F, crystallises, with the percentage composition by mass: Cu,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

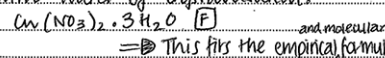


	Cu	H	N	O
mass	26.29	2.48	11.59	59.63
r.f.w	63.5	1	14	16
mol	0.414	2.48	0.828	3.73

Additional answer space if required.



A hydrated salt is made up of an anhydrous salt with water of crystallisation.



NOTE: ② The oxidation number of Cu goes from +1 to 0 in Cu_2O , and from +1 to +2 in $CuSO_4(aq)$. 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 practice in picking out and interpreting key information.



			Total	16	
3 1			$2\text{MnO}_4^- + 5\text{H}_2\text{O}_2 + 6\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{O}_2$ <p>Correctly balanced equation for $\text{MnO}_4^-/\text{H}_2\text{O}_2$ reaction but no cancelling of H^+ and/or e^- ✓</p> <p>Overall equation correct with all species cancelled ✓</p>	2	<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: $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$ (H_2O_2 is acting as both reducing and oxidising agent)</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	
3 2	i	Colourless to yellow ✓		1	<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 \rightarrow colourless.</p>
	ii	<p>Mean titre</p> $= \frac{(23.15 + 23.25)}{2} = 23.2(0) \text{ (cm}^3\text{)} \checkmark$		6	<p>Common error: Incorrect mean from all 3 titres = 23.30 cm^3</p> <p>Use ECF throughout Intermediate values for working to at</p>



		<p>Analysis of results 5 marks</p> $n(\text{Ce}^{4+}) = 23.20 \times \frac{0.0500}{1000} = 1.16 \times 10^{-3} \text{ (mol) } \checkmark$ $n((\text{COOH})_2) \text{ in } 25.0 \text{ cm}^3 = \frac{1.16 \times 10^{-3}}{2} = 5.8(0) \times 10^{-4} \text{ (mol) } \checkmark$ <p>$n((\text{COOH})_2) \text{ in } 250 \text{ cm}^3$</p> $= 5.8(0) \times 10^{-4} \times 10 = 5.8(0) \times 10^{-3} \text{ (mol) } \checkmark$ <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>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 $\div 2$ to obtain $n((\text{COOH})_2)$</p> <p>→ 1.26%: 5 marks from 23.20 → 1.27% 4 marks from 23.30</p> <p><u>Examiner's Comments</u></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
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				<p>figures; this was the most common error.</p> <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 = $\frac{23.25 + 23.15}{2} = 23.20$ ✓</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 $1.16 \times 10^{-3} \text{ mol of Ce}(\text{SO}_4)_2$</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 x Mr = $5.8 \times 10^{-3} \times 90 = 0.522 \text{ g}$</p> <p>$\frac{0.522}{82.68} \times 100 = 0.631\%$</p> <p>percentage of ethanedioic acid = 0.631% [6]</p>
		Total	7	
3 3		<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 x2 of 50 for two solutions mixed) No ÷ 2 in final step 	4	



	<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) ✓</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> $= -\frac{104.5}{2} = -52.3 \text{ OR } -52.25 \text{ kJ mol}^{-1} \checkmark$	<p>ALLOW ECF throughout</p> <p>DO NOT ALLOW less than 3 SF IGNORE units</p> <hr/> <p>ALTERNATIVE METHOD $n(\text{succinic acid}) = 0.02(00) \text{ (mol)} \checkmark$</p> <p>Energy released = 2090 (J) OR 2.090 (kJ) ✓</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 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.</p> <p>Most candidates calculated that 0.02 mol of succinic acid had reacted.</p> <p>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 $2 \times 0.02 = 0.04 \text{ mol H}_2\text{O}$. 2.09 kJ was then scaled up to give the correct ΔH value of $-52.25 \text{ kJ mol}^{-1}$. Example 1 follows this approach and each stage in the working can be followed easily.</p>
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Most candidates scaled up using 0.02 mol to obtain $-104.6 \text{ kJ mol}^{-1}$ 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 = $-52.25 \text{ kJ mol}^{-1}$.

Exemplar 13

mixture are the same as for water. $Q = mc\Delta T$

$$Q = (50 + 50) \times 4.18 \times 5 = 2090 \text{ J} = 2.09 \text{ kJ}$$

moles of acid used = $\frac{50}{1000} \times 0.4 = 0.02 \text{ mol}$

moles of NaOH used = $\frac{50}{1000} \times 1 = 0.05 \text{ mol}$

acid is the limiting reagent.

moles of H_2O formed = $0.02 \times 2 = 0.04 \text{ mol}$

$$\Delta_{\text{neut}} H = \frac{-2.09}{0.04} = -52.25$$

$\Delta_{\text{neut}} H = \dots\dots\dots -52.25 \dots\dots\dots \text{ kJ mol}^{-1}$ [4]

Exemplar 14



				$q = mc\Delta T \quad 26.5 - 21.5$ $100 \times 4.18 \times 5 = 2090 \checkmark$ $\div 1000 = 2.09$ <p>mol conc. x vol</p> $\frac{50}{1000} \times 1 = 0.05 \quad \frac{2.09}{0.05} = 41.8 \checkmark$ $\Delta_{red}H = -41.8 \text{ kJmol}^{-1} [4]$
		Total	4	
3 4		C	1 (AO2.1)	<p>Examiner's Comments</p> <p>Only the higher attaining candidates opted for the correct answer of C. Many candidates thought statement 1 was correct, (giving option A as the answer), not realising that this equation refers to acidic conditions not alkaline conditions (as mentioned in the question). Others did not recognise that the cell potential was 1.23V (obtained by calculating the difference between +0.40 and -0.83), and so opted for option B.</p>
		Total	1	
3 5	i	<p>Overall equation AND state symbols:</p> $\text{M(s)} + 2\text{HCl(aq)} \rightarrow \text{MCl}_2\text{(aq)} + \text{H}_2\text{(g)} \checkmark$ <p>STATE SYMBOLS required in overall equation ONLY</p> <p>Half equations:</p>	3 (AO 2.6x3)	<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 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 M²⁺ OR H⁺¹ for H⁺ ALLOW M - 2e⁻ → M²⁺</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>Examiner's Comments</p>



		<p>Oxidation $M \rightarrow M^{2+} + 2e^- \checkmark$</p> <p>Reduction $2H^+ + 2e^- \rightarrow H_2$ OR $H^+ + e^- \rightarrow \frac{1}{2}H_2 \checkmark$</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>2 (AO 3.3x2)</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><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>$H^+ + OH^- \rightarrow H_2O \checkmark$</p>	<p>1 (AO 2.5)</p>	<p>ALLOW multiples e.g. $2H^+ + 2OH^- \rightarrow 2H_2O$</p> <p>IGNORE state symbols, even if wrong</p> <p><u>Examiner's Comments</u></p> <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>Mean titre 1 mark $= \frac{(27.30 + 27.20)}{2} = 27.25 \text{ (cm}^3\text{)} \checkmark$</p> <p>Analysis of results 5 marks</p> <p>$n(\text{NaOH}) = 27.25 \times \frac{0.320}{1000} = 8.72 \times 10^{-3} \text{ (mol)} \checkmark$</p> <p>$n(\text{HC})$ in $25.0 \text{ cm}^3 = n(\text{NaOH})$</p> <p>$n(\text{HCl})$ in 250 cm^3 $= 8.72 \times 10^{-3} \times 10 = 8.72 \times 10^{-2} \text{ (mol)}$ \checkmark</p> <p>$n(\text{HC})$ that reacted with M $= 0.210 - 8.72 \times 10^{-2} = 0.1228 \text{ (mol)}$ \checkmark</p> <p>$n(\text{M})$ that reacted $= \frac{0.1228}{2} = 0.0614 \text{ (mol)} \checkmark$</p> <p>$A_r$ of M $= \frac{6.90}{0.0614} = 112.4$ AND M = cadmium/Cd \checkmark</p> <p>iv</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 $\rightarrow 0.06124 \rightarrow 112.7$ AND Cd: 5 marks</p> <p>No $\div 2$ to obtain $n(\text{M})$ $\rightarrow 56.2$ AND Fe (from 27.25) 5 marks</p> <p>$\rightarrow 56.3$ AND Fe (from 27.35) 4 marks</p> <p>No subtraction from 0.210 A_r of M $= \frac{6.90}{0.0614} = 112.4$ AND M = cadmium/Cd \checkmark $\rightarrow 158.2$ to 158.3 AND Tb 5 marks</p> <p>No $\times 10$ to obtain $n(\text{HCl})$ in 250 cm^3 5 marks</p>	<p>6</p> <p>(AO 2.8x5)</p> <p>(AO 3.2)</p>	<p>FULL ANNOTATIONS MUST BE USED</p> <p>-----</p> <p>--</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>ALLOW 112.2 from 0.0615 AND Cd</p> <p>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>Most candidates determined the correct mean titre of 27.25 cm^3. A few candidates did take the mean of all three</p>
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		$0.210 - 8.72 \times 10^{-3} = 0.20128$ OR 0.201 $n(M) = 0.20128/2 = 0.10064$ $A_r = 6.90/0.10064 = 68.56 \rightarrow \text{Zn}$ <i>No $\times 10$ and no $\div 2$ 4 marks</i> $0.210 - 8.72 \times 10^{-3} = 0.20128$ $A_r = 6.9/0.20128 = 34.28 \rightarrow \text{Ca}$ <i>Omitting initial titration calculation</i> <i>Zero marks</i> $0.210/2 = 0.105 \rightarrow 6.9/0.105 = 65.71$ $\rightarrow \text{Zn}$		<p>titres rather than the closest. Most calculated that 8.72×10^{-3} 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 8.72×10^{-2} mol in the 250 cm³ 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 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	
3 6		C	1 (AO 2.6)	
		Total	1	
3 7	a	Circuit Complete circuit AND voltmeter AND salt bridge linking two half-cells ✓	3 (AO 3.4x 1) (AO 1.2x	Voltmeter must be shown AND salt bridge must be labelled ALLOW small gaps in circuit If species in BOTH half cells are correct but concentration of 1 mol dm ⁻³ omitted, ALLOW 1 mark for BOTH half cells.



