Learning Activity 6.8: Problem Solving with Faraday's Law

1. Calculate the mass of products generated at each electrode if 15.0 amp flows for 7.51 minutes through molten MgF_2.

   Cathode (reduction): $Mg^{2+}{}_{(l)} + 2e^- \rightarrow Mg_{(l)}$

2. How long would an aqueous gold (III) chloride cell need to operate to plate 2.5 g of gold on a bracelet with a current of 2.5 A?

   Cathode (reduction): $Au^{3+}{}_{(aq)} + 3e^- \rightarrow Au_{(s)}$

3. If 10.0 g of sulfur is deposited on an electrode in an electrolytic cell from silver sulfide, calculate the mass of silver deposited on the other electrode.

   Cathode (reduction): $Ag^{+}{}_{(aq)} + 1e^- \rightarrow Ag_{(s)}$

4. What products would be expected at each electrode? Write the half-reactions and calculate the voltage required to electrolyze the following:

   a. aqueous copper (II) fluoride

   b. molten magnesium chloride

   c. aqueous iron (II) sulfate

   d. molten silver bromide

   e. aqueous aluminum chloride

5. How many seconds would be needed to generate 3.00 moles of electrons from 10.0 amp of current?

Check the answer key.

Lesson Summary

In this lesson, you learned about the work of Michael Faraday. You used Faraday's Law to solve various problems relating to electrolytic cells. This is the last lesson in this module as well as in the course. Congratulations on your hard work.

Learning Activity 6.8: Problem Solving with Faraday's Law

1. Calculate the mass of products generated at each electrode if 15.0 amp flows for 7.51 minutes through molten MgF_2.

   Cathode (reduction): $Mg^{2+}{}_{(l)} + 2e^- \rightarrow Mg_{(l)}$

   $$n_{e^-} = \frac{I \bullet t}{96500\ \text{C/mol}} = \frac{(15.0\ \text{A})(7.51\ \text{min} \times 60\ \text{s/min})}{96500\ \text{C/mol}} = 0.07004\ \text{moles } e^-$$

   Cathode (reduction): $Mg^{2+}{}_{(l)} + 2e^- \rightarrow Mg_{(l)}$

   $$\text{moles Mg} = 0.007004\ \text{moles } e^- \left(\frac{1\ \text{mole Mg}}{2\ \text{moles } e^-}\right) = 0.0350\ \text{moles Mg}$$

   $$m = n \bullet M = (0.0350\ \text{moles})(24.3\ \text{g/mol}) = 0.851\ \text{g Mg}$$

   Anode (oxidation): $2\,F^-{}_{(l)} \rightarrow F_{2(g)} + 2e^-$

   $$\text{moles } F_2 = 0.007004\ \text{moles } e^- \left(\frac{1\ \text{mole } F_2}{2\ \text{moles } e^-}\right) = 0.0350\ \text{moles } F_2$$

   $$m = n \bullet M = (0.0350\ \text{moles})(38.0\ \text{g/mol}) = 1.33\ \text{g } F_2$$

   0.851 g of Mg is formed at the cathode and 1.33 g of F_2 gas is formed at the anode.

2. How long would an aqueous gold (III) chloride cell need to operate to plate 2.5 g of gold on a bracelet with a current of 2.5 A?

   Cathode (reduction): $Au^{3+}{}_{(aq)} + 3e^- \rightarrow Au_{(s)}$

   $$\text{moles Au} = \frac{m}{M} = \frac{2.50\ \text{g}}{197.0\ \text{g/mol}} = 0.01269\ \text{moles Au}$$

   $$\text{moles } e^- = 0.01269\ \text{moles Au} \left(\frac{3\ \text{moles } e^-}{1\ \text{mole Au}}\right) = 0.03807\ \text{moles } e^-$$

   $$t = \frac{(96500\ \text{C/mol})\,n_{e^-}}{I} = \frac{(96500\ \text{C/mol})(0.03807\ \text{moles } e^-)}{2.5\ \text{A}} = 1469.5\ \text{s}$$

   It would require about 1470 s to plate 2.5 g of gold.

3. If 10.0 g of sulfur is deposited on an electrode in an electrolytic cell from silver sulfide, calculate the mass of silver deposited on the other electrode.

