24.2 ELECTRODE, CELL POTENTIAL & NERST EQUATION

Constructing redox equations using the relevant half-equations

1. Construct full redox equations from the following pairs of half-equations.
   a. The reaction of iodide ions with hydrogen peroxide:
   $I^- \rightarrow I_2 + e^-$
   $H_2O_2 + 2\ H^+ + 2\ e^- \rightarrow 2\ H_2O$

   b. The reaction of chloride ions with acidified manganese(IV) oxide:
   $Cl- \rightarrow Cl_2 + e^-$
   $MnO_2 + 4\ H^+ + 2e^- \rightarrow Mn^{2+} + 2\ H_2O$

   c. The reaction of acidified MnO_4^- ions with Fe^{2+} ions:
   $Fe^{2+} \rightarrow Fe^{3+} + e^-$
   $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

2. An electrochemical cell uses Ag_2O as the positive electrode and Zn as the negative electrode, immersed in an alkaline electrolyte. The overall cell reaction is shown.
   $Ag_2O + Zn + H_2O \rightarrow 2\ Ag + Zn(OH)_2$
   Complete the half-equation for the reaction at each electrode.

   at the positive electrode Ag_2O + ...

   at the negative electrode Zn + ... [2]

3. Hydrated compound J, $K_3Fe(C_2O_4)_3 \cdot xH_2O$, contains the green complex ion $[Fe(C_2O_4)_3]^{3-}$.
   The value of x can be determined by titration of a sample of J with acidified MnO_4^- ions.
   MnO_4^- ions oxidise $C_2O_4^{2-}$ ions in acidic conditions.
   $2\ MnO_4^- + 5\ C_2O_4^{2-} + 16\ H^+ \rightarrow 2\ Mn^{2+} + 10\ CO_2 + 8\ H_2O$
   Write half equations for the oxidation of $C_2O_4^-$ ions and for the reduction of MnO_4– ions.
   • oxidation of $C_2O_4^{2-}$

   .....
   • reduction of MnO_4^-

   .....[2]

Electrode potentials

4. Refer to the list of electrode potentials below to answer parts a to d.
   $Ag^+(aq) + e^-$ $Ag(s)$ voltage = +0.80 V
   $Co^{2+}(aq) + 2e^- \rightleftharpoons Co(s)$ voltage = –0.28 V
   $Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$ voltage = +0.34 V
   $Pb^{2+}(aq) + 2e^- \rightleftharpoons Pb(s)$ voltage = –0.13 V
   $Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$ voltage = –0.76 V

   a Which metal in the list is the best reducing agent?

   b Which metal ion in the list is most difficult to reduce?

   c Which metal in the list is most reactive?

   d Which metal ion in the list is the easiest to reduce?

5. Define standard electrode potential, $E^{\ominus}$, including a description of standard conditions.

……

……

…… [2]

Define standard electrode potential. Include details of the standard conditions used. [2]

6. Standard electrode potentials are measured under standard conditions.
Describe the standard conditions used in the Sn^{4+} / Sn^{2+} half-cell.

……

…… [1]

7. Standard electrode potentials can be used to compare the stability of different complex ions for a given transition element.
Table 4.1 lists electrode potentials for some electrode reactions for Fe^{3+} / Fe^{2+} systems.

electrode reaction	$E^{\ominus}$ / V
$[Fe(H_2O)_6]^{3+} + e^- \rightleftharpoons [Fe(H_2O)_6]^{2+}$	+0.77
$[Fe(CN)_6]^{3-} + e^- \rightleftharpoons [Fe(CN)_6]^{4-}$	+0.36
$[Fe(bipy)_3]^{3+} + e^- \rightleftharpoons [Fe(bipy)_3]^{2+}$	+0.96

Use relevant data from Table 4.1 to state which iron(III) complex is hardest to reduce.
Explain your choice.

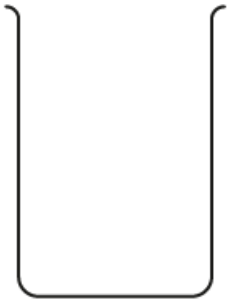
iron(III) complex .....

explanation .....

