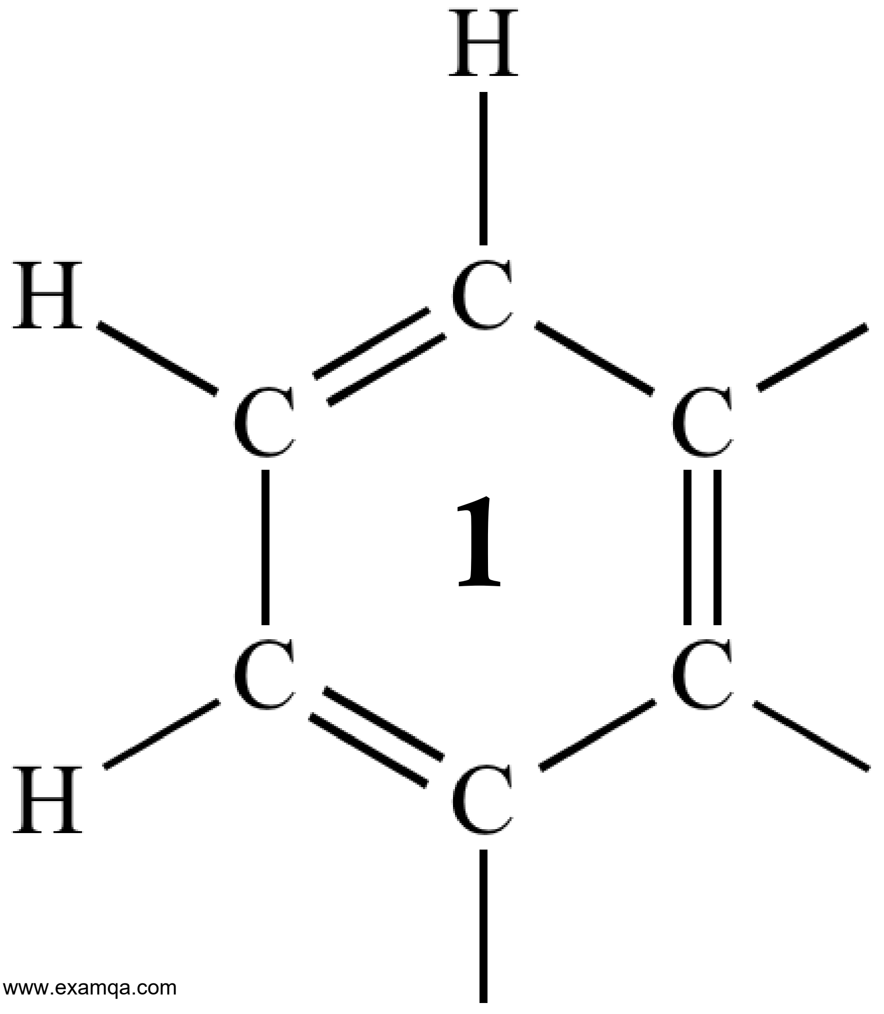
OCR A2 CHEMISTRY

# MODULE 5.5

## REDOX
## ELECTRODE POTENTIALS



**1** The table shows some standard electrode potential data.

| **Electrode half-reaction** | ***E***$^\ominus$**/V** |
|---|---|
| $Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$ | –0.76 |
| $Co^{2+}(aq) + 2e^- \longrightarrow Co(s)$ | –0.28 |
| $\frac{1}{2}O_2(g) + 2H^+ (aq) + 2e^- \longrightarrow H_2O(l)$ | +1.23 |
| $Au^+(aq) + e^- \longrightarrow Au(s)$ | +1.68 |
| $Co^{3+}(aq) + e^- \longrightarrow Co^{2+}(aq)$/p> | +1.82 |

(a) (i) Identify the weakest oxidising agent in the table.

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**(1)**

(ii) Give the conditions under which the electrode potential of the $Zn^{2+}/Zn$ electrode is –0.76 V.

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**(2)**

(b) Two half-cells, involving species in the table, are connected together to give a cell with an e.m.f. = +0.48 V.

Use data from the table to deduce the conventional representation of this cell. Write the half-equation for the reaction that occurs at the negative electrode.

Conventional representation .......................................................................................

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Half-equation .......................................................................................

**(3)**

(c) Use data from the table to identify a cobalt species that can react with water.

Write an equation for the redox reaction that occurs and identify the oxidation product in the reaction.

