Faraday's Laws (Stoichiometry)

Text p. 790 Find mass.

The amount of product of an electrolysis reaction depends on
a) mass
b) time
c) electrical current

Definitions-

1 Coulomb = 1 Ampère x 1 second
(C) (charge) (current)

1 mole e^- = 96 500 C

Example- What mass of Cu is deposited on a cathode dipped in $CuSO_4$ solution by a 5.0 A current for 30 minutes?

a) $30 \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}} = 1800 \text{ s}$

b) Coul. = $5.0 \text{ A} \times 1800 \text{ s} = 9000 \text{ C}$

c) moles e^- = $9000 \text{ C} \times \frac{1 \text{ mole}^-}{96500 \text{ C}} = 0.093 \text{ mol } e^-$

d) Use $\frac{1}{2}$ reaction (Faraday's constant)

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

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

e) $0.047 \text{ mol Cu} \times 63.55 \frac{\text{g}}{\text{mol}} = 3.0 \text{ g Cu}$ (from table)

P. 773 #5

$Cl_2 + 2Br^- \rightarrow 2Cl^- + Br_2$ V

red. $Cl_2 + 2e^- \rightarrow 2Cl^-$ 1.358

$2Br^- \rightarrow Br_2 + 2e^-$ −1.066

0.292 V

Faraday's Laws-Find Time

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

a) Moles Ag: $15.8 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} = 0.146 \text{ mol Ag}$

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

$0.146 \text{ mol Ag} \times \frac{1 \text{ mole}^-}{1 \text{ mol Ag}} = 0.146 \text{ mol } e^-$

c) $0.146 \text{ mole}^- \times \frac{96500 \text{ C}}{1 \text{ mole}^-} = 14,135 \text{ C}$

d) $\frac{14135 \text{ A} \cdot \text{s}}{25.0 \text{ A}} = 565 \text{ sec.}$ (also 14135 A-s)

Questions P. 793
#22 (time), 23 (amp), #21 (mass)