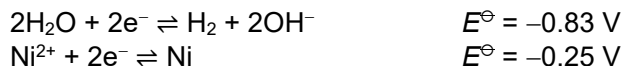


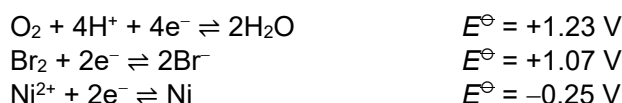
### Suggested Solutions to Self-check

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



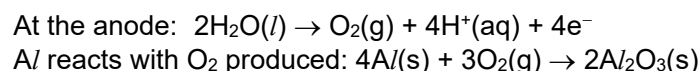
Since  $E^\ominus(\text{Ni}^{2+}/\text{Ni})$  is less negative than  $E^\ominus(\text{H}_2\text{O}/\text{H}_2)$ ,  $\text{Ni}^{2+}$  is preferentially reduced to form  $\text{Ni}(\text{s})$  at the cathode. Ni will be coated on the cathode and the cathode will increase in size / mass.

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.)



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

- 2 (a) Volume of  $\text{Al}_2\text{O}_3$  to be deposited =  $500 \times 1 \times 10^{-3} = 0.5 \text{ cm}^3$   
Mass of  $\text{Al}_2\text{O}_3$  = density  $\times$  volume =  $4.0 \times 0.5 = 2.0 \text{ g}$   
Amt of  $\text{Al}_2\text{O}_3 = 2.0 / (27.0 \times 2 + 16.0 \times 3) = 0.0196 \text{ mol}$



$$\text{Amt of e}^- \text{ required} = 4 \times \text{amt of O}_2 \text{ produced} = 4 \times \left(\frac{3}{2} \times \text{amt of Al}_2\text{O}_3\right) = 0.1176 \text{ mol}$$

$$\text{Quantity of charge needed} = 96500 \times 0.1176 = \mathbf{11300 \text{ C}}$$

Note:  $\text{Al}_2\text{O}_3$  is produced by  $\text{O}_2$  formed at anode reacting with the  $\text{Al}$  anode.  $\text{Al} \rightarrow \text{Al}^{3+}(\text{aq}) + 3\text{e}^-$  does not occur at the anode.

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

**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

$$\text{Hence mass of Ag: mass of Ni deposited} = 2 \times 107.9 : 1 \times 58.7 = \mathbf{3.68 : 1}$$

- (c) cathode:  $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu}(\text{s})$   
anode:  $2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$

$$\begin{aligned} \text{Amt of O}_2 \text{ produced} &= \frac{1}{2} \times \text{amt of Cu deposited} = \frac{1}{2} \times (0.635 / 63.5) = 0.00500 \text{ mol} \\ \text{Vol. of O}_2 \text{ produced at s.t.p.} &= 0.00500 \times 22.7 = \mathbf{0.114 \text{ dm}^3} \end{aligned}$$

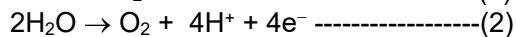
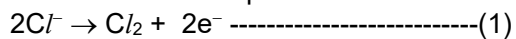
- 3 (Answer: D)

$Q = It = n_e F$ . Since  $I$  and  $F$  are constant, decreasing the time taken will decrease the amount of electrons that passed through the electrolyte.

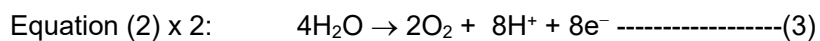
Applying Faraday's First Law, since the mass of Cu deposited is proportional to the amount of electrons passed through the electrolyte, but independent of concentration of solution and nature of anion, only statement 1 is correct.

**4 (Answer: D)**

The two relevant equations are:



Since  $V_{\text{O}_2} = 200 \text{ cm}^3$  and  $V_{\text{Cl}_2} = 100 \text{ cm}^3$ ,  $n_{\text{O}_2 \text{ produced}} = 2 \times n_{\text{Cl}_2 \text{ produced}}$ .



Comparing (1) and (3),  $n_{\text{e}^-}$  required to form  $200 \text{ cm}^3$  of  $\text{O}_2 = 4 \times n_{\text{e}^-}$  to form  $100 \text{ cm}^3$  of  $\text{Cl}_2$ .

Since  $Q = It = n_e F$ ,  $I \propto n_e$  (since  $t$  and  $F$  are constant). Hence, current used in electrolysis **2** = 4  
x current used in electrolysis **1** = 4  $I$

**5 (Answer: B)**

$n(\text{Al}) = 0.27/27 = 0.01 \text{ mol}$ .  $n(\text{e}^-) = 0.03 \text{ mol}$ .

charge per mol is  $2904/0.03 = 96800 \text{ C}$ .

electronic charge (value) =  $1.6 \times 10^{-19} \text{ C}$ .

Hence  $N_A = 96800 \div 1.6 \times 10^{-19} = 6.05 \times 10^{23}$