24. Electrochemistry

24.2 Standard electrode potentials E

Paper 4

Question Paper

1 **(g)** A fuel cell is an electrochemical cell that can be used to generate electrical energy by using oxygen to oxidise a fuel.

Ethanedioic acid, $(COOH)_2$, dissolved in an alkaline electrolyte is being investigated as a fuel.

The relevant standard electrode potentials, $E^\ominus$, for the cell are shown.

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

$$2CO_2(g) + 2e^- \rightleftharpoons C_2O_4^{2-}(aq) \qquad E^\ominus = -0.59\,V$$

Use these equations to deduce the overall cell reaction. Calculate the value of $E^\ominus_{cell}$.

overall cell reaction ....

$E^\ominus_{cell}$ = ... V [2]

2 **(a)** Define standard electrode potential, $E^{\ominus}$, including a description of standard conditions.

....

....

.... [2]

(b) **(i)** An electrochemical cell is set up to measure $E^{\ominus}$ of the $Ag^+(aq)/Ag(s)$ electrode.

Draw a labelled diagram of this electrochemical cell.

Include all necessary substances. It is **not** necessary to state conditions used.

[3]

(ii) A separate electrochemical cell is set up using a **lower** concentration of $Ag^+(aq)$ than that used in **(b)(i)**.

Suggest how the electrode potential, E, for the $Ag^+(aq)/Ag(s)$ electrode would change from its $E^{\ominus}$ value. Explain your answer.

....

....

.... [1]

3 **(c)** Table 2.1 shows electrode potentials for some electrode reactions involving manganese compounds.

Table 2.1

electrode reaction	$E^{\ominus}$/V
$Mn^{2+} + 2e^- \rightleftharpoons Mn$	–1.18
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + e^- \rightleftharpoons MnO_4^{2-}$	+0.56
$MnO_4^- + 4H^+ + 3e^- \rightleftharpoons MnO_2 + 2H_2O$	+1.67
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.52
$MnO_4^- + 2H_2O + 3e^- \rightleftharpoons MnO_2 + 4OH^-$	+0.59
$MnO_4^{2-} + 2H_2O + 2e^- \rightleftharpoons MnO_2 + 4OH^-$	+0.60
$MnO_4^{2-} + 4H^+ + 2e^- \rightleftharpoons MnO_2 + 2H_2O$	+1.70

(i) Aqueous manganate(VI) ions, MnO_4^{2-}, are unstable in acidic conditions and undergo a disproportionation reaction.

The $E^{\ominus}_{cell}$ for this reaction is +1.14 V.

Construct an overall ionic equation for this disproportionation reaction.

.... [2]

(ii) Suggest and explain how the E_{cell} value of the disproportionation reaction changes with an increase in pH.

....

....

.... [1]

4 **(c)** SO_2 dissolves in water, forming H_2SO_3.

H_2SO_3 can be oxidised under acidic conditions.

The relevant electrode reaction and its $E^{\ominus}$ value are shown.

$$SO_4^{2-} + 4H^+ + 2e^- \rightleftharpoons H_2SO_3 + H_2O \qquad E^{\ominus} = +0.17\,V$$

Four more half-equations for reactions occurring under acidic conditions, and their $E^{\ominus}$ values, are shown.

$$H_3BO_3 + 3H^+ + 3e^- \rightleftharpoons B + 3H_2O \qquad E^{\ominus} = -0.73\,V$$
$$BiO^+ + 2H^+ + 3e^- \rightleftharpoons Bi + H_2O \qquad E^{\ominus} = +0.28\,V$$
$$S + 2H^+ + 2e^- \rightleftharpoons H_2S \qquad E^{\ominus} = +0.14\,V$$
$$Sb + 3H^+ + 3e^- \rightleftharpoons SbH_3 \qquad E^{\ominus} = -0.51\,V$$

Select the oxidising agent that could oxidise H_2SO_3 to SO_4^{2-} ions under acidic conditions.

Write an equation, and give the $E^{\ominus}_{cell}$ value, for the reaction that occurs.

oxidising agent ....

equation ....