..... [1]

Standard hydrogen electrode

8. Complete the diagram to show a standard hydrogen electrode.
Label your diagram. Identify all substances. You do not need to state standard conditions. [1]



Combining two half-cells into a Galvanic Cell

9. An electrochemical cell is constructed using two half-cells.
   - an Sn^{4+}/Sn^{2+} half-cell
   - an Al^{3+}/Al half-cell

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

   ●Sn^{4+}/Sn^{2+} half-cell .....

   ● Al^{3+}/Al half-cell ..... [1]

10. 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]

11. 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.
   (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]

12. A salt bridge is used in an electrochemical cell.
   State the function of the salt bridge. Explain your answer.

   .....

   ..... [1]

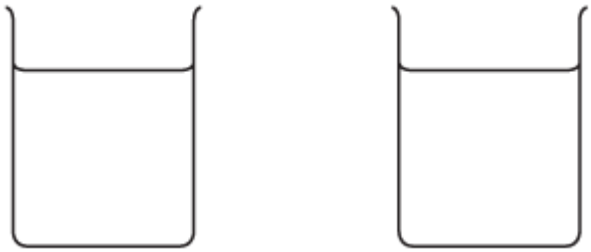
13. The equation representing the standard electrode potential, $E^\ominus$, for the reduction of $MnO_4^-(aq)$ to $Mn^{2+}(aq)$ in acid solution is given.
   $MnO_4^-(aq) + 8\ H^+(aq) + 5\ e^- \rightleftharpoons Mn^{2+}(aq) + 4\ H_2O(l)$ $E^\ominus = +1.52$ V
   Draw a diagram of the apparatus that 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]

14. 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 the conditions used. [3]

15. Draw a fully labelled diagram of the apparatus that can be used to measure the cell potential of a cell composed of a Cu(II)/Cu electrode and an Fe(III)/Fe(II) electrode. Include all necessary reactants. [3]

16. 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]
$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ half-equation 1

17. 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.



Identifying positive and Negative electrodes in a cell

18. Some electrode potentials are shown in the Table below.

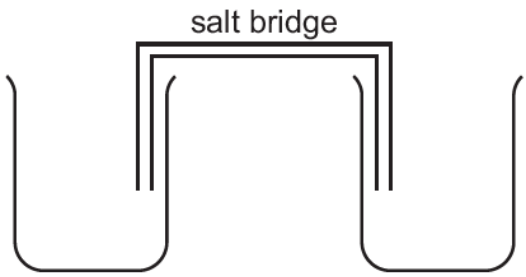
electrode reaction	$E^{\ominus}$ / V
$V^{2+} + 2e^- \rightleftharpoons V$	–1.20
$V^{3+} + e^- \rightleftharpoons V^{2+}$	–0.26
$VO^{2+} + 2H^+ + e^- \rightleftharpoons V^{3+} + H_2O$	+0.34
$VO_2^+ + 2H^+ + e^- \rightleftharpoons VO^{2+} + H_2O$	+1.00
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	–0.44
$Fe^{3+} + 3e^- \rightleftharpoons Fe$	–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

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]

19. 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]



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 the above diagram. [1]

20. (i) Complete the diagram in Fig. 2.1 to show how the standard electrode potential of the Cr^{3+}/Cr half-cell can be measured relative to that of the standard hydrogen electrode.
Identify the chemicals, conditions and relevant pieces of apparatus.

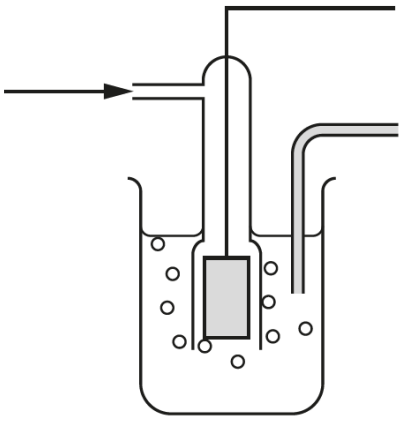


Fig. 2.1

(ii) Label Fig. 2.1 to show:
• which is the positive electrode
• the direction of electron flow in the external circuit. [1]

