Redox and Electrode Potentials

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

Redox system	Half-equation	E^θ / V
1	$CO_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons HCOOH(aq)$	−0.11
2	$HCOOH(aq) + 2H^+(aq) + 2e^- \rightleftharpoons HCHO(aq) + H2O(l)$	−0.03
3	$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
4	$MnO_4{}^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51

Table 19.1

A student warms a mixture of methanal, HCHO, and acidified potassium manganate(VII).

The student observes gas bubbles.

Explain this observation in terms of electrode potentials and equilibria.

Include overall equations in your answer.

[4]

(b). Methanoic acid, HCOOH, can be used in a fuel cell. As with all fuel cells, the fuel (HCOOH) is supplied at one electrode and the oxidant (oxygen) at the other electrode.

The standard cell potential for this fuel cell is 1.34 V.

The overall reaction is shown below.

$HCOOH + ½O_2 \rightarrow H_2O + CO_2$

Using the information in **Table 19.1**, deduce the half-equation for the reaction at the oxygen electrode, and calculate the standard electrode potential for the oxygen half-cell.

half-equation --

standard electrode potential = ----------- V

[2]

2. Sodium sulfite(IV), Na_2SO_3, is used as a preservative in some foods.
Food safety legislation allows a maximum of 850 mg Na_2SO_3 per kg of burger meat.

A chemist determines the amount of Na_2SO_3 in a sample of burger meat using a manganate(VII) titration.

Step 1 The Na_2SO_3 from 525 g of burger meat is extracted to form a solution containing SO_3^{2-}(aq) ions.

Step 2 The solution from **step 1** is made up to 250.0 cm^3 in a volumetric flask with water. 25.0 cm^3 of this diluted solution is pipetted into a conical flask.

Step 3 The pipetted solution from **step 2** is acidified with dilute sulfuric acid and then titrated with 0.0100 mol dm^{-3} potassium manganate(VII), $KMnO_4$.

$2MnO_4^-(aq) + 6H^+(aq) + 5SO_3^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 3H_2O(l) + 5SO_4^{2-}(aq)$

12.60 cm^3 of $KMnO_4$(aq) is required to reach the endpoint.

Analyse the results to determine whether the burger meat complies with food safety legislation.

[5]

3. Standard electrode potentials for four redox systems are shown in **Table 19.1**.

Redox system	**Half-equation**	***E*** $^{\theta}$ **/ V**
1	$CO_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons HCOOH(aq)$	−0.11
2	$HCOOH(aq) + 2H^+(aq) + 2e^- \rightleftharpoons HCHO(aq) + H2O(l)$	−0.03
3	$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
4	$MnO_4{}^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+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]

4. 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 cm^3 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 cm^3 with distilled water.

Stage 3
A 25.00 cm^3 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**:

$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

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 cm_3.

	1	**2**	**3**
Final reading / cm^3	27.80	37.55	32.20
Initial reading / cm^3	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 HC*l* (aq) that remains after the reaction of **M** with HC*l* (aq).
2. The amount, in mol, of HC*l* (aq) that reacted with **M**.
3. The identity of metal **M**.

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

Metal **M** = ... **[6]**

5. A student carries out two experiments in the laboratory based on succinic acid (butanedioic acid), $(CH_2COOH)_2$.

Aqueous succinic acid can be neutralised by aqueous sodium hydroxide, NaOH(aq):

$$(CH_2COOH)_2(aq) + 2NaOH(aq) \rightarrow (CH_2COONa)_2(aq) + 2H_2O(l)$$

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

The student follows this method:

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

Temperature readings

Maximum temperature of mixture/°C	26.5
Initial temperature of both solutions/°C	21.5

Calculate a value for the enthalpy change of neutralisation, $\Delta_{neut}H$, 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.

$\Delta_{neut}H$ = ________________________________ kJ mol^{-1} **[4]**

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

Some standard electrode potentials are shown below.

$2H^+(aq) + O_2(g) + 2e^-$	$\rightleftharpoons$	$H_2O_2(aq)$	$E^\ominus = +0.68\ V$
$H_2O_2(aq) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$2H_2O(l)$	$E^\ominus = +1.77\ V$
$VO^{2+}(aq) + 2H^+(aq) + e^-$	$\rightleftharpoons$	$V^{3+}(aq) + H_2O(l)$	$E^\ominus = +0.34\ V$
$MnO_4^-(aq) + 8H^+(aq) + 5e^-$	$\rightleftharpoons$	$Mn^{2+}(aq) + 4H_2O(l)$	$E^\ominus = +1.51\ V$

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

-- **[2]**

7. Ethanedioic acid, $(COOH)_2$, is present in rhubarb leaves.

A student carries out a redox titration using aqueous cerium(IV) sulfate, $Ce(SO_4)_2(aq)$, to determine the percentage, by mass, of ethanedioic acid in rhubarb leaves.

