

## Quantitative Electrolysis: The Math (Or How I Learned to Stop Worrying and Love Dimensional Analysis)

In an electrolytic cell, the amount of product made is related to the number of electrons transferred. Essentially, the electrons are a reactant. To solve, use the formula for current from the periodic table and follow the following steps:

**1.** Convert your time,  $t$ , into seconds.

$s$  (seconds)



**2.** Multiply your time (sec) by the current (Amp) to solve for the charge ( $q$ ) in Coulombs.

$$I = \frac{q}{t} \rightarrow q = t \cdot I$$

$\times A$  (amp =  $\frac{C}{sec}$ )

$I$  = current (amperes, A)  
 $q$  = charge (coulombs, C)  
 $t$  = time (sec)

$$\text{Amp} = \frac{C}{sec}$$



**3.** Divide by Faraday's constant (given on the formula chart) to convert Coulombs into moles of electrons.

$$q / F = \text{mol } e^-$$

Faraday's constant,  $F$ , is the charge on 1 mole of electrons.

$$F = 96,485 \text{ C/mol } e^-$$

Use  $F = 100,000$  (1E5) for multiple choice!

$$\times \frac{1}{F} = \frac{\text{mol } e^-}{C}$$

\*Note: a "faraday" can be used to refer to the number of moles of electrons being transferred:

- "3 faradays"  $\rightarrow$  3 moles of  $e^-$  of transferred
- "0.25 faradays"  $\rightarrow$  0.25 moles of  $e^-$  of transferred



**4.** Use the appropriate half-reaction to determine the moles of the metal formed.

For example,



This indicates for every 3 moles of electrons gained, 1 mole of solid gold is formed.

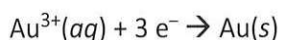
$$\rightarrow \times \frac{1 \text{ mol metal}}{\# \text{ mole } e^-}$$



**5.** Use molar mass to convert from moles to grams of metal.

$$\times \frac{g}{\text{mol}}$$

Example: Gold can be plated out of a solution containing  $\text{Au}^{3+}$  according to the half-reaction:



What mass of gold (in grams) is plated by a 25 minute flow of 5.5 A current?

Solution: 
$$25 \text{ min} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{5.5 \text{ C}}{1 \text{ s}} \times \frac{1 \text{ mole e}^{-}}{96,485 \text{ C}} \times \frac{1 \text{ mol Au}}{3 \text{ mole e}^{-}} \times \frac{196.97 \text{ g Au}}{1 \text{ mol Au}} = 5.6 \text{ g Au}$$

time (sec)  $\times$  current (Amp)  $\times$   $\frac{1}{F}$   $\times$   $\frac{\text{solid metal}}{\text{mol e}^{-}}$   $\times$  molar mass (g/mol)

You will need to be able to do two basic calculations for quantitative electrolysis:

- Given **time (sec)** and **current (A)**, calculate **mass (g)**.
- Given **mass (g)** and **current (A)**, calculate **time required (sec)**.

Of course, there are endless varieties of these two calculation types we can give you! 😊

Quick Trick to remember the order of steps to calculate mass of a metal produced (given time and current):

Are you **SAFE? MMM.**

Sec  $\uparrow$  amp  $\uparrow$   $F$   $\uparrow$  mol  $\uparrow$  molar mass

Going backwards?  
MMM, FASE  $\hat{\uparrow}$

**Let's Try!**

- If 3.30 faraday of charge is passed through a solution of  $\text{Al}_2(\text{SO}_4)_3$ , what mass of aluminum is deposited?

$$3.30 \text{ F} \Rightarrow 3.30 \text{ mol e}^{-} \times \frac{1 \text{ mol Al}}{3 \text{ mol e}^{-}} \times \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} = 29.7 \text{ g Al}$$

- How long must a current of 5.00 A be applied to a solution of  $\text{AgCN}$  to produce 10.5 g silver metal?

$$10.5 \text{ g Ag} \times \frac{1 \text{ mol Ag}}{107.87 \text{ g Ag}} \times \frac{1 \text{ mole e}^{-}}{1 \text{ mol Ag}} \times \frac{96,485 \text{ C}}{1 \text{ mole e}^{-}} \times \frac{1 \text{ sec}}{5.00 \text{ C}} = 1,880 \text{ Sec (or 31.3 min)}$$

- Copper may be used for electroplating, starting with a solution of  $\text{Cu}(\text{NO}_3)_2(\text{aq})$ .

- If a current of 10.0 amp is applied to the  $\text{Cu}(\text{NO}_3)_2(\text{aq})$  solution for 60.0 minutes, what mass of copper will be plated out? (Assume excess  $\text{Cu}(\text{NO}_3)_2(\text{aq})$ ).

$$60.0 \text{ min} \times \frac{60 \text{ Sec}}{1 \text{ min}} \times \frac{10.0 \text{ C}}{1 \text{ Sec}} \times \frac{1 \text{ mole e}^{-}}{96,485 \text{ C}} \times \frac{1 \text{ mol Cu}}{2 \text{ mole e}^{-}} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 11.9 \text{ g Cu}$$

- How many moles of electrons must be transferred in this reaction to produce 5.16 g of copper metal?

$$5.16 \text{ g Cu} \times \frac{1 \text{ mol Cu}}{63.55 \text{ g Cu}} \times \frac{2 \text{ mole e}^{-}}{1 \text{ mol Cu}} = 0.162 \text{ mole e}^{-}$$