

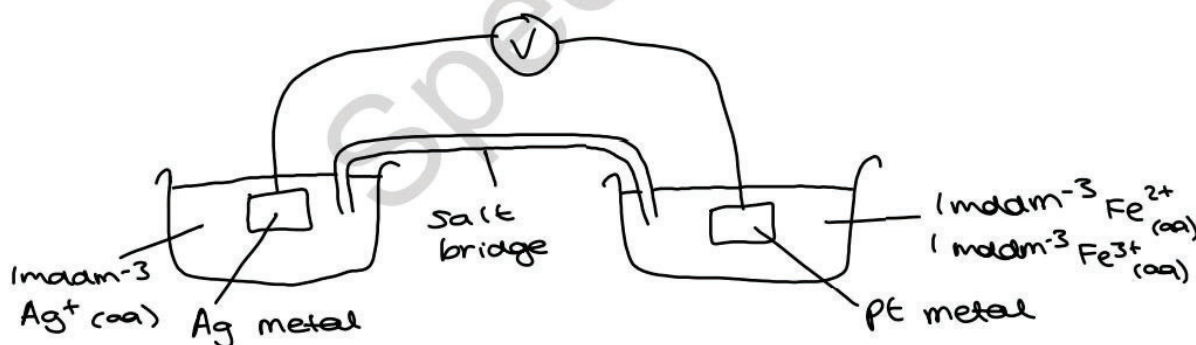
1. Table 21.1 below gives the standard electrode potentials for seven redox systems. You need to use this information to answer the questions below.

Redox system	Equation	$E^\circ/V$
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	$\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightleftharpoons \text{Fe}^{2+}(\text{aq})$ <i>no solid</i>	+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.



standard conditions: 298K, 100kPa

.....

.....

..... [4]

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

redox system 4  
shifts to the right

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

Ag/Ag<sup>+</sup> becomes more positive so, E<sub>cell</sub>.....  
becomes less positive.....

[2]

- (b) From Table 21.1, predict the oxidising agent(s) that will not oxidise Fe<sup>2+</sup>(aq) to Fe<sup>3+</sup>(aq).

Ce<sup>3+</sup>, Zn<sup>2+</sup>

E<sup>o</sup> will need to be  
less than +0.77V.....

[1]

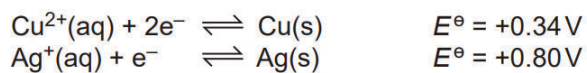
- (c) An aqueous solution of iron(II) bromide is mixed with an excess of acidified solution containing manganate(VII) ions.

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

Mn<sup>2+</sup>, H<sub>2</sub>O, Fe<sup>3+</sup>, Br<sub>2</sub>.....

[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 1.14V.  $+0.8 - 0.34 = +0.46\text{V}$
- 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

C



↑  
more solid produced

[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.

oxidising state

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

+7

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 ..... 1 mol dm<sup>-3</sup> solutions .....  
 ..... 298K, 25°C .....  
 .....

[4]

(ii) Predict the standard cell potential of this cell.

