Suggested Solutions to Self-check for Tut 22b Electrochemistry 2

1 At the cathode(–), 2 reduction reactions are possible:

$2H_2O + 2e^- \rightleftharpoons H_2 + 2OH^-$ $\quad E^\ominus = -0.83$ V
$Ni^{2+} + 2e^- \rightleftharpoons Ni$ $\quad E^\ominus = -0.25$ V

Since $E^\ominus(Ni^{2+}/Ni)$ is less negative than $E^\ominus(H_2O/H_2)$, Ni^{2+} is preferentially reduced to form Ni(s) at the cathode. Ni will be coated on the cathode and the cathode will increase in size / mass.

At the anode(+), 3 oxidation reactions are possible. (Note: Oxidation reactions are in the reverse direction, and the starting materials that are required are on the right-hand side.)

$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$ $\quad E^\ominus = +1.23$ V
$Br_2 + 2e^- \rightleftharpoons 2Br^-$ $\quad E^\ominus = +1.07$ V
$Ni^{2+} + 2e^- \rightleftharpoons Ni$ $\quad E^\ominus = -0.25$ V

Since $E^\ominus(Ni^{2+}/Ni)$ is the least positive, Ni metal is preferentially oxidised to form Ni^{2+} at the anode. The anode dissolves.

2 **(a)** Volume of Al_2O_3 to be deposited = 500 x 1 x 10^{-3} = 0.5 cm^3
Mass of Al_2O_3 = density x volume = 4.0 x 0.5 = 2.0 g
Amt of Al_2O_3 = 2.0 / (27.0 x 2 + 16.0 x 3) = 0.0196 mol

At the anode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$
Al reacts with O_2 produced: $4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

Amt of e^- required = 4 x amt of O_2 produced = 4 x ($\frac{3}{2}$ x amt of Al_2O_3) = 0.1176 mol
Quantity of charge needed = 96500 x 0.1176 = **11300 C**

Note: Al_2O_3 is produced by O_2 formed at anode reacting with the Al anode. $Al \rightarrow Al^{3+}(aq) + 3e^-$ does not occur at the anode.

(b) $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$
$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

For cells in series, the current passed must be the same.
For the same current passed, amt of Ag: amt of Ni deposited = 2 : 1
Hence mass of Ag: mass of Ni deposited = 2 x 107.9 : 1 x 58.7 = **3.68 : 1**

(c) cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
anode: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$

Amt of O_2 produced = ½ x amt of Cu deposited = ½ x (0.635 / 63.5) = 0.00500 mol
Vol. of O_2 produced at s.t.p. = 0.00500 x 22.7 = **0.114 dm^3**

3 **(Answer: D)**
$Q = It = n_eF$. Since I and F are constant, decreasing the time taken will decrease the amount of electrons that passed through the electrolyte.

Applying Faraday's First Law, since the mass of Cu deposited is proportional to the amount of electrons passed through the electrolyte, but independent of concentration of solution and nature of anion, only statement 1 is correct.

4 **(Answer: D)**

The two relevant equations are:

$2Cl^- \rightarrow Cl_2 + 2e^-$ ----------(1)

$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ ----------(2)

Since V_{O2} = 200 cm^3 and V_{Cl2} = 100 cm^3, n_{O2} produced = 2 x n_{Cl2} produced.

Equation (2) x 2: $4H_2O \rightarrow 2O_2 + 8H^+ + 8e^-$ ----------(3)

Comparing (1) and (3), n_{e^-} required to form 200 cm^3 of O_2 = 4 x n_{e^-} to form 100 cm^3 of Cl_2.
Since Q = It = n_eF, $I \propto n_e$ (since t and F are constant). Hence, current used in electrolysis **2** = 4 x current used in electrolysis **1** = 4 I

5 (**Answer: B)**

n(Al) = 0.27/27 = 0.01 mol. $n(e^-)$ = 0.03 mol.
charge per mol is 2904/0.03 = 96800 C.
elementary charge (value) = 1.6×10^{-19} C.
Hence $N_A = 96800 \div 1.6 \times 10^{-19} = 6.05 \times 10^{23}$