$E^{\ominus}_{cell}$ = ...V

[3]

5 **(a)** Define standard cell potential, $E^{\ominus}_{cell}$.

....

.... [1]

(b) An electrochemical cell is set up to measure $E^{\ominus}_{cell}$ of a cell consisting of an Fe^{3+}/Fe^{2+} half-cell and a Cl_2/Cl^- half-cell.

Draw a labelled diagram of this electrochemical cell.

Include all necessary substances. It is **not** necessary to state conditions used.

[3]

(c) The cell reaction for the electrochemical cell in **(b)** is shown.

$$Cl_2 + 2Fe^{2+} \rightarrow 2Fe^{3+} + 2Cl^- \qquad E^{\ominus}_{cell} = +0.59\,V$$

Calculate $\Delta G^{\ominus}$, in $kJ\,mol^{-1}$, for this cell reaction.

$\Delta G^{\ominus}$ = ... $kJ\,mol^{-1}$ [2]

(d) Another experiment is set up using the same electrochemical cell.

In this experiment the Fe^{2+} concentration is 0.15 mol dm^{-3}. All other concentrations remain at their standard values.

The Nernst equation is shown.

$$E = E^{\ominus} + (0.059/z)\log\frac{[\text{oxidised species}]}{[\text{reduced species}]}$$

(i) Use the Nernst equation to calculate the electrode potential, E, for the Fe^{3+}/Fe^{2+} half-cell in this experiment.

[$E^{\ominus}$: Fe^{3+}/Fe^{2+} = +0.77 V]

E = ... V [1]

(ii) Use your answer to **(d)(i)** to calculate E_{cell} for this electrochemical cell.

E_{cell} = ... V [1]

6 Hydrogen peroxide is a liquid at 298 K. It is moderately stable under room conditions but will decompose quickly if a catalyst is added.

reaction 1 $2H_2O_2(l) \rightarrow 2H_2O\ (l) + O_2\ (g)$

(e) The $E^\ominus$ values for two electrode reactions are given.

$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$ $E^\ominus = +1.77\,V$

$Cr^{3+} + e^- \rightleftharpoons Cr^{2+}$ $E^\ominus = -0.41\,V$

(i) An electrochemical cell is constructed with the following half-cells (electrodes):

- an acidified solution of H_2O_2, a platinum wire
- Cr^{2+} mixed with Cr^{3+}, a platinum wire.

Identify the positive half-cell and calculate the standard cell potential, $E^\ominus_{cell}$.

positive half-cell .... $E^\ominus_{cell}$ = ... V

[1]

(ii) Calculate the value of $\Delta G^\ominus$ for the cell reaction that occurs, per mole of H_2O_2.

$\Delta G^\ominus$ = ... $kJ\,mol^{-1}$ [2]

(f) The $E^\ominus$ values for two electrode reactions are given.

$$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O \qquad E^\ominus = +1.77\,V$$

$$Co^{3+} + e^- \rightleftharpoons Co^{2+} \qquad E^\ominus = +1.82\,V$$

An electrochemical cell is constructed with the following half-cells.

half-cell 1 an acidified solution of H_2O_2 under standard conditions, a platinum wire

half-cell 2 a solution containing 0.020 mol dm^{-3} Co^{3+} and 2.0 mol dm^{-3} Co^{2+}, a platinum wire

(i) Use the Nernst equation to calculate the value of E, the electrode potential of half-cell 2 under these conditions.

E = ... V [2]

(ii) Write an equation for the cell reaction that occurs in this cell under these conditions.

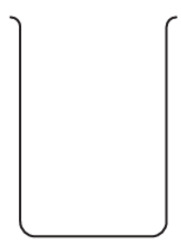
.... [1]

7 Some electrode potentials are shown in Table 3.1.

Table 3.1

electrode reaction	$E^{\ominus}$/V
$V^{2+} + 2e^- \rightleftharpoons V$	–1.20
$V^{3+} + e^- \rightleftharpoons V^{2+}$	–0.26
$VO^{2+} + 2H^+ + e^- \rightleftharpoons V^{3+} + H_2O$	+0.34
$VO_2^+ + 2H^+ + e^- \rightleftharpoons VO^{2+} + H_2O$	+1.00
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	–0.44
$Fe^{3+} + 3e^- \rightleftharpoons Fe$	–0.04
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
$ClO^- + H_2O + 2e^- \rightleftharpoons Cl^- + 2OH^-$	+0.89

(a) (i) Complete the diagram to show a standard hydrogen electrode.
Label your diagram. Identify all substances. You do **not** need to state standard conditions.