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

$$2Ce^{4+}(aq) + (COOH)_2(aq) \rightarrow 2Ce^{3+}(aq) + 2CO_2(g) + 2H^+(aq)$$

$Ce^{4+}(aq)$ ions have a yellow colour. $Ce^{3+}(aq)$ 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} $Ce(SO_4)_2$ 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/cm³	24.30	47.80	23.65
Initial reading/cm³	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]**

8. 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.

Step 2: 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 KI(aq).

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

$IO_3^-(aq) + 6H^+(aq) + 5I^-(aq) \rightarrow 3I_2(aq) + 3H_2O(l)$

Step 3: The resulting mixture is titrated with 0.150 mol dm^{-3} $Na_2S_2O_3(aq)$.

$2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$

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 ±0.05 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]**

9(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.

$Zn^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Zn(s)$	$E^\ominus = -0.76\,V$
$Cr^{3+}(aq) + e^-$	$\rightleftharpoons$	$Cr^{2+}(aq)$	$E^\ominus = -0.42\,V$
$Ni^{2+}(aq) + 2e^-$	$\rightleftharpoons$	$Ni(s)$	$E^\ominus = -0.25\,V$
$I_2(aq) + 2e^-$	$\rightleftharpoons$	$2I^-(aq)$	$E^\ominus = +0.54\,V$
$Fe^{3+}(aq) + e^-$	$\rightleftharpoons$	$Fe^{2+}(aq)$	$E^\ominus = +0.77\,V$
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^-$	$\rightleftharpoons$	$2Cr^{3+}(aq) + 7H_2O(l)$	$E^\ominus = +1.33\,V$
$H_2O_2(aq) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$2H_2O(l)$	$E^\ominus = +1.78\,V$

Table 21.1

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

$$Ni(s) + I_2(aq) \rightarrow Ni^{2+}(aq) + 2I^-(aq)$$

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 $Cr_2O_7^{2-}$ 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%.

Identify **C−F** by formulae or structures, as appropriate.

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

10. 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

iii. Construct the equation for this reaction.

Equation

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

$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$

This reaction is catalysed by manganese dioxide, $MnO_2(s)$.

Standard electrode potentials are shown below.

$O_2(g) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$H_2O_2(aq)$	$E^\theta = +0.70$ V
$MnO_2(s) + 4H^+(aq) + 2e^-$	$\rightleftharpoons$	$Mn^{2+}(aq) + 2H_2O(l)$	$E^\theta = +1.51$ V
$H_2O_2(g) + 2H^+(aq) + 2e^-$	$\rightleftharpoons$	$2H_2O(l)$	$E^\theta = +1.78$ V

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]

12(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	$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
2	$SO_4^{2-}(aq) + 2H^+(aq) + 2e^- \rightleftharpoons SO_3^{2-}(aq) + H_2O(l)$	+0.17
3	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
4	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51

Table 22.1

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]

(b). 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]**

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

Strongest reducing agent

Strongest oxidising agent

[2]

13. 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]

14(a). When heated with dilute acid, $MnO_4^{2-}(aq)$ ions disproportionate into MnO_4^- and MnO_2.

i. Balance the equation for this disproportionation reaction.

...... $MnO_4^{2-}(aq)$ + ... $H^+(aq) \rightleftharpoons$ ...... $MnO_4^-(aq)$ + ... $MnO_2(s)$ + ... $H_2O(l)$

[1]

ii. Although $MnO_4^{2-}(aq)$ ions disproportionate in acidic conditions, $MnO_4^{2-}(aq)$ 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.

$2H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq) + H_2(g) \quad E^\ominus = -0.83\,V$

$2H^+(aq) + 2e^- \rightleftharpoons H_2(g) \quad E^\ominus = 0.00\,V$

$O_2(g) + 2H_2O(l) + 4e^- \rightleftharpoons 4OH^-(aq) \quad E^\ominus = +0.40\,V$

$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) \quad E^\ominus = +1.23\,V$

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]

15(a). Hydrated copper(II) methanoate, $Cu(HCOO)_2{\bullet}xH_2O$, is a copper salt.

A student carries out the procedure below to prepare $Cu(HCOO)_2{\bullet}xH_2O$ and to determine the value of *x* in its formula.

