1. **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°/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

(a) (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.

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.... **[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.

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.... **[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.

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.... **[2]**

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

$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ $\quad E^\ominus = +0.34\,V$

$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$ $\quad E^\ominus = +0.80\,V$

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 $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$.

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]**

3. 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

(a) 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]**

4. This question is about reactions of hydrogen peroxide, H_2O_2.

(a) Hydrogen peroxide, H_2O_2, iodide ions, I^-, and acid, H^+, react as shown in the equation below.

$$H_2O_2(aq) + 2I^-(aq) + 2H^+(aq) \rightarrow I_2(aq) + 2H_2O(l)$$

A student carries out several experiments at the same temperature, using the initial rates method, to determine the rate constant, *k*, for this reaction.

The results are shown below.

Experiment	**Initial concentrations**			**Rate**
	$[H_2O_2(aq)]$ / mol dm^{-3}	**$[I^-(aq)]$ / mol dm^{-3}**	**$[H^+(aq)]$ / mol dm^{-3}**	**/ 10^{-6} mol dm^{-3} s^{-1}**
1	0.0100	0.0100	0.100	2.00
2	0.0100	0.0200	0.100	4.00
3	0.0200	0.0100	0.100	4.00
4	0.0200	0.0100	0.200	4.00

(i) Determine the rate equation and calculate the rate constant, *k*, including units.

k = ... units .... **[3]**

(ii) The rate constant, *k*, for this reaction is determined at different temperatures, *T*.

Explain how the student could determine the activation energy, E_a, for the reaction graphically using values of *k* and *T*.

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.... **[3]**

(b) 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.

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

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

You answer should include relevant equations.

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.... **[4]**

(c) Peroxycarboxylic acids are organic compounds with the COOOH functional group.

Peroxyethanoic acid, CH_3COOOH, is used as a disinfectant.

(i) Suggest the structure for CH_3COOOH.

The COOOH functional group must be clearly displayed.

[1]

(ii) Peroxyethanoic acid can be prepared by reacting hydrogen peroxide with ethanoic acid. This is a heterogeneous equilibrium.

$$H_2O_2(aq) + CH_3COOH(aq) \rightleftharpoons CH_3COOOH(aq) + H_2O(l) \quad K_c = 0.37\,dm^3\,mol^{-1}$$

A $250\,cm^3$ equilibrium mixture contains concentrations of $0.500\,mol\,dm^{-3}$ $H_2O_2(aq)$ and $0.500\,mol\,dm^{-3}$ $CH_3COOH(aq)$.

Calculate the amount, in mol, of peroxyethanoic acid in the equilibrium mixture.

amount = ... mol **[3]**

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

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

$^1/_2O_2(g) + H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq)$ $E^\ominus = +0.40\,V$

What is the equation for the overall cell reaction?

A $H_2(g) + 4OH^-(aq) \rightarrow 3H_2O(l) + {}^1/_2O_2(g)$

B $3H_2O(l) + {}^1/_2O_2 \rightarrow H_2(g) + 4OH^-(aq)$

C $H_2O(l) \rightarrow H_2(g) + {}^1/_2O_2(g)$

D $H_2(g) + {}^1/_2O_2(g) \rightarrow H_2O(l)$

Your answer [ ] **[1]**

6. 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) Complete the electron configuration of

a Ni atom: $1s^2$ ....

a Ni^{2+} ion: $1s^2$ .... **[2]**

(b) 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]**

(c) 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.

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.... **[4]**

(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.

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]**

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Additional answer space if required.

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7. This question looks at reactions of hydrogen peroxide and of cobalt(II) ions.

(a) Aqueous hydrogen peroxide decomposes as shown in **equation 2.1**.

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g) \qquad \textbf{Equation 2.1}$$

The reaction is catalysed by manganese(IV) oxide, MnO_2.

A student investigates the decomposition of a hydrogen peroxide solution as outlined below.

- The student adds 50.00 cm^3 of $H_2O_2(aq)$ to a conical flask.
- The student adds a small spatula measure of MnO_2 and quickly connects the flask to a gas syringe.
- The student measures the volume of oxygen every 200 seconds.

Results

Time/s	Volume of O_2/cm^3
0	0
200	15
400	28
600	36
800	41
1000	46
1200	48
1400	50

(i) Process the results as outlined below.

- On page 5, plot a graph of **volume of O_2** against **time**.
- Use your graph to find the rate of the reaction, in $cm^3\,s^{-1}$, at t = 500 s.

Show your working on the graph and in the space below.

rate = ... $cm^3\,s^{-1}$ **[5]**

(ii) The student allows the reaction in **equation 2.1** to proceed until no more gas is evolved. The volume of O_2 in the syringe is now 55 cm^3, measured at RTP.

Calculate the initial concentration of the H_2O_2.

Give your answer to **two** significant figures.

initial concentration of H_2O_2 = ... $mol\,dm^{-3}$ **[3]**

(b) 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]**

(c) Cobalt(II) forms complex ions with water ligands and with chloride ligands.

- With water ligands, cobalt(II) forms a pink octahedral complex ion, $[Co(H_2O)_6]^{2+}$.
- With chloride ligands, cobalt(II) forms a blue tetrahedral complex ion.

A student dissolves cobalt(II) sulfate in water in a boiling tube. A pink solution forms.

Experiment 1
The student places the boiling tube in a water bath at 100 °C.
Concentrated hydrochloric acid is added dropwise.
The colour of the solution changes from pink to blue.

Experiment 2
The student places the boiling tube from **experiment 1** in an ice/water bath at 0 °C.
The colour of the solution changes from blue to pink.

(i) Write the equilibrium equation for the reaction that takes place when the colour of the solution changes.

.... **[1]**

(ii) Explain the observations and predict whether the formation of the blue colour is exothermic or endothermic.

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.... **[2]**

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

	$E^{\ominus}$/V
$H_2O(l) + e^- \rightleftharpoons OH^-(aq) + \frac{1}{2}H_2(g)$	−0.83
$H^+(aq) + e^- \rightleftharpoons \frac{1}{2}H_2(g)$	0.00
$\frac{1}{2}O_2(g) + H_2O(l) + 2e^- \rightleftharpoons 2OH^-(aq)$	+0.40
$\frac{1}{2}O_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons H_2O(l)$	+1.23

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

1 The reaction at the positive electrode is: $\frac{1}{2}O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O(l)$.

2 The overall cell reaction is: $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$.

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]**

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

Redox system	Half-equation	$E^\ominus$/V
1	$CO_2(g) + 2H^+(aq) + 2e^- \rightleftharpoons HCOOH(aq)$	−0.11
2	$HCOOH(aq) + 2H^+(aq) + 2e^- \rightleftharpoons HCHO(aq) + H_2O(l)$	−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) 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]

(b) 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.

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.... **[4]**

(c) 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.

$$HCOOH + \frac{1}{2}O_2 \rightarrow H_2O + CO_2$$

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]