	<p>Half cells</p> <p>Ag AND Ag⁺ AND 1 mol dm⁻³ solution ✓</p> <p>Pt AND H⁺ AND MnO₄⁻ AND Mn²⁺ AND 1 mol dm⁻³ /equimolar solution ✓</p>	<p>1) (AO 1.2x1)</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 MnO₄⁻/Mn²⁺ cell.</p>
<p>b</p>	<p>Comparison of E values</p> <p>E of redox system 4 (MnO₄⁻ /Mn²⁺) is more positive/less negative than E of redox systems 2 (HCOOH/HCHO) OR 1 (CO₂/HCOOH)✓</p> <p>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>	<p>4 (AO 3.1x 2)</p> <p>(AO 3.2x 2)</p>	<p>IGNORE higher/lower</p> <p>ALLOW Overall E_{reaction} = (+)1.54V OR (+)1.62V</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><u>Examiner's Comments</u></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</p>




				the E 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.
	c	$2\text{H}^+ + \frac{1}{2}\text{O}_2 + 2\text{e}^- \rightarrow \text{H}_2\text{O} \checkmark$ $1.34 + (-0.11) = (+)1.23 \text{ (V)} \checkmark$	2 (AO 2.6) (AO 2.2x1)	IGNORE state symbols ALLOW multiples
		Total	9	
3 8		$n(\text{MnO}_4^-)$ in titration $= 0.01 \times \frac{12.6}{1000} = 1.26 \times 10^{-4} \checkmark$ $n(\text{SO}_3^{2-})$ in 25.0 cm ³ $= 1.26 \times 10^{-4} \times 2.5 = 3.15 \times 10^{-4} \text{ (mol)}$ \checkmark $n(\text{SO}_3^{2-})$ in 250 cm ³ $= 10 \times 3.15 \times 10^{-4} = 3.15 \times 10^{-3} \text{ (mol)}$ \checkmark mass Na ₂ SO ₃ in 525 g meat $= 3.15 \times 10^{-3} \times 126.1 = 0.397 \text{ (g)} \checkmark$ mass Na ₂ SO ₃ in 1 kg of meat $= 0.397215 \times \frac{1000}{525} = 0.7566 \text{ g OR } 756.6 \text{ mg}$ AND less than the maximum permitted level OR AW \checkmark	5 (AO 1.2x1) (AO 2.8x3) (AO 3.2x1)	ALLOW 3 SF or more throughout ALLOW ECF throughout Calculator = 0.397215 g ALLOW within range: 756 to 757 mg ALLOW 0.397 g < 0.446 g per 525 g meat. <u>Examiner's Comments</u> 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.
		Total	5	
3 9		C	1 (AO1.2)	
		Total	1	
4 0	a i	Complete circuit with voltmeter AND labelled salt bridge linking two half-cells \checkmark	3 (AO1.2x3)	Half cells can be drawn in either order Half cells must show electrodes dipping into solutions ALLOW small gaps in circuit



		<p>Cr electrode in Cr³⁺ ✓</p> <p>Pt electrode in MnO₄⁻ AND H⁺ AND Mn²⁺ ✓</p>		<p>IGNORE any stated concentrations</p> <p>IGNORE state symbols</p> <p>In salt bridge, ALLOW any stated ion that may be present, e.g. Cr³⁺, MnO₄⁻, Mn²⁺, H⁺</p> <p>Examiner's Comments Most candidates were able to draw a circuit containing a voltmeter and the chromium half-cell but very few candidates included the H⁺ in the MnO₄⁻/Mn²⁺ cell.</p>
	ii	$5\text{Cr} + 3\text{MnO}_4^- + 24\text{H}^+ \rightarrow 5\text{Cr}^{3+} + 3\text{Mn}^{2+} + 12\text{H}_2\text{O} \checkmark$	1 (AO2.6)	<p>IGNORE state symbols</p> <p>ALLOW multiples</p> <p>Examiner's Comments Most candidates were able to construct the equation for the overall cell reaction and to use oxidation numbers to explain disproportionation.</p>
b	i	<p>Mn is oxidised from +6 (in MnO₄²⁻) to +7 (in MnO₄⁻) ✓</p> <p>Mn is reduced from +6 (in MnO₄²⁻) to +4 (in MnO₂) ✓</p>	2 (AO2.1x2)	<p>IGNORE '6' (signs required)</p> <p>ALLOW after number, e.g. 5+</p> <p>ALLOW 1 mark for correct oxidation numbers but not linked to oxidation/reduction.</p> <p>IGNORE any reference to electron loss/gain (even if wrong)</p> <p>Examiner's Comments Most candidates were able to construct the equation for the overall cell reaction and to use oxidation numbers to explain disproportionation.</p>
	ii	<p><i>Explanation using E° values</i> (E° of) system 3 (MnO₄⁻/MnO₄²⁻) is less positive / more negative than system 5 (MnO₄²⁻/MnO₂) ✓</p> <p><i>Equilibrium shift related to E° values</i> system 3 (MnO₄⁻/MnO₄²⁻) shifts left AND system 5 (MnO₄²⁻/MnO₂) shifts right ✓</p>	2 (AO3.1x2)	<p>IGNORE 'lower/higher'</p> <p>ALLOW reverse argument: System 5 more positive than system 3, etc Must be comparative</p> <p>ALLOW response in terms of E_{cell} E = (+)1.14 V for system 5 – system 3</p> <p>Shift dependent on systems 3 and 5 correctly identified</p> <p>Examiner's Comments Only a few candidates were able to link the electrode potentials and equilibrium</p>



				<p>shifts. Many candidates did not clarify which redox system they were referring to and only stated one of the ions. Often only one equilibrium shift was mentioned, and it was not linked to any redox system. Centres should advise candidates to use the redox system numbers found within the question to aid communication. Candidates should avoid using the terms 'higher/lower' to compare the E cell values.</p> <p style="text-align: center;">  OCR support </p> <p>Further guidance on electrode potentials can be found in our delivery guide: Delivery Guide for OCR AS/A Level Chemistry A</p>	
	c	i	$\text{H}_2 + 2\text{OH}^- \rightarrow 2\text{H}_2\text{O} + 2\text{e}^- \checkmark$	1 (AO2.6)	<p>ALLOW multiples ALLOW $\text{H}_2 + 2\text{OH}^- - 2\text{e}^- \rightarrow 2\text{H}_2\text{O}$ ALLOW equation with equilibrium sign</p>
		ii	$(0.40 - 1.23 =) -0.83 \text{ (V)} \checkmark$	1 (AO1.2)	
		iii	Fuel reacts with oxygen/oxidant to give electrical energy/voltage \checkmark	1 (AO1.1)	<p>ALLOW named fuel. e.g. hydrogen/H_2; ethanol; methanol, etc</p> <p>ALLOW fuel cell requires <u>continuous</u> supply of fuel AND oxygen/an oxidant OR fuel cell operates <u>continuously</u> as long as a fuel AND oxygen/an oxidant are added</p> <p>IGNORE 'reactants' 'products' and comments about pollution and efficiency</p>
			Total	11	
4 1	a		<p>Formula: $\text{CuCO}_3 \checkmark$</p> <p>$\text{CuCO}_3 + 2\text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{CO}_2 + \text{H}_2\text{O} \checkmark$</p>	2 (AO1.2) (AO2.6)	<p>IGNORE state symbols ALLOW formula within equation.</p> <p>ALLOW other copper(II) compounds which can react with nitric acid to form a gas e.g. CuS, CuSO_3 for mark 1, with correct equation for mark 2. e.g. $\text{CuSO}_3 + 2\text{HNO}_3 \rightarrow \text{Cu}(\text{NO}_3)_2 + \text{SO}_2 + \text{H}_2\text{O}$</p>
	b		$2\text{Cu}^{2+}(\text{aq}) + 4\text{I}^-(\text{aq}) \rightarrow 2\text{CuI}(\text{s}) + \text{I}_2(\text{aq})$	1 (AO2.6)	<p>ALLOW multiples State symbols are required</p>