Step 1
The student prepares $Cu(HCOO)_2{\bullet}xH_2O$ by reacting a copper compound with aqueous methanoic acid to form $Cu(HCOO)_2$(aq) and allowing the solvent to evaporate.

Step 2
The student dissolves 2.226 g of $Cu(HCOO)_2{\bullet}xH_2O$ 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 KI(aq).

The Cu^{2+}(aq) ions react to form a precipitate of copper(I) iodide and I_2(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} $Na_2S_2O_3$ (aq).

$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$

23.5 cm^3 0.0420 mol dm^{-3} $Na_2S_2O_3$ (aq) is required to reach the end point.

Complete the electron configuration of copper in

$Cu(HCOO)_2{\bullet}xH_2O$: $1s^2$ ....

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 $Cu(HCOO)_2$(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 $Cu(HCOO)_2 \cdot xH_2O$.

Show your working.

[5]

16. 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, $[Au(CN)_2]^-$, with a bond angle of 180°. Hydroxide ions are also formed.

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

[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.
$CN^-(aq) + ClO^-(aq) + 2H_2O(l) \rightarrow NH_4^+(aq) + CO_3^{2-}(aq) + Cl^-(aq)$

The oxidation half-equation is:
$CN^- + 3H_2O \rightarrow NH_4^+ + CO_3^{2-} + 2H^+ + 2e^-$

Construct the reduction half-equation.

State symbols are **not** required.

[1]

17(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*⊖/V**
1	$Al^{3+}(aq) + 3e^- \rightleftharpoons Al(s)$	−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]**

(c). Explain what is meant by the terms *oxidising agent* and *reducing agent*.

Illustrate your answer by predicting reactions from redox systems **1**–**6** in which Cr^{3+}(aq) behaves as an oxidising agent and as a reducing agent.

Include overall equations and explain why you have made your predictions.

State symbols are **not** required.

[6]

18. 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]

19(a). Chromite is the main ore of chromium. The chromium-containing compound in chromite is $Fe(CrO_2)_2$. 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.

$$2Fe(CrO_2)_2 + 7Na_2O_2 \rightarrow 2NaFeO_2 + 4Na_2CrO_4 + 2Na_2O$$

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 mol dm^{-3} sodium thiosulfate, $Na_2S_2O_3(aq)$ to estimate the I_2 present:

$$I_2(aq) + 2S_2O_3^{2-}(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$$

The average titre of $Na_2S_2O_3(aq)$ 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.

[2]

(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:

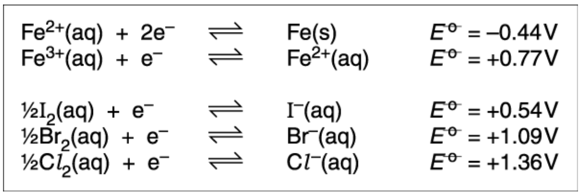
- CrO_4^{2-} ions, are reduced to chromium(III) ions, Cr^{3+}
- I^- ions are oxidised to iodine, I_2
- Construct an overall equation for the redox reaction and write half equations for the oxidation and reduction.

Overall equation:

Half equations:

[3]

20. Some electrode potentials for ions are shown below.



$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s)$ $E^\ominus = -0.44\,V$

$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$ $E^\ominus = +0.77\,V$

$\frac{1}{2}I_2(aq) + e^- \rightleftharpoons I^-(aq)$ $E^\ominus = +0.54\,V$

$\frac{1}{2}Br_2(aq) + e^- \rightleftharpoons Br^-(aq)$ $E^\ominus = +1.09\,V$

$\frac{1}{2}Cl_2(aq) + e^- \rightleftharpoons Cl^-(aq)$ $E^\ominus = +1.36\,V$

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 $I_2(aq)$, $Br_2(aq)$ and $Cl_2(aq)$.

Explain your predictions using the electrode potential data above.

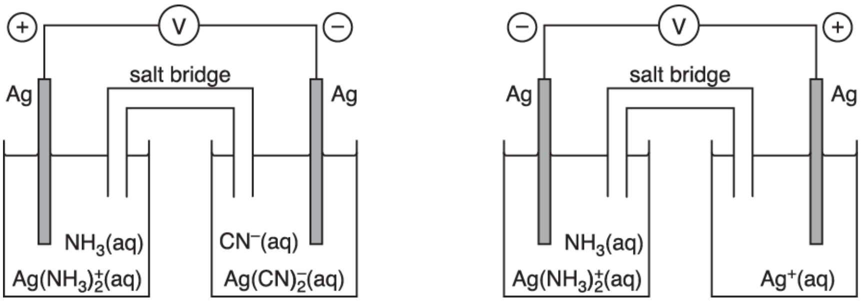
[3]

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

C	$Ag(NH_3)_2{}^+(aq) + e^- \rightleftharpoons Ag(s) + 2NH_3(aq)$
D	$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$
E	$Ag(CN)_2{}^-(aq) + e^- \rightleftharpoons Ag(s) + 2CN^-(aq)$

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.