   Cathode (reduction): $Ag^{+}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$

   Find the moles of sulfur deposited:

   $$\text{moles S} = \frac{\text{m}}{\text{molar mass}} = \frac{10.0\text{ g}}{32.1\text{ g/mol}} = 0.3115\text{ moles S deposited}$$

   Find moles of e^{-} needed:

   $S^{2-} \rightarrow S + 2e^{-}$

   $$\text{moles } e^{-} = 0.3115\text{ moles S}\left(\frac{2\text{ moles }e^{-}}{1\text{ mole S}}\right) = 0.623\text{ moles }e^{-}$$

   Find the moles of silver deposited from 0.623 moles of e^{-}:

   $Ag^{+}_{(aq)} + 1e^{-} \rightarrow Ag_{(s)}$

   0.623 moles Ag produced

   m = nm (molar mass) = (0.623 moles)(107.9 g/mol) = 67.2 g

   67.2 g of silver is deposited.

4. What products would be expected at each electrode? Write the half reactions and calculate the voltage required to electrolyze the following:

   a. aqueous copper (II) fluoride

      Species present: Cu^{2+}, F^{-} and H_2O

      At the cathode, either Cu^{2+} or H_2O will be reduced.

      $Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)} \quad E° = +0.34V$

      $2\,H_2O_{(l)} + 2e^{-} \rightarrow 2\,OH^{-}_{(aq)} + H_{2(g)} \quad E° = -0.83V$

      At anode, either F^{-} or H_2O will be oxidized.

      $2\,F^{-}_{(aq)} \rightarrow F_{2(g)} + 2e^{-} \quad E° = -2.87V$

      $H_2O_{(l)} \rightarrow 2\,H^{+}_{(aq)} + \tfrac{1}{2}O_{2(g)} + 2e^{-} \quad E° = -1.23V$

      From the half-cell potentials, we can predict

      Cathode: copper

      Anode: oxygen gas and H^{+} ions

      $E°_{cell} = E°_{ox} + E°_{red} = -1.23V + (+0.34) = -0.89V$

      At least 0.89 V is needed.

b. molten magnesium chloride

Cathode: $Mg^{2+}_{(aq)} + 2e^- \rightarrow Mg_{(s)}$ $E° = -2.37V$
Anode: $2\ Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$ $E° = -1.36V$
$E°_{cell} = E°_{ox} + E°_{red} = -1.36V + (-2.37V) = -3.73V$
At least 3.73V is needed.

c. aqueous iron (II) sulfate

Species: Fe^{2+}, SO_4^{2-}, H_2O
Cathode: $Fe^{2+}_{(aq)} + 2e^- \rightarrow Fe_{(s)}$ $E° = -0.44V$
Anode: $H_2O_{(l)} \rightarrow 2\ H^+_{(aq)} + \frac{1}{2}O_{2(g)} + 2e^-$ $E° = -1.23V$
$E°_{cell} = E°_{ox} + E°_{red} = -1.23V + (-0.44V) = -1.67V$
Minimum voltage required is 1.67V.

d. molten silver bromide

Cathode: $Ag^+_{(l)} + 1e^- \rightarrow Ag_{(l)}$ $E° = +0.80V$
Anode: $2\ Br^-_{(l)} \rightarrow Br_{2(g)} + 2e^-$ $E° = -1.06V$
$E°_{cell} = E°_{ox} + E°_{red} = -1.06V + (+0.80V) = -0.26V$
At least 0.26V is needed.

e. aqueous aluminum chloride

Species: Al^{3+}, Cl^-, H_2O
Cathode: $2\ H_2O_{(l)} + 2e^- \rightarrow 2\ OH^-_{(aq)} + H_{2(g)}$ $E° = -0.83V$
Anode: $H_2O_{(l)} \rightarrow 2\ H^+_{(aq)} + \frac{1}{2}O_{2(g)} + 2e^-$ $E° = -1.23V$
$E°_{cell} = E°_{ox} + E°_{red} = -1.23V + (-0.83V) = -2.06V$
Minimum voltage required is 2.06V.

5. How many seconds would be needed to generate 3.00 moles of electrons from 10.0 amp of current?

$$t = \frac{n_{e^-} \times 96500\ C/mol}{I} = \frac{(3.00\ moles\ e^-)(96500\ C/mol)}{10.0\ A} = 28950\ s$$

28950 s are needed.