	c	<p>starch (solution) AND blue-black to colourless ✓</p>	1 (AO1.2)	<p>ALLOW blue OR black OR purple for colour of mixture</p> <p>ALLOW blue colour disappears (to colourless) IGNORE 'clear' IGNORE 'colorimetry'</p>	
	d	<p>FIRST CHECK THE ANSWER ON ANSWER LINE If answer = 1.35 award 4 marks</p> <p>-----</p> $n(\text{S}_2\text{O}_3^{2-}) = 0.0200 \times \frac{26.55}{1000}$ $= 5.31 \times 10^{-4} \text{ (mol) } \checkmark$ $n(\text{I}_2) = 2.655 \times 10^{-4} \text{ OR}$ $n(\text{Cu}^{2+}) = 5.31 \times 10^{-4} \text{ (mol) } \checkmark$ $m(\text{Cu/Cu}^{2+}) \text{ in ore} = 63.5 \times 5.31 \times 10^{-4}$ $= 0.0337 \dots \text{ (g) } \checkmark$ $\text{percentage} = \frac{0.0337 \dots}{2.50} \times 100$ $= 1.35 \text{ (\%)} \checkmark \text{ (3SF required)}$	4 (AO2.8x5)	<p>FULL ANNOTATIONS MUST BE USED</p> <p>-----</p> <p>ALLOW ECF throughout</p> <p>If 1:2 ratio for $\text{I}_2:\text{Cu}^{2+}$ not used check ratio in b) and allow ECF</p> <p>IGNORE rounding errors after 3 SF</p> <p>Calculator: 0.0337185 ALLOW 3 SF (0.0337) up to calculator value</p> <p>ECF dependent on the use of a calculated mass of Cu/Cu^{2+}</p> <p>Examiner's Comments This was a percentage by mass calculation. Many candidates correctly calculated the number of moles of thiosulphate and used the ratio to calculate the moles of iodine. Many candidates then either linked this, incorrectly, to the number of moles of iodide or used an incorrect mole ratio to find the moles of Cu^{2+}. Multiplication by the relative formula mass of copper was well understood and some candidates gained error carried forward marks for their calculation of the percentage.</p>	
	e	i	Lower AND smaller titre ✓	1 (AO3.4)	ALLOW less I_2 produced / less Cu^{2+} reacts
		ii	The same AND burette measures by difference ✓	1 (AO3.4)	ALLOW AW
	f		<p>Any two of the following:</p> <p>Make up a (standard solution) from Step 2 to a stated volume (e.g. 250 cm³)</p> <p>OR</p> <p>Repeat titrations</p>	2 (AO3.4x2)	



		<p>AND Take mean of concordant/closest titres/ identify anomalies</p> <p>OR</p> <p>lower $[S_2O_3]^{2-}$ to increase titre volume (to reduce the percentage error).</p> <p>OR</p> <p>higher $[S_2O_3]^{2-}$ so not to refill the burette.</p> <p>OR</p> <p>Use a 3 dec place balance (to reduce the percentage error).</p>		
		Total	12	
4 2	a	<p>6 marking points → 5 MAX</p> <p>-----</p> <p>--</p> <p>ALLOW labels 1, 2 and 3; A, B and C, etc, provided that meaning is clear</p> <p>-----</p> <p>--</p> <p>Oxidising agent AND equation</p> <p>$Cr_2O_7^{2-}$ is oxidising agent with C_2H_5OH /oxidises C_2H_5OH ✓</p> <p>$Cr_2O_7^{2-} + 8H^+ + 3C_2H_5OH \rightarrow 2Cr^{3+} +$ $7H_2O + 3CH_3CHO$ ✓</p> <p>Explanation for $Cr_2O_7^{2-}/Cr^{3+}$ and CH_3CHO/C_2H_5OH</p> <p>E for $Cr_2O_7^{2-}/Cr^{3+}$ is more +ve /higher /greater</p> <p>OR</p> <p>$E_{cell} = (+)1.527 V$ + sign not required</p> <p>OR</p> <p>$Cr_2O_7^{2-}/Cr^{3+}$ equilibrium shifts right ✓</p> <p>Reducing agent AND equation</p> <p>Cr^{3+} is reducing agent with FeO_4^{2-} /reduces FeO_4^{2-} ✓</p> <p>$2Cr^{3+} + 2H^+ + 2FeO_4^{2-} \rightarrow C_2O_7^{2-} +$ $H_2O + 2Fe^{3+}$ ✓</p> <p>Explanation for $Cr_2O_7^{2-}/Cr^{3+}$ and FeO_4^{2-}</p>	<p>5 (AO2.5) (AO2.6) (AO2.6) (AO2.5) (AO2.6) (AO2.6)</p>	<p>ALLOW reverse argument (ORA) throughout</p> <p>For equations, ALLOW multiples In equations, ALLOW \rightleftharpoons for \rightarrow</p> <p>ALLOW $Cr_2O_7^{2-}$ is oxidising agent if linked to C_2H_5OH as reactant in equation</p> <p>ALLOW Cr^{6+} for $Cr_2O_7^{2-}$ ALLOW $Cr_2O_7^{2-}$ is reduced by C_2H_5OH</p> <p>In explanation, look for CONs between 'OR' statements</p> <p>ALLOW Cr^{3+} is reducing agent if clearly linked to FeO_4^{2-} as reactant in equation</p> <p>ALLOW Fe^{6+} for FeO_4^{2-} ALLOW Cr^{3+} is oxidised by FeO_4^{2-}</p> <p>In explanation, look for CONs between 'OR' statements</p> <p>-----</p> <p>Note on equations There are 2 marks for the equations with H^+, H_2O and e^- cancelled down</p>



	<p>/Fe³⁺ <i>E</i> for Cr₂O₇²⁻/Cr³⁺ is less +ve (<i>E</i>) / lower /smaller OR <i>E</i>_{cell} = (+)0.87 V + sign not required OR Cr₂O₇²⁻/Cr³⁺ equilibrium shifts left ✓</p>		<p>ALLOW 1 mark for 2 'correct' equations where H⁺, H₂O and e⁻ have NOT all been cancelled down.</p> <p><i>e.g.</i> 1 mark from 2 uncancelled equations Cr₂O₇²⁻ + 14H⁺ + 3C₂H₅OH → 2Cr³⁺ + 6H⁺ + 7H₂O + 3CH₃CHO</p> <p>2Cr³⁺ + 2H⁺ + 2FeO₄²⁻ + 6e⁻ → Cr₂O₇²⁻ + H₂O + 2Fe³⁺ + 6e⁻</p> <p><u>Examiner's Comments</u></p> <p>Most candidates were able to obtain some credit for this problem. The equations proved to be among the easier marks, and it was encouraging to see how many candidates were able to use the <i>E</i> data to help them combine the half equations correctly.</p> <p>Most then cancelled down the common species present on both sides of the equation: H⁺, H₂O and e⁻.</p> <p>Explanations were generally sound, either by comparing the <i>E</i> values, stating the direction of equilibrium shift, or by using cell potentials.</p> <p>The hardest marks proved to be the identification of the chromium ions that were the oxidising and reduction agents. Many candidates just referred to 'chromium' despite the chromium half equation containing two chromium ions, Cr₂O₇²⁻ and Cr³⁺.</p> <p>Candidates who had written correct equations sometimes assigned the wrong chromium ion as the oxidising or reducing agent. More successful responses constructed correct equations and stated that Cr₂O₇²⁻ was the oxidising agent as it oxidised C₂H₅OH and that Cr³⁺ was the reducing agent as it reduced FeO₄²⁻.</p>
b	<p>5 H₂S + 2 MnO₄⁻ + 6 H⁺ → 2 Mn²⁺ + 5 S + 8 H₂O OR 40 H₂S + 16 MnO₄⁻ + 48 H⁺ → 16 Mn²⁺ +</p>	<p>2 (AO3.2)</p>	<p>ALLOW multiples <i>e.g.</i> 2½ H₂S + MnO₄⁻ + 3 H⁺ → Mn²⁺ + 2½ S + 4 H₂O</p>



