Q1. The electrons transferred in redox reactions can be used by electrochemical cells to provide energy.

Some electrode half-equations and their standard electrode potentials are shown in the table below.

Half-equation	***E*⊖/V**
$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$	+1.33
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	–0.44
$Li^+(aq) + e^- \rightarrow Li(s)$	–3.04

(a) Describe a standard hydrogen electrode.

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(4)

(b) A conventional representation of a lithium cell is given below.
This cell has an e.m.f. of +2.91 V

$Li(s) \mid Li^+(aq) \parallel Li^+(aq) \mid MnO_2(s), LiMnO_2(s) \mid Pt(s)$

Write a half-equation for the reaction that occurs at the positive electrode of this cell.

Calculate the standard electrode potential of this positive electrode.

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(2)

(c) Suggest what reactions occur, if any, when hydrogen gas is bubbled into a solution containing a mixture of iron(II) and iron(III) ions. Explain your answer.

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(2)

(d) A solution of iron(II) sulfate was prepared by dissolving 10.00 g of $FeSO_4.7H_2O$ (M_r = 277.9) in water and making up to 250 cm^3 of solution. The solution was left to stand, exposed to air, and some of the iron(II) ions became oxidised to iron(III) ions. A 25.0 cm^3 sample of the partially oxidised solution required 23.70 cm^3 of 0.0100 mol dm^{-3} potassium dichromate(VI) solution for complete reaction in the presence of an excess of dilute sulfuric acid.

Calculate the percentage of iron(II) ions that had been oxidised by the air.

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(6)
(Total 14 marks)

Q2. In this question consider the data below.

	$E^\ominus$ / V
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	−0.13

The e.m.f. of the cell Ag(s) | $Ag^+(aq)$ || $Pb^{2+}(aq)$ | Pb(s) is

A 0.93 V

B 0.67 V

C −0.67 V

D −0.93 V

(Total 1 mark)

Q3. Nickel–cadmium cells are used to power electrical equipment such as drills and shavers.
The electrode reactions are shown below.

$$NiO(OH) + H_2O + e^- \rightarrow Ni(OH)_2 + OH^- \qquad E^\ominus = +0.52\ V$$

$$Cd(OH)_2 + 2e^- \rightarrow Cd + 2OH^- \qquad E^\ominus = –0.88\ V$$

(a) Calculate the e.m.f. of a nickel–cadmium cell.

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(1)

(b) Deduce an overall equation for the reaction that occurs in the cell when it is used.

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(2)

(c) Identify the oxidising agent in the overall cell reaction and give the oxidation state of the metal in this oxidising agent.

Oxidising agent ....

Oxidation state ....

(2)
(Total 5 marks)

Q4. (a) The term oxidation was used originally to describe a reaction in which a substance gained oxygen. The oxygen was provided by the oxidising agent. Later the definition of oxidation was revised when the importance of electron transfer was recognised.

An aqueous solution of sulfur dioxide was reacted in separate experiments as follows.

Reaction **1** with HgO

$$H_2O + SO_2 + HgO \rightarrow H_2SO_4 + Hg$$

Reaction **2** with chlorine

$$2H_2O + SO_2 + Cl_2 \rightarrow H_2SO_4 + 2HCl$$

(i) In Reaction **1**, identify the substance that donates oxygen and therefore is the oxidising agent.

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(ii) Show, by writing a half-equation, that this oxidising agent in reaction **1** is an electron acceptor.

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(iii) Write a half-equation for the oxidation process occurring in reaction **2**.

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(iv) Write a half-equation for the reduction process occurring in reaction **2**.

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(4)

(b) Use the standard electrode potential data given in the able below to answer the questions which follow.

			E / V
$V^{3+}(aq) + e^{-}$	$\rightarrow$	$V^{2+}(aq)$	–0.26
$SO_4^{2-}(aq) + 4H^{+}(aq) + 2e^{-}$	$\rightarrow$	$H_2SO_3(aq) + H_2O(l)$	+0.17
$VO^{2+}(aq) + 2H^{+}(aq) + e^{-}$	$\rightarrow$	$V^{3+}(aq) + H_2O(l)$	+0.34
$Fe^{3+}(aq) + e^{-}$	$\rightarrow$	$Fe^{2+}(aq)$	+0.77
$VO_2^{+}(aq) + 2H^{+}(aq) + e^{-}$	$\rightarrow$	$VO^{2+}(aq) + H_2O(l)$	+1.00
$MnO_4^{-}(aq) + 8H^{+}(aq) + 5e^{-}$	$\rightarrow$	$Mn^{2+}(aq) + 4H_2O(l)$	+1.52

Each of the above can be reversed under suitable conditions

(i) An excess of potassium manganate(VII) was added to a solution containing $V^{2+}(aq)$ ions. Determine the vanadium species present in the solution at the end of this reaction. State the oxidation state of vanadium in this species and write a half-equation for its formation from $V^{2+}(aq)$.

Vanadium species present at the end of the reaction ....

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Oxidation state of vanadium in the final species ....

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Half-equation ....

