

Applications of Electrochemistry – Faraday’s Law

Faraday’s Law: the amount of a substance produced or consumed in an electrolysis reaction is directly proportional to the quantity of electricity that flows through the circuit.

Electrical measurements:

Electric current – flow of electrons through an external circuit, I
 - measured in units of *Ampere* (A)

Electric charge – quantity of electricity, Q
 - the product of the current flowing through a circuit by the amount of time it flows
 - measured in units of *Coulomb* (C)

Coulomb: the quantity of electricity that flows through a circuit in one second if the current is one ampere ie. $1 A = 1C/s$

This relationship can also be expressed mathematically by:

$$\text{Charge (coulomb)} = \text{current (in ampere)} \times \text{time (in seconds)}$$

or

$$Q = It$$

Example:

Find the quantity of electricity that results from a current of 3.50 A flowing for 6.00 min.

$$Q = It$$

$$Q = 3.50 \text{ A} \times 6.00 \text{ min} \times (60 \text{ s} / 1 \text{ min})$$

$$Q = 1.26 \times 10^3 \text{ C}$$

Electrochemical Stoichiometry:

The equation for an electrode reaction relates the number of moles of substance consumed or produced during the reaction to the quantity of charge that passes through the cell

$\text{Ag}^+_{(\text{aq})} + 1\text{e}^- \rightarrow \text{Ag}_{(\text{s})}$ In order to produce one mole of solid silver one mole of electrons must be picked up by one mole of silver ions.

$\text{Au}^{3+}_{(\text{aq})} + 3\text{e}^- \rightarrow \text{Au}_{(\text{s})}$ To produce one mole of solid gold three moles of electrons must be picked up by one mole of gold ions.

- To perform stoichiometric calculations for electrochemical and electrolytic cells we also need to know the charge on one mole of electrons.
- The charge on one mole of electrons is one Faraday (1 F) or 96500 C
- Normally, the rounded number **96 500 C/mol** is used in calculations involving Faraday’s constant. Note that this number only has *three* significant figures.

$$\text{Moles of electrons, } n_e = Q/F \quad \text{ie. } n = \text{charge} / \text{charge per mole of electrons}$$

- Since $Q = It$ then $n_e = (It)/F$

$$\text{Ie. } n_e = \text{current(A)} \times \text{time} / \text{charge per mole of electrons}$$

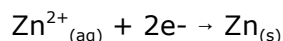
To Calculate the Mass of an Electrolysis Product:

Steps:

1. Find the total moles of electrons that passed through the circuit using $n_e = (I t)/F$ or $n_e = Q/F$
2. Find the total moles of metal produced using the half-reaction and mole ratio $n_R = n_G \times R/G$
3. Convert the moles of metal to mass using $m = nM$

Example 1:

Calculate the mass of zinc plated onto the cathode of an electrolytic cell by a current of 750 mA in 3.25 h.



$$I = 750 \text{ mA} = 0.750 \text{ A}$$

$$t = 3.25 \text{ h} \times (60 \text{ min}/1 \text{ h}) \times (60 \text{ s} / 1 \text{ min}) = 1.17 \times 10^4 \text{ s}$$

$$F = 96\,500 \text{ C/mol}$$

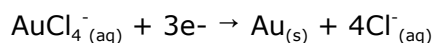
$$\begin{aligned} \text{Step 1: } n_e &= (I t) / F = (0.750 \text{ A} \times 1.17 \times 10^4 \text{ s}) / 96\,500 \text{ C/mol} \\ &= 0.0909 \text{ mol e}^- \quad (\text{this is the total moles of electrons}) \end{aligned}$$

$$\begin{aligned} \text{Step 2: } & \qquad \qquad \qquad \mathbf{G} \quad \mathbf{R} \\ & \text{Zn}^{2+}_{(\text{aq})} + 2\text{e}^- \rightarrow \text{Zn}_{(\text{s})} \\ n_R &= n_G \times R/G = 0.0909 \text{ mol e}^- \times 1/2 \\ n_{\text{Zn}_{(\text{s})}} &= 0.0455 \text{ mol of Zn produced} \end{aligned}$$

$$\begin{aligned} \text{Step 3: } m &= nM \\ m &= 0.0455 \text{ mol} \times 65.39 \text{ g/mol} \\ m &= \mathbf{2.98 \text{ g}} \end{aligned}$$

Example 2

How many grams of gold will be deposited when 5000 C pass through a solution by the following equation.