$$0.77 - (-0.76) = 1.53V$$

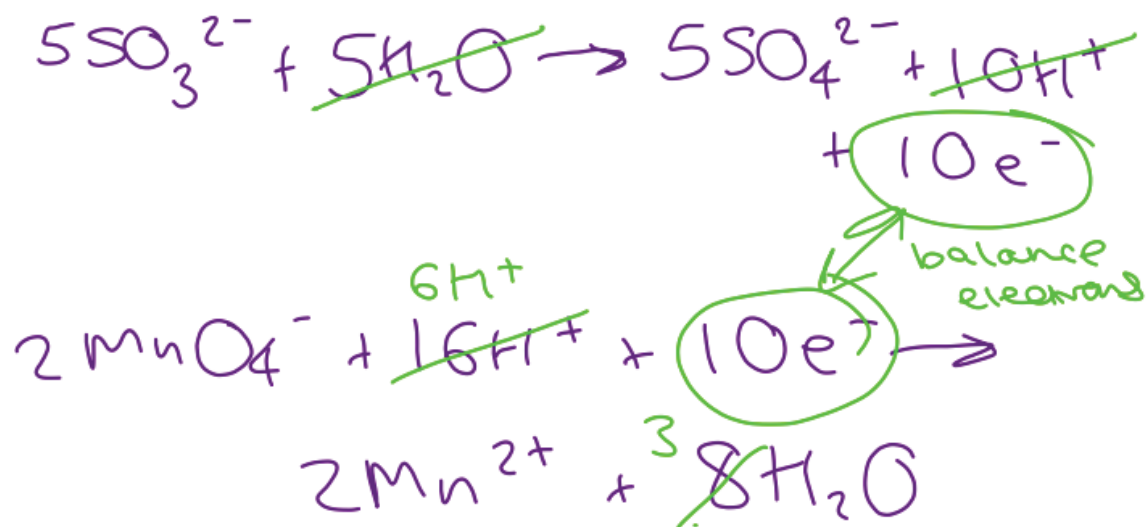
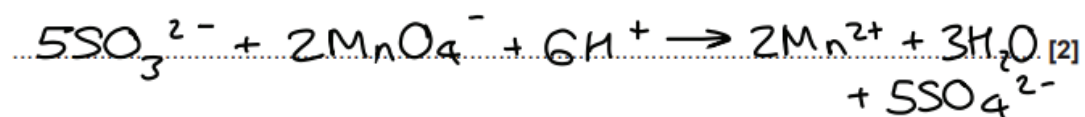
standard cell potential = ..... 1.53 ..... V [1]

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

Strongest reducing agent ..... Zn .....  
 Strongest oxidising agent .....  $\text{MnO}_4^-$  .....  
*oxidation*  
*reduction* [2]

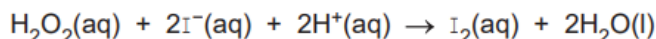
(c) Electrode potentials can be used to predict the feasibility of reactions.

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



4. This question is about reactions of hydrogen peroxide,  $\text{H}_2\text{O}_2$ .

(a) Hydrogen peroxide,  $\text{H}_2\text{O}_2$ , iodide ions,  $\text{I}^-$ , and acid,  $\text{H}^+$ , react as shown in the equation below.



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 / $10^{-6} \text{ mol dm}^{-3} \text{ s}^{-1}$
	$[\text{H}_2\text{O}_2(\text{aq})]$ / $\text{mol dm}^{-3}$	$[\text{I}^-(\text{aq})]$ / $\text{mol dm}^{-3}$	$[\text{H}^+(\text{aq})]$ / $\text{mol dm}^{-3}$	
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

*Handwritten notes on the table:*  
 - Above the table: "1<sup>st</sup> order" with an arrow pointing to the [I<sup>-</sup>] column, and "0<sup>th</sup> order" with an arrow pointing to the [H<sup>+</sup>] column.  
 - In the table: Each concentration cell has a circled header. Arrows with multipliers point from these cells to the rate column:  
 - [H<sub>2</sub>O<sub>2</sub>] → x1 (for exp 1, 2) and x2 (for exp 3, 4)  
 - [I<sup>-</sup>] → x2 (for exp 2) and ÷2 (for exp 3)  
 - [H<sup>+</sup>] → x1 (for exp 1, 2, 3) and x2 (for exp 4)  
 - Rate column: x2 (for exp 2), x1 (for exp 3), x1 (for exp 4)

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

$$\text{rate} = k [\text{H}_2\text{O}_2]^1 [\text{I}^-]^1 [\text{H}^+]^0$$

$$\text{rate} = k [\text{H}_2\text{O}_2] [\text{I}^-]$$

$$\frac{2 \times 10^{-6}}{0.01 \times 0.01} = 0.02$$

$$\frac{\text{mol dm}^{-3} \text{ s}^{-1}}{\text{mol dm}^{-3} \times \text{mol dm}^{-3}}$$

↑

$$k = 0.02 \text{ units mol}^{-1} \text{ dm}^3 \text{ s}^{-1}$$

(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  $T$ .