Cobalt species ..........................................................................................

Equation ...................................................................................................

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Oxidation product .......................................................................................

**(3)**

(d) Use data from the table to explain why gold jewellery is unreactive in moist air.

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**(2)**
**(Total 11 marks)**

**2** A representation of a hydrogen–oxygen fuel cell that operates in alkaline conditions is

$$Pt|H_2|H_2O||O_2|OH^-|Pt$$

(a) (i) Write a half-equation for the reaction that occurs at each electrode.
Use the half-equations to deduce an overall equation for the cell.

Half-equation at positive electrode ........................................................

Half-equation at negative electrode .......................................................

Overall equation ........................................................................................

**(3)**

(ii) State and explain the effect, if any, of increasing the pressure of oxygen on the e.m.f of this cell.

Effect on e.m.f. .........................................................................................

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

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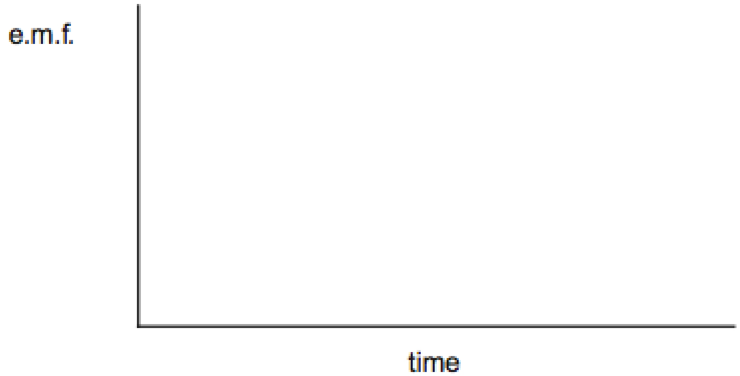
**(2)**

(b) Complete the diagram to show how the e.m.f. of a hydrogen–oxygen fuel cell changes with time.




**(1)**

(c) (i) Suggest the effect, if any, on the e.m.f. of this cell if the surface area of each platinum electrode is increased.

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**(1)**

(ii) State the main environmental advantage of using a hydrogen–oxygen fuel cell to power a car.

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**(1)**

(d) Suggest why the use of a hydrogen–oxygen fuel cell might not be carbon-neutral.

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**(1)**
**(Total 9 marks)**

**3** Fuel cells are an increasingly important energy source for vehicles. Standard electrode potentials are used in understanding some familiar chemical reactions including those in fuel cells.

The following table contains some standard electrode potential data.

| **Electrode half-equation** | **$E^{\ominus}$ / V** |
|---|---|
| $F_2 + 2e^- \longrightarrow 2F^-$ | +2.87 |
| $Cl_2 + 2e^- \longrightarrow 2Cl^-$ | +1.36 |
| $O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$ | +1.23 |
| $Br_2 + 2e^- \longrightarrow 2Br^-$ | +1.07 |
| $I_2 + 2e^- \longrightarrow 2I^-$ | +0.54 |
| $O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$ | +0.40 |
| $SO_4^{2-} + 4H^+ + 2e^- \longrightarrow SO_2 + 2H_2O$ | +0.17 |
| $2H^+ + 2e^- \longrightarrow H_2$ | 0.00 |
| $4H_2O + 4e^- \longrightarrow 4OH^- + 2H_2$ | −0.83 |

(a) A salt bridge was used in a cell to measure electrode potential.

Explain the function of the salt bridge.

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**(2)**

(b) Use data from the table above to deduce the halide ion that is the weakest reducing agent.

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**(1)**

(c) Use data from the table to justify why sulfate ions should **not** be capable of oxidising bromide ions.

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**(1)**

(d) Use data from the table to calculate a value for the EMF of a hydrogen–oxygen fuel cell operating under alkaline conditions.

EMF = .............................. V

**(1)**

(e) There are two ways to use hydrogen as a fuel for cars. One way is in a fuel cell to power an electric motor, the other is as a fuel in an internal combustion engine.

Suggest the major advantage of using the fuel cell.

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**(1)**
**(Total 6 marks)**

**4** The following cell has an EMF of +0.46 V.

$$Cu \mid Cu^{2+} \parallel Ag^{+} \mid Ag$$

Which statement is correct about the operation of the cell?