Calculating $E^{\ominus}_{cell}$

21. Define the term standard cell potential, $E^{\ominus}_{cell}$.

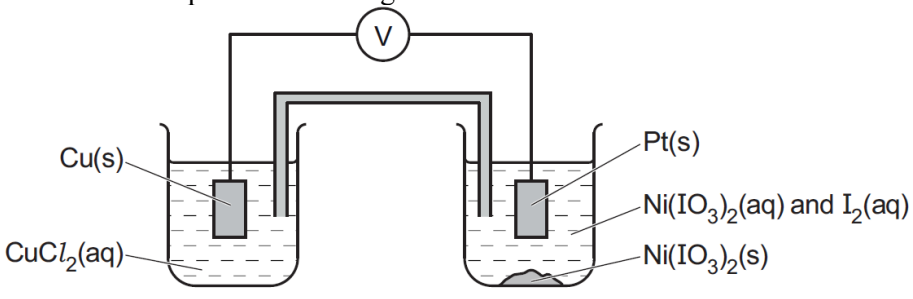
……

……

…… [2]

Define the term *standard cell potential*. [2]

22. An electrochemical cell is set up as shown in Fig.



The relevant standard electrode potentials, $E^{\ominus}$, for this electrochemical cell are shown.
$IO_3^-(aq) + 6H^+(aq) + 5e^- \rightleftharpoons 1/2\ I_2(aq) + 3H_2O(l)$ $E^{\ominus} = +1.19$ V
$Cu^{2+}(aq) + 2e- \rightleftharpoons Cu(s)$ $E^{\ominus} = +0.34$ V
(i) Use this information to calculate the value of $E^{\ominus}_{cell}$. State which electrode is positive.

$E^{\ominus}_{cell}$ = ... positive electrode is ..... [1]

23. $H_2PO_2^-$ is a strong reducing agent. It can be used to reduce metal cations without the need for electrolysis.
equation 1 $HPO_3^{2-} + 2H_2O + 2e^- \rightleftharpoons H_2PO_2^- + 3OH^-$ $E^\ominus = -1.57$ V
The Cr^{3+}/Cr half-cell has a standard electrode potential of –0.74 V.
An electrochemical cell consists of an alkaline $HPO_3^{2-}/H_2PO_2^-$ half-cell and a Cr^{3+}/Cr half-cell.
Calculate the standard cell potential, $E^\ominus_{cell}$.

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

24. 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.

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

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

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

overall cell reaction .....

$E^\ominus_{cell}$ = ... (ans = 0.99) ..... V [2]

25. Table 3.1 lists relevant electrode potentials for some electrode reactions.

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

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]

26. $H_2PO_2^-$ reduces Ni^{2+} to Ni in alkaline conditions.
Use equation 1 to construct the ionic equation for this reaction.
equation 1 $HPO_3^{2-} + 2H_2O + 2e^- \rightleftharpoons H_2PO_2^- + 3OH^-$

……..... [1]

27. 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.
$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ half-equation 1
Use these equations to deduce the half-equation for the reduction of carbon dioxide in this process. [2]

28. The decomposition of hydrogen peroxide, H_2O_2, is catalysed by acidified IO_3^-.
H_2O_2 reduces acidified IO_3^- as shown.
$5H_2O_2 + 2H^+ + 2IO_3^- \rightarrow I_2 + 5O_2 + 6H_2O$
This reaction is followed by the oxidation of I_2 by H_2O_2.

half-equation	$E^\ominus$ / V
$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$	+1.77
$IO_3^- + 6H^+ + 5e^- \rightleftharpoons 1/2\ I_2 + 3H_2O$	+1.19
$O_2 + 2H^+ + 2e^- \rightleftharpoons H_2O_2$	+0.68

(i) Use the data to show that the separate reactions of H_2O_2 with IO_3^- and with I_2 are both feasible under standard conditions.
In your answer, give the equation for the reaction of H_2O_2 with I_2.

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.....

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..... [3]

(ii) Write the overall equation for the decomposition of H_2O_2 catalysed by acidified IO_3^-.