[1]

(ii) An electrochemical cell is set up using an Fe^{3+}/Fe^{2+} electrode and a standard hydrogen electrode.

Identify the positive electrode in the electrochemical cell and the direction of electron flow in the external circuit.

positive electrode ....

Electrons flow from the .... electrode to the .... electrode.
[1]

(b) The vanadium-containing species in the electrode reactions given in Table 3.1 are V, V^{2+}, V^{3+}, VO^{2+} and VO_2^{+}.

(i) Identify **one** vanadium-containing species that does **not** react with Fe^{2+} ions under standard conditions.

Use data from Table 3.1 to explain your answer.

....

.... [1]

(ii) Identify **all** the vanadium-containing species that will react with Fe^{2+} ions under standard conditions.

.... [1]

(iii) Write an equation for **one** of the possible reactions identified in **(ii)**.

.... [1]

(c) Another electrochemical cell is set up using an Fe^{3+}/Fe^{2+} electrode and an alkaline ClO^{-}/Cl^{-} electrode.

The concentration of Fe^{3+} is 1000 times greater than the concentration of Fe^{2+} in the Fe^{3+}/Fe^{2+} electrode. All other conditions are standard.

(i) Use the Nernst equation to calculate the *E* value of the Fe^{3+}/Fe^{2+} electrode.

Show your working.

E = ... V [2]

(ii) Write an equation for the reaction that occurs in the cell, under these conditions.

.... [1]

(d) Another electrochemical cell is set up using an Fe^{2+}/Fe electrode and an alkaline ClO^{-}/Cl^{-} electrode under standard conditions.

Calculate the value of $\Delta G^{\ominus}$ for the cell.

$\Delta G^{\ominus}$ = ... $kJ\,mol^{-1}$ [3]

8 **(c)** **(i)** Define standard electrode potential, $E^\ominus$.

....

....

.... [1]

(ii) A salt bridge is used in an electrochemical cell.

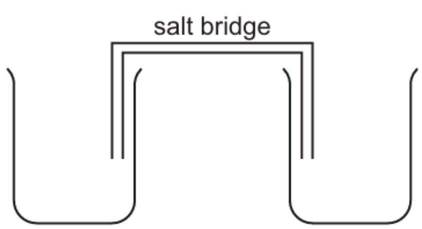
State the function of the salt bridge. Explain your answer.

....

.... [1]

(iii) Complete the diagram of the apparatus that can be used to measure the $E^\ominus$ of the $Cr_2O_7^{2-}(aq), H^+(aq)/Cr^{3+}(aq)$ electrode against the standard hydrogen electrode.

Your diagram should be fully labelled to identify all apparatus, substances and conditions.



[3]

(iv) The $E^\ominus$ of the $Cr_2O_7^{2-}(aq), H^+(aq)/Cr^{3+}(aq)$ electrode is +1.33 V.

Label the negative electrode and the direction of electron flow in the external circuit when the current flows in your diagram in **(c)(iii)**. [1]

(d) Table 3.1 lists relevant electrode potentials for some electrode reactions for use in **(d)(i)** and **(d)(ii)**.

Table 3.1

electrode reaction	$E^\ominus$/V
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	+1.33
$CH_3CHO + 2H^+ + 2e^- \rightleftharpoons CH_3CH_2OH$	–0.61
$CH_3COOH + 2H^+ + 2e^- \rightleftharpoons CH_3CHO + H_2O$	–0.94
$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23

(i) Ethanal is oxidised to ethanoic acid in the presence of $Cr_2O_7^{2-}$ ions.