**A** Metallic copper is oxidised by $Ag^+$ ions. ⬭

**B** The silver electrode has a negative polarity. ⬭

**C** The silver electrode gradually dissolves to form $Ag^+$ ions. ⬭

**D** Electrons flow from the silver electrode to the copper electrode via an external circuit. ⬭

**(Total 1 mark)**

**5** The table contains some standard electrode potential data.

| **Electrode half-equation** | ***E*⦵ / V** |
|---|---|
| $F_2 + 2e^- \longrightarrow 2F^-$ | +2.87 |
| $Au^+ + e^- \longrightarrow Au$ | +1.68 |
| $2HOCl + 2H^+ + 2e^- \longrightarrow Cl_2 + 2H_2O$ | +1.64 |
| $Cl_2 + 2e^- \longrightarrow 2Cl^-$ | +1.36 |
| $O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$ | +1.23 |
| $Ag^+ + e^- \longrightarrow Ag$ | +0.80 |
| $Fe^{3+} + e^- \longrightarrow Fe^{2+}$ | +0.77 |
| $2H^+ + 2e^- \longrightarrow H_2$ | 0.00 |
| $Fe^{2+} + 2e^- \longrightarrow Fe$ | −0.44 |

(a) In terms of electrons, explain the meaning of the term **oxidising agent**.

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**(1)**

(b) Identify the weakest oxidising agent in the table.
Explain your choice.

Weakest oxidising agent ..........................................................................................

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

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**(2)**

(c) Write the conventional representation of the cell used to measure the standard electrode potential for the $Ag^{+}$ / Ag electrode.

State the conditions necessary when measuring this value.

Conventional representation ........................................................................

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Conditions ...........................................................................................

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**(4)**

(d) Use data from the table to explain, in terms of redox, what happens when a soluble gold(I) compound containing $Au^{+}$ ions is added to water.

State what you would observe.

Write an equation for the reaction that occurs.

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

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Observation ..........................................................................................

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Equation ...............................................................................................

**(4)**

The table is repeated below to help you answer these questions.

| **Electrode half-equation** | ***E*$^{\ominus}$ / V** |
|---|---|
| $F_2 + 2e^- \longrightarrow 2F^-$ | +2.87 |
| $Au^+ + e^- \longrightarrow Au$ | +1.68 |
| $2HOCl + 2H^+ + 2e^- \longrightarrow Cl_2 + 2H_2O$ | +1.64 |
| $Cl_2 + 2e^- \longrightarrow 2Cl^-$ | +1.36 |
| $O_2 + 4H^+ + 4e^- \longrightarrow 2H_2O$ | +1.23 |
| $Ag^+ + e^- \longrightarrow Ag$ | +0.80 |
| $Fe^{3+} + e^- \longrightarrow Fe^{2+}$ | +0.77 |
| $2H^+ + 2e^- \longrightarrow H_2$ | 0.00 |
| $Fe^{2+} + 2e^- \longrightarrow Fe$ | −0.44 |

(e) A cell is made by connecting $Fe^{2+}$ / Fe and $Ag^+$ / Ag electrodes with a salt bridge.

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

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Answer ......................................................................

**(1)**

(ii) Suggest why potassium chloride would **not** be suitable for use in the salt bridge of this cell.

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**(1)**

(f) Use data from the table to explain what happens when a solution of iron(II) chloride is exposed to the air.

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**(2)**
**(Total 15 marks)**

**6** A biocide is a chemical that kills bacteria. A biocide is added to prevent the growth of bacteria in the water used in vases of flowers. Household bleach contains aqueous chlorine and can be used as the biocide. The concentration of chlorine in vase water decreases with time. It was decided to investigate the rate of this decrease.

The following experimental method was used to determine the concentration of chlorine in vase water at different times.

- A sample of vase water was taken.
- An excess of potassium iodide solution was added to the sample.
- The chlorine in the sample oxidised the $I^-$ ions to $I_2$
- The iodine was titrated with sodium thiosulfate ($Na_2S_2O_3$) solution.
- These steps were repeated using further samples taken from the vase water at hourly intervals.

(a) Suggest **two** reasons why the concentration of chlorine in the vase water decreases with time.

Reason 1 ..........................................................................................................

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Reason 2 ..........................................................................................................

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**(2)**

(b) Suggest why this sampling technique has no effect on the rate at which the concentration of chlorine in the vase water decreases.

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**(1)**

(c) Why was it important to use an **excess** of potassium iodide solution?

