Answers to exam-style questions

Topic 9

1 A

2 C

3 C

4 C

5 C

6 C

7 C

8 B

9 C

10 D

11 a Oxidation is the loss of electrons. **[1]**

b i +7; O is the more electronegative element and has an oxidation number of −2; 4 × −2 = −8, so the oxidation number of Mn must be +7 in order to cancel out all but one of the negative charges and leave an overall charge of −1. **[1]**

ii Fe^{2+}, because it has been oxidised. **[1]**

iii Separate into its two half-equations:

$MnO_4^-(aq) + H^+(aq) \rightarrow Mn^{2+}(aq) + H_2O(l)$

$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq)$

Each is balanced separately by following the procedure given on page **376** or they can be looked up in the IB Chemistry data booklet:

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

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

The number of electrons is balanced by multiplying the bottom equation by 5 and then the equations can be recombined:

$MnO_4^-(aq) + 8H^+(aq) + 5Fe^{2+}(aq) \rightarrow Mn^{2+}(aq) + 4H_2O(l) + 5Fe^{3+}(aq)$ **[2]**

c i Number of moles of $KMnO_4$ = $(21.50/1000) \times 5.00 \times 10^{-3} = 1.08 \times 10^{-4}$ mol **[1]**

ii Number of moles of Fe^{2+} is five times the number of moles of MnO_4^- from the balanced equation for the reaction, so no. of moles of $Fe^{2+} = 5 \times 1.08 \times 10^{-4} = 5.38 \times 10^{-4}$ mol **[1]**

iii Number of moles of Fe^{2+} in 250.0 cm^3 solution = $10 \times 5.38 \times 10^{-4} = 5.38 \times 10^{-3}$ mol

This is the number of moles of iron in five iron tablets.

The mass of iron in 5 iron tablets is:

$5.38 \times 10^{-3} \times 55.85 = 3.00 \times 10^{-1}$ g

So the mass of iron in one iron tablet is:

$3.00 \times 10^{-1}/5 = 6.00 \times 10^{-2}$ g **[3]**

iv 6.00×10^{-2} g is 60.0 mg of iron per tablet; the manufacturer's claim seems to be wrong. However, some solution would be lost in the filtering process and there could have been other systematic errors in the experiment. **[2]**

12 a A species that donates electrons (to another substance) and is itself oxidised. **[1]**

b i $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ **[2]**

ii Copper is less reactive than lead; it is unable to displace lead ions from solution. **[1]**

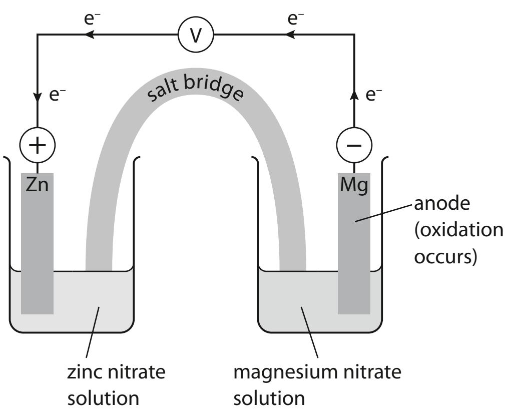
iii magnesium (most reactive) > zinc > lead > copper (least reactive)

Zinc can displace copper and lead from solution but cannot displace magnesium, so zinc must be more reactive than lead and copper but less reactive than magnesium. Copper cannot displace lead, therefore must be less reactive than lead. **[1]**

iv Magnesium is the strongest reducing agent; it will displace the other metals from solution. This means that magnesium will reduce Zn^{2+} to Zn, Pb^{2+} to Pb and Cu^{2+} to Cu. Zinc is able to reduce only lead and copper ions, lead can reduce only copper ions and copper cannot reduce the ions of any of these metals. **[3]**

v $Mg(s) + Pb^{2+}(aq) \rightarrow Mg^{2+}(aq) + Pb(s)$ **[2]**

c i



[3]

ii $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

The zinc ion gains electrons; this is reduction. **[2]**

13 a In solid sodium chloride, the ions are fixed in position in the lattice structure (they can only vibrate); when the sodium chloride is melted the ions are free to move around. **[2]**

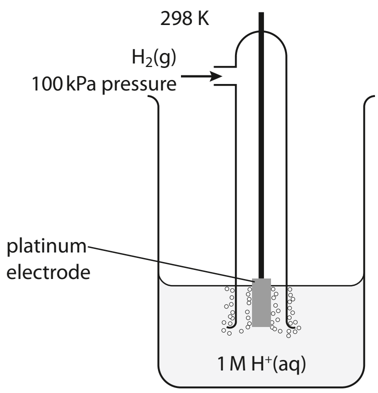
b At the anode, the product is chlorine:

$2Cl^- \rightarrow Cl_2 + 2e^-$

At the cathode, the product is sodium:

$Na^+ + e^- \rightarrow Na$ **[3]**

14 a



[5]

b i $Mn^{2+}(aq) + 2e^- \rightarrow Mn(s)$ $-1.18\ V$

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $+0.77\ V$

The more negative electrode potential is reversed: $1.18 + 0.77 = +1.95\ V$.

