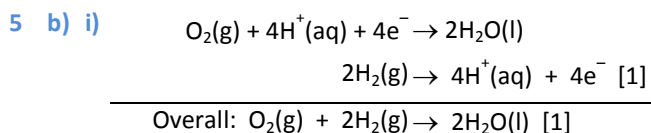


Pages 113–115

- 1 a) i) Nitrogen oxidised from -3 to 0 . [1] Chlorine reduced from 0 to -1 . [1]
 ii) Copper oxidised from $+1$ to $+2$ [1] and reduced from $+1$ to 0 . [1]
 iii) Oxygen oxidised from -2 to 0 . Nitrogen reduced from $+5$ to $+3$. [1]
- b) i) $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 14\text{H}^+(\text{aq}) + 6\text{e}^- \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 7\text{H}_2\text{O}(\text{l})$ [1]
 $\text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^+(\text{aq}) + 2\text{e}^-$ [1]
 ii) $\text{Cr}_2\text{O}_7^{2-}(\text{aq}) + 8\text{H}^+(\text{aq}) + 3\text{SO}_3^{2-}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 4\text{H}_2\text{O}(\text{l}) + 3\text{SO}_4^{2-}(\text{aq})$ [1]
 iii) The yellow/orange dichromate solution [1] turns green as chromium(III) ions form. [1]
 iv) 3 [1]
- 2 a) i) $\text{Cl}_2(\text{aq}) + 2\text{I}^-(\text{aq}) \rightarrow 2\text{Cl}^-(\text{aq}) + \text{I}_2(\text{aq})$ [1]
 ii) $2\text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{I}_2(\text{aq}) \rightarrow \text{S}_4\text{O}_6^{2-}(\text{aq}) + 2\text{I}^-(\text{aq})$ [2]
- b) Starch [1]
- c) On mixing the swimming pool water with excess potassium iodide the colourless solutions react to give a dark yellow-brown solution. [1] During the titration the colour of the solution becomes paler. Eventually it turns pale yellow. [1] On adding starch the solution turns blue-black. At the end-point the black colour disappears to give a colourless solution. [1]
- d) Amount of sodium thiosulfate added to reach the end-point
 $= 0.0194 \text{ dm}^3 \times 0.0050 \text{ mol dm}^{-3}$
 $= 0.000\ 097 \text{ mol}$ [1]
 This reacted with the iodine displaced by the chlorine in 20 cm^3 of the water. [1]
 From the equations $2\text{S}_2\text{O}_3^{2-}(\text{aq}) \equiv \text{Cl}_2(\text{aq})$, so the amount of chlorine in the 20 cm^3 water
 $= 0.5 \times 0.000\ 097 \text{ mol}$ [1]
 So the concentration of chlorine in the water
 $= \frac{0.5 \times 0.000\ 097 \text{ mol}}{0.020 \text{ dm}^3} = 0.002\ 42 \text{ mol dm}^{-3}$ [1]
- 3 a) $\text{Fe}^{3+}(\text{aq}) + \text{e}^- \rightarrow \text{Fe}^{2+}(\text{aq})$ [1] reduction
 $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$ [1] oxidation [1]
- b) Fe is reduced from $+3$ to $+2$ [1]
 Cu is oxidised from 0 to $+2$ [1]
- c) $\text{Cu}(\text{s}) | \text{Cu}^{2+}(\text{aq}) :: \text{Fe}^{3+}(\text{aq}), \text{Fe}^{2+}(\text{aq}) | \text{Pt}(\text{s})$ [2]
 Cell e.m.f. $= E_{\text{cell}}^{\ominus} = +0.77 - 0.34 = +0.43 \text{ V}$ [1]
- 4 a) Using filter paper soaked in saturated $\text{KNO}_3(\text{aq})$ or a U-tube of agar gel containing saturated $\text{KNO}_3(\text{aq})$. [1]
- b) i) $\text{Cu}(\text{s}) \rightarrow \text{Cu}^{2+}(\text{aq}) + 2\text{e}^-$ [1]
 ii) $\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$ [1]
- c) $2\text{Ag}^+(\text{aq}) + \text{Cu}(\text{s}) \rightarrow 2\text{Ag}(\text{s}) + \text{Cu}^{2+}(\text{aq})$ [1]

- d) $E_{\text{cell}}^{\ominus} = -0.34 \text{ V} + 0.80 \text{ V} [1] = +0.46 \text{ V} [1]$
- e) At the silver electrode [1] where Ag^+ ions gain electrons. Silver changes oxidation number from +1 to zero. [1]
- f) Any two of the following points for [2]:
- temperature was not standard
 - concentration(s) were not standard
 - conditions were not standard.



ii) From the hydrogen half-cell [1]

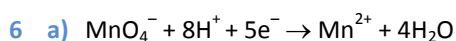
- b) Fuel cells differ from storage cells in that:
- instead of recharging the cells, the reactants are fed into the cell continuously [1]
 - the products are drawn off continuously [1]
 - the voltage in the cell remains constant so long as fuel and oxygen are supplied. [1]
- c) Fuel cells produce electric power from the chemical free energy of the fuel and oxygen [1] much more efficiently. [1]

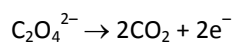
Fuel cells do not pollute the atmosphere with pollutants from combustion [1] such as oxides of nitrogen and particulates; and they do not pollute rivers with warm water. [1]

- d) This question assesses a student's ability to show a coherent and logically structured answer with linkages and fully sustained line of reasoning. Assess the quality of the answer taking into account both the key points made (*up to 4 marks*) and the logic and coherence of the discussion (*up to 2 marks*).