Construct the ionic equation for the oxidation of ethanal to ethanoic acid using dichromate(VI) in acid conditions. Calculate the $E^\ominus_{cell}$ for this reaction.

ionic equation ....

$E^\ominus_{cell}$ = ... V
[2]

(ii) In an ethanol-oxygen fuel cell, $CH_3CH_2OH(l)$ and $O_2(g)$ are in contact with two inert electrodes immersed in an acidic solution.

The cell reaction for the oxidation of ethanol by oxygen is shown.

$$2CH_3CH_2OH + O_2 \rightarrow 2CH_3COOH + 2H_2O \qquad E^\ominus_{cell} = +2.01\,V$$

Calculate $\Delta G^\ominus$, in $kJ\,mol^{-1}$, for the oxidation of ethanol by oxygen.

$\Delta G^\ominus$ = ... $kJ\,mol^{-1}$ [2]

9 Data should be selected from Table 3.1 in order to answer some parts of this question.

Table 3.1

electrode reaction	$E^\ominus$ / V
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$2HOCl + 2H^+ + 2e^- \rightleftharpoons Cl_2 + 2H_2O$	+1.64
$ClO^- + H_2O + 2e^- \rightleftharpoons Cl^- + 2OH^-$	+0.89
$Sn^{4+} + 2e^- \rightleftharpoons Sn^{2+}$	+0.15
$Sn^{2+} + 2e^- \rightleftharpoons Sn$	–0.14
$V^{2+} + 2e^- \rightleftharpoons V$	–1.20
$V^{3+} + e^- \rightleftharpoons V^{2+}$	–0.26
$VO^{2+} + 2H^+ + e^- \rightleftharpoons V^{3+} + H_2O$	+0.34
$VO_2^+ + 2H^+ + e^- \rightleftharpoons VO^{2+} + H_2O$	+1.00

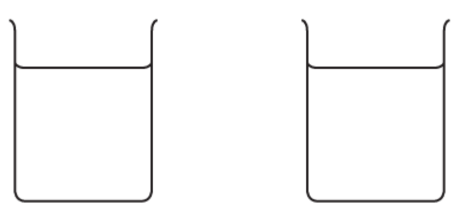
(a) Standard electrode potentials are measured under standard conditions.

(i) Describe the standard conditions used in the Sn^{4+}/Sn^{2+} half-cell.

....

.... [1]

(ii) Complete the diagram below to show how $E^\ominus(Sn^{4+}/Sn^{2+})$ can be measured experimentally. Your diagram should be fully labelled to identify all apparatus and substances.



[3]

(iii) Equal volumes of 1.0 mol dm^{-3} Sn^{2+}(aq) and 1.0 mol dm^{-3} Cl^{-}(aq) are mixed.

Use relevant E° values to explain whether a reaction occurs between these two ions.

....

....

....

.... [2]

(iv) Equal volumes of 1.0 mol dm^{-3} of Sn^{2+}(aq) and acidified 1.0 mol dm^{-3} VO^{2+}(aq) are mixed.

Write an equation for the reaction that takes place in the resulting mixture.

.... [2]

10 Data should be selected from Table 3.1 in order to answer some parts of this question.

Table 3.1

electrode reaction	$E^{\ominus}$/V
$Mg^{2+} + 2e^- \rightleftharpoons Mg$	–2.38
$Mn^{2+} + 2e^- \rightleftharpoons Mn$	–1.18
$Mn^{3+} + e^- \rightleftharpoons Mn^{2+}$	+1.49
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + e^- \rightleftharpoons MnO_4^{2-}$	+0.56
$MnO_4^- + 4H^+ + 3e^- \rightleftharpoons MnO_2 + 2H_2O$	+1.67
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.52

(a) An electrochemical cell can be constructed from a Mg^{2+}/Mg half-cell and a MnO_4^-/Mn^{2+} half-cell. The standard cell potential of this cell can be calculated using the standard electrode potentials of the two half-cells.

(i) Define standard electrode potential. Include details of the standard conditions used.

