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) $Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$ [1]

$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$ [1]

ii) $Cr_2O_7^{2-}(aq) + 8H^+(aq) + 3SO_3^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 4H_2O(l) + 3SO_4^{2-}(aq)$ [1]

iii) The yellow/orange dichromate solution [1] turns green as chromium(III) ions form. [1]

iv) 3 [1]

2 a) i) $Cl_2(aq) + 2I^-(aq) \rightarrow 2Cl^-(aq) + I_2(aq)$ [1]

ii) $2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow S_4O_6^{2-}(aq) + 2I^-(aq)$ [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

$= 0.0194\ dm^3 \times 0.0050\ mol\ dm^{-3}$

= 0.000 097 mol [1]

This reacted with the iodine displaced by the chlorine in 20 cm^3 of the water. [1]

From the equations $2S_2O_3^{2-}(aq) \equiv Cl_2(aq)$, so the amount of chlorine in the 20 cm^3 water

$= 0.5 \times 0.000\,097$ mol [1]

So the concentration of chlorine in the water

$$= \frac{0.5 \times 0.000\,097\ mol}{0.020\ dm^3} = 0.002\,42\ mol\ dm^{-3}$$ [1]

3 a) $Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ [1] reduction

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ [1] oxidation [1]

b) Fe is reduced from +3 to +2 [1]

Cu is oxidised from 0 to +2 [1]

c) $Cu(s)|Cu^{2+}(aq)⁞⁞Fe^{3+}(aq), Fe^{2+}(aq)|Pt(s)$ [2]

Cell e.m.f. = $E_{cell}^{\ominus} = +0.77 - 0.34 = +0.43$ V [1]

4 a) Using filter paper soaked in saturated $KNO_3(aq)$ or a U-tube of agar gel containing saturated $KNO_3(aq)$. [1]

b) i) $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ [1]

ii) $Ag^+(aq) + e^- \rightarrow Ag(s)$ [1]

c) $2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$ [1]

d) $E_{cell}^{\ominus}$ = – 0.34 V + 0.80 V [1] = +0.46 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.

5 b) i)

$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$
$$2H_2(g) \rightarrow 4H^+(aq) + 4e^- \quad [1]$$

Overall: $O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$ [1]

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.

6 a) $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$

$\Rightarrow$ 5 mol $C_2O_4^{2-} \equiv$ 2 mol MnO_4^- [1]

Concentration of $C_2O_4^{2-} = \dfrac{7.445 \text{ g dm}^{-3}}{134 \text{ g mol}^{-1}}$ = 0.055 56 mol dm^{-3} [1]

$\Rightarrow$ Amount of $C_2O_4^{2-}$ reacting = $\dfrac{25.00}{1000}$ dm^3 × 0.055 56 mol dm^{-3} = 0.001 389 mol dm^{-3} [1]

Amount of MnO_4^- in 28.85 cm^3 solution = $\dfrac{2 \times 0.001\,389}{5}$ mol [1]

$\Rightarrow$ Concentration of the $KMnO_4$ solution = $\dfrac{2 \times 0.001\,389}{5 \times 0.028\,85}$ = 0.019 26 mol dm^{-3} [1]

b) The half-equations for the reaction are:

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and $(Fe^{2+} \rightarrow Fe^{3+} + e^-) \times 5$

$\Rightarrow$ 5 mol $Fe^{2+} \equiv$ 1 mol MnO_4^- [1]

Amount of MnO_4^- reacting = $\dfrac{26.75}{1000}$ dm^3 × 0.0200 mol dm^{-3} = $\dfrac{0.535}{1000}$ mol [1]

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

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

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

c) The half-equations for the reaction are:

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$ and $(Fe^{2+} \rightarrow Fe^{3+} + e^-) \times 5$

$\Rightarrow$ 5 mol $Fe^{2+} \equiv$ 1 mol MnO_4^-

Zinc reduces iron(III) to iron(II): $Zn \rightarrow Zn^{2+} + 2e^-$ [1]

In the first titration:

Amount of MnO_4^- reacting = $\dfrac{18.00}{1000}$ dm^3 × 0.0200 mol dm^{-3} = $\dfrac{0.360}{1000}$ mol

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

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

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

Amount of MnO_4^- reacting = $\dfrac{22.50}{1000}$ dm^3 × 0.0200 mol dm^{-3} = $\dfrac{0.450}{1000}$ mol

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

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

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

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

7 a) 1 mol $S_2O_3^{2-}$ ≡ 1 mol electrons [1]

Amount of $S_2O_3^{2-}$(aq) needed in the titration

= 0.0180 dm^3 × 0.0760 mol dm^{-3} = 0.001 37 mol [1]

Amount of copper ions in the pipette volume of solution

$$= \frac{3.405\ \text{g}}{249.5\ \text{g mol}^{-1}} \times 0.1 = 0.001\ 36\ \text{mol [1]}$$

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) $2Cu^{2+}(aq) + 4I^{-}(aq) \rightarrow 2CuI(s) + I_2(aq)$

$I_2(aq) + 2S_2O_3^{2-}aq) \rightarrow 2I^{-}aq) + S_4O_6^{2-}aq)$

2 mol $S_2O_3^{2-}$ ≡ 1 mol I_2 ≡ 2 mol Cu^{2+} [1]

Amount of $S_2O_3^{2-}$ reacting = $\frac{22.5}{1000}$ dm^3 × 0.14 mol dm^{-3}

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

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

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

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

8 a) i) $Fe(s) + I_2(aq) \rightarrow Fe^{2+}(aq) + 2I^{-}(aq)$ or $Cd(s) + I_2(aq) \rightarrow Cd^{2+}(aq) + 2I^{-}(aq)$ [1]

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

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

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

b) i) Labels on the diagram should point to

- Fe(s) electrode, in Fe^{2+}(aq) solution, [1] 1 mol dm^{-3} [1]
- Pt electrode [1], in a solution containing I_2(aq) and I^{-}(aq), both at 1 mol dm^{-3} [1]
- salt bridge. [1]

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

iii) $E_{cell}^{\ominus}$ = + 0.54 – (– 0.44) = + 0.98 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 Fe^{2+}(aq) [1], which would make the Fe^{2+}|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[Cu^{2+}(aq)] = 0, [Cu^{2+}(aq)] = 1.0 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[Cu^{2+}(aq)]. [1]

The equation for a linear equation takes the general form: $y = mx + c$

In this instance this can be written as: $E = \text{constant} \times \log [Cu^{2+}(aq)] + E^{\ominus}$ [2]