		<p>$5 \text{ S}_8 + 64 \text{ H}_2\text{O}$</p> <p>Any FIVE correct species ✓</p> <p>Correct balanced equation ✓</p>		<p>$20 \text{ H}_2\text{S} + 8 \text{ MnO}_4^- + 24 \text{ H}^+$ $\rightarrow 8 \text{ Mn}^{2+} + 2\frac{1}{2} \text{ S}_8 + 32 \text{ H}_2\text{O}$</p> <p>IGNORE extra species containing: Mn, H, S and O ONLY BUT ALLOW KMnO_4 on LHS, forming K^+ on RHS</p> <p>IGNORE electrons</p> <p>IGNORE state symbols</p> <p><u>Examiner's Comments</u></p> <p>Candidates needed to interpret the information provided and to use this as the basis for their redox equation. The clue of a yellow product proved to be very difficult to interpret as being sulphur. The equation then required H^+ to be added as a reactant ('acidified' in the information) and H_2O as the other product. Balancing required use of oxidation numbers.</p> <p>Candidates found this equation very difficult and relatively few correct equations were seen. The mark scheme did allow 1 mark for any correct five species but the correct equation proved to be challenging in demand.</p>
		Total	7	
4 3		<p>FIRST CHECK ANSWER ON THE ANSWER LINE If answer = 38 (mg) award 4 marks</p> <p>-----</p> <p>--</p> <p>$n(\text{I}_2) = 22.50 \times \frac{9.60 \times 10^{-4}}{1000} = 2.16 \times 10^{-5} \text{ (mol) } \checkmark$</p> <p>$n(\text{vitamin C})$ in 250 cm^3 volumetric flask $= 10 \times 2.16 \times 10^{-5} = 2.16 \times 10^{-4} \text{ (mol) } \checkmark$ $M(\text{Vitamin C: C}_6\text{H}_8\text{O}_6) = 176 \text{ OR } (12 \times 6) + (1 \times 8) + (16 \times 6) \text{ Seen anywhere}$ Mass vitamin C in 150 cm^3 of orange $= 2.16 \times 10^{-4} \times 176.0 = 0.038016 \text{ g}$ $= 38 \text{ (mg) } \checkmark$ 2 SF or more</p>	4 (4 xAO2.8)	<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> <p>-----</p> <p>---</p> <p>COMMON ERRORS: 22.81 mg scaling by 150/250 → 3 marks FINAL MARK LOST BY SCALING</p>



Determine the mass, in mg, of vitamin C in a 150 cm³ serving of the orange juice.

$$\begin{aligned}
 0.0275 \times 9.60 \times 10^{-4} &= 2.64 \times 10^{-5} \text{ mol} \quad \checkmark & C_6H_8O_6 \\
 2.16 \times 10^{-5} \times 10 &= 2.16 \times 10^{-4} \text{ mol} \quad \checkmark & = 176 \quad \checkmark \\
 \frac{2.16 \times 10^{-4}}{0.250} &= 8.64 \times 10^{-4} \text{ mol dm}^{-3} \\
 8.64 \times 10^{-4} \times 0.15 &= 1.296 \times 10^{-4} \text{ mol} \\
 1.296 \times 10^{-4} \times 176 &= 0.0228096 \text{ g} \\
 0.0228 \times 1000 &= 22.8 \quad \times
 \end{aligned}$$

42.24 mg using 25.0 cm³ instead of 22.50 → 3 marks

25.34 mg using 25.0 cm³ **AND** scaling by 150/250 instead of 22.50

→ 2 marks

63.36 mg scaling by 250/150 → 3 marks

Examiner's Comments

This question was a standard titration calculation, set in a practical context. As with Question 6 (c) (i), this assessed one of the important principles encountered in A Level Chemistry. Success required three main stages:

- Calculation of the amount, in mol, of I₂ used in the titration of 25.0 cm³ of the diluted orange juice.
- Determination of the amount in mol of vitamin C in the 250 cm³ solution (effectively scaling by ×10)
- Use of this value with the molar mass of vitamin C to determine the mass of vitamin C, in mg, in the 150 cm³ serving of the orange juice.

Most candidates calculated the amount of I₂ (and vitamin C) in the titre as 2.16 × 10⁻⁵ mol. The next scaling stage by ×10 to 2.16 × 10⁻⁴ mol introduced a problem. Many candidates were distracted by the '150 cm³' serving of orange juice and they scaled further by a factor of 150/250 to give 1.296 × 10⁻⁴ mol.

Candidates scaling corrected by ×10 usually used the vitamin C molar mass of 176 g mol⁻¹ to determine the correct mass of 38 mg in the orange juice serving.

The common incorrect mass of 22.8 mg



				<p>resulted from the extra scaling by 150/250 described above. This could still be given 3 of the available 4 marks as only one error had been made.</p> <p>Some candidates worked out an incorrect molar mass for vitamin C, with 174 often seen. Provided that this value was used with a correct method, ECF could still be applied for the final mass.</p>	
			Total	4	
4 4			C	1(AO2.8)	<p><u>Examiner's Comments</u></p> <p>The correct answer C required candidates to calculate the moles of manganate (VII) ions and then to use the stoichiometry of the equation to calculate the moles of H₂O₂ before calculating the concentration. Option A provided a distractor with the incorrect ratio. Option B was obtained without the reference to the molar ratio.</p>
			Total	1	
4 5		i	<p>Ca fizzes faster AND Ca dissolves/disappears more quickly ✓</p>	1 (AO2.3)	<p>CARE Both needed for 1 mark.</p> <p>ORA ALLOW AW</p> <p>IGNORE finishes first IGNORE more bubbles (need idea of rate) IGNORE exothermic</p> <p><u>Examiner's Comments</u></p> <p>Very few candidates made two valid statements where both clearly indicated an idea of relative rate – in almost all cases one of the descriptions would be about quantity of gas rather than rate of gas production. Some candidates identified a precipitate being formed, colour change, or gave a general answer of the reaction happening quicker.</p>
		ii	<p>Oxidation $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^- \checkmark$ Reduction $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$</p>	2 (AO2.6x2)	<p>In half equations, ALLOW the use of e for e⁻</p> <p>ALLOW $\text{Mg} - 2\text{e}^- \rightarrow \text{Mg}^{2+}$</p>




			<p>OR $\text{H}^+ + \text{e}^- \rightarrow \frac{1}{2}\text{H}_2 \checkmark$</p>		<p>IGNORE state symbols even is wrong BUT half equations MUST only have species that change.</p> <p>For charges on half equations,</p> <p>ALLOW Mg^{+2} for Mg^{2+}</p> <p>OR H^{+1} for H^+</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>Some candidates coped well with this question which was based on the AS part of the specification and gained both marks. More candidates gained 1 mark through writing one half equation, usually the oxidation of magnesium. Common errors were for chlorine to featuring in the reduction half equation and the lack of electrons in their answers. Very few candidates mixed up the oxidation and reduction equations.</p>
			Total	3	
4 6		i	<p>$\text{Fe}^{2+} + 2\text{OH}^- \rightarrow \text{Fe}(\text{OH})_2 \checkmark$</p>	1(AO2.6)	<p>IGNORE state symbols, even if wrong</p> <p>ALLOW</p> <p>$[\text{Fe}(\text{H}_2\text{O})_6]^{2+} + 2\text{OH}^- \rightarrow \text{Fe}(\text{OH})_2(\text{H}_2\text{O})_4 + 2\text{H}_2\text{O}$</p> <p>OR</p> <p>$[\text{Fe}(\text{H}_2\text{O})_6]^{2+} + 2\text{OH}^- \rightarrow \text{Fe}(\text{OH})_2 + 6\text{H}_2\text{O}$</p> <p><u>Examiner's Comments</u></p> <p>Most students scored this mark, although several gave no response.</p>