Points to make in the answer:

- Advantage that a fuel cell is more efficient as a power source than a combustion engine.
- Advantage that there is no local air pollution (no CO_2 , no particulates, no oxides of nitrogen) because the only product is steam.
- Advantage, in the future, that it may become practicable to make the hydrogen using replenishable energy sources.
- Disadvantage that hydrogen is currently made from fossil fuels.
- Disadvantage there are many technical difficulties associated with the handling, transport and storage of hydrogen for use in vehicles.
- Disadvantage that hydrogen fuel cells are more costly than combustion engines and have a shorter life.
- Disadvantage that the public may not be willing to accept the use of hydrogen as a fuel because of perceptions of its flammability and the danger of explosions.





$$\Rightarrow 5 \text{ mol C}_2\text{O}_4^{2-} \equiv 2 \text{ mol MnO}_4^- [1]$$

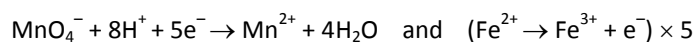
$$\text{Concentration of C}_2\text{O}_4^{2-} = \frac{7.445 \text{ g dm}^{-3}}{134 \text{ g mol}^{-1}} = 0.05556 \text{ mol dm}^{-3} [1]$$

$$\Rightarrow \text{Amount of C}_2\text{O}_4^{2-} \text{ reacting} = \frac{25.00}{1000} \text{ dm}^3 \times 0.05556 \text{ mol dm}^{-3} = 0.001389 \text{ mol dm}^{-3} [1]$$

$$\text{Amount of MnO}_4^- \text{ in } 28.85 \text{ cm}^3 \text{ solution} = \frac{2 \times 0.001389}{5} \text{ mol} [1]$$

$$\Rightarrow \text{Concentration of the KMnO}_4 \text{ solution} = \frac{2 \times 0.001389}{5 \times 0.02885} = 0.01926 \text{ mol dm}^{-3} [1]$$

b) The half-equations for the reaction are:



$$\Rightarrow 5 \text{ mol Fe}^{2+} \equiv 1 \text{ mol MnO}_4^- [1]$$

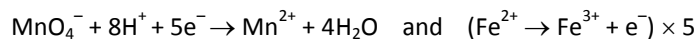
$$\text{Amount of MnO}_4^- \text{ reacting} = \frac{26.75}{1000} \text{ dm}^3 \times 0.0200 \text{ mol dm}^{-3} = \frac{0.535}{1000} \text{ mol} [1]$$

$$\Rightarrow \text{Amount of Fe}^{2+} \text{ reacting} = \frac{0.535}{1000} \times 5 \text{ mol} = \frac{2.675}{1000} \text{ mol} [1]$$

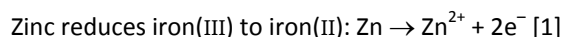
$$\text{So the mass of iron in the ore} = \frac{2.675}{1000} \text{ mol} \times 55.8 \text{ g mol}^{-1} = 0.15 \text{ g} [1]$$

$$\Rightarrow \text{Percentage iron in the iron ore} = \frac{0.15 \text{ g}}{1.34 \text{ g}} \times 100 = 11.2\% [1]$$

c) The half-equations for the reaction are:



$$\Rightarrow 5 \text{ mol Fe}^{2+} \equiv 1 \text{ mol MnO}_4^-$$



In the first titration:

$$\text{Amount of MnO}_4^- \text{ reacting} = \frac{18.00}{1000} \text{ dm}^3 \times 0.0200 \text{ mol dm}^{-3} = \frac{0.360}{1000} \text{ mol}$$

$$\Rightarrow \text{Amount of Fe}^{2+} \text{ in } 25.00 \text{ cm}^3 \text{ of the solution} = \frac{0.360}{1000} \times 5 \text{ mol}$$

$$\text{Concentration of Fe}^{2+} \text{ in the solution} = \frac{0.360}{1000 \times 0.0250} \times 5 \text{ mol} = 0.0720 \text{ mol dm}^{-3} [1]$$