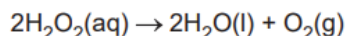
plot a graph of  $\ln k$  against  
 $1/T$  and measure the  
 gradient.

$$E_a = \text{gradient} \times R$$

$$\ln k = \frac{-E_a}{RT} + \ln A$$

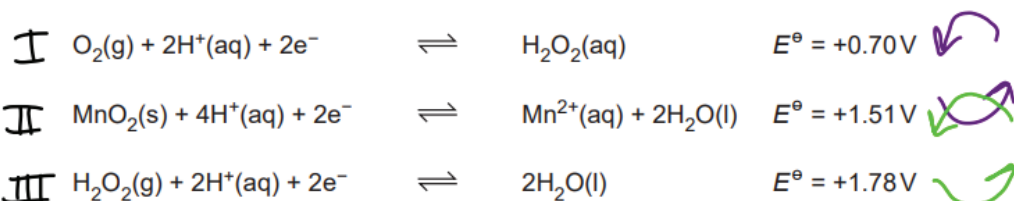
[3]

(b) Solutions of hydrogen peroxide decompose slowly into water and oxygen:



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

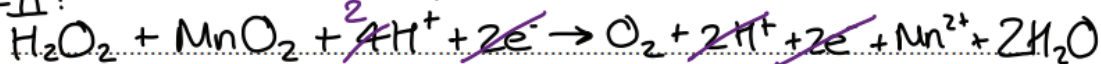
Standard electrode potentials are shown below.



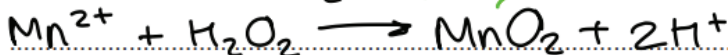
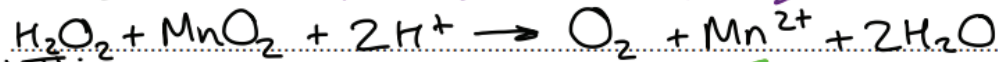
Using the electrode potentials, explain how  $\text{MnO}_2$  is able to act as a catalyst for the decomposition of hydrogen peroxide.

You answer should include relevant equations.

I+II:



II+III:



I more -ve E than II so I moves

left. II more -ve E than III moves left.

$\text{MnO}_2$  is regenerated so acts  
as a catalyst.

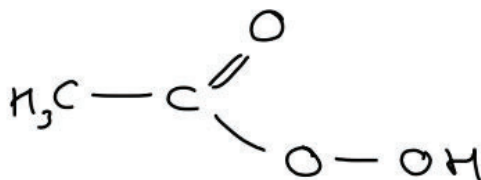
[4]

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

Peroxyethanoic acid,  $\text{CH}_3\text{COOOH}$ , is used as a disinfectant.

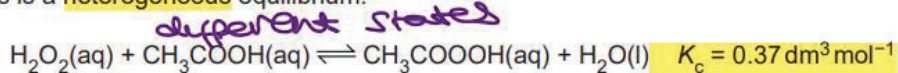
(i) Suggest the structure for  $\text{CH}_3\text{COOOH}$ .

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  $250 \text{ cm}^3$  equilibrium mixture contains concentrations of  $0.500 \text{ mol dm}^{-3} \text{ H}_2\text{O}_2(\text{aq})$  and  $0.500 \text{ mol dm}^{-3} \text{ CH}_3\text{COOH}(\text{aq})$ .

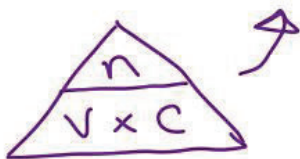
Calculate the amount, in mol, of peroxyethanoic acid in the equilibrium mixture.

*squares brackets means concentration*

$$0.37 = K_c = \frac{[\text{CH}_3\text{COOOH}]}{[\text{H}_2\text{O}_2][\text{CH}_3\text{COOH}]} = \frac{[\text{CH}_3\text{COOOH}]}{[0.5][0.5]}$$

$$0.37 \times [0.5][0.5] = [\text{CH}_3\text{COOOH}] = 0.0925 \text{ mol dm}^{-3}$$

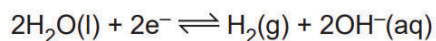
$$0.0925 \times 250 \times 10^{-3} = 0.023125 \text{ mol}$$



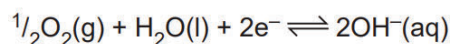
amount = ..... 0.023 ..... mol [3]



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

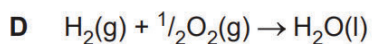
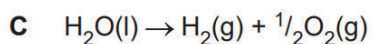
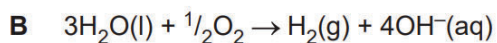
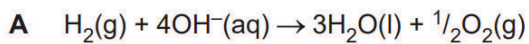


$$E^\ominus = -0.83\text{V}$$



$$E^\ominus = +0.40\text{V}$$

What is the equation for the overall cell reaction?