(ii) The cell represented below was set up under standard conditions.

$$Pt|H_2SO_3(aq),SO_4^{2-}(aq),H^{+}(aq)||Fe^{3+}(aq),Fe^{2+}(aq)|Pt$$

Calculate the e.m.f. of this cell and state, with an explanation, how this e.m.f. will change if the concentration of $Fe^{3+}(aq)$ ions is increased.

Cell e.m.f. ....

Change in cell e.m.f. ....

Explanation ....

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(7)

(c) Consider the cell below

$$- \quad Pt|H_2(g)|H^+(aq)||O_2(g)|OH^-(aq)|Pt \quad +$$

(i) Using half-equations, deduce an overall equation for the cell reaction.

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(ii) State how, if at all, the e.m.f. of this cell will change if the surface area of each platinum electrode is doubled.

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(3)

(d) Currently, almost all hydrogen is produced by the high-temperature reaction between methane, from North Sea gas, and steam. Give one economic and one environmental disadvantage of this method of producing hydrogen.

Economic disadvantage ....

Environmental disadvantage ....

(2)

(e) Hydrogen can also be produced by the electrolysis of acidified water using electricity produced using solar cells. Give one reason why this method is not used on a large scale.

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(1)
(Total 17 marks)

Q5. Use the standard electrode potential data given in the table below, where appropriate, to answer the questions which follow.

	$E^{\ominus}/V$
$V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$	–0.26
$SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow H_2SO_3(aq) + H_2O$	+0.17
$VO^{2+}(aq) + 2H^+(aq) + e^- \rightarrow V^{3+}(aq) + H_2O(l)$	+0.34
$O_2(g) + 2H^+(aq) + 2e^- \rightarrow H_2O_2(aq)$	+0.68
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.77
$VO_2^+(aq) + 2H^+(aq) + e^- \rightarrow VO^{2+}(aq) + H_2O(l)$	+1.00
$2IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(aq) + 6H_2O(l)$	+1.19
$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$	+1.52

Each of the above can be reversed under suitable conditions.

(a) The cell represented below was set up under standard conditions.

Pt | H_2SO_3(aq), SO_4^{2-} (aq), || Fe^{3+}(aq), Fe^{2+}(aq) |Pt

(i) Calculate the e.m.f. of this cell.

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(ii) Write a half-equation for the oxidation process occurring at the negative electrode of this cell.

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(2)

(b) The cell represented below was set up under standard conditions.

Pt | H_2O_2(aq), O_2(g) || IO_3^- (aq), I_2(aq) |Pt

(i) Write an equation for the spontaneous cell reaction.

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(ii) Give **one** reason why the e.m.f. of this cell changes when the electrodes are connected and a current flows.

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(iii) State how, if at all, the e.m.f. of this standard cell will change if the surface area of each platinum electrode is doubled.

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(iv) State how, if at all, the e.m.f. of this cell will change if the concentration of IO_3^- ions is increased. Explain your answer.

Change, if any, in e.m.f. of cell ....

Explanation ....

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(7)

(c) An excess of acidified potassium manganate(VII) was added to a solution containing V^{2+}(aq) ions. Use the data given in the table to determine the vanadium species present in the solution at the end of this reaction. State the oxidation state of vanadium in this species and write a half-equation for its formation from V^{2+}(aq).

Vanadium species present at end of reaction ....

Oxidation state of vanadium in final species ....

Half-equation ....

(3)
(Total 12 marks)

Q6. Use the data in the table below, where appropriate, to answer the questions which follow.

Standard electrode potentials			$E^{\ominus}$ / V
$Fe^{3+}(aq) + e^-$	$\longrightarrow$	$Fe^{2+}(aq)$	+0.77
$Cl_2(g) + 2e^-$	$\longrightarrow$	$2Cl^-(aq)$	+1.36
$2BrO_3^-(aq) + 12H^+(aq) + 10e^-$	$\longrightarrow$	$Br_2(aq) + 6H_2O(l)$	+1.52
$O_3(g) + 2H^+(aq) + 2e^-$	$\longrightarrow$	$O_2(g) + H_2O(l)$	+2.08
$F_2O(g) + 2H^+(aq) + 4e^-$	$\longrightarrow$	$2F^-(aq) + H_2O(l)$	+2.15

Each of the above can be reversed under suitable conditions.

(a) (i) Identify the most powerful reducing agent in the table.

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(ii) Identify the most powerful oxidising agent in the table.

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(iii) Identify **all** the species in the table which can be oxidised in acidic solution by $BrO_3^-(aq)$.

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(4)

(b) The cell represented below was set up.

$Pt|Fe^{2+}(aq), Fe^{3+}(aq) || BrO_3^-(aq), Br_2(aq)|Pt$

(i) Deduce the e.m.f. of this cell.

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(ii) Write a half-equation for the reaction occurring at the negative electrode when current is taken from this cell.

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(iii) Deduce what change in the concentration of Fe^{3+}(aq) would cause an increase in the e.m.f. of the cell. Explain your answer.

Change in concentration ....

Explanation ....

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(6)
(Total 10 marks)