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:

$$n_{e^-} = \frac{It}{F}$$ * 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** (9.65×10^4 C/mol) ← found on pg. 3 of data booklet

watch correct units!

*- 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

1. 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

$$Cu^{2+} + \underset{\text{"have"}}{2e^-} \longrightarrow \underset{\text{"want"}}{Cu_{(s)}}$$

$I = 12.0\,A$

$t = 40.0\,min \times \left(\frac{60\,s}{1\,min}\right)$

$t = 2400\,s$

$m = ?$

$M = 63.55\,g/mol$

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

② $n_{Cu_{(s)}} = 0.2984\ldots\,mol \times \left(\frac{1}{2}\right) = 0.1492\ldots\,mol$

③ $m = Mn = (63.55\,g/mol)(0.1492\ldots\,mol) = 9.483108\ldots\,g$

$\boxed{m = 9.48\,g}$

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

$$\underset{+1\ -1}{2\ NaCl_{(aq)}} + \underset{+1\ -2}{2\ H_2O_{(l)}} \rightarrow \underset{+1\ -2\ +1}{2\ NaOH_{(aq)}} + \underset{0}{Cl_{2(g)}} + \underset{0}{H_{2(g)}}$$

balanced redox rxn!

If 500g of $Cl_{2(g)}$ needs to be produced in 24.00hrs, what is the required current?

* need half-rxn! *

$$2Cl^-_{(aq)} \longrightarrow \underset{\text{"have"}}{Cl_{2(g)}} + \underset{\text{"want"}}{2e^-}$$

$m = 500g$

$M = 70.90 g/mol$

$t = 24.00 hr \times \left(\frac{3600s}{1 hr.}\right)$

$t = 86400 s$

$I = ?$

① $m = nM \rightarrow n_{Cl_2} = \frac{m}{M} = \frac{500g}{70.90g/mol} = 7.0521... mol$

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

③ $n_{e^-} = \frac{It}{F} \rightarrow I = \frac{n_{e^-}F}{t} = \frac{(14.1043... mol)(9.65\times10^4 C/mol)}{86400 s}$

$I = 15.753... A$

$\boxed{I = 15.8 A}$

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

Practice Problem

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

$$Cl_{2(aq)} \quad + \quad 2\ KI_{(aq)} \quad \rightarrow \quad I_{2(s)} \quad + \quad 2\ KCl_{(aq)}$$

   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 $Mg_{(s)}/Mg^{2+}_{(aq)}$ and $Cu_{(s)}/Cu^{2+}_{(aq)}$ half-cells operates under standard conditions. The cell delivers 0.22A for 31.6 hours.

   a. Will the mass of the magnesium metal increase or decrease? **[decrease]**

   b. Find the change in mass of the magnesium metal. **[3.15 g]**