..... [1]

29. Data should be selected from the given table in order to answer some parts of this question.

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

(ii) Equal volumes of 1.0 mol dm^{-3} Sn^{2+}(aq) and 1.0 mol dm^{-3} Cl^-(aq) are mixed.
Use relevant E$^{\ominus}$ 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} VO2+(aq) are mixed.
Write an equation for the reaction that takes place in the resulting mixture.

..... [2]

30. An electrochemical cell is constructed using the electrodes shown in Table 1.4.

electrode	half-equation	$E^{\ominus}$/V
1	$AgCl(s) + e^- \rightleftharpoons Ag(s) + Cl^-(aq)$	+0.222
2	$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.342

(i) Calculate the standard cell potential, $E^{\ominus}_{cell}$.
Construct an equation for the overall cell reaction.

$E^{\ominus}_{cell}$ = ... (Ans = + 0.120) V

equation ..... [2]

31. 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 value are shown.

$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$ E = +0.17 V

Four more half-equations for reactions occurring under acidic conditions, and their E o– values, are shown.

$H_3BO_3 + 3H^+ + 3e^- \rightleftharpoons B + 3H_2O$ $E^\ominus$ = –0.73 V
$BiO^+ + 2H^+ + 3e^- \rightleftharpoons Bi + H_2O$ $E^\ominus$ = +0.28 V
$S + 2H^+ + 2e^- \rightleftharpoons H_2S$ $E^\ominus$ = +0.14 V
$Sb + 3H^+ + 3e^- \rightleftharpoons SbH_3$ $E^\ominus$ = –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]

32. Fe^{2+} ions can be oxidised to Fe^{3+} ions under alkaline conditions by suitable oxidising agents.
The half-equation for the reduction of Fe^{3+} under alkaline conditions and its $E^\ominus$ value are shown.

$Fe(OH)_3 + e^- \rightarrow Fe(OH)_2 + OH^-$ $E^\ominus$ = – 0.56 V

Four more half-equations for reactions under alkaline conditions, and their E o values, are shown.

$Al(OH)_4^- + 3e^- \rightarrow Al + 4OH^-$ $E^\ominus$ = –2.35 V
$ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$ $E^\ominus$ = +0.89 V
$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$ $E^\ominus$ = +0.40 V
$Zn(OH)_4^{2-} + 2e^- \rightarrow Zn + 4OH^-$ $E^\ominus$ = –1.22 V

Select two oxidising agents that can oxidise Fe^{2+} ions to Fe^{3+} ions under alkaline conditions.
Write an equation, and give the E cell value, for each of the two reactions that occur.

oxidising agent 1: .....

equation: .....

E_{cell} = ... V

oxidising agent 2: .....

equation: .....

$E^\ominus_{cell}$ = ... V [4]

33. H_2O_2 can act as an oxidising agent or a reducing agent when reacting with species that contain manganese. Table 2.2 shows electrode potentials, $E^{\ominus}$, for some electrode reactions.

Table 2.2

electrode reaction	$E^{\ominus}$/ V
$MnO_2 + 2H_2O + 2e^- \rightleftharpoons Mn(OH)_2 + 2OH^-$	–0.04
$MnO_2 + 4H^+ + 2e^- \rightleftharpoons Mn^{2+} + 2H_2O$	+1.22
$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$	+1.78
$H_2O_2 + OH– + 2e^- \rightleftharpoons 3OH^-$	+0.88
$O2 + 2H^+ + 2e^- \rightleftharpoons H_2O_2$	+0.68

Use only the species listed in Table 2.2 to suggest:
• one reaction in which H_2O_2 acts as an oxidising agent and
• one reaction in which H_2O_2 acts as a reducing agent.
Include the value of the standard cell potential, $E^{\ominus}_{cell}$, and an overall equation for each reaction.

H_2O_2 acting as an oxidising agent .....

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.....

H_2O_2 acting as a reducing agent .....

.....

.....