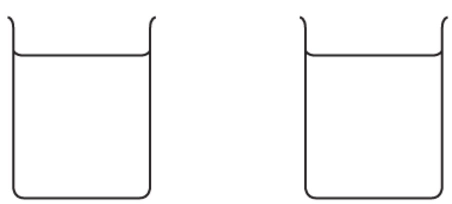
....

....

....

.... [2]

(ii) Complete the diagram below to show an electrochemical cell constructed from a Mg^{2+}/Mg half-cell and a MnO_4^-/Mn^{2+} half-cell.
Label your diagram.



[3]

(iii) Use a positive (+) sign and a negative (–) sign to identify the polarity of each of the two electrodes in your diagram.
Use an arrow and the symbol 'e' to show the direction of electron flow in the external circuit. [1]

(iv) Calculate the standard cell potential, $E^{\ominus}_{cell}$, of this cell.

$E^{\ominus}_{cell}$ = ... V [1]

(v) Construct an equation for the cell reaction.

.... [1]

(vi) Predict how the cell reaction will change, if at all, when the solution in the Mg^{2+} / Mg half-cell is diluted by the addition of a large volume of water. Explain your answer.

....

....

.... [1]

11 **(e)** The $[Cr_2(O_2CCH_3)_4(H_2O)_2]$ complex reacts with aqueous acid to form Cr^{2+}(aq) ions.

Cr^{2+}(aq) ions react with O_2(aq) under acidic conditions. Cr^{3+}(aq) ions are formed.

Use the *Data Booklet* to answer the following questions.

(i) Construct an ionic equation for the reaction of Cr^{2+}(aq) with O_2(aq) under acidic conditions.

.... [2]

(ii) Calculate $E^\ominus_{cell}$ for the reaction in **(e)(i)**.

$E^\ominus_{cell}$ = ... V [1]

12 **(a)** **(i)** Define the term *standard electrode potential*.

....

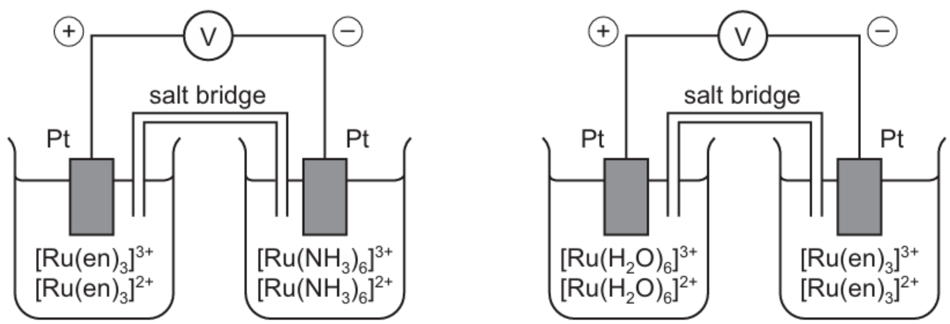
....

.... [2]

Three redox systems, **A**, **B** and **C**, are shown. The ligand 1,2-diaminoethane, $H_2NCH_2CH_2NH_2$, is represented by en.

A	$[Ru(H_2O)_6]^{3+} + e^- \rightleftharpoons [Ru(H_2O)_6]^{2+}$
B	$[Ru(NH_3)_6]^{3+} + e^- \rightleftharpoons [Ru(NH_3)_6]^{2+}$
C	$[Ru(en)_3]^{3+} + e^- \rightleftharpoons [Ru(en)_3]^{2+}$

Two electrochemical cells are set up to compare the standard electrode potentials, $E^\ominus$, of three half-cells. The diagrams show the relative potential of each electrode.



(ii) Use this information to complete the table by adding the labels **A**, **B** and **C** to deduce the order of $E^\ominus$ for the three half-cells.