Your answer

**D**



[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.

<i>stage 2</i>	$Zn^{2+}(aq) + 2e^{-}$	$\rightleftharpoons$	$Zn(s)$	$E^{\ominus} = -0.76V$	<i>Cell with more negative <math>E^{\ominus}</math> undergoes oxidation</i>
	$Cr^{3+}(aq) + e^{-}$	$\rightleftharpoons$	$Cr^{2+}(aq)$	$E^{\ominus} = -0.42V$	
	$Ni^{2+}(aq) + 2e^{-}$	$\rightleftharpoons$	$Ni(s)$	$E^{\ominus} = -0.25V$	
<i>stage 1</i>	$I_2(aq) + 2e^{-}$	$\rightleftharpoons$	$2I^{-}(aq)$	$E^{\ominus} = +0.54V$	
	$Fe^{3+}(aq) + e^{-}$	$\rightleftharpoons$	$Fe^{2+}(aq)$	$E^{\ominus} = +0.77V$	
	$Cr_2O_7^{2-}(aq) + 14H^{+}(aq) + 6e^{-}$	$\rightleftharpoons$	$2Cr^{3+}(aq) + 7H_2O(l)$	$E^{\ominus} = +1.33V$	
	$H_2O_2(aq) + 2H^{+}(aq) + 2e^{-}$	$\rightleftharpoons$	$2H_2O(l)$	$E^{\ominus} = +1.78V$	

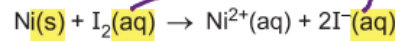
Table 21.1

(a) Complete the electron configuration of

a Ni atom:  $1s^2 \dots 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

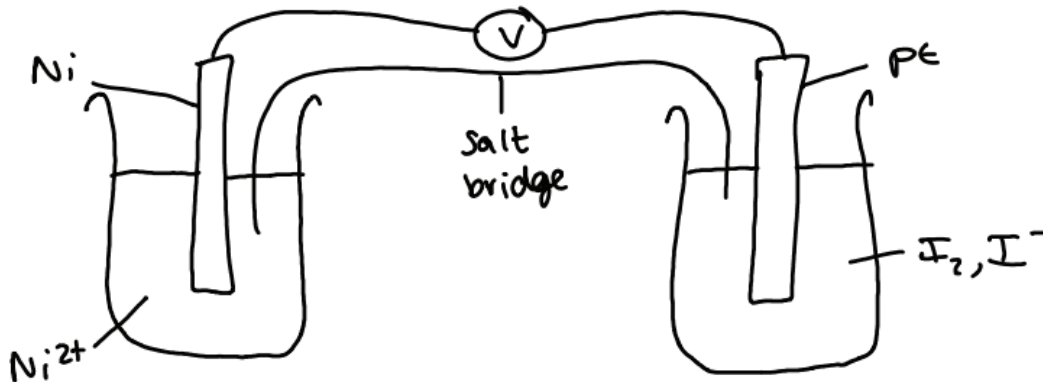
a Ni<sup>2+</sup> ion:  $1s^2 \dots 2s^2 2p^6 3s^2 3p^6 3d^8$  [2]

(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  $1 \text{ mol dm}^{-3}$ ,  $298 \text{ K}$

[4]

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

$$0.54 - (-0.25) = 0.79 \text{ V}$$

most positive - most negative

standard cell potential = ..... 0.79 ..... V [1]

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

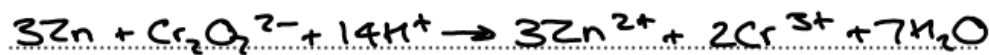
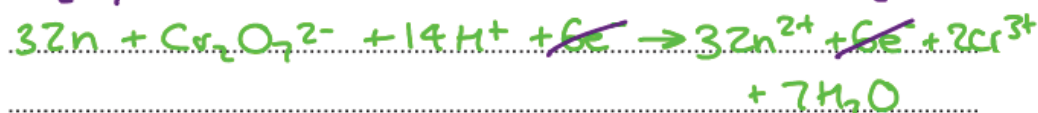
- (i) Write the overall equation for the oxidation of
- $\text{Fe}^{2+}$
- by acidified
- $\text{H}_2\text{O}_2$
- .



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

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

Include overall equations for the reactions which occur.