..... [4]

34. Co^{2+} and Co^{3+} both form complexes with $edta^{4–}$.

half-equation	E o / V
$Co^{3+} + e^- \rightarrow Co^{2+}$	+1.82
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.23
$[Co(edta)]– + e^- \rightarrow [Co(edta)]^{2-}$	+0.38
$Co^{2+} + 2e^- \rightarrow Co$	–0.28

Use the data in the table to predict what happens, if anything, when separate aqueous solutions of Co^{3+} and $[Co(edta)]^-$ are left to stand in the air.

aqueous solution of Co^{3+}

.....

…….....

aqueous solution of $[Co(edta)]^-$

…….....

.....[3]

35. Table 2.1 lists relevant electrode potentials for some electrode reactions.

electrode reaction	E o / V
$Mn^{2+} + 2e^- \rightarrow Mn$	–1.18
$Cl_2 + 2e- \rightarrow 2Cl^-$	+1.36
$2HOCl + 2H^+ + 2e^- \rightarrow Cl_2 + 2H_2O$	+1.64
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	+1.67

Suggest the formula of the manganese species formed when $Mn^{2+}(aq)$ reacts with Cl_2.
State the type of reaction.

formula of manganese species formed .....

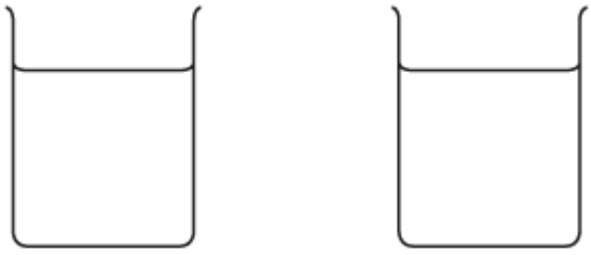
type of reaction ..... [1]

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

electrode reaction	E⊖ / 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.

(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.



(ii) 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]

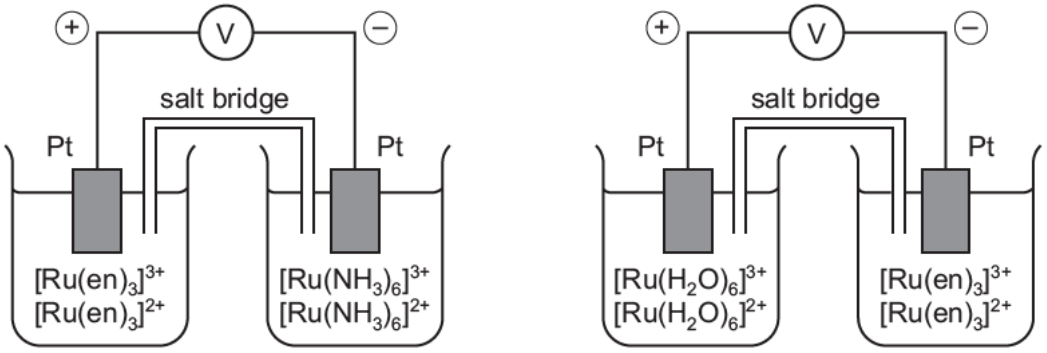
(iv) Equal volumes of 1.0 mol dm^{-3} of $Sn^{2+}(aq)$ and acidified 1.0 mol dm^{-3} $VO2^{+}(aq)$ are mixed.
Write an equation for the reaction that takes place in the resulting mixture.

..... [2]

37. 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.



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	

38. Table 1.1 lists standard electrode potentials, E ⊖, for some electrode reactions of chromium and other species.
Table 1.1

electrode reaction	E ⊖ / V
$Cr^{2+}(aq) + 2e^{+}$ Cr(s)	– 0.91
Cr3+(aq) + e– Cr2+(aq)	– 0.41
$Cr(OH)_3(s)$ + e– $Cr(OH)_2(s) + OH–(aq)$	–1.26
$Cr_2O72–(aq) + 14H+(aq)$ + 6e– $2Cr3+(aq) + 7H_2O(l)$	+1.33
$CrO_4^{2–}(aq) + 4H_2O(l)$ + 3e– $Cr(OH)_3(s) + 5OH–(aq)$	– 0.13
$O2(g) + 4H+(aq)$ + 4e– $2H_2O(l)$	+1.23
$O2(g) + 2H_2O(l)$ + 4e– 4OH–(aq)	+0.40

Explain fully, with reference to relevant data in Table 1.1, why solutions of $Cr^{2+}(aq)$ need to be stored in an atmosphere of argon instead of in air.