	<p>Explanation of the brown precipitate</p> <p>The brown ppt is Fe(OH)₃</p> <p>OR</p> <p>Fe(OH)₂ loses electrons/ Fe(OH)₂ oxidised ✓</p> <p>Comparison of E values</p> <p>(E of) Fe/Redox system 1 is more negative/less positive</p> <p>(than E of O₂/redox system 2)</p> <p>OR</p> <p>(E of) O₂/Redox system 2 is more positive/less negative</p> <p>(than E of Fe/redox system 1) ✓</p> <p>ii Equilibrium shift</p> <p>More negative/less positive OR Fe system OR Redox system</p> <p>1 shifts left</p> <p>OR</p> <p>More Positive/less negative OR O₂ system OR Redox system</p> <p>2 shifts right ✓</p> <p>Equation</p> $4\text{Fe}(\text{OH})_2(\text{s}) + \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) \rightarrow 4\text{Fe}(\text{OH})_3(\text{s}) \checkmark$	<p>4(AO3.1x4)</p>	<p>ORA</p> <p>ALLOW Fe²⁺ is oxidised to Fe³⁺</p> <p>ALLOW Fe</p> <p>ALLOW E_{cell} is (+) 0.96V</p> <p>IGNORE 'lower/higher'</p> <p>For equilibrium shift</p> <p>ALLOW E_{cell} is +ve therefore the reaction is feasible.</p> <p>OR</p> <p>Direction of half equation correctly written.</p> <p>ALLOW multiples</p> <p>ALLOW equilibrium</p> <p>IGNORE state symbols, even if wrong</p> <p>DO NOT ALLOW uncanceled species</p> <p><u>Examiner's Comments</u></p> <p>Although a spread of marks across the full available range was seen, a good proportion of candidates gained 3 or 4 marks. Most candidates were able to produce a balanced equation, but candidates should take care cancelling out any species present on both side of the equation, e.g. the hydroxide ions. A common error within the formula of iron (III) hydroxide was to place the number of hydroxide ions within the brackets, e.g. Fe(OH₃). Candidates are advised to read the instructions contained within the equation and to use or comment on all the data presented. When commenting on electrode potentials, candidates should avoid the use of higher/lower as these phrases are meaningless due to the negative signs involved.</p>
	<p>Total</p>	<p>5</p>	



4 7			C	1 (AO 2.6)	Examiner's Comments Most candidates answered this question correctly with C.			
			Total	1				
4 8			B	1 (AO 2.1)	<p>Examiner's Comments</p> <p>This question was for the most part answered correctly with B. Errors came from not recognising the reaction is endothermic and therefore its equilibrium would shift to the left when the temperature decreases, ruling out option 2.</p> <p> Assessment for learning</p> <p>Practice multiple choice questions can improve the skill in solving and identifying the distractors. Exposure to this type of question style will decrease the time taken over each question. These can often form the basis of end of topic tests.</p> <p>Multiple choice question quizzes can be found via the resource-finder on Teach Cambridge and there are instructions on how to use the online versions of the multiple choice quizzes.</p>			
			Total	1				
4 9	a	i	Colourless to (pale) pink	1 (AO 1.1)	<p>ALLOW <u>Pale</u> purple</p> <p>DO NOT ALLOW purple</p> <p>Examiner's Comments</p> <p>Only a few candidates were given this mark. Some had the colour change inverted and most stated a variety of different colours.</p>			
		ii	<table border="1" data-bbox="248 1989 608 2051"> <tbody> <tr> <td>12.65</td> <td>12.95</td> <td>12.75</td> </tr> </tbody> </table>	12.65	12.95	12.75	2 (AO 2.8 x 2)	<p>Examiner's Comments</p> <p>Almost all candidates calculated the titres correctly. A significant number</p>
12.65	12.95	12.75						



		<p>✓</p> $\frac{12.65+12.75}{2} = 12.7(0) \text{ cm}^3 \checkmark$		<p>used all three titre values to derive their mean value.</p> <p> OCR support</p> <p>Links to the legacy coursework tasks and PAG practice question sets can be found on Teach Cambridge and can help students prepare for practical-based questions like this one. Exam hints for students is useful to share with candidates.</p>
	iii	<p>FIRST CHECK THE ANSWER ON ANSWER LINE if answer = 6.35×10^{-3} award 3 marks</p> <p>-----</p> <p>$n(\text{MnO}_4^-)$ in titration $= (0.00250 \times \frac{12.7}{1000})$ $= 3.175 \times 10^{-5} \checkmark$</p> <p>$n(\text{Fe}^{2+})$ in 25.0 cm³ $= (3.175 \times 10^{-5} \times 5)$ $= 1.5875 \times 10^{-4} \text{ (mol)} \checkmark$</p> <p>$[\text{Fe}^{2+}] = (1.5875 \times 10^{-4} \div 0.025)$ OR $(1.5875 \times 10^{-4} \times 40)$ $= 6.35 \times 10^{-3} \text{ (mol dm}^{-3}\text{)} \checkmark$</p>	<p>3 (AO 2.8 x 3)</p>	<p>ALLOW ECF from incorrect titre in 21 a ii) for 3 marks e.g. Titre of 12.78 cm³ gives 6.39×10^{-3}</p> <p>-----</p> <p>ALLOW 3 SF or more throughout</p> <p>ALLOW ECF throughout</p> <p>ALLOW $n(\text{Fe}^{2+})$ in 250 cm³ = 1.5875×10^{-3} (mol) so $[\text{Fe}^{2+}]$ in 25 cm³ $= 1.5875 \times 10^{-3} \div 0.25 = 6.35 \times 10^{-3}$</p> <p>Common errors for 2 marks</p> <p>2.46×10^{-2} (volumes transposed) 1.25×10^{-2} (same volume used twice) 1.27×10^{-3} (no x 5) 2.54×10^{-4} ($\div 5$)</p> <p><u>Examiner's Comments</u></p> <p>Candidates made good progress with this calculation, many gaining 2 or 3 marks, including error carried forward from incorrect titres. Common errors included, in various combinations: transposing volumes, not using the stoichiometry of the equation and using the same volume twice.</p>
b		<p>System $1/E^\ominus(\text{Zn})$ is more negative/less</p>	<p>4 (AO 3.1 x 1) (AO 3.4 x 1)</p>	<p>ALLOW ORA throughout IGNORE larger/smaller/greater/less throughout</p>



		<p>positive than <i>system 2</i> / $E^\ominus(\text{Fe}^{3+})$ ✓</p> <p>Eqm 2 shifts to right AND Eqm 1 shifts to left OR Zinc reduces iron(III) ions (to iron(II)) OR $\text{Zn} + 2\text{Fe}^{3+} \rightarrow \text{Zn}^{2+} + 2\text{Fe}^{2+}$ ✓</p> <p><i>System 1</i> / $E^\ominus(\text{Zn})$ is more negative than <i>system 3</i> / $E^\ominus(\text{MnO}_4^-)$ ✓</p> <p>Eqm 3 shifts to right AND Eqm 1 shifts to left OR (If unfiltered), MnO_4^- oxidise zinc OR $2\text{MnO}_4^- + 5\text{Zn} + 16\text{H}^+ \rightarrow 2\text{Mn}^{2+} + 5\text{Zn}^{2+} + 8\text{H}_2\text{O}$ ✓</p>	<p>(AO 3.1 × 1) (AO 3.4 × 1)</p>	<p>ALLOW $E^\ominus = (+)1.53(\text{V})$ ALLOW comparison if Fe system is identified</p> <p>ALLOW $E^\ominus = (+) 2.27(\text{V})$ ALLOW comparison if MnO_4^- is identified</p> <p><u>Examiner's Comments</u></p> <p>Most candidates were able to successfully explain Step 1, scoring the first 2 marks. When comparing electrode potentials, candidates should avoid the use of higher/lower as these phrases are ambiguous due to the negative signs involved. Describing them as 'more negative' or 'more positive' is clearer.</p> <p>Candidates are advised to read the instructions contained within the question and to use or comment on all the data presented. Very few candidates linked Step 2 to the reducing effect of zinc to manganate(VII) ions hence the need for filtration. A few candidates explained Step 2 in terms of the warming/cooling and zinc crystallising rather than explaining the redox chemistry given in the table.</p>
		Total	10	
50		<p>$3\text{V}^{3+} + \text{Cr}_2\text{O}_7^{2-} + 2\text{H}^+ \rightarrow 3\text{VO}_2^+ + 2\text{Cr}^{3+} + \text{H}_2\text{O}$</p> <p>ALL reactant and product species correct ✓</p> <p>Correct balancing (of correct equation) AND cancelling of species ✓</p>	<p>2 (AO 2.5) (AO 2.6)</p>	<p>IGNORE Balancing and electrons for first mark</p> <p>DO NOT ALLOW electrons in final answer</p> <p><u>Examiner's Comments</u></p> <p>Very few candidates were able to produce the balanced overall equation; a few had the correct reactants and</p>



					products but not balanced. Candidates are advised to look for the information contained within the question. The formulas were given, and it was stated that the solution was acidified, leaving only water to be identified. Some candidates approached this through two half equations whereas others used oxidation numbers to balance their equations.
			Total	2	
5 1	a	i	$\text{Au} + 4 \text{HCl} \rightarrow 4 \text{H}^+ + \text{AuCl}_4^- + 3 \text{e}^- \checkmark$	1 (AO 1.2)	<p>Examiner's Comments</p> <p>Most candidates added '4' before HCl and H⁺, and 3 before e⁻ to gain this mark. Where an error was made, it invariably was with the number of electrons, usually 4e⁻.</p>
		ii	<p>Formulae</p> <p>X = NO ✓</p> <p>Z = H₂O ✓</p> <p>Equation Independent from ID of X and Z</p> <p>$\text{HNO}_3 + 3 \text{H}^+ + 3 \text{e}^- \rightarrow \text{NO} + 2 \text{H}_2\text{O}$ OR $\text{NO}_3^- + 4 \text{H}^+ + 3 \text{e}^- \rightarrow \text{NO} + 2 \text{H}_2\text{O} \checkmark$</p> <p>CHECK BELOW ANSWER SPACE FOR RESPONSE</p>	3 (AO 3.1 x3)	<p>If X and Z in wrong order award 1 out of 2 formula marks i.e. X = H₂O and Z = NO 1 mark</p> <p>ALLOW multiples</p> <p>Examiner's Comments</p> <p>Almost all candidates identified X and Y as NO and H₂O respectively, but the equation proved to be much more testing. Some candidates were careless, showing NO and H₂O the wrong way round (credited with 1 out of these 2 marks) or with charges.</p> <p>For the equation, candidates needed to consider the oxidation number change of N from +5 to +2, This should have naturally led to 3e⁻ being added on the left-hand side. Many candidates omitted the electrons entirely. Some did add 3e⁻ but on the right. This suggests that candidates would benefit with practising the construction of half equations.</p>
	b		Level 3 (5–6 marks)	6 (AO 3.1 x3)	Indicative scientific points may include:



(AO 3.2
x3)

- Reaches a comprehensive conclusion to determine all **three** correct formulae of **D, E AND F**
- **AND** constructs most equations with few errors

There is a well-developed line of reasoning which is clear and logically structured.

The information presented is relevant and substantiated.

Level 2 (3–4 marks)

- Reaches a comprehensive conclusion to determine **two** correct formulae of **D, E AND F**
- **AND** constructs some equations with some errors

There is a line of reasoning presented with some structure.

The information presented is relevant and supported by some evidence.

Level 1 (1–2 marks)

- Determines a correct formula for **one** of **D, E AND F**
- **AND** provides some evidence to support the formula

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

0 mark No response or no response worthy of credit.

EQUATIONS SHOULD BE USED TO INFORM THE COMMUNICATION STRAND

See next page for details

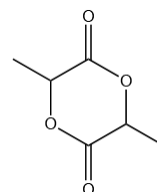
CHECK TOP OF QUESTION FOR RESPONSES

IGNORE CONNECTIVITY FOR F

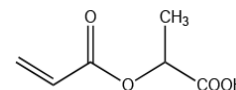
SUMMARY

Identify of D, E and F

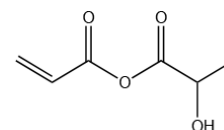
- **D:** $\text{NiSO}_4 \cdot 6\text{H}_2\text{O}$
OR $\text{NiSO}_4(\text{H}_2\text{O})_6$ **OR** $\text{NiSO}_{10}\text{H}_{12}$
- **E:** SO_2
- **F:** **Cyclic diester**



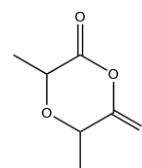
OR unsaturated ester/acid



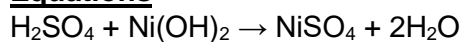
OR unsaturated acid anhydride



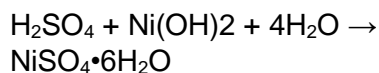
OR cyclic acid anhydride



Equations

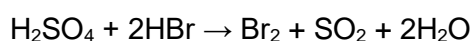


OR



For equation

ALLOW $\text{NiSO}_4 \cdot 6\text{H}_2\text{O}$ **OR** $\text{NiSO}_4(\text{H}_2\text{O})_6$





Setting the level

For Level 3 (5–6 marks),

- All 3 identified: **D, E and F**
- Most equations

For Level 2 (3–4 marks),

- 2 identified from **D, E and F**
- 2 equations

For Level 1 (1–2 marks),

- 1 identified from **D, E and F**
- Evidence

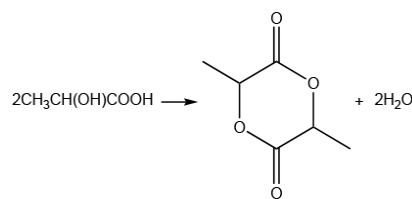
Evidence to support a formula for Level 1

Molar ratios of D

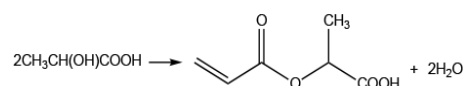
Ni	:	S	:	O	:	H	
<u>22.33</u>	:	<u>12.20</u>	:	<u>60.87</u>	:	<u>4.60</u>	
58.7	:	32.1	:	16.0	:	1.0	
0.38	:	0.38	:	3.80	:	4.60	
1	:	1	:	10	:	12	OR NiSO ₁₀ H ₁₂

Molar mass of E

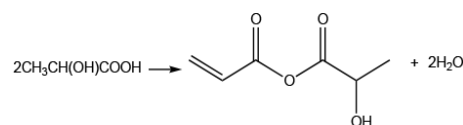
$$\text{Molar mass} = 2.67 \times 24 = 64(.08) \text{ g mol}^{-1}$$



OR



OR



If structure of **F** is shown, **ALLOW** equation using molecular formulae, e.g. $2\text{C}_3\text{H}_6\text{O}_3 \rightarrow \text{C}_6\text{H}_8\text{O}_4 + 2\text{H}_2\text{O}$

Examiner's Comments

This level of response question required candidates to interpret three pieces of information to identify 3 unknown chemicals, linked by three reactions of sulfuric acid. Levels were assigned based on identifying the three unknowns and writing equations for the reactions. The 3 reactions were tiered in difficulty with the cyclic structure for compound **F** being the most difficult.

The question discriminated extremely well, with comparatively few candidates not scoring any marks.

Exemplar 4



				<p>the solve. They initially went for a straight chain structure and then worked out that there must be a cyclic structure. They finally show the cyclic diester as a conventional structural (based on glucose?), and with a balanced equation.</p> <p>This was an excellent response of a high-ability candidate and shows what a well-prepared candidate is capable of achieving.</p>
			Total	10
5 2	a	<p>Complete circuit AND voltmeter AND labelled salt bridge linking two half-cells ✓</p> <p>Pt AND Fe²⁺ AND Fe³⁺ ✓</p> <p>Pt AND H₂ AND H⁺ AND delivery system for H₂ gas ✓</p> <p>Standard conditions 1 mol dm⁻³ AND Temperature: 298 K / 25 °C AND Pressure: 1 atm / 100 kPa/101 kPa ✓</p>	4	<p>Electrodes / salt bridge must at least touch the surface of solutions ALLOW small gaps in circuit wires</p> <p>ALLOW half-cells drawn on either side</p> <p>ALLOW a formula of a strong acid for H⁺</p> <p><u>For standard conditions:</u></p> <p>Can be awarded if all quoted on standard condition line or in labelled diagram.</p> <p>ALLOW 1M</p> <p>ALLOW equimolar solutions for Fe²⁺ AND Fe³⁺ only. i.e. need 1 mol dm⁻³ for [H⁺]</p> <p>IGNORE H₂SO₄ in diagram unless concentration is stated with a value other than 0.5 moldm⁻³</p> <p>DO NOT ALLOW if any concentration is incorrect</p> <p><u>Examiner's Comments</u></p> <p>Successful candidates drew a complete circuit and voltmeter with the labelled salt bridge dipped into the two solutions. Both cells had Pt as the electrodes. One cell contained 1 moldm⁻³ Fe²⁺ and Fe³⁺ where the other had a delivery mechanism for H₂ (at 1 atm) and H⁺ (at</p>



				<p>1 mol dm⁻³). Standard temperature of 298K (or 25°C) was stated.</p> <p>Common errors included: not have solutions in the beakers, used Fe or Fe²⁺ electrodes, the hydrogen cell was missing and/or the H₂ with a device for adding it. A few candidates suggested two cells with only Fe²⁺ one side and only Fe³⁺ on the other.</p>
b	i	Li + CoO ₂ → LiCoO ₂ ✓	1	<p>ALLOW ⇌</p> <p>DO NOT ALLOW uncanceled species</p> <p><u>Examiner's Comments</u></p> <p>Most candidates were successful in constructing this equation. Common errors included the backwards reaction and some with unbalanced Li⁺. Candidates should also check their equations, as a few included an erroneous C in the formula e.g. LiCoCO₂.</p>
	ii	<p>Cell potentials:</p> <p>(E°) = 1.23 - 0.00 OR 1.23 (V)</p> <p>OR (redox system 6 – redox system 3) = 1.23 (V)</p> <p>AND</p> <p>(E°) = 0.40 – (–0.83) = 1.23 (V)</p> <p>OR (redox system 4 – redox system 2) = 1.23 (V)</p> <p>✓</p> <p>Acidic <i>Cell equation</i></p> <p>(2x) H₂ ⇌ 2H⁺ + 2e⁻</p> <p>AND</p> <p>O₂ + 4H⁺ + 4e⁻ ⇌ 2H₂O</p>	3	<p>IGNORE state symbols throughout</p> <p>ALLOW multiples</p> <p>ALLOW ⇌</p> <p>Overall equation AND</p> <p>with evidence of working: e.g. half-equations written out / combined but not cancelled / with crossings out OR <u>System 6</u> goes forward / reduced OR <u>system 3</u> goes backwards / oxidised</p> <p>ALLOW multiples</p> <p>Overall equation AND</p> <p>with evidence of working: e.g. half-equations written out / combined but not cancelled / with crossings out OR <u>System 4</u> goes forward / reduced OR <u>system 2</u> goes backwards oxidised</p> <p>ALLOW 1 mark for both equations with</p>



