

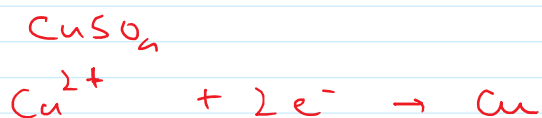
Question

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:

$$Q = It = 1.5 \text{ A} \times 10 \text{ min} = 1.5 \text{ A} \times 10 \times 60 \text{ s} = 900 \text{ C}$$

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



By stoichiometry:

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

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

$$n_{\text{Cu}} = 0.00466 \text{ mol}$$

$$W_{\text{Cu}} = n_{\text{Cu}} \times M_{\text{Cu}} = 0.00466 \times 63.5$$

$$= 0.296 \text{ g}$$

Alternate:

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

$$W_{\text{Cu}} = \frac{E_{\text{Cu}} It}{96500} = \frac{63.5 \times 1.5 \times 10 \times 60}{2 \times 96500}$$

$$= 0.296 \text{ g.}$$

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

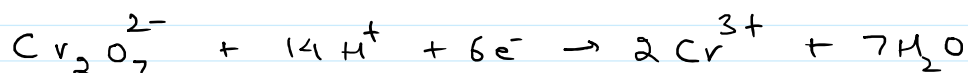
$$Q = It = 0.5A \times 2 \text{ hrs} = 0.5A \times 2 \times 3600 \text{ s} = 3600 \text{ C}$$

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

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

Question

Consider the reaction:



What is the quantity of electricity in coulombs needed to reduce 1 mol of $\text{Cr}_2\text{O}_7^{2-}$?

Ans: By stoichiometry

$$\frac{n_{e^-}}{6} = \frac{n_{\text{Cr}_2\text{O}_7^{2-}}}{1}$$

$$\begin{aligned} n_{e^-} &= 6 \times n_{\text{Cr}_2\text{O}_7^{2-}} \\ &= 6 \times 1 \\ &= 6 \text{ mol} \end{aligned}$$

$$\begin{aligned} Q &= n_{e^-} \times \text{charge on 1 mol } e^- \\ &= 6 \times 96500 \\ &= 579,000 \text{ C} \end{aligned}$$

Question

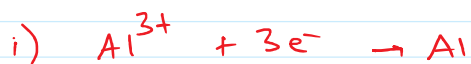
How much charge is required for the following reductions

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

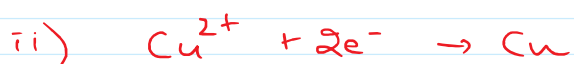


By stoichiometry.

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

$$\begin{aligned} n_{e^-} &= 3 \times n_{\text{Al}^{3+}} \\ &= 3 \times 1 \\ &= 3 \text{ mol} \end{aligned}$$

$$\begin{aligned} Q &= n_{e^-} \times \text{charge on } 1 \text{ mole } e^- \\ &= 3 \times 96500 \\ &= 289,500 \text{ C} \end{aligned}$$

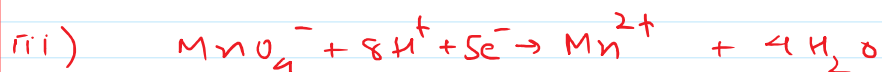


By stoichiometry

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

$$\begin{aligned} n_{e^-} &= 2 \times n_{\text{Cu}^{2+}} \\ &= 2 \times 1 \\ &= 2 \text{ mol} \end{aligned}$$

$$\begin{aligned} Q &= n_{e^-} \times \text{Charge on } 1 \text{ mole } e^- \\ &= 2 \times 96500 \\ &= 193,000 \text{ C} \end{aligned}$$



By stoichiometry

$$\frac{n_{e^-}}{5} = \frac{n_{\text{MnO}_4^-}}{1}$$

$$\begin{aligned} n_{e^-} &= 5 \times n_{\text{MnO}_4^-} \\ &= 5 \times 1 \\ &= 5 \text{ mol} \end{aligned}$$

$$\begin{aligned}
 Q &= n_{e^-} \times \text{charge on } 1 \text{ mol } e^- \\
 &= 5 \times 96500 \\
 &= 482,500 \text{ C}
 \end{aligned}$$

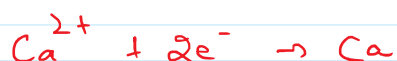
Question

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

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



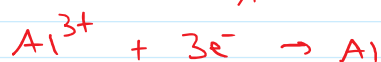
By stoichiometry:

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

$$\begin{aligned}
 n_{e^-} &= 2 \times n_{\text{Ca}} \\
 &= 2 \times 0.5 \\
 &= 1 \text{ mol}
 \end{aligned}$$

$$\begin{aligned}
 Q &= n_{e^-} \times \text{charge on } 1 \text{ mol } e^- \\
 &= 1 \times 1 \text{ F} \\
 &= 1 \text{ F}
 \end{aligned}$$

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



By stoichiometry

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

$$\begin{aligned}
 n_{e^-} &= 3 \times n_{\text{Al}} \\
 &= 3 \times \frac{40}{27}
 \end{aligned}$$

$$= \frac{40}{a} \text{ mol}$$

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

$$= \frac{40}{a} \times 1 \text{ F}$$

$$= 4.44 \text{ F}$$

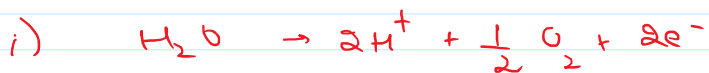
Question

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.



By stoichiometry.

$$\frac{n_{e^-}}{2} = \frac{n_{\text{H}_2\text{O}}}{1}$$

$$n_{e^-} = 2 \times n_{\text{H}_2\text{O}}$$

$$= 2 \times 1$$

$$= 2 \text{ mol}$$

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

$$= 2 \times 96500 \text{ C}$$

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



By stoichiometry

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

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

$$= 1 \text{ mol}$$

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

$$= 1 \times 96500 \text{ C}$$

$$= 96500 \text{ C}$$

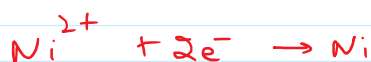
Question

A solution of $\text{Ni}(\text{NO}_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Answer

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

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



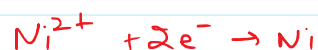
By stoichiometry

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

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

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

Alternate



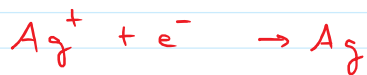
$$E_{\text{Ni}} = \frac{M_{\text{Ni}}}{n \text{ factor}} = \frac{58.7}{2}$$

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

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:



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

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

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

$$t = 864 \text{ s}$$

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.

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

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

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

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

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