.....

.....

..... [2]

(iii) A green precipitate of $Cr(OH)_3$ forms when aqueous NaOH is added to aqueous $CrCl_3$.
Construct an ionic equation for the formation of this green precipitate.

..... [1]

(iv) Green $Cr(OH)_3(s)$ dissolves in alkali to form a yellow solution containing $CrO_4^{2-}(aq)$ when it is left in air for several hours.
Use Table 1.1 to construct an ionic equation for this reaction.

..... [2]

39. $LiFePO_4$ can be used in lithium-ion rechargeable batteries.
When the cell is charging, lithium reacts with a graphite electrode to form LiC_6.
When the cell is discharging, the half-equations for the two processes that occur are as follows.
anode half-equation $LiC_6 \rightarrow 6C + Li^+ + e^-$
cathode half-equation $Li^+ + FePO_4 + e^- \rightarrow LiFePO_4$
(i) State one possible advantage of developing cells such as lithium-ion rechargeable batteries.

..... [1]

(i) Use the cathode half-equation to determine the change, if any, in oxidation states of lithium and iron at the cathode during discharging. metal change in oxidation state during discharging from to lithium iron [1]

metal	change in oxidation state during discharging	
	from	to
lithium		
iron		

(ii) Write the equation for the overall reaction that occurs when this cell is discharging.

..... [1]

Variation of electrode potential with concentration

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

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

41. $H_2PO_2^-$ is a strong reducing agent. It can be used to reduce metal cations without the need for electrolysis.
equation 1 $HPO_3^{2-} + 2H_2O + 2e^- \rightleftharpoons H_2PO_2^- + 3OH^-$ $E^\ominus = -1.57$ V
(i) In an experiment, an alkaline $HPO_3^{2-}/H_2PO_2^-$ half-cell is constructed with $[H_2PO_2^-] = 0.050$ mol dm^{-3}.
All other ions are at their standard concentration.
Predict how the value of E of this half-cell differs from its $E^\ominus$ value.
Explain your answer.

.....

.....

.....

..... [2]

42. Copper metal can be oxidised by acidified $KMnO_4$. The relevant half-equations and their standard electrode potentials, $E^{\ominus}$, are shown.
$Cu^{2+} + 2e^- \rightleftharpoons Cu$ $E^{\ominus} = +0.34$ V
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$ $E^{\ominus} = +1.52$ V
(i) A MnO_4^-/Mn^{2+} electrode is constructed using 0.0020 mol dm^{-3} MnO_4^-, 1.0 mol dm^{-3} $Mn^{2+}$and 1.0 mol dm^{-3} H+. The temperature used is 298 K.
Use the Nernst equation to show that the E value for this MnO_4–/Mn^{2+} electrode is +1.49 V. [2]

(ii) An electrochemical cell is constructed using a standard Cu^{2+}/Cu electrode and the MnO_4^-/Mn^{2+} electrode described in (d)(i).
Calculate the value of $E^{\ominus}_{cell}$.

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

(iii) Write an equation for the reaction taking place in the electrochemical cell described in (d)(ii).

..... [1]

(iv) Complete the sentences for the electrochemical cell described in (d)(ii).

The ..... electrode is the negative electrode. Electrons flow from the ..... electrode to

the ..... electrode when the cell is in use. [1]

43. The relevant standard electrode potentials, $E^{\ominus}$, for this electrochemical cell are shown.
$IO_3^-(aq) + 6H^+(aq) + 5e^- \rightleftharpoons 1/2\ I_2(aq) + 3H_2O(l)$ $E^{\ominus} = +1.19$ V
$Cu^{2+}(aq) + 2e– \rightleftharpoons Cu(s)$ $E^{\ominus} = +0.34$ V
(i) Use this information to calculate the value of $E^{\ominus}_{cell}$. State which electrode is positive.