$E^\ominus$  of Zn is more negative than  $E^\ominus$  of  $\text{Cr}_2\text{O}_7^{2-}$  so Zn system shifts left. [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.

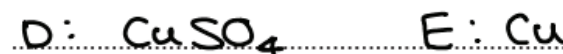
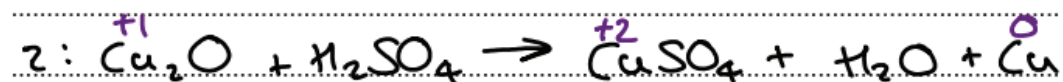
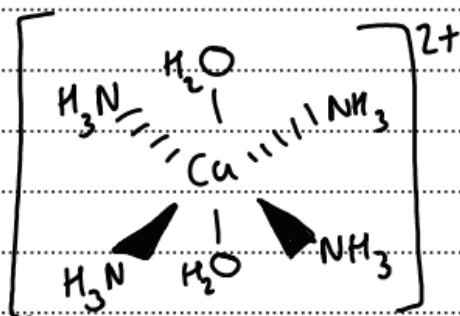
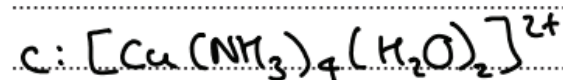
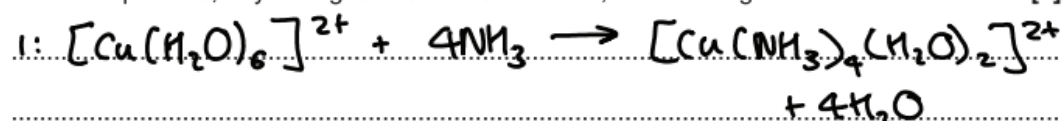
A hydrated salt, F, crystallises, with the percentage composition by mass:

Cu, 26.29%; H, 2.48%; N, 11.59%; O, 59.63%.

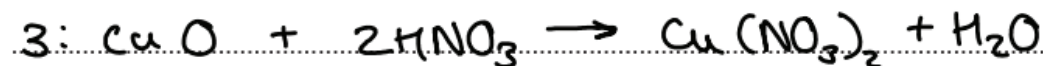
Identify C–F by formulae or structures, as appropriate.

Include equations, any changes in oxidation number, and working.

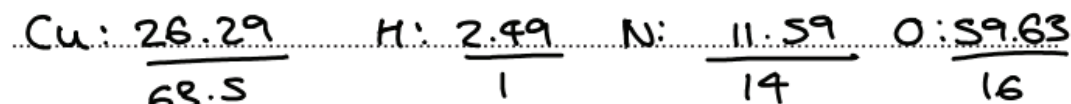
[6]



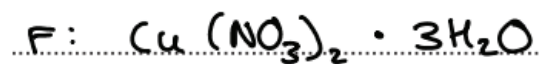
Cu oxidation states:  $+1 \rightarrow +2 + 0$



Additional answer space if required.

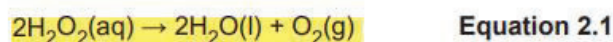


$$\begin{array}{cccc} = 0.41 & = 2.49 & = 0.82 & = 3.72 \\ \text{divide each of these numbers by the smallest (0.41)} & & & \\ \frac{0.41}{0.41} = 1 & \frac{2.49}{0.41} = 6 & \frac{0.82}{0.41} = 2 & \frac{3.72}{0.41} = 9 \end{array}$$



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.