In the second titration, after reduction of iron(III) to iron(II):

$$\text{Amount of MnO}_4^- \text{ reacting} = \frac{22.50}{1000} \text{ dm}^3 \times 0.0200 \text{ mol dm}^{-3} = \frac{0.450}{1000} \text{ mol}$$

$$\Rightarrow \text{Amount of Fe}^{2+} \text{ and Fe}^{3+} \text{ in } 25.00 \text{ cm}^3 \text{ of the solution} = \frac{0.450}{1000} \times 5 \text{ mol} [1]$$

$$\text{Concentration of Fe}^{2+} \text{ and Fe}^{3+} \text{ in the solution} = \frac{0.450}{1000 \times 0.0250} \times 5 \text{ mol} = 0.0900 \text{ mol dm}^{-3}$$

$$\text{Total concentration of Fe}^{2+} \text{ and Fe}^{3+} = 0.0900 \text{ mol dm}^{-3} [1]$$

Concentration of $\text{Fe}^{3+} = (0.0900 - 0.0720) \text{ mol dm}^{-3} = 0.0180 \text{ mol dm}^{-3}$ [1]

7 a) $1 \text{ mol S}_2\text{O}_3^{2-} \equiv 1 \text{ mol electrons}$ [1]

Amount of $\text{S}_2\text{O}_3^{2-}(\text{aq})$ needed in the titration

$$= 0.0180 \text{ dm}^3 \times 0.0760 \text{ mol dm}^{-3} = 0.00137 \text{ mol}$$
 [1]

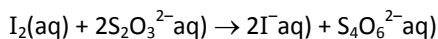
Amount of copper ions in the pipette volume of solution

$$= \frac{3.405 \text{ g}}{249.5 \text{ g mol}^{-1}} \times 0.1 = 0.00136 \text{ mol}$$
 [1]

So there is a change of 1 mol of electrons per mol of copper(II) ions when they oxidise iodide ions. [1]

Copper(II) is reduced to copper(I). [1]

b) $2\text{Cu}^{2+}(\text{aq}) + 4\text{I}^{-}(\text{aq}) \rightarrow 2\text{CuI}(\text{s}) + \text{I}_2(\text{aq})$



$2 \text{ mol S}_2\text{O}_3^{2-} \equiv 1 \text{ mol I}_2 \equiv 2 \text{ mol Cu}^{2+}$ [1]

$$\text{Amount of S}_2\text{O}_3^{2-} \text{ reacting} = \frac{22.5}{1000} \text{ dm}^3 \times 0.14 \text{ mol dm}^{-3}$$

$$= \frac{3.15}{1000} \text{ mol}$$
 [1]

$$\text{So the amount of Cu}^{2+} \text{ reacting} = \frac{3.15}{1000} \text{ mol}$$
 [1]

$$\Rightarrow \text{Mass of copper reacting} = \frac{3.15}{1000} \text{ mol} \times 63.5 \text{ g mol}^{-1} = 0.20 \text{ g}$$
 [1]

$$\text{Percentage of copper in the alloy} = \frac{0.20 \text{ g}}{0.275 \text{ g}} \times 100 = 72.7\%$$
 [1]

8 a) i) $\text{Fe}(\text{s}) + \text{I}_2(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + 2\text{I}^{-}(\text{aq})$ or $\text{Cd}(\text{s}) + \text{I}_2(\text{aq}) \rightarrow \text{Cd}^{2+}(\text{aq}) + 2\text{I}^{-}(\text{aq})$ [1]

because each of these reactions would have a large positive $E_{\text{cell}}^{\ominus}$. [1]

ii) $\text{Fe}(\text{s}) + \text{Cd}^{2+}(\text{aq}) \rightarrow \text{Fe}^{2+}(\text{aq}) + \text{Cd}(\text{s})$ [1]

because this produces a very small positive $E_{\text{cell}}^{\ominus}$. [1]

b) i) Labels on the diagram should point to

- Fe(s) electrode, in $\text{Fe}^{2+}(\text{aq})$ solution, [1] 1 mol dm^{-3} [1]
- Pt electrode [1], in a solution containing $\text{I}_2(\text{aq})$ and $\text{I}^{-}(\text{aq})$, both at 1 mol dm^{-3} [1]
- salt bridge. [1]

ii) Arrow from Fe electrode to Pt electrode and labelled electron flow. [1]

$$\text{iii) } E_{\text{cell}}^{\ominus} = +0.54 - (-0.44) = +0.98 \text{ V}$$

[1] for magnitude, [1] for sign to correspond to that of the right-hand electrode

iv) There would be a greater tendency for Fe(s) to form $\text{Fe}^{2+}(\text{aq})$ [1], which would make the $\text{Fe}^{2+}|\text{Fe}$ electrode more negative [1], thus increasing the magnitude of the cell e.m.f. [1]

- c) i) +0.34 V [1]
- ii) When $\log[\text{Cu}^{2+}(\text{aq})] = 0$, $[\text{Cu}^{2+}(\text{aq})] = 1.0 \text{ mol dm}^{-3}$ [1] and the value of E is the standard electrode potential. [1]
- iii) There is a linear relationship between the electrode potential and $\log[\text{Cu}^{2+}(\text{aq})]$. [1]
The equation for a linear equation takes the general form: $y = mx + c$
In this instance this can be written as: $E = \text{constant} \times \log [\text{Cu}^{2+}(\text{aq})] + E^\ominus$ [2]