		<p>AND</p> $2 \text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \checkmark$ <p>Alkaline <i>Cell equation</i></p> $(2\times) \text{H}_2 + 2\text{OH}^- \rightleftharpoons 2\text{H}_2\text{O} + 2\text{e}^-$ <p>AND</p> $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \rightleftharpoons 4\text{OH}^-$ <p>AND</p> $2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} \checkmark$		<p>uncancelled species.</p> <p>ALLOW 1 mark for <u>System 6</u> / reduced goes forward and <u>system 3</u> goes backwards oxidised AND <u>System 4</u> / goes forward / reduced and <u>system 2</u> / goes backwards / oxidised</p> <p><u>Examiner's Comments</u></p> <p>Most candidates were able to successfully show the overall equations were the same. Most combined the two half equations and then cancelled down to give the overall equations. A few candidates chose to describe the redox nature or use the equilibrium shifts scoring the first 2 marks. Some candidates, incorrectly, deduced the equation as $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$. Some candidates gave the 1.23V value without showing any calculation or did not refer to this part of the question at all. This was one of the most common questions where candidates omitted to provide a response on the paper.</p>
		Total	8	
5 3	i	$\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^- \checkmark$ <p>OR $\text{Fe}^{2+} - \text{e}^- \rightarrow \text{Fe}^{3+}$</p> $\text{C}_2\text{O}_4^{2-} \rightarrow 2\text{CO}_2 + 2\text{e}^- \checkmark$ <p>OR $\text{C}_2\text{O}_4^{2-} - 2\text{e}^- \rightarrow 2\text{CO}_2$</p>	2	<p>For both half-equations,</p> <p>ALLOW multiples</p> <p>ALLOW e for e⁻</p> <p>IGNORE state symbols.</p> <p>ALLOW $\text{C}_2\text{O}_4^{2-} \rightarrow \text{C}_2\text{O}_4^- + \text{e}^-$</p> $\text{C}_2\text{O}_4^{2-} \rightarrow \text{C}_2\text{O}_4 + 2\text{e}^-$ $2\text{H}_2\text{O} + \text{C}_2\text{O}_4^{2-} \rightarrow 2\text{CO}_3^{2-} + 4\text{H}^+ + 2\text{e}^-$ <p>ALLOW $2\text{C}_2\text{O}_4^{2-} \rightarrow \text{C}_4\text{O}_8^{2-} + 2\text{e}^-$</p> <p><u>Examiner's Comments</u></p> <p>Most candidates were given a mark for</p>



				<p>one of their half-equations.</p> <p>The half-equation for the oxidation of iron(II) ions was correct more often than that for ethanedioate ions. Strangely, a significant number of half-equations started with Fe metal instead of Fe²⁺ ions. The mark scheme allowed for variations in the oxidation of ethanedioate ions, with several alternative equations being given from the correct equation forming CO₂ receiving a mark.</p>																
	ii	<table border="1"> <thead> <tr> <th></th> <th>1</th> <th>2</th> <th>3</th> </tr> </thead> <tbody> <tr> <td>Final reading/ cm³</td> <td>23.55</td> <td>45.40</td> <td>22.75</td> </tr> <tr> <td>Initial reading/ cm³</td> <td>1.90</td> <td>23.55</td> <td>1.20</td> </tr> <tr> <td>Titre / cm³</td> <td>21.65</td> <td>21.85</td> <td>21.55</td> </tr> </tbody> </table> <p>Readings recorded to accuracy of burette</p> <p>All readings recorded to two decimal places with the last figure either 0 or 5</p> <p>AND</p> <p>Final and initial readings in correct rows ✓</p> <p>Correct titres</p> <p>All 3 titres correct to 2 DP: ✓✓</p> <p>2 titres correct to 2 DP: ✓</p>		1	2	3	Final reading/ cm ³	23.55	45.40	22.75	Initial reading/ cm ³	1.90	23.55	1.20	Titre / cm ³	21.65	21.85	21.55	3	<p>Examiner's Comments</p> <p>Most candidates were given all 3 marks for their titration tables. One mark was allocated to the readings, the other 2 for the correct titres.</p> <p>Candidates are expected to record burette readings to 2 decimal places with the last figure being 0 or 5, reflecting the accuracy of the burette, which is stated in the question. Two issues were seen:</p> <ul style="list-style-type: none"> Some candidates ignored the terminal 0 in their readings, e.g. 1.9 for 1.90 cm³. A small number of candidates recorded the initial and final readings the wrong way round despite these terms being used above each burette diagram. <p>Titres were usually correct.</p>
	1	2	3																	
Final reading/ cm ³	23.55	45.40	22.75																	
Initial reading/ cm ³	1.90	23.55	1.20																	
Titre / cm ³	21.65	21.85	21.55																	
	iii	<p>FIRST, CHECK THE ANSWER ON ANSWER LINE</p> <p>IF % error = 0.46, award 1 mark</p> <p>-----</p>	1	<p>Check Titres from 3b(ii) at top of response</p> <p>-----</p> <p>-----</p>																



	<p>-----</p> $\frac{2 \times 0.05}{21.65} \times 100 = 0.46 (\%) \checkmark \quad \text{2 DP minimum}$ <p>Calculator value: 0.46189...</p>		<p>ALLOW % error from ANY of the 3 titres from 3b(ii) OR from the mean titre</p> <p>DO NOT ALLOW 0.50%</p> <p>Examiner's Comments</p> <p>Candidates are much better at calculating percentage uncertainties than in previous series.</p> <p>Almost all candidates realised that the uncertainty in the titre resulted from two readings and that the overall uncertainty in the titre would be $2 \times 0.05 = 0.1 \text{ cm}^3$, resulting in a percentage uncertainty of 0.46%. Comparatively few candidates ignored the '2' and gave 0.23% as their answer.</p> <p>Some candidates worked out the percentage uncertainty from the mean titre but this was still given marks.</p>
	<p>Level 3 (5-6 marks) Analyses the results to calculate the correct amount of MnO_4^- using the correct mean titre from the candidate's titres AND Obtains correct value of x as 2 <i>There is a well-developed line of reasoning which is clear and logically structured.</i> <i>The information presented is relevant and substantiated.</i></p> <p>iv Level 2 (3-4 marks) Analyses titration results to determine an amount of MnO_4^- from a mean titre of the candidate's titres AND amount of FeC_2O_4 in 25.0 cm^3 OR 250 cm^3 OR uses a mass of FeC_2O_4 to obtain a value of x with few errors <i>There is a line of reasoning presented with some structure.</i> <i>The information presented is relevant and supported by some evidence.</i></p>	<p>6</p>	<p>*For mean titre, Check Titres from 3b(ii) at top of response*</p> <p>Indicative scientific points may include:</p> <p>Mean titre and $n(\text{MnO}_4^-)$</p> $\text{Mean titre} = \frac{(21.65 + 21.55)}{2} = 21.6(0) \text{ (cm}^3\text{)}$ $n(\text{MnO}_4^-) = 0.0200 \times \frac{21.6(0)}{1000} = 4.32 \times 10^{-4} \text{ (mol)}$ <p>Amount of FeC_2O_4 in mol</p> $n(\text{FeC}_2\text{O}_4) \text{ in } 25.0 \text{ cm}^3 = 5/3 \times n(\text{MnO}_4^-)$ $= 7.2(0) \times 10^{-4} \text{ (mol)}$ $n(\text{FeC}_2\text{O}_4) \text{ in } 250 \text{ cm}^3 = 7.2(0) \times 10^{-3} \text{ (mol)}$ <p>Value of x (final answer)</p> $\text{Molar mass } \text{FeC}_2\text{O}_4 \cdot x\text{H}_2\text{O} = \frac{1.295}{7.2(0) \times 10^{-3}}$ $= 179.9$ $\text{Molar mass of } x\text{H}_2\text{O} = 179.9 - 143.8 = 36(\dots)$ $x = 36/18 = 2$



Level 1 (1-2 marks)

Analyses results to determine an amount of MnO_4^- from the candidate's titres

OR

Analyses the information to obtain values of $n(\text{MnO}_4^-)$ and $n(\text{FeC}_2\text{O}_4)$ with some errors.