$E^\ominus$	redox system
most negative	
↑	
least negative	

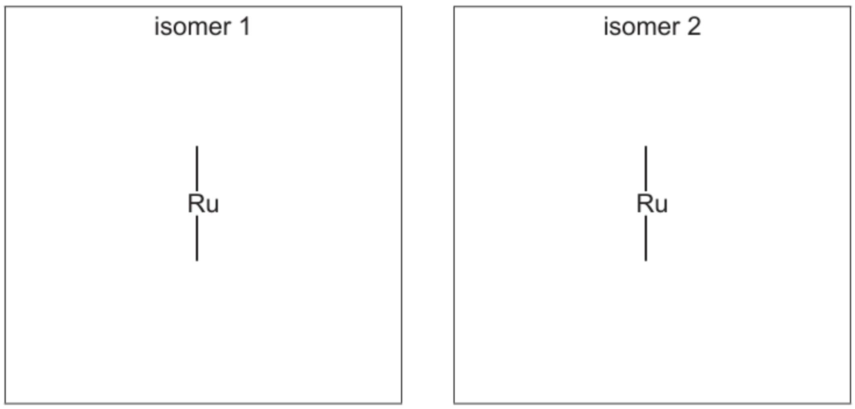
[1]

(iii) The complex $[Ru(en)_3]^{3+}$ shows stereoisomerism. The ligand en is bidentate.

Draw three-dimensional diagrams to show the two isomers of $[Ru(en)_3]^{3+}$. Represent the ligand en by using 

.

Name the type of stereoisomerism.



type of stereoisomerism ....

[3]

(b) (i) An electrochemical cell consists of a Br_2/Br^- half-cell and a Ag^+/Ag half-cell, under standard conditions.

Use the *Data Booklet* to calculate the $E^{\ominus}_{cell}$. Deduce the direction of electron flow in the wire through the voltmeter between these two half-cells.

$E^{\ominus}_{cell}$ = ... V

direction of electron flow from .... to .... [1]

(ii) Water is added to the Ag^+/Ag half-cell in **(b)(i)**.

Suggest the effect of this addition on the E_{cell}. Place a tick (✓) in the appropriate box.

less positive	no change	more positive

Explain your answer.

....

....

.... [2]

13 An excess of sodium iodide is added to a solution of copper(II) sulfate. Iodine and a white precipitate of copper(I) iodide are formed.

(a) Write an equation for the reaction that occurs.

.... [1]

(c) Use suitable $E^\ominus$ values from the *Data Booklet* to predict whether iodide ions can reduce Cu^{2+} to Cu^+ under standard conditions. Explain your answer.

....

....

.... [2]

(d) An excess of sodium iodide is added to copper(II) sulfate solution. Copper(I) iodide forms as a precipitate. After precipitation, $[Cu^+]$ is much lower than $1.0\,mol\,dm^{-3}$.

Use this information and your answer to **(c)** to explain how the relevant electrode potentials change and hence why I^- ions can reduce Cu^{2+} ions.

....

....

....

....

.... [2]

14 An excess of aqueous ammonia is added to a solution containing the complex ion $[Co(H_2O)_6]^{2+}$.

$$[Co(H_2O)_6]^{2+} + 6NH_3 \rightleftharpoons [Co(NH_3)_6]^{2+} + 6H_2O$$

(d) Oxygen can oxidise $[Co(NH_3)_6]^{2+}$ to $[Co(NH_3)_6]^{3+}$ under standard conditions in alkaline solutions.

$$[Co(NH_3)_6]^{3+} + e^- \rightleftharpoons [Co(NH_3)_6]^{2+} \qquad E^\ominus = +0.10\,V$$

(i) Use this information and the *Data Booklet* to calculate the $E^\ominus_{cell}$ value for this oxidation of $[Co(NH_3)_6]^{2+}$.

....

....

$E^\ominus_{cell}$ = ... V [1]

(ii) Write an ionic equation for this oxidation of $[Co(NH_3)_6]^{2+}$.

.... [1]

(iii) Predict, by selecting suitable data from the *Data Booklet*, whether oxygen can oxidise $Co^{2+}(aq)$ in **acidic** solution, in the absence of ammonia. Explain your answer.

....

