

- 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.

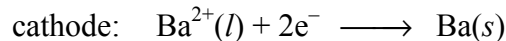
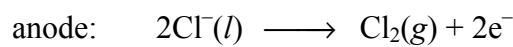
- 45 **The half-reaction at an electrode is: $\text{Mg}^{2+}(\text{molten}) + 2\text{e}^- \rightarrow \text{Mg}(\text{s})$. Calculate the number of grams of magnesium that can be produced by supplying 1.00 F to the electrode.**

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

- 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:



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

(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.

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

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:

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

- 49 **One of the half-reactions for the electrolysis of water is: $2\text{H}_2\text{O}(l) \rightarrow \text{O}_2(g) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$. 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

$$n = \frac{PV}{RT} = \frac{\left(755 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}}\right)(0.076 \text{ L})}{(0.0821 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(298 \text{ K})} = 3.1 \times 10^{-3} \text{ mol O}_2$$

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

$$(3.1 \times 10^{-3} \text{ mol O}_2) \times \frac{4 F}{1 \text{ mol O}_2} = 0.012 F$$

- 57 **The passage of a current of 0.750 A for 25.0 min deposited 0.369 g of copper from a CuSO₄ 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:

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