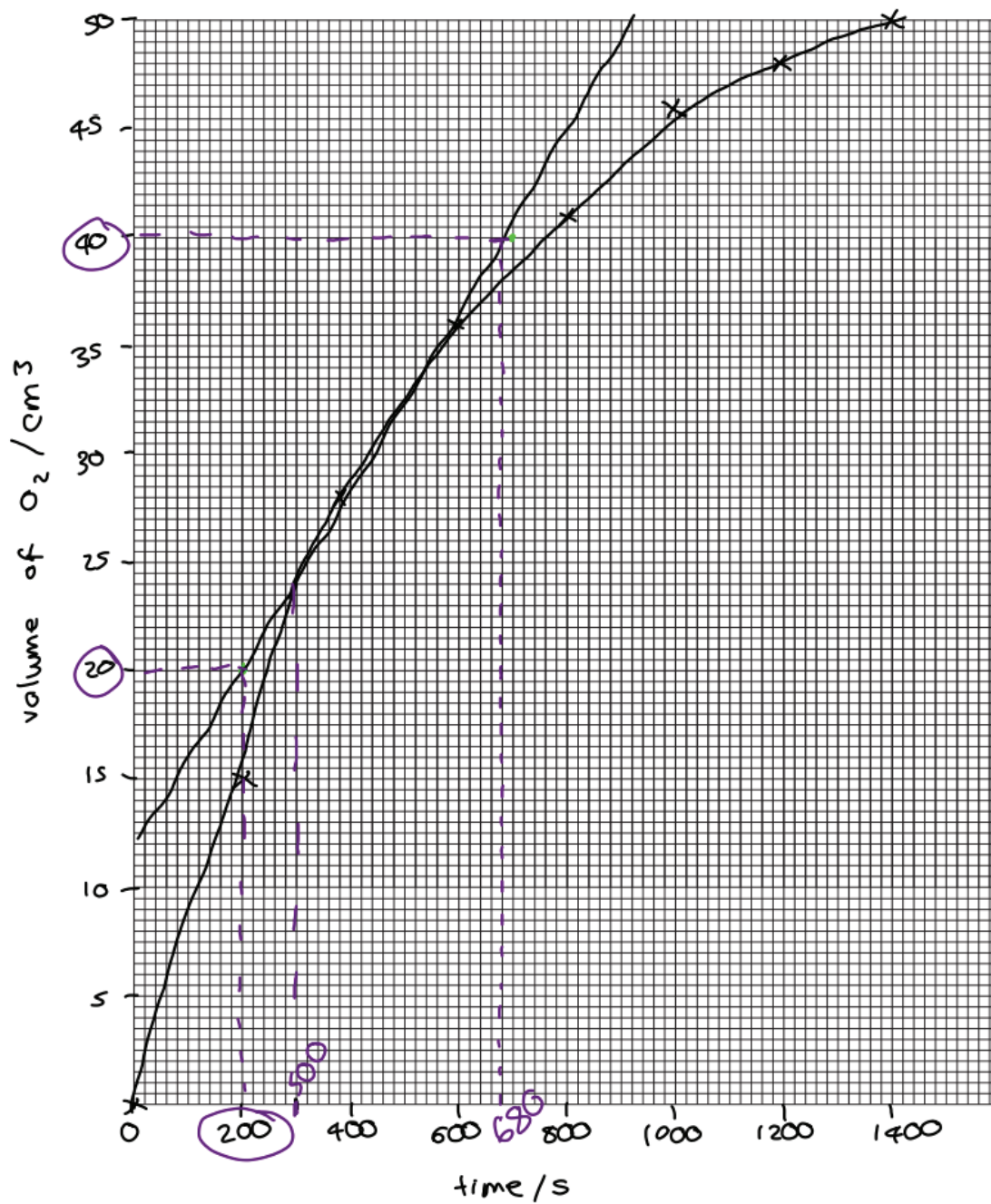
- On page 5, plot a graph of **volume of  $\text{O}_2$**  against **time**.
- Use your graph to find the **rate of the reaction**, in  $\text{cm}^3\text{s}^{-1}$ , at  $t = 500 \text{ s}$ .

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

$$\frac{40 - 20}{680 - 200} = 0.042 \text{ cm}^3\text{s}^{-1}$$

*tangent at 500 s*

rate = 0.042 .....  $\text{cm}^3\text{s}^{-1}$  [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**.

↑  
volume = mol  
molar volume  
( $24000\text{ cm}^3$ )

$$\frac{55}{24000} = 2.29 \times 10^{-3} \text{ mol}$$

$$2.29 \times 10^{-3} \times 2 = 4.58 \times 10^{-3} \text{ mol of } H_2O_2$$

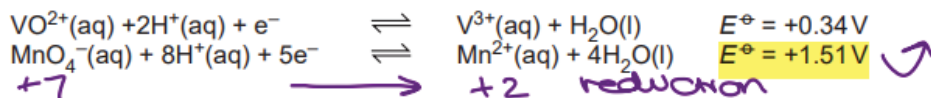
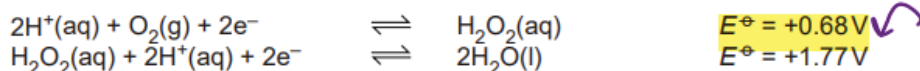
$$4.58 \times 10^{-3} \div 50 \times 10^{-3} = 0.092 \text{ mol dm}^{-3}$$