.... [2]

15 The overall reaction for photosynthesis is shown.

$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$

Water is oxidised in this process according to the following half-equation.

$$2H_2O \rightarrow O_2 + 4H^+ + 4e^- \qquad \text{half-equation 1}$$

(a) **(i)** Use these equations to deduce the half-equation for the reduction of carbon dioxide in this process.

[2]

(ii) Draw a fully labelled diagram of the apparatus that should be used to measure the standard electrode potential, $E^\ominus$, of O_2(g) in half-equation 1 under standard conditions. Include all necessary chemicals.

[4]

(iii) For the cell drawn in **(a)(ii)**, use the *Data Booklet* to calculate the $E^\ominus_{cell}$ and deduce which electrode is positive.

$E^\ominus_{cell}$ = ... V

identity of the positive electrode = ...

[1]

16 **(a)** Manganese(IV) oxide, MnO_2, catalyses the decomposition of hydrogen peroxide, H_2O_2, as shown.

$$2H_2O_2(aq) \xrightarrow{MnO_2} 2H_2O(l) + O_2(g)$$

The mechanism involves the formation of the intermediate species, Mn^{2+}, in the first step which is subsequently used up in the second step.

State and use relevant electrode potentials, $E^\ominus$, to construct **two** equations to show how MnO_2 can catalyse this reaction.

....

....

....

....

....

equation 1 ....

equation 2 ....

[3]

17 **(f)** The half-equation for the reduction of iodate(V) ions is shown.

$$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O \qquad E^\ominus = +1.19\,V$$

Use data from the *Data Booklet* to predict whether a reaction is feasible when aqueous solutions of acidified iodate(V) ions and bromide ions are mixed. Explain your answer.

....

.... [1]

18 **(a)** **(i)** Define the term *standard cell potential*.

....

....

.... [2]

An electrochemical cell is set up to measure the standard electrode potential of a cell, $E^{\ominus}_{cell}$, made of a Co^{3+}/Co^{2+} half-cell and a Cl_2/Cl^- half-cell.

(ii) Complete the table with the substance used to make the electrode in each of these half-cells.

half-cell	electrode
Co^{3+}/Co^{2+}	
Cl_2/Cl^-	

[1]

(iii) Use data from the *Data Booklet* to calculate the $E^{\ominus}_{cell}$.

$E^{\ominus}_{cell}$ = ... V [1]

(iv) Write the equation for the overall cell reaction.

.... [1]

(b) A fuel cell is an electrochemical cell that can be used to generate electrical energy.
A methanol-oxygen fuel cell can be used as an alternative to a hydrogen-oxygen fuel cell. When the cell operates, the carbon atoms in the methanol molecules are converted into carbon dioxide.

$$CH_3OH + H_2O \rightarrow CO_2 + 6H^+ + 6e^-$$

Calculate the volume of CO_2, in cm^3, formed when a current of 2.5A is delivered by the cell for 30 minutes. Assume the cell is operated at room conditions.

volume of CO_2 = ... cm^3 [2]

19 **(c)** The equation representing the standard electrode potential, $E^\ominus$, for the reduction of $MnO_4^-(aq)$ to $Mn^{2+}(aq)$ in acid solution is given.

$$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightleftharpoons Mn^{2+}(aq) + 4H_2O(l) \qquad E^\ominus = +1.52\,V$$

(i) Draw a diagram of the apparatus that would be used to measure the $E^\ominus$ value of this half-cell. Your diagram should be fully labelled to identify all apparatus, substances and conditions.

[4]

(ii) Use the *Data Booklet* to identify a substance that could be used to oxidise Mn^{2+} ions to MnO_4^- ions under standard conditions.

Write an equation for the reaction.

....

....

.... [2]

20 **(c)** The halogens chlorine, bromine and iodine differ in their strengths as oxidising agents. These strengths are indicated by the $E^\ominus$ values for these halogens.

(i) Give the $E^\ominus$ values for chlorine, bromine and iodine acting as oxidising agents.

.... [1]

(ii) Deduce which of chlorine, bromine and iodine will react with a solution of $Sn^{2+}(aq)$ under standard conditions.

Explain your answer. Include a relevant equation in your explanation.

....

....