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**(1)**

(d) Use the following standard electrode potential data to explain why $I_2$ oxidises $S_2O_3^{2-}$ under standard conditions.

$$\frac{1}{2}I_2 + e^- \longrightarrow I^- \qquad E^\ominus = +0.54\ V$$

$$\frac{1}{2}S_4O_6^{2-} + e^- \longrightarrow S_2O_3^{2-} \qquad E^\ominus = +0.09\ V$$

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**(1)**

(e) Deduce an ionic equation for the reaction between $I_2$ and $S_2O_3^{2-}$

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**(1)**

**(Total 6 marks)**

**7** The table below shows some standard electrode potential data.

| | $\boldsymbol{E^\ominus}$ **/ 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 the table.

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**(2)**

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

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(c) Identify the weakest reducing agent from the species in the table.

Explain how you deduced your answer.

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.

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

**(1)**

(ii) Explain your answer to part (i) in terms of the change in the oxidation state of this atom.

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

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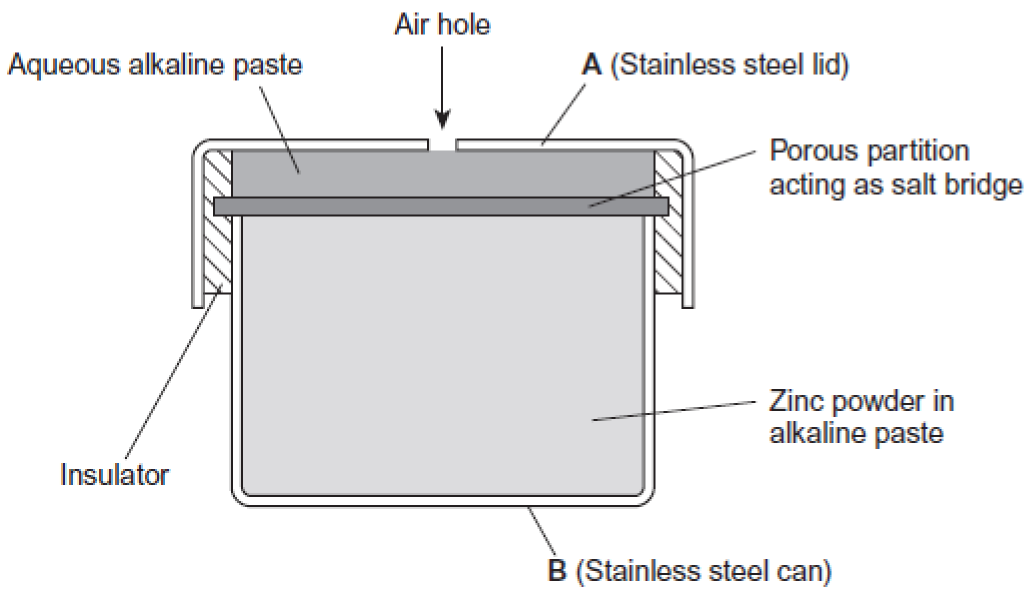
(f) The table is repeated to help you answer this question.

| | $E^{\ominus}$ / 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 the table 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 the following figure.




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

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

**(1)**

(ii) Use half-equations from the table to construct an overall equation for the cell reaction.

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**(1)**

(iii) Identify which of **A** or **B**, in the figure, is the positive electrode. Give a reason for your answer.

Positive electrode ..........................................................................

Reason ...........................................................................................

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**(2)**

(iv) Suggest **one** reason, other than cost, why this type of cell is **not** recharged.

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**(1)**
**(Total 14 marks)**

**8** Hydrogen–oxygen fuel cells are used to provide electrical energy for electric motors in vehicles.

(a) In a hydrogen–oxygen fuel cell, a current is generated that can be used to drive an electric motor.

(i) Deduce half-equations for the electrode reactions in a hydrogen–oxygen fuel cell.

Half-equation 1 ............................................................................

Half-equation 2 ............................................................................

**(2)**

(ii) Use these half-equations to explain how an electric current can be generated.

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**(2)**

(b) Explain why a fuel cell does **not** need to be recharged.

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**(1)**

(c) To provide energy for a vehicle, hydrogen can be used either in a fuel cell or in an internal combustion engine.

Suggest the main advantage of using hydrogen in a fuel cell rather than in an internal combustion engine.

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**(1)**

(d) Identify **one** major hazard associated with the use of a hydrogen–oxygen fuel cell in a vehicle.

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**(1)**
**(Total 7 marks)**