7.4 $Cu^{2+}(aq)$ and $H^+(aq)$

7.5 Anode: $2H_2O(\ell) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
Cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

7.6 Anode: oxygen   Cathode: copper (metal)

7.7 $H^+(aq)$ and $SO_4^{2-}(aq)$

7.8 $H_2SO_4(aq)$ sulfuric acid

8.1 Not spontaneous
$E^0_{cell} < 0$/energy source needed in external circuit.

8.2 Anode
Ag metal electrode is oxidised at the anode to Ag^+ ions.

8.3 Decrease.

8.4 $Ag \rightarrow Ag^+ + e^-$

8.5 Negative
The watch that is plated, is always at the cathode, where Ag^+ ions are reduced to Ag metal.

8.6 No
Only a silver nitrate solution can be used.
Silver sulphate is an insoluble salt and can therefore not dissolve in water to form an electrolyte.

8.7 Ag^+ cations

8.8 The rate at which the Ag metal is oxidised to Ag^+ at the anode, is equal to the rate at which the Ag^+ is reduced to Ag metal at the cathode. Therefore, the number of Ag^+ ions in the solution remains constant, in other words, the Ag^+ ion concentration remains constant.

Exercise 29: Page 389

1.1 Anode: positive   Cathode: negative

1.2 Anode: $2C\ell^- \rightarrow C\ell_2 + 2e^-$ Cathode: $Mg^{2+} + 2e^- \rightarrow Mg$

1.3 Melted magnesium chloride

2.1 A non-spontaneous electrochemical process during which electrical current flows through an electrolyte between two electrodes by the movement of ions, and that leads to the formation of new products through oxidation and reduction.

2.2

direct current power source

anode +
impure
copper

cathode -
pure copper

$CuSO_4$(aq) electrolyte

2.3 $Cu \rightarrow Cu^{2+} + 2e^-$

2.4 $Cu^{2+} + 2e^- \rightarrow Cu$
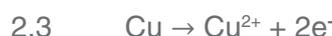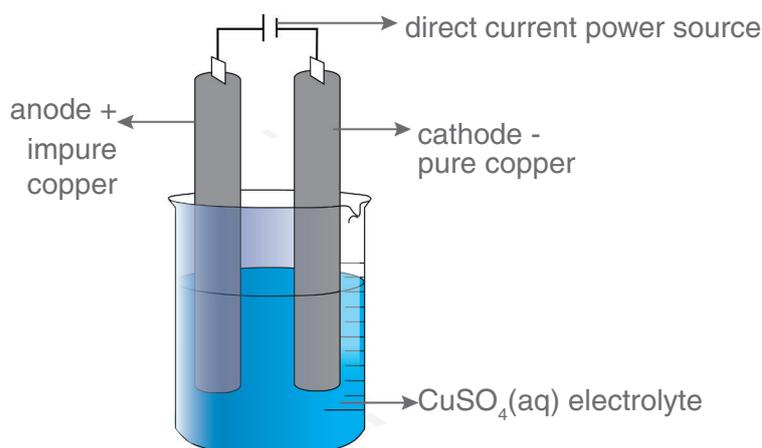
2.5 The rate at which the Cu^{2+}(aq) forms at the anode due to the oxidation of Cu(s), is equal to the rate at which Cu^{2+}(aq) is reduced to Cu(s) at the cathode. The number of Cu^{2+}(aq) ions in the solution therefore remains constant, and therefore the Cu^{2+}(aq) concentration remains constant.

2.6 According to the Table of Standard Reduction Potentials, Cu is a stronger reducing agent than both Pt and Ag. Therefore, Cu will be oxidised and not Pt or Ag.

3.1 Silver nitrate solution, $AgNO_3$(aq)

3.2 Silver, Ag

3.3 At the anode Ag(s) is oxidised to form the Ag^+(aq) ions that move to the cathode. The piece of silver metal that serves as the anode, gradually becomes lighter, as the solid atoms change into ions in solution.
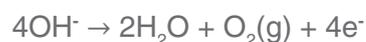
3.4 Ag^+(aq) $+ e^- \rightarrow Ag$

3.5 First cover the plastic with a layer of graphite, which is a good conductor. The metal that is plated then precipitates on top of the graphite layer or plastic. OR
First paint a layer of special conducting paint over the object. OR
Object is etched with chromic acid, then covered with a layer that contains Pd and Sn salts and is then electroplated.

4.1 (Saturated) sodium chloride solution, NaCℓ(aq), brine

4.2 X: saturated NaCℓ solution

Y: lower concentration NaCℓ solution

4.3 Sodium hydroxide solution, NaOH(aq), caustic soda

4.4.1 Anode, positive

4.4.2 $2Cℓ^-$(aq) $\rightarrow Cℓ_2$(g) $+ 2e^-$

4.4.3 O_2(g), oxygen

4.4.4 $2H_2O \rightarrow O_2$(g) $+ 4H^+ + 4e^-$ or $4OH^- \rightarrow 2H_2O + O_2$(g) $+ 4e^-$

4.5.1   Hydrogen

4.5.2 $2H_2O + 2e^- \rightarrow H_2(g) + 2OH^-$

4.6 Na^+ ions are weak oxidising agents, $H^+(aq)$ or $H_2O(\ell)$ is a stronger oxidising agent. Therefore the favoured product will be $H_2(g)$ and not Na metal. The cations of very reactive metals like Na, are not readily reduced.

5.1   Aluminium is much more reactive than iron (Aℓ is a strong reducing agent), therefore the bonds in aluminium oxide are stronger than in the iron oxides and the bonds do not break easily. More energy is required to break the bonds in $A\ell_2O_3$.

5.2   High costs of extraction of aluminium from the ore, mainly due to the high costs of electricity.

5.3   Cathode

5.4 $A\ell^{3+}(\ell) + 3e^- \rightarrow A\ell(s)$

5.5   Oxygen, $O_2(g)$

5.6   Anode; oxidation ALWAYS occurs at the anode. The O^{2-} anions move to the anode where it is oxidised to form oxygen.

5.7   Carbon (graphite)

5.8   The electrodes must be unreactive – carbon (graphite) or platinum. Cost of platinum electrodes is unrealistic/uneconomical – therefore carbon is used.

5.9   The oxygen, $O_2(g)$, that forms through oxidation at the anodes (+ electrodes), reacts with the carbon anodes due to high temperature, to form carbon(IV) oxide/carbon dioxide, $CO_2(g)$. The anodes gradually burn away and must be replaced.

5.10   High cost of extraction of aluminium to make new cans, impact of high costs of electricity for this extraction, shortage of electricity in SA.

5.11   According to the Table of Standard Reduction Potentials, aluminium is a strong reducing agent. Therefore aluminium metal is readily oxidised to $A\ell^{3+}$ ions, which could react with other substances in food.

6.1   Diaphragm cell: the separation between half cells is a diaphragm, since it allows both positive (Na^+) and negative ions ($C\ell^-$) through.

6.2   A: $Na^+(aq)$
       B: $C\ell^-(aq)$

6.3   D: Chlorine gas
       E: Hydrogen gas

6.4   Sodium hydroxide ($NaOH(aq)$)