<u>Electrolysis problem (Faraday's Law)</u>

<u>What's it look like</u>

You are given: ***mass of material*** used up or deposited and asked for the time or strength of a current used to produce these results ***or*** conversely, you are given the ***time and strength of current*** and asked how much material is used up or deposited.

Ex: How long will it take to plate out 1.0 kg Al from aqueous Al^{3+} using a current of 100.0 A?

<u>Concept behind it</u>

Electrolysis is a process where electricity is used to drive a redox reaction with a negative cell potential (sort of the opposite of a battery where electricity is produced letting a redox reaction with a positive cell potential proceed). The process is used for recharging batteries, production of certain substances (like aluminum), and electroplating. The chemistry problem you'll be given is a redox half-reaction stoich problem with a little dimensional analysis thrown in to convert current from amps to moles of electrons.

<u>How to tackle it</u>

If they give you mass of material, do the stoich (using the redox half-reaction) to find the moles of electrons needed. Then do a little dimensional analysis to convert that into strength or duration of current using Faraday's constant (96,500 coulombs = 1 mol of electrons) and the definition of an amp (1 A = 1 C per second). If they give you the current (both strength and duration), do the opposite: use dimensional analysis to find the moles of electrons, and then use the stoich (on the redox half-reaction) to determine the mass of the material.

<u>Detailed steps</u>

1) Write the half-reaction for the material in question.
2) Convert grams of material to moles of material.
3) Using stoich (coefficients), determine the number of moles of electrons involved in the process.
4) Using dimensional analysis, calculate the time or amperage required to move that many electrons.

Example (continued from above):

<u>Half-reaction</u>: $Al^{3+} + 3e^- \rightarrow Al$

<u>Stoichiometry:</u>

$$1.0\ x10^3\ g\ Al \cdot \frac{1\ \text{mol Al}}{26.98\ \text{g Al}} = 37.06449\ mol\ Al$$

$$37.06449\ mol\ Al \cdot \frac{3\ \text{mol e}^-}{1\ \text{mol Al}} = 111.19348\ mol\ e^-$$

<u>Dimensional analysis (Finding time or amps from mol of e⁻)</u>

$$x\ s \cdot \frac{100.0\ \text{C}}{1\ \text{s}}\ \frac{1\ \text{mol}}{96{,}500\ \text{C}} = 111.19348\ mol\ e^- \qquad (100.0\ C/s = 100.0\ A)$$

x = 107,301 s or 1788.36 min or 29.8060 hr

= 110,000 s or 1800 min or 30. hr