Question

A solution of $CuSO_4$ is electrolysed for 10 minutes with a current of 1.5 amperes. What mass of copper is deposited at the cathode?

Answer:

$Q = It = 1.5\,A \times 10\,min = 1.5\,A \times 10 \times 60\,s = 900\,C$

$n_{e^-} = \frac{Q}{\text{charge on 1 mol } e^-} = \frac{900}{96500}$ mol

$CuSO_4$

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

By stoichiometry

$\frac{n_{Cu}}{1} = \frac{n_{e^-}}{2}$

$= \frac{1}{2} \times \frac{900}{96500}$

$n_{Cu} = 0.00466$ mol

$W_{Cu} = n_{Cu} \times M_{Cu} = 0.00466 \times 63.5$
$= 0.296\,g$

Alternate:

$E_{Cu} = \frac{M_{Cu}}{n \text{ factor}} = \frac{63.5}{2}$

$W_{Cu} = \frac{E_{Cu}\,I\,t}{96500} = \frac{63.5 \times 1.5 \times 10 \times 60}{2 \times 96500}$

$= 0.296\,g.$

Question

If a current of 0.5 A flows through a metallic wire for 2 hours, then how many electrons would flow through the wire?

Answer

$Q = It = 0.5A \times 2\ hrs = 0.5A \times 2 \times 3600\ s = 3600\ C$

$n_{e^-} = \frac{Q}{\text{charge on 1mol } e^-} = \frac{3600}{96500}$ mol

$N_{e^-} = n_{e^-} \times N_A = \frac{3600}{96500} \times 6.022 \times 10^{23} = 2.25 \times 10^{22}\ e^-$

Question

Consider the reaction:

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

What is the quantity of electricity in coulombs needed to reduce 1 mol of $Cr_2O_7^{2-}$?

Ans: By stoichiometry

$\frac{n_{e^-}}{6} = \frac{n_{Cr_2O_7^{2-}}}{1}$

$n_{e^-} = 6 \times n_{Cr_2O_7^{2-}}$

$= 6 \times 1$

$= 6$ mol

$Q = n_{e^-} \times$ charge on 1mol e^-

$= 6 \times 96500$

$= 579,000$ C

Question

How much charge is required for the following reductions

i) 1 mol of Al^{3+} to Al ?

ii) 1 mol of Cu^{2+} to Cu ?

iii) 1 mol of MnO_4^- to Mn^{2+} ?

Answer

i) $Al^{3+} + 3e^- \rightarrow Al$

By stoichiometry.

$$\frac{n_{e^-}}{3} = \frac{n_{Al^{3+}}}{1}$$

$$n_{e^-} = 3 \times n_{Al^{3+}}$$

$$= 3 \times 1$$

$$= 3 \text{ mol}$$

$Q = n_{e^-} \times$ charge on 1mol e^-

$$= 3 \times 96500$$

$$= 289,500 \text{ C}$$

ii) $Cu^{2+} + 2e^- \rightarrow Cu$

By stoichiometry

$$\frac{n_{e^-}}{2} = \frac{n_{Cu^{2+}}}{1}$$

$$n_{e^-} = 2 \times n_{Cu^{2+}}$$

$$= 2 \times 1$$

$$= 2 \text{ mol}$$

$Q = n_{e^-} \times$ charge on 1mol e^-

$$= 2 \times 96500$$

$$= 193,000 \text{ C}$$

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

By stoichiometry

$$\frac{n_{e^-}}{5} = \frac{n_{MnO_4^-}}{1}$$

$$n_{e^-} = 5 \times n_{MnO_4^-}$$

$$= 5 \times 1$$

$$= 5 \text{ mol}$$

$Q = n_{e^-} \times$ charge on 1 mol e^-

$= 5 \times 96500$

$= 482,500\,C$

Question

How much electricity in terms of Faraday is required to produce

i) 20.0 g of Ca from molten $CaCl_2$?

ii) 40.0 g of Al from molten Al_2O_3?

