Faraday's Laws-Stoichiometry

The amount of product of an electrolysis reaction depends on
i) mass
ii) time
iii) electrical current

Definitions- 1 Coulomb = 1 Ampère/sec.
(charge) (current)
e^-

Faraday's Constant) 1 mol e^- = 96 500 C

Example- What mass of Cu is deposited on a cathode dipped in $\overset{+2}{Cu}SO_{4(aq)}$ by a 5.0 A for 30 min.

sec. a) $30\ \cancel{min.} \times \frac{60\ s}{1\ \cancel{min.}} = 1800\ sec.$

C b) Coulombs = 5.0 A × 1800 s = 9000 C

moles e^- c) $9000\ \cancel{C} \times \frac{1\ mole^-}{96\ 500\ \cancel{C}} = 0.093\ mol\ e^-$

mol Cu ½ rxn. d) $Cu^{2+} + 2e^- \longrightarrow Cu_{(s)}$

e) $0.093\ \cancel{mol\ e^-} \times \frac{1\ mol\ Cu}{2\ \cancel{mole^-}} = 0.047\ mol\ Cu$

Mass Cu f) $0.047\ \cancel{mol\ Cu} \times \frac{63.55\ g\ Cu}{1\ \cancel{mol\ Cu}} = 3.0\ g\ Cu$

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Faraday's Laws-Time

How much time is needed to deposit 15.8 g Ag using a 25.0 A current?

a) $Ag^+ + e^- \longrightarrow Ag_{(s)}$

$15.8\,\cancel{g\,Ag} \times \frac{1\text{ mole- Ag}}{107.87\,\cancel{g.Ag}} = 0.146\text{ mol Ag}$

b) $0.146\,\cancel{\text{mol Ag}} \times \frac{1\text{ mol }e^-}{1\,\cancel{\text{mol }Ag^+}} = 0.146\text{ mol }e^-$

c) $0.146\,\cancel{\text{mol }e^-} \times \frac{96500\text{C}}{1\,\cancel{\text{mol }e^-}} = 14135\text{ C}$

also

14135 A-sec.

d) $\frac{14135\,\cancel{A}\text{-s}}{25.0\,\cancel{A}} = 565\text{ sec.}$

Questions-
p. 793
#22, 23, 21 ← mass
↑ Time, ↑ current