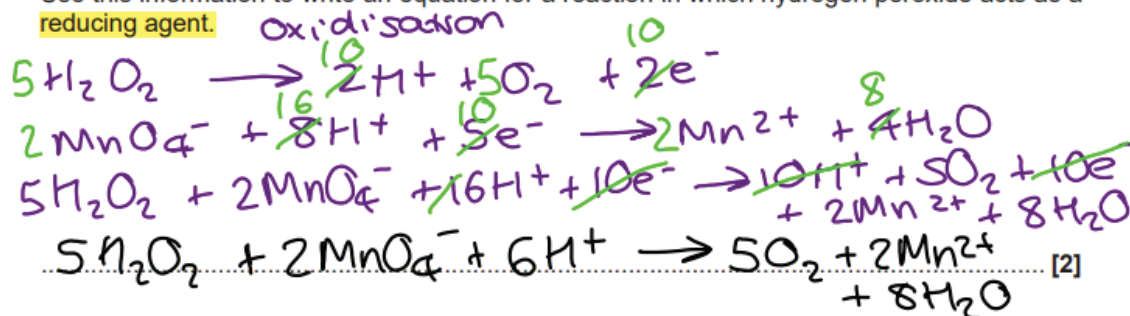
initial concentration of  $H_2O_2$  = .....  $0.092$  .....  $\text{mol dm}^{-3}$  [3]

- (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**.





(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.

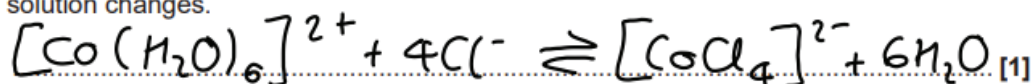
The colour of the solution changes from pink to blue.

#### 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.



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

*equilibrium shifts right at  $100^\circ\text{C}$   
and shifts left at  $0^\circ\text{C}$  so  
endothermic.*

[2]

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

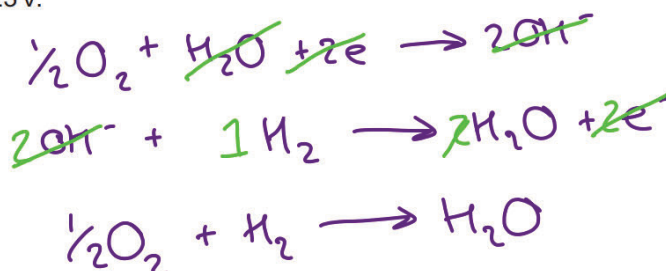
	$E^\circ / \text{V}$
$\text{H}_2\text{O}(\text{l}) + \text{e}^- \rightleftharpoons \text{OH}^-(\text{aq}) + \frac{1}{2}\text{H}_2(\text{g})$	-0.83
$\text{H}^+(\text{aq}) + \text{e}^- \rightleftharpoons \frac{1}{2}\text{H}_2(\text{g})$	0.00
$\frac{1}{2}\text{O}_2(\text{g}) + \text{H}_2\text{O}(\text{l})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

acid fuel cell

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

- 1 ~~✓~~ The reaction at the positive electrode is:  $\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(\text{aq}) + 2\text{e}^- \rightarrow \text{H}_2\text{O}(\text{l})$ . *positive electrode for one reaction but not other*
- 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

C

[1]

