Consider electrolysis of aqueous silver ion.

\[ \text{Ag}^{+} (aq) + e^- \rightarrow \text{Ag(s)} \]

1 mol e- \rightarrow 1 mol Ag

If we could measure the moles of e-, we could know the quantity of Ag formed.

But how to measure moles of e-?

Current = charge passing

\[ \text{I (amps)} = \frac{\text{coulombs}}{\text{seconds}} \]

But how is charge related to moles of electrons?

Charge on 1 mol e- =

\[ (1.60 \times 10^{-19} \text{C/e-})(6.02 \times 10^{23} \text{e-/mol}) \]

= 96,500 C/mol e- = 1 Faraday

1.50 amps flow through a Ag⁺(aq) solution for 15.0 min. What mass of Ag metal is deposited?

Solution

(a) Calculate charge

\[ \text{Coulombs} = \text{amps x time} \]

\[ = (1.5 \text{ amps})(15.0 \text{ min})(60 \text{ s/min}) = 1350 \text{ C} \]

(b) Calculate moles of e-

\[ 1350 \text{ C} \times \frac{1 \text{ mol e-}}{96,500 \text{ C}} = 0.0140 \text{ mol e-} \]

(c) Calculate quantity of Ag

\[ 0.0140 \text{ mol Ag} = 1.51 \text{ g Ag} \]

The anode reaction in a lead storage battery is

\[ \text{Pb(s) + HSO}_4^- (aq) \rightarrow \text{PbSO}_4(s) + H^+(aq) + 2e^-} \]

If a battery delivers 1.50 amp, and you have 454 g of Pb, how long will the battery last?

Solution

(a) Charge = 454 g Pb = 2.19 mol Pb
(b) Calculate moles of e-

\[ 2.19 \text{ mol Pb} \times \frac{2 \text{ mol e-}}{1 \text{ mol Pb}} = 4.38 \text{ mol e-} \]

(c) Calculate charge

\[ 4.38 \text{ mol e-} \times 96,500 \text{ C/mol e-} = 423,000 \text{ C} \]

\[ = \frac{1 \text{ Faraday}}{1.50 \text{ amp}} = 282,000 \text{ s} \]

About 78 hours