# CHAPTER 20 ELECTRODE POTENTIALS & ELECTROCHEMICAL CELLS

**1** Hydrogen – oxygen fuel cells can operate in acidic or in alkaline conditions but commercial cells use porous platinum electrodes in contact with concentrated aqueous potassium hydroxide. The table below shows some standard electrode potentials measured in acidic and in alkaline conditions.

| **Half-equation** | $E^\ominus$ **/ V** |
|---|---|
| $O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$ | +1.23 |
| $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$ | +0.40 |
| $2H^+(aq) + 2e^- \longrightarrow H_2(g)$ | 0.00 |
| $2H_2O(l) + 2e^- \longrightarrow 2OH^-(aq) + H_2(g)$ | –0.83 |

(a) State why the electrode potential for the standard hydrogen electrode is equal to 0.00 V.

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*(1 mark)*

(b) Use data from the table to calculate the e.m.f. of a hydrogen – oxygen fuel cell operating in alkaline conditions.

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*(1 mark)*

(c) Write the conventional representation for an alkaline hydrogen – oxygen fuel cell.

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*(2 marks)*

(d) Use the appropriate half-equations to construct an overall equation for the reaction that occurs when an alkaline hydrogen – oxygen fuel cell operates. Show your working.

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*(2 marks)*

(e) Give **one** reason, other than cost, why the platinum electrodes are made by coating a porous ceramic material with platinum rather than by using platinum rods.

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*(1 mark)*

(f) Suggest why the e.m.f. of a hydrogen–oxygen fuel cell, operating in acidic conditions, is exactly the same as that of an alkaline fuel cell.

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*(1 mark)*

(g) Other than its lack of pollution, state briefly the main advantage of a fuel cell over a re-chargeable cell such as the nickel–cadmium cell when used to provide power for an electric motor that propels a vehicle.

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*(1 mark)*

(h) Hydrogen–oxygen fuel cells are sometimes regarded as a source of energy that is carbon neutral. Give **one** reason why this may **not** be true.

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*(1 mark)*

**2** Where appropriate, use the standard electrode potential data in the table below to answer the questions which follow.

| | $E^\ominus$/V |
|---|---|
| $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ | −0.76 |
| $V^{3+}(aq) + e^- \rightarrow V^{2+}(aq)$ | −0.26 |
| $SO_4^{2-}(aq) + 2H^+(aq) + 2e^- \rightarrow SO_3^{2-}(aq) + H_2O(l)$ | +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 |
| $Cl_2(aq) + 2e^- \rightarrow 2Cl^-(aq)$ | +1.36 |

(a) From the table above select the species which is the most powerful reducing agent.

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*(1 mark)*

(b) From the table above select

(i) a species which, in acidic solution, will reduce $VO_2^+(aq)$ to $VO^{2+}(aq)$ but will **not** reduce $VO^{2+}(aq)$ to $V^{3+}(aq)$,

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(ii) a species which, in acidic solution, will oxidise $VO^{2+}(aq)$ to $VO_2^+(aq)$.

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*(2 marks)*

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

$Pt|Fe^{2+}(aq), Fe^{3+}(aq)||Tl^{3+}(aq), Tl^+(aq)|Pt$ Cell e.m.f. = + 0.48 V

(i) Deduce the standard electrode potential for the following half-reaction.

$$Tl^{3+}(aq) + 2e^- \rightarrow Tl^+(aq)$$

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

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*(3 marks)*

**3** **Table 3** shows some standard electrode potential data.

**Table 3**

| | $E^{\ominus}$ / V |
|---|---|
| $ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$ | –1.25 |
| $Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$ | –0.44 |
| $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$ | +0.40 |
| $2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$ | +1.64 |

**(a)** Give the conventional representation of the cell that is used to measure the standard electrode potential of iron as shown in **Table 3**.

**[2 marks]**

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**(b)** With reference to electrons, give the meaning of the term **reducing agent**.

**[1 mark]**

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**(c)** Identify the weakest reducing agent from the species in **Table 3**.

Explain how you deduced your answer.

**[2 marks]**

Species ......................................................................................................................

Explanation ................................................................................................................

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**(d)** When HOCl acts as an oxidising agent, one of the atoms in the molecule is reduced.

**(i)** Place a tick (✓) next to the atom that is reduced.

**[1 mark]**

| Atom that is reduced | Tick (✓) |
|---|---|
| H | |
| O | |
| Cl | |

**(ii)** Explain your answer to Question **5 (d) (i)** in terms of the change in the oxidation state of this atom.

**[1 mark]**

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**(e)** Using the information given in **Table 3**, deduce an equation for the redox reaction that would occur when hydroxide ions are added to HOCl

**[2 marks]**

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**(f)** **Table 3** is repeated to help you answer this question.

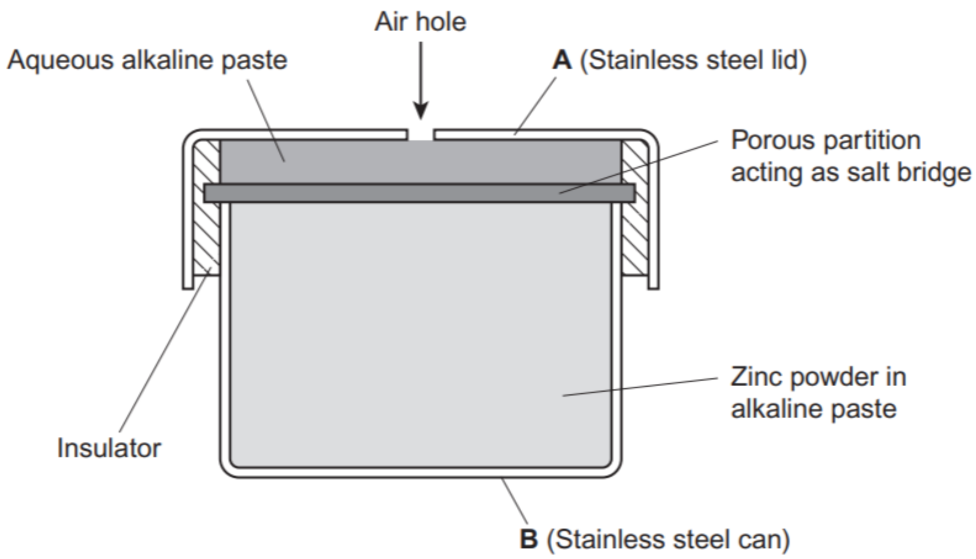
**Table 3**

| | $E^{\ominus}$ / V |
|---|---|
| $ZnO(s) + H_2O(l) + 2e^- \longrightarrow Zn(s) + 2OH^-(aq)$ | –1.25 |
| $Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$ | –0.44 |
| $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$ | +0.40 |
| $2HOCl(aq) + 2H^+(aq) + 2e^- \longrightarrow Cl_2(g) + 2H_2O(l)$ | +1.64 |

The half-equations from **Table 3** that involve zinc and oxygen are simplified versions of those that occur in hearing aid cells.

A simplified diagram of a hearing aid cell is shown in **Figure 1**.

**Figure 1**




**(i)** Use data from **Table 3** to calculate the e.m.f. of this cell.

**[1 mark]**

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Answer = ..................................