.... [3]

(iii) An excess of chlorine is added to a solution of acidified $Mn^{2+}(aq)$ under standard conditions.

Give the formula of the product of this reaction that contains manganese.

.... [1]

(d) An electrochemical cell can be made by connecting an Fe^{3+}/Fe^{2+} half-cell to an $S_2O_8^{2-}/SO_4^{2-}$ half-cell under standard conditions.

(i) Calculate the standard cell potential of this electrochemical cell.

$E^\ominus_{cell}$ = ... V [1]

(ii) State the material that should be used as the electrode in each half-cell.

in the Fe^{3+}/Fe^{2+} half-cell ....

in the $S_2O_8^{2-}/SO_4^{2-}$ half-cell ....
[1]

(iii) Describe **one** change to each half-cell that would **increase** the value of the cell potential. The temperature should remain at 298 K.

Fe^{3+}/Fe^{2+} half-cell ....

....

$S_2O_8^{2-}/SO_4^{2-}$ half-cell ....

....
[1]

21 Chlorate(V) ions are powerful oxidising agents.

(c) A lithium-iodine electrochemical cell can be used to generate electricity for a heart pacemaker. The cell consists of a lithium electrode and an inert electrode immersed in body fluids. When current flows lithium is oxidised and iodine is reduced.

(i) Use the *Data Booklet* to write half-equations for the reactions taking place at the two electrodes. Hence write the overall equation for when a current flows.

- ....
- ....

overall equation .... [2]

(ii) Use the *Data Booklet* to calculate the $E^{\ominus}_{cell}$ for this cell.

$E^{\ominus}_{cell}$ = ... V [1]

(iii) A current of 2.5×10^{-5} A is drawn from this cell.

Calculate the time taken for 0.10 g of lithium electrode to be used up. Assume the current remains constant throughout this period.

time = ... s [3]

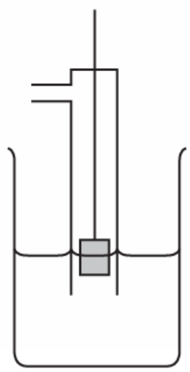
22 **(d)** Define the term *standard electrode potential, $E^\ominus$*.

....

....

.... [1]

(e) **(i)** Complete and **label** the diagram to show how the standard electrode potential, $E^\ominus$, of $Ag^+(aq)/Ag(s)$ could be measured under **standard conditions**.



[4]

(ii) Use the *Data Booklet* to label the diagram in **(e)(i)** to show

- which is the positive electrode,
- the direction of electron flow in the external circuit when a current flows.

[1]

23 An electrochemical cell is constructed using two half-cells.

- a Br_2/Br^- half-cell
- an Mn^{3+}/Mn^{2+} half-cell

(a) State the material used for the electrode in each half-cell.

Br_2/Br^- half-cell ....

Mn^{3+}/Mn^{2+} half-cell ....
[1]

(b) The cell is operated at 298 K.

The Br_2/Br^- half-cell has standard concentrations.

The Mn^{3+}/Mn^{2+} half-cell has $[Mn^{3+}] = 0.500\,mol\,dm^{-3}$ and $[Mn^{2+}] = 0.100\,mol\,dm^{-3}$.

(i) Use the Nernst equation to calculate the electrode potential, E, of the Mn^{3+}/Mn^{2+} half-cell under these conditions.

E = ... V [2]

(ii) Calculate the E_{cell} under these conditions.

E_{cell} = ... V [1]

(iii) Write an equation for the overall cell reaction that occurs.

.... [2]

24 Many copper compounds, such as $CuSO_4$ and $Cu(NO_3)_2$ contain Cu^{2+} ions. Aqueous solutions of this ion contain the $[Cu(H_2O)_6]^{2+}$ complex ion, in which water behaves as a monodentate ligand.

(d) When chlorine gas is bubbled into $FeSO_4$(aq) the colour of the solution changes from pale green to yellow.

Use data from the *Data Booklet* to explain this observation. Include an equation in your answer.

Reference to electron movement between orbitals is not needed.

....

....

.... [2]