



## Learning Activity 6.8: Problem Solving with Faraday's Law

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1. Calculate the mass of products generated at each electrode if 15.0 amp flows for 7.51 minutes through molten  $\text{MgF}_2$ .  
Cathode (reduction):  $\text{Mg}^{2+}(\text{l}) + 2\text{e}^- \rightarrow \text{Mg}(\text{l})$
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):  $\text{Au}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Au}(\text{s})$
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):  $\text{Ag}^+(\text{aq}) + 1\text{e}^- \rightarrow \text{Ag}(\text{s})$
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  $\text{MgF}_2$ .

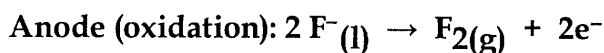


$$n_{\text{e}^-} = \frac{I \cdot t}{96500 \text{ C/mol}} = \frac{(15.0 \text{ A})(7.51 \text{ min} \times 60 \text{ s/min})}{96500 \text{ C/mol}} = 0.07004 \text{ moles e}^-$$



$$\text{moles Mg} = 0.07004 \text{ moles e}^- \left( \frac{1 \text{ mole Mg}}{2 \text{ moles e}^-} \right) = 0.0350 \text{ moles Mg}$$

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

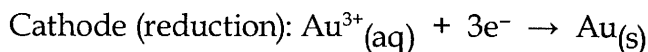


$$\text{moles F}_2 = 0.07004 \text{ moles e}^- \left( \frac{1 \text{ mole F}_2}{2 \text{ moles e}^-} \right) = 0.0350 \text{ moles F}_2$$

$$m = n \cdot M = (0.0350 \text{ moles})(38.0 \text{ g/mol}) = 1.33 \text{ g F}_2$$

0.851 g of Mg is formed at the cathode and 1.33 g of  $\text{F}_2$  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?



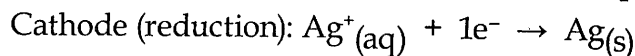
$$\text{moles Au} = \frac{m}{M} = \frac{2.50 \text{ g}}{197.0 \text{ g/mol}} = 0.01269 \text{ moles Au}$$

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

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

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

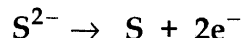
3. If 10.0 g of sulfur is deposited on an electrode in an electrolytic cell from silver sulfide, calculate the mass of silver deposited on the other electrode.



Find the moles of sulfur deposited:

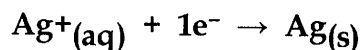
$$\text{moles S} = \frac{m}{\text{molar mass}} = \frac{10.0 \text{ g}}{32.1 \text{ g/mol}} = 0.3115 \text{ moles S deposited}$$

Find moles of  $\text{e}^-$  needed:



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

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



0.623 moles Ag produced

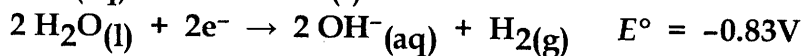
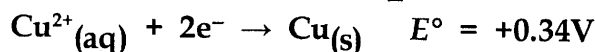
$$m = nm (\text{molar mass}) = (0.623 \text{ moles})(107.9 \text{ g/mol}) = 67.2 \text{ g}$$

67.2 g of silver is deposited.

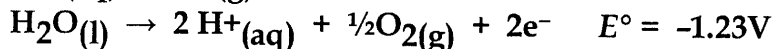
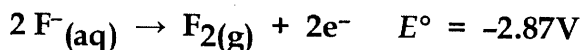
4. What products would be expected at each electrode? Write the half reactions and calculate the voltage required to electrolyze the following:
- a. aqueous copper (II) fluoride

Species present:  $\text{Cu}^{2+}$ ,  $\text{F}^-$  and  $\text{H}_2\text{O}$

At the cathode, either  $\text{Cu}^{2+}$  or  $\text{H}_2\text{O}$  will be reduced.



At anode, either  $\text{F}^-$  or  $\text{H}_2\text{O}$  will be oxidized.



From the half-cell potentials, we can predict

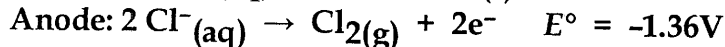
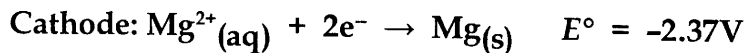
Cathode: copper

Anode: oxygen gas and  $\text{H}^+$  ions

$$E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}} = -1.23\text{V} + (+0.34) = -0.89\text{V}$$

At least 0.89 V is needed.

b. molten magnesium chloride

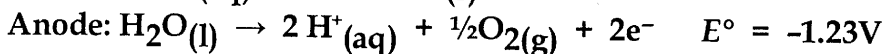
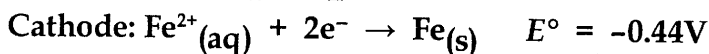


$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = -1.36\text{V} + (-2.37\text{V}) = -3.73\text{V}$$

At least 3.73V is needed.

c. aqueous iron (II) sulfate

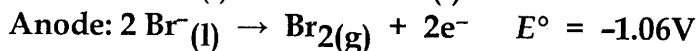
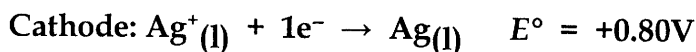
Species:  $\text{Fe}^{2+}$ ,  $\text{SO}_4^{2-}$ ,  $\text{H}_2\text{O}$



$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = -1.23\text{V} + (-0.44\text{V}) = -1.67\text{V}$$

Minimum voltage required is 1.67V.

d. molten silver bromide

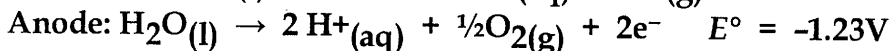


$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = -1.06\text{V} + (+0.80\text{V}) = -0.26\text{V}$$

At least 0.26V is needed.

e. aqueous aluminum chloride

Species:  $\text{Al}^{3+}$ ,  $\text{Cl}^{-}$ ,  $\text{H}_2\text{O}$



$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{ox}} + E^{\circ}_{\text{red}} = -1.23\text{V} + (-0.83\text{V}) = -2.06\text{V}$$

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?

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

28950 s are needed.