$E^{\ominus}_{cell}$ = ... positive electrode is ..... [1]

(ii) Suggest how the measured E_{cell} of this cell compares to the $E^{\ominus}_{cell}$ under standard conditions.
Explain your answer.

.....

..... [1]

The Nernst Equation

44. Some electrode potentials are shown in the Table below.

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) 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]

(b) 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 = ...(ans = 0.947)..... V [2]

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

..... [1]

45. The $E^{\ominus}$ values for two electrode reactions are given.
$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$ $E^{\ominus} = +1.77$ V
$Co^{3+} + e^- \rightleftharpoons Co^{2+}$ $E^{\ominus} = +1.82$ 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]

46. An electrochemical cell is constructed using the electrodes shown in Table 1.4.

electrode	half-equation	$E^{\ominus}$/V
1	$AgCl(s) + e^- \rightleftharpoons Ag(s) + Cl^-(aq)$	+0.222
2	$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.342

(ii) In a different experiment, electrode 1 is set up using a saturated solution of KCl.
Saturated KCl (aq) contains 36.0 g of KCl per 100 cm^3 of solution at 298 K.
The Nernst equation for electrode 1 is:

$$E = E^{\ominus} + \frac{0.059}{z} \log \frac{1}{[Cl^-(aq)]}$$

Calculate the electrode potential, E, of electrode 1 under these conditions.

E = ...(Ans = +0.182)..... V [3]

47. The cell is operated at 298 K.
The Al^{3+}/Al half-cell has standard concentrations.
The Sn^{4+}/Sn^{2+} half-cell has $[Sn^{4+}]$ = 0.300 mol dm^{-3} and $[Sn^{2+}]$ = 0.150 mol dm^{-3}.
(i) Use the Nernst equation to calculate the electrode potential, E, of the Sn^{4+}/Sn^{2+} half-cell under these conditions.

E = ...(ans = 0.159)..... V [2]

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

E_{cell} = ...(Ans = 1.82)..... V [1]

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

……..... [2]

48. An electrochemical cell is constructed using two half-cells.
●● a Br_2/Br^- half-cell
●● an Mn^{3+}/Mn^{2+} half-cell
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 = ...(and = 1.53)..... V [2]

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

E_{cell} = ...(ans = 0.46)..... V [1]

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

……..... [2]

49. 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.

..... 1.36 1.07 0.54 ..... [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]

50. 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.

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 .....

.....

Changes in pH

51. Table 2.1 shows electrode potentials for some electrode reactions involving manganese compounds.

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]

52. Table 3.1 lists relevant electrode potentials for some electrode reactions.

Table 3.1

electrode reaction	E o / 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

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.

$CH_3CH_2OH + O_2 \rightarrow CH_3COOH + H_2O$ $E_{cell}^{\ominus} = +2.01$ V

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

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

53. Some electrode potentials are shown in the Table below.

electrode reaction	E⊖ / V
$V^{2+} + 2e- \rightarrow V$	–1.20
$V^{3+} + e- \rightarrow V^{2+}$	–0.26
$VO^{2+} + 2H^+ + e^- \rightarrow V^{3+} + H_2O$	+0.34
$VO_2+ + 2H^+ + e^- \rightarrow VO2^+ + H_2O$	+1.00
$Fe^{2+} + 2e^- \rightarrow Fe$	–0.44
$Fe^{3+} + 3e^- \rightarrow Fe$	–0.04
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77
$2H^+ + 2e^- \rightarrow H_2$	0.00
$Cl\ O^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$	+0.89

An 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.

$\Delta G^{\ominus}$ = ... (ans = -257) ..... kJ mol^{-1} [3]

(ii) Predict whether this reaction becomes more or less feasible at a higher temperature.
Explain your answer.

The reaction becomes ..... feasible.

explanation .....

.....[1]

54. The $E^{\ominus}$ values for two electrode reactions are given.
$H_2O_2 + 2H^+ \rightleftharpoons + 2e^- \; 2H_2O \; E^{\ominus} = +1.77$ 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* o

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

55. A fuel cell is an electrochemical cell that can be used to generate electrical energy by using oxygen to oxidise a fuel.
Methanoic acid, HCOOH, is being investigated as a fuel in fuel cells.
When the cell operates, HCOOH is oxidised to carbon dioxide.
The half-equation for the reaction at the cathode is: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$.
In this fuel cell, the overall cell reaction is the same as that for the complete combustion of HCOOH.
(i) Deduce the half-equation for the reaction at the anode.