The standard cell potential is $+1.95\ V$. **[1]**

ii $2Fe^{3+}(aq) + Mn(s) \rightarrow 2Fe^{2+}(aq) + Mn^{2+}(aq)$ **[2]**

iii The anode is the electrode at which oxidation occurs; Mn is oxidised, so this is the anode. **[2]**

iv From the Mn electrode to the platinum electrode (Fe^{2+}/Fe^{3+} half-cell). **[1]**

v The salt bridge completes the circuit; allows ions to flow into/out of the half-cells and prevents a build-up of charge. **[2]**

vi $\Delta G^{\ominus} = -nFE^{\ominus} = -2 \times 96\,500 \times 1.95$

$= -376\,000\ J\ mol^{-1}$

$= -376\ kJ\ mol^{-1}$

$\Delta G^{\ominus}$ is negative; the reaction is spontaneous. **[3]**

15 a MnO_4^- is the strongest oxidising agent. It has the most positive electrode potential and therefore the strongest tendency to gain electrons/be reduced. **[2]**

b From the standard electrode potentials it can be seen that $Cr_2O_7^{2-}$ will oxidise Br^- ions but not Cl^- ions. $Cr_2O_7^{2-}$ has a more positive standard electrode potential than Br_2 and a less positive electrode potential than Cl_2. It is thus a stronger oxidising agent than Br_2 but weaker than Cl_2. This could also be explained by working out the cell potentials. The cell potential for the reaction with Br^- ions is positive (+0.24 V), but the cell potential for the reaction with Cl^-, is negative (-0.03 V). A positive cell potential indicates a spontaneous reaction.

To work out the overall equation for the reaction between $Cr_2O_7^{2-}$ and Br^- the half-equation for Br_2/Br^- must be reversed because the bromide ions are oxidised:

$2Br^-(aq) \rightarrow Br_2(l) + 2e^-$

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

The top equation must be multiplied by 3 so that the electrons balance:

$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6Br^-(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3Br_2(l)$ **[3]**

c $2Cl^-(aq) + Br_2(l) \rightarrow Cl_2(g) + 2Br^-(aq)$

Using the standard electrode potentials we can work out that the cell potential for this reaction is: $-1.36 + 1.09 = -0.27\ V$.

The cell potential is negative, so the reaction will not be spontaneous and the sign of $\Delta G^{\ominus}$ will be positive. **[2]**

d i HCOOH: +2 HCHO: 0

The oxidation number of C decreases, so HCOOH is reduced. **[3]**

ii There is the same number of C atoms on both sides. The H atoms must then be balanced by adding another H^+ to the left hand side so that there are 2 H^+ on that side. The O atoms are balanced. The total charge on the left hand side is now 1+ but there is no charge on the right hand side, therefore an extra electron must be added to the left hand side.

$HCOOH(aq) + 2H^+(aq) + 2e^- \rightarrow HCHO(aq) + H_2O(l)$ **[1]**

iii MnO_4^- will oxidise methanal to methanoic acid (MnO_4^- has a more positive electrode potential).

$MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$

$HCHO(aq) + H_2O(l) \rightarrow HCOOH(aq) + 2H^+(aq) + 2e^-$

To balance the number of electrons, the top equation must be multiplied by 2 and the bottom equation by 5:

$2MnO_4^-(aq) + 16H^+(aq) + 5HCHO(aq) + 5H_2O(l) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 5HCOOH(aq) + 10H^+(aq)$

H^+ ions and H_2O molecules can then be cancelled from each side:

$2MnO_4^-(aq) + 6H^+(aq) + 5HCHO(aq) \rightarrow 2Mn^{2+}(aq) + 3H_2O(l) + 5HCOOH(aq)$ **[2]**

iv The reaction mixture may need to be heated as the reaction may have a high activation energy. **[1]**

16 **a** Because the solution is concentrated, chlorine will be formed at the anode, rather than oxygen.
At the anode the product is chlorine gas:
$2Cl^- \rightarrow Cl_2 + 2e^-$
At the cathode the product is hydrogen gas:
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$
Hydrogen is produced at the cathode rather than sodium because sodium is a reactive metal and so it is more difficult to reduce sodium ions to sodium than it is to reduce water to hydrogen. **[4]**

b **i** Oxygen gas is formed at the anode:
$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ **[2]**

ii The electrolyte will become less blue as the copper ions, which are responsible for the blue colour, are removed at the cathode.
The electrolyte will become more acidic as H^+ ions are produced at the anode. The electrolyte will become sulfuric acid because copper ions are removed from the solution and replaced by H^+ ions. **[2]**

c **i** Anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ **[2]**

ii The colour of the electrolyte will not change – it will remain the same shade of blue – as the concentration of copper ions in the solution will remain constant. Copper ions are produced at the anode at the same rate as they are used up at the cathode. **[2]**