Homework #19-6: Electrolysis (Section 19.8 in book-- pg. 834 #43, 46, 49, 56, 80, 89)

19.43 What is the difference between a galvanic cell and an electrolytic cell?

Galvanic (Voltaic) Cell:

Electrolytic Cell:

19.46 Consider the electrolysis of molten barium chloride, BaCl₂.

(a) Write the half-reactions and determine the $E^\circ$ for a standard cell.

\[
\begin{align*}
\text{anode:} & \quad E^\circ_{\text{red}} = \\
\text{cathode:} & \quad E^\circ_{\text{red}} = \\
\text{E}^\circ \text{ for a standard cell} = 
\end{align*}
\]

(b) How many grams of barium metal can be produced by supplying 0.50 A for 30 min?

19.49 One of the half-reactions for the electrolysis of water is: \(2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+(aq) + 4e^-\). If 0.076 L of O₂ is collected at 25°C and 755 mm Hg, how many coulombs of electricity had to pass through the solution?

19.56 In the 1950’s, 60’s and 70’s, most car bumpers were plated with chromium to give a shiny, mirror-like finish. Chromium plating can be accomplished by the electrolysis of a suspended car bumper in an acidified dichromate solution.

a) The following (unbalanced) reduction half-reaction takes place in the cell. Balance the reduction half-reaction.  \((Hint: \ \text{Balance Cr, then O, then H, than electrons.})\)

\[
\begin{align*}
\text{Cr}_2\text{O}_7^{2-}(aq) + \quad \text{H}^+(aq) + \quad e^- & \quad \rightarrow \quad \text{Cr}(s) + \quad \text{H}_2\text{O}(l)
\end{align*}
\]

b) How long would it take (in hours) for a car bumper to be plated with 18 g of chromium if the electrolytic cell carries a current of 25.0 amps.
19.80 An acidified aqueous solution (has H⁺) was electrolyzed using copper electrodes. A constant current of 1.18 A is applied for 1520 seconds causing the anode to lose 0.584 g of Cu when Cu is oxidized to Cu²⁺.

(a) Write the half-reaction that occurs at the anode, the half-reaction that occurs at the cathode and the balanced overall reaction. Then calculate \( E^\circ \) of a standard cell.

\[
\begin{align*}
\text{anode:} & \quad E^\circ_{\text{red}} = \\
\text{cathode:} & \quad E^\circ_{\text{red}} = \\
\text{Overall Rxn} & \quad E^\circ_{\text{cell}} =
\end{align*}
\]

(b) What gas is produced at the cathode? _____ What volume of this gas would be collected if the gas was collected at STP?

(c) Calculate the quantity of electricity (in coulombs) used in this cell.

(d) How many electrons were used? 1 electron has a charge of \( 1.6022 \times 10^{-19} \text{C} \).

(e) Calculate Avogadro’s number (# electrons/1 mole e⁻) based on the data in this question

19.89 Industrially, copper is purified by electrolysis. The impure copper acts as the anode, and the cathode is made of pure copper. The electrodes are immersed in a CuSO₄ solution. During electrolysis, copper at the anode enters the solution as Cu²⁺ while Cu²⁺ ions are reduced at the cathode.

(a) Write the half-cell reactions and the overall rxn for the cell.

Anode (impure Cu):

Cathode (pure Cu):

Overall rxn:

(b) Suppose the impure copper anode was contaminated with Zn and Ag. Explain what happens to these impurities during electrolysis. (Hint: Use a reduction chart to determine whether either Zn or Ag is more easily oxidized than Cu. Then determine if there are now any new reduction reactions that would occur.)