..... [1]

(ii) Calculate the volume, in cm^3, of oxygen used when a current of 3.75 A is delivered by the cell for 40.0 minutes. Assume the cell operates at room conditions.

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

56. **(a)** An electrochemical cell is set up to measure E ocell 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]

(b) The cell reaction for the electrochemical cell in **(a)** is shown.
$Cl_2 + 2Fe^{2+} \rightleftharpoons 2Fe^{3+} + 2Cl^-$ E o cell = +0.59 V
Calculate $\Delta G^{\ominus}$, in kJ mol^{-1}, for this cell reaction.

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

57. 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.

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

(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 **(i)** to calculate E_{cell} for this electrochemical cell.

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

58. Complete the diagram in Fig. 2.1 to show how the standard electrode potential of the Cr^{3+}/Cr half-cell can be measured relative to that of the standard hydrogen electrode.
Identify the chemicals, conditions and relevant pieces of apparatus.

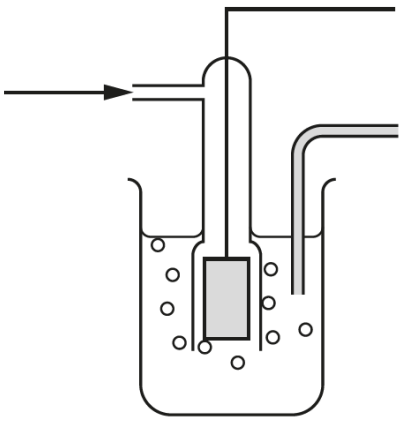


Fig. 2.1

(iv) Label Fig. 2.1 to show:
• which is the positive electrode
• the direction of electron flow in the external circuit. [1]

(v) $H_2PO_2^-$ reduces Ni^{2+} to Ni in alkaline conditions.
Use equation 1 to construct the ionic equation for this reaction.
equation 1 $HPO_3^{2-} + 2H_2O + 2e^- \rightleftharpoons H_2PO_2^- + 3OH^-$

……..... [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.

.....

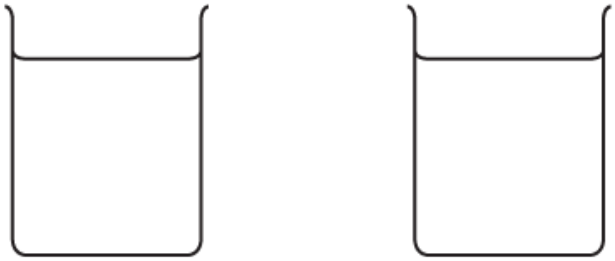
.....

..... [2]

59. Data should be selected from Table 3.1 in order to answer some parts of this question.
Table 3.1
electrode reaction *E* o / V
Mg2+ + 2e– Mg –2.38
Mn2+ + 2e– Mn –1.18
Mn3+ + e– Mn2+ +1.49
MnO2 + 4H+ + 2e– Mn2+ + 2H2O +1.23
MnO4– + e– MnO42– +0.56
MnO4– + 4H+ + 3e– MnO2 + 2H2O +1.67
MnO4– + 8H+ + 5e– Mn2+ + 4H2O +1.52
(a) An electrochemical cell can be constructed from a Mg2+ / Mg half-cell and a MnO4– / Mn2+ half-cell. The standard cell potential of this cell can be calculated using the standard electrode potentials of the two half-cells.

(i) Complete the diagram below to show an electrochemical cell constructed from a Mg2+ / Mg half-cell and a MnO4– / Mn2+ 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 Mg2+ / Mg half-cell is diluted by the addition of a large volume of water. Explain your answer.

.....

.....

..... [1]

60. **(a)** 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 cm3, formed when a current of 2.5 A is delivered by the cell for 30 minutes. Assume the cell is operated at room conditions.

volume of CO_2 = ... cm3 [2]

***... ***