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 or no response worthy of credit.

Credit other correct methods,

e.g. For value of x

$$\text{Mass of FeC}_2\text{O}_4 = 7.2(0) \times 10^{-3} \times 143.8 = 1.03536 \text{ g}$$

$$\text{Mass of H}_2\text{O} = 1.295 - 1.035 = 0.25964 \text{ g}$$

$$n(\text{H}_2\text{O}) = \frac{0.25964}{18} = 0.0144 \text{ mol}$$

$$x = \frac{0.0144}{7.2 \times 10^{-3}} = 2$$

Responses using 25.0 cm^3 rather than the titres are limited to **Level 1**

For communication, a typical 'logical structure' would label most calculation steps in response

e.g.

Communication strand met

$$\begin{aligned} \text{KMnO}_4 \text{ mean titre: } & \frac{21.65 + 21.55}{2} = 21.6 \text{ cm}^3 \\ \text{or } n(\text{KMnO}_4) &= \frac{21.6 \times 10^{-3}}{1000} \times 0.02 \\ &= 4.32 \times 10^{-4} \\ n(\text{FeC}_2\text{O}_4 \cdot x\text{H}_2\text{O}) &= \frac{4.32 \times 10^{-4}}{3} \times 5 \\ &= 7.2 \times 10^{-4} \text{ in } 25 \text{ cm}^3 \\ n &= \frac{7.2 \times 10^{-4}}{25} \text{ in } 250 \text{ cm}^3 \\ n &= \frac{\text{mass}}{\text{M}_r} \\ \text{M}_r &= \frac{\text{mass}}{n} = \frac{1.295}{7.2 \times 10^{-3}} = 179.86 \end{aligned}$$

Communication strand not met

$$\begin{aligned} \frac{21.60}{1000} \times 0.02 &= 4.32 \times 10^{-4} \text{ mol} \\ \left(\frac{4.32 \times 10^{-4}}{3} \right) \times 5 &= 7.2 \times 10^{-4} \text{ mol in } 25 \text{ cm}^3 \\ (7.2 \times 10^{-4}) \times 10 &= 7.2 \times 10^{-3} \text{ mol in } 250 \text{ cm}^3 \\ \frac{1.295}{7.2 \times 10^{-3}} &= 179.86 \end{aligned}$$

Examiner's Comments

This unstructured titration problem was assessed by Level of Response (LOR).

Candidates answered this stock titration



calculation well. The key stages are listed below:

Determination of the mean titre from the two closest titres in the candidate's response to Question 3

1. (b) (ii). If the titres were correct, this would be 21.60 cm^3 , the mean of 21.55 cm^3 and 21.65 cm^3 .

2. Calculation of the number of moles of MnO_4^- as $4.32 \times 10^{-4} \text{ mol}$.

Calculation of the number of moles of FeC_2O_4 in 25 cm^3 as $5/3 \times 4.32 \times$

3. $10^{-4} = 7.20 \times 10^{-4} \text{ mol}$. Scaling up this number of moles by 10 for the moles in 250 cm^3 as $7.20 \times 10^{-3} \text{ mol}$.

4. Determination of the number of waters of crystallisation.

Most candidates completed Stages 1 and 2 correctly, securing a minimum of a Level 1 response.

A significant number of candidates then completed Stage 3, with most including scaling by 10 to secure a minimum of Level 2.

Many of these candidates compared the moles of FeC_2O_4 to the mass of $\text{FeC}_2\text{O}_4 \cdot x\text{H}_2\text{O}$ used to determine the value of x as 2. Such candidates would have reached Level 3.

The communication strand of the LOR mark was determined by the clarity of the response, particularly whether the numbers in the calculation had been labelled. Unfortunately, many 'correct' responses had not done this. Over half the candidates were given 5 or 6 marks for this stock calculation.




				<p>Exemplar 2</p> <p>Analyse the student's results to find the number of waters of crystallisation, x, in the hydrated iron(II) ethanedioate, FeC₂O₄·xH₂O.</p> <p>4.32 Mean titre: $\frac{21.65 + 21.55}{2} = 21.60 \text{ cm}^3$ [6]</p> <p>$\frac{21.60}{1000} \times 0.02 = 4.32 \times 10^{-4} \text{ mol}$</p> <p>$1.08 \times \frac{5}{2} = 2.7 \text{ mol}$</p> <p>$4.32 \times 10^{-4} \times \frac{5}{2} = 1.08 \times 10^{-3} \text{ mol}$</p> <p>$\times \frac{250}{25} = 0.0072$</p> <p>$\frac{1.295}{0.0072} = 180$</p> <p>179.86 180 $180 - 143.8 = 36.06$</p> <p>$\frac{36}{18} = 2$</p> <p>$\therefore x = 2 \therefore \text{FeC}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$</p> <p>This exemplar demonstrates the hazards of not communicating well.</p> <p>The candidate has followed the four stages described above and the chemistry behind the titration analysis is correct. However, the response is mostly a page of numbers, with no explanation about what the numbers mean.</p> <p>The response is clearly at Level 3 but it is impossible to credit the communication strand, receiving a total of 5 marks.</p>
		Total	12	
5 4	i	$4\text{PH}_3 + 8\text{O}_2 \rightarrow \text{P}_4\text{O}_{10} + 6\text{H}_2\text{O} \checkmark$	1	<p>ALLOW multiples</p> <p>ALLOW $2\text{PH}_3 + 4\text{O}_2 \rightarrow \text{P}_2\text{O}_5 + 3\text{H}_2\text{O}$</p> <p>IGNORE state symbols, even if wrong</p> <p><u>Examiner's Comments</u></p> <p>Candidates found this question quite challenging, with only about one-third writing a correct equation. The question gave the reactants and products with only the formula of phosphorus(V) oxide having to be worked out.</p> <p>The actual reaction does produce P₄O₁₀ but P₂O₅ was shown in almost all equations, and this was acceptable.</p>



				<p>Various incorrect formulae were seen for phosphorus(V) oxide including PO, PO₂, P₅O, HPO, etc. Unfortunately a significant number of candidates could not balance the equation, despite using correct formulae.</p>
		<p>6AgNO₃ + (1)PH₃ + 3H₂O → 6Ag + (1)H₃PO₃ + 6HNO₃ ✓</p> <p>Ag is reduced from +1 to 0 ✓</p> <p>ii P is oxidised from -3 to +3 ✓</p> <p>IGNORE oxidation numbers written around equation <i>Treat as rough working</i></p> <p>IGNORE reference to electrons <i>Question states oxidation numbers</i></p>	3	<p>ALLOW equation with '1' omitted, i.e. 6AgNO₃ + PH₃ + 3H₂O → 6Ag + H₃PO₃ + 6HNO₃ ✓ BUT DO NOT ALLOW '0'</p> <p>ALLOW 1 mark for BOTH correct oxidation number changes with 'reduced' and 'oxidised' omitted</p> <p>OR 'oxidised and reduced the wrong way round</p> <p>+ signs required for +1 and +3</p> <p>For oxidation numbers, ALLOW 1+, 3- and 3+</p> <p>Examiner's Comments</p> <p>This question generated a wide range of responses, testing many important chemical skills.</p> <p>Candidates often used oxidation numbers correctly to show that silver is reduced and phosphorus oxidised, with silver being the easier. Hydrogen was sometimes incorrectly chosen for oxidation.</p> <p>The oxidation number change of +1 to 0 for silver was usually correct although +9 and +11 were common errors for silver in AgNO₃, presumably by choosing the oxidation number of nitrogen as -3 or -5.</p> <p>Candidates usually recognised that phosphorus started with an oxidation number of -3 but the oxidation number of +5 was a common error in H₃PO₃.</p> <p>Balancing the equation was the most difficult part of this question with</p>



			<p>numbers being added almost at random. It is easier to balance equations for redox reactions by balancing the oxidation number changes first.</p> <p> Assessment for learning</p> <p>Ag⁺ and NO₃⁻ are among the common ions that students should know (see also Question 4 (c) (i)). For a NO₃⁻ ion to have a charge of 1-, the oxidation number of nitrogen must be +5. By choosing -5, the charge on NO₃ would be -11 and silver would have an oxidation number of +11. This is completely unrealistic and should be rejected as it points to a serious error.</p> <p>The specification states the following: <i>2.1.5 (a) rules for assigning and calculating oxidation number for atoms in elements, compounds and ions.</i></p> <p>This section will have been studied at the start of the two-year course and forms part of the backbone of chemical literacy.</p> <p>For success in chemistry, the ions should be learnt and the rules for assigning oxidation numbers need to be mastered.</p>
		Total	4