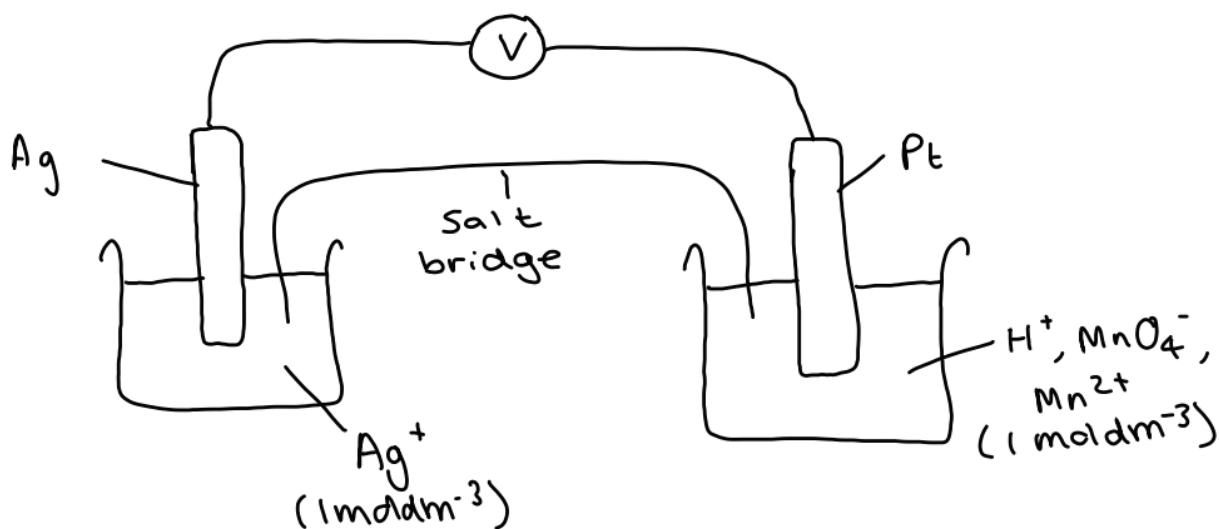
9. Standard electrode potentials for four redox systems are shown in **Table 19.1**.

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]

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

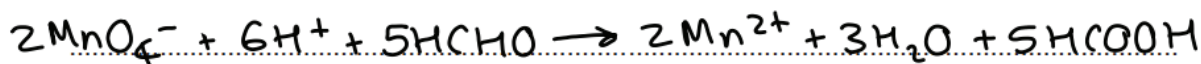
The student observes gas bubbles.

Explain this observation in terms of electrode potentials and equilibria.

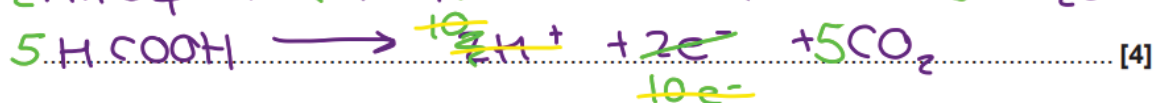
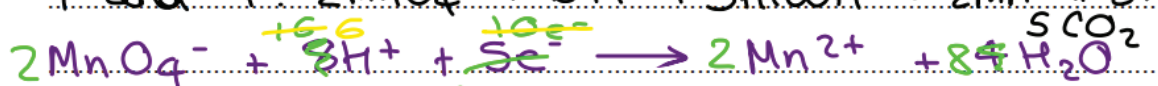
Include overall equations in your answer.

$E^\ominus$  of redox system 4 is more positive than  $E^\ominus$  of redox system 2 and 1. More negative systems 2 and 1 shift left

2 and 4:



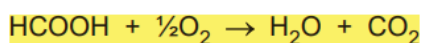
1 and 4:  $2\text{MnO}_4^- + 6\text{H}^+ + 5\text{HCOOH} \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} + 5\text{CO}_2$



- (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.

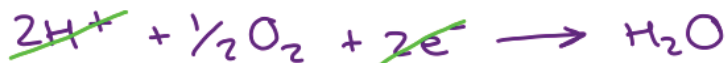
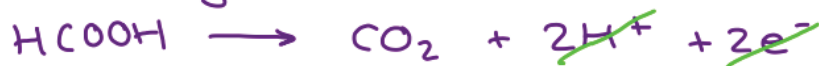
The standard cell potential for this fuel cell is 1.34 V.

The overall reaction is shown below.

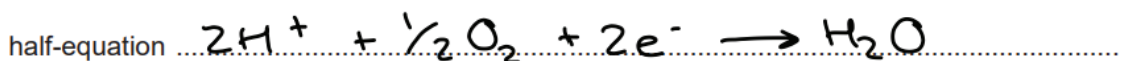


Using the information in Table 19.1, deduce the half-equation for the reaction at the oxygen electrode, and calculate the standard electrode potential for the oxygen half-cell.

redox system 1:



$$1.34 + (-0.11) = +1.23 \text{ V}$$



standard electrode potential = +1.23 V

[2]