$$Q = 5000 \text{ C}$$

$$F = 96\,500 \text{ C/mol}$$

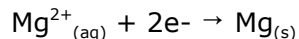
$$\begin{aligned} \text{Step 1: } n_e &= Q/F \\ n &= \frac{5000 \text{ C}}{96\,500 \text{ C/mol}} \\ n &= 0.0518 \text{ moles of electrons} \end{aligned}$$

$$\begin{aligned} \text{Step 2: } & \qquad \qquad \qquad \mathbf{G} \quad \mathbf{R} \\ & \text{AuCl}_4^{-}_{(\text{aq})} + 3\text{e}^- \rightarrow \text{Au}_{(\text{s})} + 4\text{Cl}^{-}_{(\text{aq})} \\ n_R &= n_G \times R/G = 0.0518 \text{ mol e}^- \times 1/3 \\ n_{\text{Au}_{(\text{s})}} &= 0.0173 \text{ mol of Au}_{(\text{s})} \text{ will be produced} \end{aligned}$$

$$\begin{aligned} \text{Step 3: } m &= nM \\ m &= 0.0173 \text{ mol} \times 196.97 \text{ g/mol} \\ m &= \mathbf{3.41 \text{ g}} \end{aligned}$$

Example 3:

Calculate the moles of Mg produced when a current of 60 amps is passed through a magnesium chloride solution for 4.00 h.



$$I = 60 \text{ A}$$

$$t = 4.00 \text{ h}$$

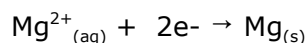
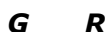
$$F = 96\,500 \text{ C/mol}$$

Step 1: $n_e = It/F$

$$n = \frac{(60 \text{ A})(1.44 \times 10^4 \text{ s})}{96\,500 \text{ C/mol}}$$

$$n = 8.95 \text{ mol of e}^-$$

Step 2:



$$n_R = n_G \times R/G = 8.95 \text{ mol e}^- \times 1/2$$

$$n_{\text{Mg}_{(\text{s})}} = 4.48 \text{ mol of Mg produced}$$

Step 3:

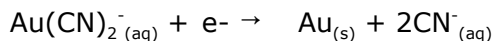
$$m = nM$$

$$m = 4.48 \text{ mol} \times 24.32 \text{ g/mol}$$

$$\mathbf{m = 109 \text{ g}}$$

Example 3:

How many hours would it take for a 3.50 A current to electroplate 129 g of gold according to the following equation.



$$I = 3.50 \text{ A}$$

$$m = 129 \text{ g}$$

$$F = 96\,500 \text{ C/mol}$$

$$t = ?$$

To solve for time, rearrange the equation $n = It/F$ such that $t = nF/I$ and convert mass to moles.

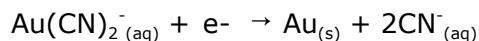
Step 1: *Convert mass of gold to moles of gold*

$$\mathbf{n = m/M}$$

$$n = 129 / 196.97 \text{ g/mol}$$

$$n = 0.655 \text{ mol of gold}$$

Step 2: *Convert moles of gold to moles of electrons*



$$n_R = n_G \times R/G = 0.655 \text{ mol e}^- \times 1/1$$

$$n = 0.655 \text{ mol e}^-$$

Step 3: *Determine the time*

$$t = n_e F / I$$

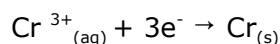
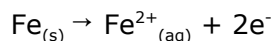
$$t = \frac{0.655 \text{ mol} \times 96\,500 \text{ C/mol}}{3.50 \text{ A}}$$

$$t = 1.81 \times 10^5 \text{ s}$$

$$t = \mathbf{5.02 \text{ h}}$$

ELECTROCHEMISTRY WORKSHEET #7

1. Calculate the minimum time, in seconds, required to deposit 40.0 g of copper at the cathode of an electrolysis cell containing $\text{CuSO}_{4(\text{aq})}$ using a current of 20 000 mA.
2. How many minutes does it take to plate 0.925 g of silver onto the cathode of an electrolytic cell using a current of 1.55 A?
3. The nickel anode in an electrolytic cell decreases in mass by 1.20 g in 35.5 min. The oxidation half-reaction converts nickel atoms to nickel (III) ions. What is the constant current?
4. The following two half-reactions take place in an electrolytic cell with an iron anode and a chromium cathode.



During the process, the mass of the iron anode decreases by 1.75 g.
Find the change in mass of the chromium cathode.

5. A student wishes to set up an electrolytic cell to plate copper onto a belt buckle. Predict the length of time it will take to plate out 2.5 g of copper from a copper (II) nitrate solution using a 2.5 A current. At which electrode should the belt buckle be attached?
6. Determine the mass of chlorine produced when a 200 A current flows for 24.0 h through a cell containing molten sodium chloride. At which electrode is the chlorine produced?
7. A trophy company is setting up a nickel plating cell using an electrolyte containing nickel (III) ions. Predict the current required to produce metal at a rate of 5.00 g/min.
8. In the electrolysis of a molten group II metal chloride 2.50 A of constant current is passed through a cell for 1.28 hours. Use the information provided below to determine the identity of the group II metal.

Mass of cathode before application of current:	25.720 g
Mass of cathode after application of current:	30.949 g