Answer

i) $n_{Ca} = \frac{w_{Ca}}{M_{Ca}} = \frac{20}{40} = 0.5$ mol

$Ca^{2+} + 2e^- \rightarrow Ca$

By stoichiometry:

$\frac{n_{e^-}}{2} = \frac{n_{Ca}}{1}$

$n_{e^-} = 2 \times n_{Ca}$

$= 2 \times 0.5$

$= 1$ mol

$Q = n_{e^-} \times$ charge on 1 mol e^-

$= 1 \times 1\,F$

$= 1\,F$

ii) $n_{Al} = \frac{w_{Al}}{M_{Al}} = \frac{40}{27}$ mol

$Al^{3+} + 3e^- \rightarrow Al$

By stoichiometry

$\frac{n_{e^-}}{3} = \frac{n_{Al}}{1}$

$n_{e^-} = 3 \times n_{Al}$

$= 3 \times \frac{40}{27}$

$= \frac{40}{9}$ mol

$Q = n_{e^-} \times$ charge on 1mole e^-

$= \frac{40}{9} \times 1 F$

$= 4.44 F$

Question

How much electricity is required in coulomb for the oxidation of

i) 1 mol of H_2O to O_2 ?

ii) 1 mol of FeO to Fe_2O_3 ?

Answer.

i) $H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$

By stoichiometry.

$\frac{n_{e^-}}{2} = \frac{n_{H_2O}}{1}$

$n_{e^-} = 2 \times n_{H_2O}$

$= 2 \times 1$

$= 2$ mol

$Q = n_{e^-} \times$ charge on 1mol e^-

$= 2 \times 96500C$

$= 193,000 C$

ii) $2FeO + H_2O \rightarrow Fe_2O_3 + 2H^+ + 2e^-$

By stoichiometry

$\frac{n_{e^-}}{2} = \frac{n_{FeO}}{2}$

$n_{e^-} = n_{FeO}$

$= 1$ mol

$Q = n_{e^-} \times$ charge on 1mol e^-

$= 1 \times 96500\ C$

$= 96500\ C$

Question

A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Answer

$Q = It = 5A \times 20\ min = 5A \times 20 \times 60\ s = 6000\ C$

$$n_{e^-} = \frac{Q}{\text{charge on 1 mol } e^-} = \frac{6000}{96500}\ mol$$

$$Ni^{2+} + 2e^- \rightarrow Ni$$

By stoichiometry

$$\frac{n_{Ni}}{1} = \frac{n_{e^-}}{2}$$

$$n_{Ni} = \frac{1}{2} \times \frac{6000}{96500}\ mol$$

$$w_{Ni} = n_{Ni} \times M_{Ni} = \frac{6000}{96500 \times 2} \times 58.7 = 1.82 g$$

Alternate

$Ni^{2+} + 2e^- \rightarrow Ni$ $\quad E_{Ni} = \frac{M_{Ni}}{n\ factor} = \frac{58.7}{2}$

$$w_{Ni} = \frac{E_{Ni}\ It}{96500} = \frac{58.7 \times 5 \times 20 \times 60}{2 \times 96500} = 1.82 g$$

Question:

Three electrolytic cells A, B, C containing solutions of $ZnSO_4$, $AgNO_3$ and $CuSO_4$ respectively are connected in series. A steady current of 1.5 A was passed through them until 1.45 g of silver deposited at the cathode of cell B. How long did the current flow? What mass of copper and zinc were deposited?

Answer:

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

$$E_{Ag} = \frac{M_{Ag}}{n\ factor} = \frac{108}{1} = 108\,g$$

$$W_{Ag} = \frac{E_{Ag}\, I\, t}{96500}$$

$$1.45 = \frac{108 \times 1.5 \times t}{96500}$$

$$t = 864\,s$$

Since cells are connected in series same amount of charge flows through them, hence number of equivalents deposited for each substance is same. This is in accordance with Faraday's second law of electrolysis.

$$n_{eq}(Cu) = n_{eq}(Zn) = n_{eq}(Ag)$$

$$\frac{W_{Cu}}{E_{Cu}} = \frac{W_{Zn}}{E_{Zn}} = \frac{W_{Ag}}{E_{Ag}}$$

$$\frac{W_{Cu}}{M_{Cu}/(n\ factor\ of\ Cu)} = \frac{W_{Zn}}{M_{Zn}/(n\ factor\ of\ Zn)} = \frac{W_{Ag}}{M_{Ag}/(n\ factor\ of\ Ag)}$$

$$\frac{W_{Cu}}{63.5/2} = \frac{W_{Zn}}{65.4/2} = \frac{1.45}{108}$$

$W_{Cu} = 0.426\,g$, $W_{Zn} = 0.439\,g$