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]

(b). 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:

$$C_2H_5OH + 3H_2O \rightarrow 2CO_2 + 12H^+ + 12e^-$$

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

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

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.

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

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

Oxidation:
....

Reduction:
....

22(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, FeO_4^{2-}(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, I_2(aq).

$$BaFeO_4(s) + 8H^+(aq) + 3I^-(aq) \rightarrow Fe^{3+}(aq) + Ba^{2+}(aq) + 1\tfrac{1}{2}I_2(aq) + 4H_2O(l)$$

Step 4

The student determines the amount of I_2 formed by carrying out a titration with aqueous sodium thiosulfate, $Na_2S_2O_3$(aq).

$$2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$$

26.4 cm^3 of 0.100 mol dm^{-3} $Na_2S_2O_3$(aq) are required to reach the end point.

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

[2]

(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 × 10^{-3} 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]

23(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**
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

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

Water is then added to the $Cu^{2+}(aq)$ / Cu(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.

$$Al(s) + 4OH^-(aq) \rightarrow Al(OH)_4^-(aq) + 3e^-$$

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

i. The standard electrode potential for the O_2 / OH^- 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]

24(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	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l)$	+1.51
2	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
3	$Br_2(aq) + 2e^- \rightleftharpoons 2Br^-(aq)$	+1.09
4	$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80
5	$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$	+0.77
6	$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	−0.76
7	$Ce^{3+}(aq) + 3e^- \rightleftharpoons Ce(s)$	−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 $Fe^{2+}(aq)$ to $Fe^{3+}(aq)$.

[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]

25(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 $CuSO_4 \cdot xH_2O$.

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 KI(aq).

$$2Cu^{2+}(aq) + 4I^-(aq) \rightarrow 2CuI(s) + I_2(aq)$$

Step 3 The resulting mixture is titrated with 0.120 mol dm^{-3} $Na_2S_2O_3$(aq).

$$2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$$

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 = ... cm^3 **[1]**

ii. The uncertainty in each burette reading is ±0.05 cm^3.

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.

$Fe^{2+}(aq) + 2e^- \rightleftharpoons Fe(s) \quad E^\ominus = -0.44\ V$

$2H^+ (aq) + 2e^- \rightleftharpoons H2(g) \quad E^\ominus = 0.00\ V$

$Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq) \quad E^\ominus = +0.77\ V$

$O_2(g) + 4H^+(aq) + 4e^- \rightleftharpoons 2H_2O(l) \quad E^\ominus = +1.23\ V$

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightleftharpoons 2Cr^{3+}(aq) + 7H_2O(l) \quad E^\ominus = +1.33\ V$

$Cl_2(g) + 2e^- \rightleftharpoons 2Cl^-(aq) \quad E^\ominus = +1.36\ V$

$H_2O_2(aq) + 2H^+(aq) + 2e^- \rightleftharpoons 2H_2O(l) \quad E^\ominus = +1.78V$

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]

26. 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.

Equation 1 $NH_2OH + Fe^{3+} \longrightarrow Fe^{2+} + ½N_2 + H^+ + H_2O$

Equation 2 $NH_2OH + 2Fe^{3+} \longrightarrow 2Fe^{2+} + ½N_2O + 2H^+ + ½H_2O$

Equation 3 $NH_2OH + 3Fe^{3+} \longrightarrow 3Fe^{2+} + NO + 3H^+$

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 4.32×10^{-2} 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 mol dm^{-3} Fe^{3+}(aq).
Stage 2 Boil the mixture for 5 minutes and allow to cool.
Stage 3 Titrate the cooled mixture with 2.00×10^{-2} mol dm^{-3} $KMnO_4$(aq).

a. Determine the minimum volume of 0.0400 mol dm^{-3} Fe^{3+}(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 $KMnO_4(aq)$ is required to reach the end point.

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

$$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O(l)$$

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]

27. 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.

…… ClO^- + …… H_2O + …… e^- → …… Cl^- + …… OH^-

…… Fe_2O_3 + …… OH^- → …… FeO_4^{2-} + …… H_2O + …… e^-

overall
equation:

[3]

28. 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]**

END OF QUESTION PAPER