Chapter 20 – Answers to end of chapter questions

1  a  Oxidation is the loss of electrons. Reduction is the gain of electrons.

   b  i  $\text{Mg} \rightarrow \text{Mg}^{2+} + 2e^-$  oxidation

   $\text{Cl}_2 + 2e^- \rightarrow 2\text{Cl}^-$  reduction

   ii  $\text{MnO}_2 + 4\text{H}^+ + 2e^- \rightarrow \text{Mn}^{2+} + 2\text{H}_2\text{O}$  reduction

   $2\text{Cl}^- \rightarrow \text{Cl}_2 + 2e^-$  oxidation

   iii  $\text{H}_2\text{O}_2 + 2\text{H}^+ + 2e^- \rightarrow 2\text{H}_2\text{O}$  reduction

   $2\text{I}^- \rightarrow \text{I}_2 + 2e^-$  oxidation

   iv  $2\text{Fe}^{3+} + 2e^- \rightarrow 2\text{Fe}^{2+}$  reduction

   $2\text{I}^- \rightarrow \text{I}_2 + 2e^-$  oxidation

c

2  a  The diagram should show the two half cells linked by a salt bridge and connected between their electrodes by a switch and/or voltmeter. The first half-cell should show a Pt electrode in an aqueous solution containing 1 mol dm$^{-3}$ H$_2$O$_2$ and 1 mol dm$^{-3}$ H$^+$ ions at 298 K. The second half-cell should show a Pt electrode in an aqueous solution containing 1 mol dm$^{-3}$ of both Sn$^{2+}$ and Sn$^{4+}$ ions at 298 K.

   b  Electrons flow in the external circuit from the Sn$^{2+}$/Sn$^{4+}$ half-cell towards the H$_2$O$_2$/H$^+$ half-cell.

   c  When the cell produces an electric current

   $\text{H}_2\text{O}_2(\text{aq}) + 2\text{H}^+(\text{aq}) + 2e^- \rightarrow 2\text{H}_2\text{O}(\text{l})$ and

   $\text{Sn}^{2+}(\text{aq}) \rightarrow \text{Sn}^{4+}(\text{aq}) + 2e^-$

   d  $E^\circ_{\text{cell}} = +1.77 \text{ V} - 0.15 \text{ V} = +1.62 \text{ V}$

   e  i  cell potential would increase

   ii  cell potential would decrease.
3  a  \( \text{Fe(s)} \rightarrow \text{Fe}^{2+} (aq) + 2e^- \)  
   \( \text{H}_2\text{O(l)} + \frac{1}{2}\text{O}_2(g) + 2e^- \rightarrow 2\text{OH}^- (aq) \)  
   Dissolved \( \text{O}_2 \) is reduced to \( \text{OH}^- \) ions. 
   Then \( \text{Fe}^{2+} (aq) + 2\text{OH}^- (aq) \rightarrow \text{Fe(OH)}_2(s) \)  
   \( \text{Fe(OH)}_2(s) \) is precipitated. 
   \( \text{Fe(OH)}_2(s) + \text{OH}^- (aq) \rightarrow \text{Fe(OH)}_3(s) + e^- \)  
   \( \text{Fe(OH)}_2 \) is oxidised to \( \text{Fe(OH)}_3 \) 
   \( \text{H}_2\text{O(l)} + \frac{1}{2}\text{O}_2(g) + 2e^- \rightarrow 2\text{OH}^- (aq) \)  
   More \( \text{O}_2 \) is reduced to \( \text{OH}^- \) 
   Finally, \( \text{Fe(OH)}_3 \) decomposes to form rust. 
   \( 2\text{Fe(OH)}_3(s) \rightarrow \text{Fe}_2\text{O}_3(s) + 3\text{H}_2\text{O(l)} \)

b  If the Zn coating is pierced, the Zn and Fe metals form a cell in aqueous solution and 
   the Zn will be more likely to go into solution as \( \text{Zn}^{2+} \) ions than \( \text{Fe} \) as \( \text{Fe}^{2+} \). 
   If the Sn coating is pierced, the Sn and Fe metals also form a cell in aqueous solution, 
   but in this case the \( \text{Fe} \) will go into solution as \( \text{Fe}^{2+} \) ions, so making rusting more likely. 

c  Sn is much less reactive than Zn, so is less likely to contaminate food. Zn is much 
   cheaper than Sn, so is used to galvanise buckets.

