Ch 19.8: Electrolysis
Homework #19-6: pg. 834 #43, 45, 46, 49, 57

43 **What is the difference between a galvanic cell (such as a Daniell cell) and an electrolytic cell?**
Galvanic Cell: A spontaneous process that produces electricity
Electrolytic Cell: A nonspontaneous process that requires an input of electricity to occur.

45 **The half-reaction at an electrode is: Mg^{2+} (molten) + $2e^-$ → Mg(s). Calculate the number of grams of magnesium that can be produced by supplying 1.00 *F* to the electrode.**

$$\textbf{Mass Mg} = 1.00\ \boldsymbol{F} \times \frac{1\text{ mol Mg}}{2\text{ mol e}^-} \times \frac{24.31\text{ g Mg}}{1\text{ mol Mg}} = \textbf{12.2 g Mg}$$

46 **Consider the electrolysis of molten barium chloride, $BaCl_2$.**
(a) Write the half-reactions.
The only active species present in molten $BaCl_2$ are Ba^{2+} and Cl^-. The electrode reactions are:

anode: $2Cl^-(l) \longrightarrow Cl_2(g) + 2e^-$
cathode: $Ba^{2+}(l) + 2e^- \longrightarrow Ba(s)$

This cathode half-reaction tells us that 2 moles of e^- are required to produce 1 mole of Ba(*s*).

(b) How many grams of barium metal can be produced by supplying 0.50 A for 30 min?
First, we calculate the coulombs of electricity that pass through the cell.

$$\frac{0.50\text{ C}}{1\text{ s}} \times \frac{60\text{ s}}{1\text{ min}} \times 30\text{ min} = 9.0 \times 10^2\ C$$

Stoichiometrically, for every mole of Ba formed at the cathode, 2 moles of electrons are needed. The grams of Ba produced at the cathode are:

$$\textbf{? g Ba} = (9.0 \times 10^2\ C) \times \frac{1\text{ mol e}^-}{96{,}500\text{ C}} \times \frac{1\text{ mol Ba}}{2\text{ mol e}^-} \times \frac{137.3\text{ g Ba}}{1\text{ mol Ba}} = \textbf{0.64 g Ba}$$

49 **One of the half-reactions for the electrolysis of water is: $2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$. If 0.076 L of O_2 is collected at 25°C and 755 mm Hg, how many faradays of electricity had to pass through the solution?**
Find the amount of oxygen using the ideal gas equation

$$\boldsymbol{n} = \frac{\boldsymbol{PV}}{\boldsymbol{RT}} = \frac{\left(755\text{ mmHg} \times \dfrac{1\text{ atm}}{760\text{ mmHg}}\right)(0.076\text{ L})}{(0.0821\text{ L}\cdot\text{atm/K}\cdot\text{mol})(298\text{ K})} = 3.1 \times 10^{-3}\text{ mol }O_2$$

Since the half-reaction shows that one mole of oxygen requires four faradays of electric charge, we write

$$(3.1 \times 10^{-3}\text{ mol }O_2) \times \frac{4\ F}{1\text{ mol }O_2} = \textbf{0.012}\ \boldsymbol{F}$$

57 The passage of a current of 0.750 A for 25.0 min deposited 0.369 g of copper from a $CuSO_4$ solution. From this information, calculate the molar mass of copper.

The quantity of charge passing through the solution is:

$$\frac{0.750\text{ C}}{1\text{ s}} \times \frac{60\text{ s}}{1\text{ min}} \times \frac{1\text{ mol e}^-}{96500\text{ C}} \times 25.0\text{ min} = 1.17 \times 10^{-2}\text{ mol e}^-$$

Since the charge of the copper ion is +2, the number of moles of copper formed must be:

$$(1.17 \times 10^{-2}\text{ mol e}^-) \times \frac{1\text{ mol Cu}}{2\text{ mol e}^-} = 5.85 \times 10^{-3}\text{ mol Cu}$$

The units of molar mass are grams per mole. The molar mass of copper is:

$$\frac{0.369\text{ g}}{5.85 \times 10^{-3}\text{ mol}} = \mathbf{63.1\ g/mol}$$