Electrochemistry- Electrolysis

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10) $Li^+ Br^-$

reduction $Li^+ + e^- \longrightarrow Li$

oxidation $2Br^- \longrightarrow 2e^- + Br_2$

Faraday's Laws- Stoichiometry

The ~~amount~~ mass of product Text p. 790
of an electrolysis reaction is affected by,

- time
- electrical current

Definitions-

1. 1 Coulomb (charge) = 1 Ampère (current) x 1 second (time)
2. 1 mole ___ = ___ g (periodic table)
3. Faraday's constant
   1 mole e^- = 96 500 Coulombs
4. 1 min. = 60 seconds

Example 1 – Mass

What mass of $Cu_{(s)}$ is deposited on a cathode dipped in $CuSO_{4(aq)}$ by a 5.0A current for 30 minutes?

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

b. Coulombs $= 5.0 \text{ A} \times 1800 \text{ s}$

$= 9000 \text{ C}$

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

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

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

e. $0.047 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol}} = 3.0 \text{ g Cu}$

Example 2- Time

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

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

b. $Ag^+ + e^- \longrightarrow Ag$

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

c. $0.146\text{ mol } e^- \times \frac{96500\text{ C}}{1\text{ mol } e^-} = 14135\text{ C}$ also $14135\text{ A} \times \text{s}$

d. $\frac{14135\text{ A} \times \text{s}}{25.0\text{ A}} = 565\text{ seconds}$

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Electrochemistry Unit Outline

1. Draw a working galvanic cell & predict voltage and e^- flow.
2. Predict whether a redox rxn is spontaneous or not.
3. Write equations showing electrolysis of a molten salt.
4. Solve Faraday's Laws questions (stoichiometry)