

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	$E^\ominus/V$
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**

- (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  $\text{Fe}^{2+}(\text{aq})$  to  $\text{Fe}^{3+}(\text{aq})$ .

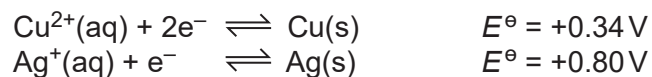
..... [1]

- (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.

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.....  
..... [2]

2. 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]

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	$E^\circ/V$
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]**





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

Peroxyethanoic acid, CH<sub>3</sub>COOOH, is used as a disinfectant.

(i) Suggest the structure for CH<sub>3</sub>COOOH.

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.



A 250 cm<sup>3</sup> equilibrium mixture contains concentrations of 0.500 mol dm<sup>-3</sup> H<sub>2</sub>O<sub>2</sub>(aq) and 0.500 mol dm<sup>-3</sup> CH<sub>3</sub>COOH(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.



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]**

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.

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

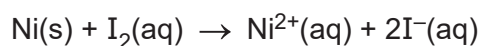
**Table 21.1**

(a) Complete the electron configuration of

a Ni atom:  $1\text{s}^2$  .....

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

(b) 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]





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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**.



The reaction is catalysed by manganese(IV) oxide,  $\text{MnO}_2$ .

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

- The student adds  $50.00 \text{ cm}^3$  of  $\text{H}_2\text{O}_2(\text{aq})$  to a conical flask.
- The student adds a small spatula measure of  $\text{MnO}_2$  and quickly connects the flask to a gas syringe.
- The student measures the volume of oxygen every 200 seconds.

### Results

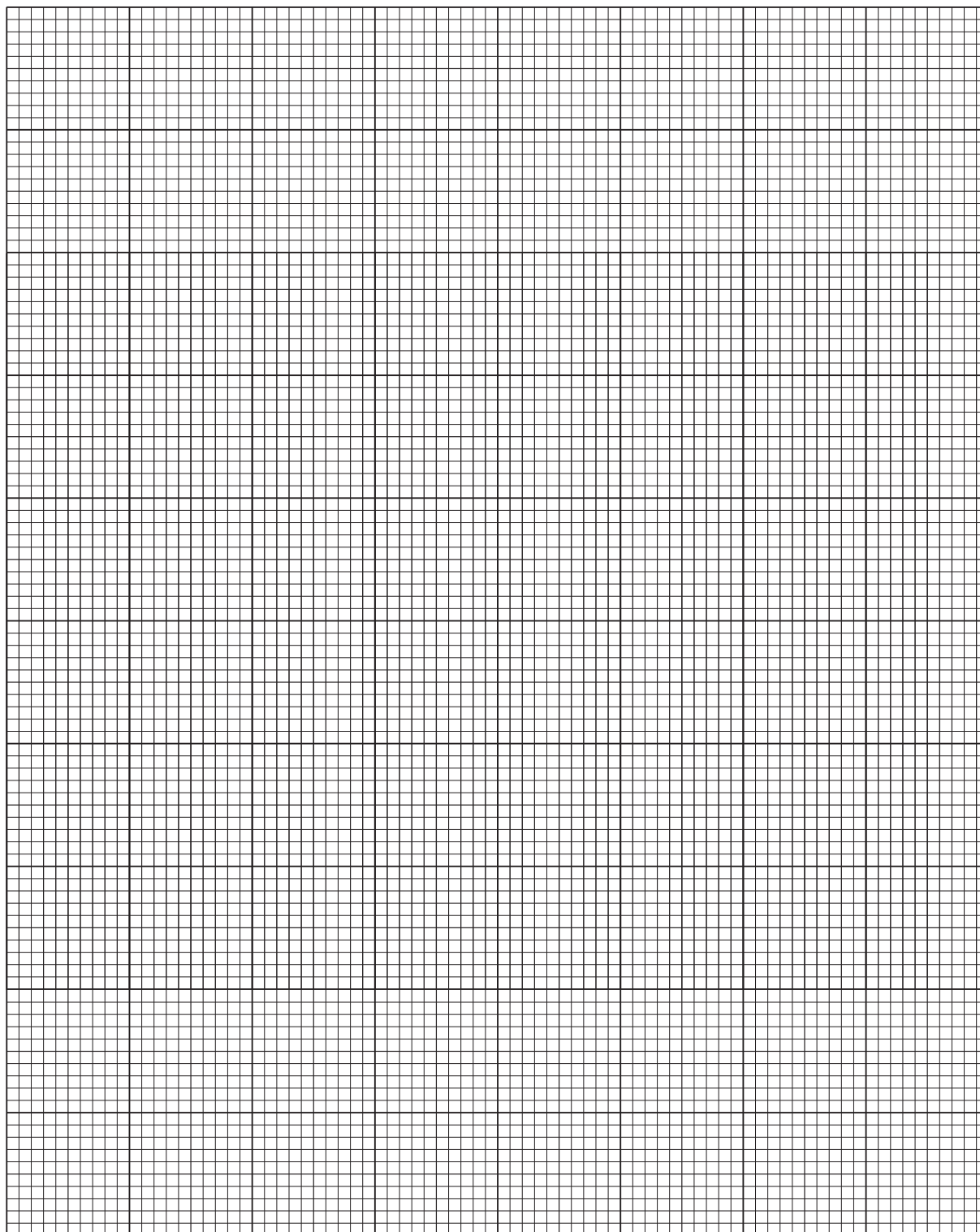
Time/s	Volume of $\text{O}_2/\text{cm}^3$
0	0
200	15
400	28
600	36
800	41
1000	46
1200	48
1400	50

(i) Process the results as outlined below.

- On page 5, plot a graph of **volume of  $\text{O}_2$**  against **time**.
- Use your graph to find the rate of the reaction, in  $\text{cm}^3 \text{ s}^{-1}$ , at  $t = 500 \text{ s}$ .

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

rate = .....  $\text{cm}^3 \text{ s}^{-1}$  [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  $55\text{ cm}^3$ , measured at RTP.

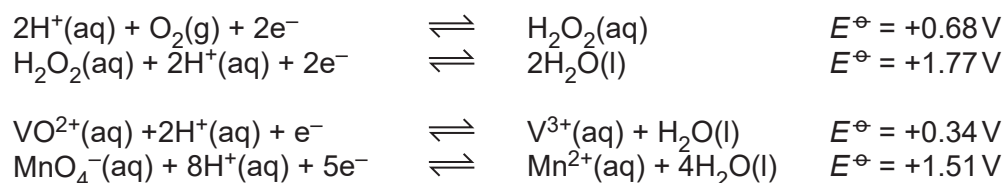
Calculate the initial concentration of the  $H_2O_2$ .

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

initial concentration of  $H_2O_2 = \dots\dots\dots\text{ mol dm}^{-3}$  [3]

- (b) Hydrogen peroxide can act as an oxidising agent or as a reducing agent.

Some standard electrode potentials are shown below.



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,  $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ .
- 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  $100^\circ\text{C}$ .

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  $0^\circ\text{C}$ .

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.

.....  
.....  
.....  
.....  
..... [2]

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

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

- 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	$\text{CO}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightleftharpoons \text{HCOOH}(\text{aq})$	-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) 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]