4  a  \( 2\text{H}^+(aq) + 2e^- \rightarrow \text{H}_2(g) \)

b  i  \( \text{Ce}^{4+}(aq) \)  
   ii  \( \text{Zn}(s) \)

c  \( \text{Fe}^{3+}(aq) \) and \( \text{Ce}^{4+}(aq) \) 
   \( 2\text{I}^-(aq) + 2\text{Fe}^{3+}(aq) \rightarrow \text{I}_2(aq) + 2\text{Fe}^{2+}(aq) \)
   \( 2\text{I}^-(aq) + 2\text{Ce}^{4+}(aq) \rightarrow \text{I}_2(aq) + 2\text{Ce}^{3+}(aq) \)

d  i  \( +0.77 + 0.44 = +1.21 \text{ V} \)
   ii  \( \text{Fe}^{3+}(aq) + e^- \rightarrow \text{Fe}^{2+}(aq) \) and \( \text{Fe}(s) \rightarrow \text{Fe}^{2+}(aq) + 2e^- \)
   iii  From the \( \text{Fe}/\text{Fe}^{2+} \) half-cell to the \( \text{Fe}^{3+}/\text{Fe}^{2+} \) half-cell

e  \( 2\text{Fe}^{3+}(aq) + \text{Fe}(s) \rightarrow 3\text{Fe}^{2+}(aq) \)

5  a  i  \( E = +0.08 + \frac{0.059}{1} \log[0.01] = -0.038 \)
   ii  \( E = +0.08 + \frac{0.059}{1} \log[0.05] = -0.032 \)

b  Even though the concentration in (ii) is 5 times greater than the concentration in (i), 
   the values of \( E \) are very similar. A big difference in concentration has made a small 
   difference in the value of \( E \)

6  a  \( E = +0.08 + \frac{0.059}{1} \log[0.01] = -0.038 \)
   b  \( E = +0.08 + \frac{0.059}{1} \log[0.05] = -0.032 \)

7  a  Cathode – lead; anode – bromine
   b  Cathode: \( \text{Pb}^{2+} + 2e^- \rightarrow \text{Pb} \)
      Anode: \( 2\text{Br}^- \rightarrow \text{Br}_2 + 2e^- \)
c \( (0.20 \text{ A} \times (30 \times 60) \text{ s}) = 360 \text{ C} \)

d i Number of moles Pb = \( \frac{360 \text{ C}}{193000 \text{ C/mol}} \) = 0.00187 mol

Mass of this amount of Pb = 0.00187 mol \times 207 \text{ g mol}^{-1} = 0.387 \text{ g}

ii Number of moles of Br\(_2\) = 0.00187 mol

Mass of this amount of Br\(_2\) = 0.00187 mol \times (2 \times 79.9) \text{ g mol}^{-1} = 0.299 \text{ g}

e Time = \( \frac{30 \text{ min \times 10 g}}{0.387 \text{ g}} \) = 775 min

8 a Anode (+) \( \text{2Cl}^- (l) \rightarrow \text{Cl}_2 (g) + 2e^- \)

Cathode (−) \( \text{2Na}^+(l) + 2e^- \rightarrow \text{2Na}(l) \)

b 96 500 C

c 1 kg Na = 1000 g Na = \( \frac{1000}{23} \) mol Na

\( \frac{1000}{23} \) mol. Na requires \( \frac{1000}{23} \times 96 500 \text{ C} \)

But \( Q \) (coulombs) = \( I \) (amps) \times \( t \) (secs)

\( : \quad \frac{1000}{23} \times 96 500 = 100 \times 1000 \times t \)

So \( t = \frac{965}{23} = 42.0 \text{ seconds} \)