

Cell Stoichiometry and Faraday's Law

- The amount of electrons that pass through a cell will determine the mass of the substances reacted or produced at the electrodes
- Michael Faraday investigated the relationship between electrical current (amount of electrons) and the electrochemical changes (change in mass or concentration) that occur at the anode and cathode
 - He found that the mass of an element produced or consumed at an electrode is directly proportional to the current and the time the cell operates. This is known as **Faraday's Law**.
 - Faraday's Law can be summarized by the following equation:

$$n_{e^-} = \frac{It}{F} \quad * \text{ memorize!}$$

where n_{e^-} is the moles of electrons (mol)

I is current measured in amperes (A)

t is time measured in seconds (s)

F is **Faraday's constant** ($9.65 \times 10^4 \text{ C/mol}$)

} watch correct units!

← found on pg. 3 of data booklet

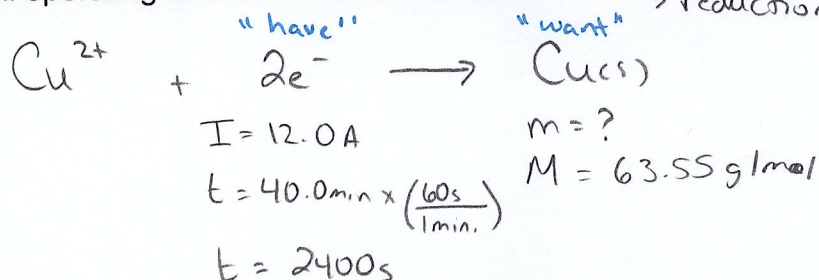
- * Faraday's Law applies to both voltaic and electrolytic cells

- In order for us to use Faraday's law, we need a reaction that actually shows the electrons.

Important! * Half-reactions are the only reactions that show electrons, therefore all cell stoichiometry questions are based off half-reactions and NOT the balanced redox reaction.

EXAMPLES

- What is the mass of copper deposited at the cathode of a copper electro-refining cell operating at 12.0 A for 40.0 minutes? → reduction half-rxn



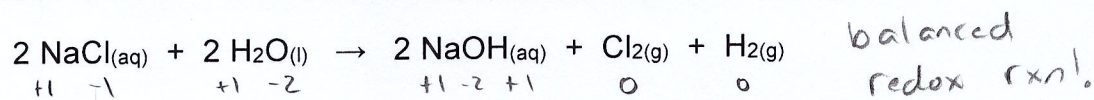
$$\textcircled{1} \quad n_{e^-} = \frac{It}{F} = \frac{(12.0 \text{ A})(2400 \text{ s})}{9.65 \times 10^4 \text{ C/mol}} = 0.2984 \dots \text{ mol}$$

$$\textcircled{2} \quad n_{\text{Cu(s)}} = 0.2984 \dots \text{ mol} \times \left(\frac{1}{2}\right) = 0.1492 \dots \text{ mol}$$

$$\textcircled{3} \quad m = M_n = (63.55 \text{ g/mol})(0.1492 \dots \text{ mol}) = 9.483108 \dots \text{ g}$$

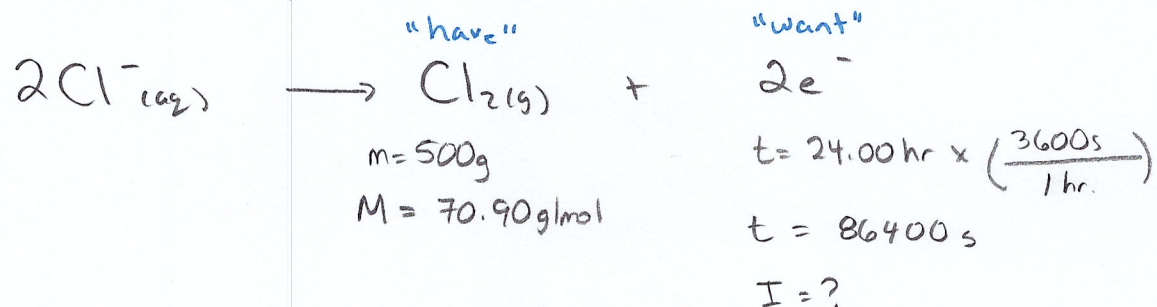
$m = 9.48 \text{ g}$

2. In the chloralkali industry, large quantities of $\text{Cl}_2(\text{g})$ are produced using an electrolytic process that is described below by the following chemical equation.



If 500g of $\text{Cl}_2(\text{g})$ needs to be produced in 24.00hrs, what is the required current?

* need half-rxn! *



$$\textcircled{1} \quad m = nM \rightarrow n_{\text{Cl}_2} = \frac{m}{M} = \frac{500\text{g}}{70.90 \text{ g/mol}} = 7.0521... \text{ mol}$$

$$\textcircled{2} \quad n_{\text{e}^{-}} = 7.0521... \text{ mol} \times \left(\frac{2}{1} \right) = 14.1043... \text{ mol}$$

$$\textcircled{3} \quad n_{\text{e}^{-}} = \frac{It}{F} \rightarrow I = \frac{n_{\text{e}^{-}} F}{t} = \frac{(14.1043... \text{ mol})(9.65 \times 10^4 \text{ C/mol})}{86400 \text{ s}}$$

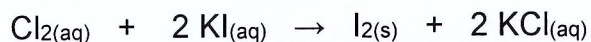
$$I = 15.753... \text{ A}$$

$$I = 15.8 \text{ A}$$

Now try pg. 516 #9-12 & Practice Problem

Practice Problem

1. The following reaction describes the process that takes place in an electrochemical cell.



When 33.0g of iodine solid is produced by the cell, the cell needs to operate for _____ minutes at 5.66A? **[73.9 minutes]**

2. A voltaic cell using $\text{Mg}_{(\text{s})}/\text{Mg}^{2+}_{(\text{aq})}$ and $\text{Cu}_{(\text{s})}/\text{Cu}^{2+}_{(\text{aq})}$ half-cells operates under standard conditions. The cell delivers 0.22A for 31.6 hours.
- Will the mass of the magnesium metal increase or decrease? **[decrease]**
 - Find the change in mass of the magnesium metal. **[3.15 g]**