

### Lesson 3 –electrolysis-Faraday’s Laws

[Click](#) for a detailed look at Faraday.

Faraday’s first law states that the mass of metal deposited at the electrode of an electrolytic cell is directly proportional to the electrical charge passed through the cell.

Charge (coulombs) = current (amps) X time (seconds)

$$\Rightarrow Q = I \times t$$

The charge of one mol of electrons is 96,500 C. An amount of charge of 96,500 C is known as a Faraday, hence  $1F = 96,500 \text{ C}$  and represents one mol of electrons.

This expression enables us to calculate the mole of electrons delivered in an electrolytic cell and to complete some very useful stoichiometric calculations using balanced redox half equations.

Example 1 At the cathode of an electrolytic cell silver metal is deposited according to the half equation  $\text{Ag}^+(\text{aq}) + e \Rightarrow \text{Ag}(\text{s})$ . A steady current of 20.0 Amps was applied for 30.0 minutes. What mass of silver was deposited?

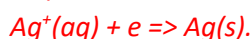
*Step 1 Calculate the charge in C*

$$\Rightarrow Q = I \times t = 20.00 \times 30.0 \times 60 = 36000\text{C}$$

*Step 2 Calculate the mol of electrons*

$$\Rightarrow 3600 / 96500 = 3.73 \times 10^{-1} \text{ mol}$$

*Step 3 Calculate the mol of silver deposited*



*According to the stoichiometry for every one mol of electrons delivered one mol of silver metal is deposited.*

$$\Rightarrow 3.73 \times 10^{-1} \text{ mol of silver is deposited.}$$

*Step 4 Calculate the mass of silver metal.*

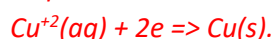
$$\Rightarrow \text{mass} = n_{\text{Silver}} \times 107.9 = 3.73 \times 10^{-1} \times 107.9 = 40.3 \text{ grams}$$

Example 2 At the cathode of an electrolytic cell 50.0 grams copper metal needs to be deposited according to the half equation  $\text{Cu}^{2+}(\text{aq}) + 2e \Rightarrow \text{Cu}(\text{s})$ . How long, in minutes, should a steady current of 20.0 Amps be applied for?

*Step 1 Calculate the mol of copper metal that needs to be deposited.*

$$\Rightarrow n_{\text{copper}} = 50.0 / 63.5 = 0.7874$$

*Step 2 Calculate the mol of electrons needed.*



*For every mol of copper deposited two mol of electrons are needed.*

$$\Rightarrow 2 \times 0.7874 = 1.575$$

*Step 3 Calculate the charge that 1.575 mol of electrons represents.*

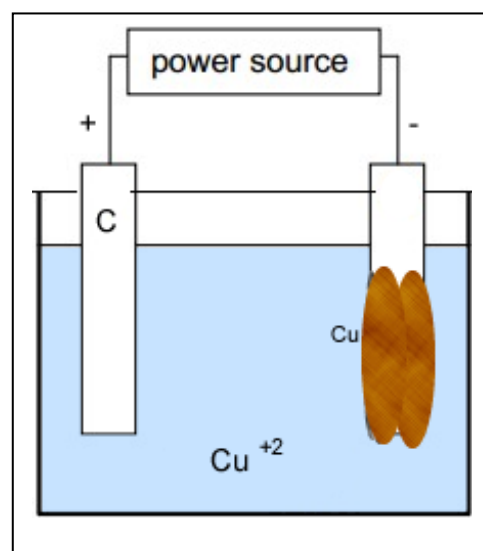
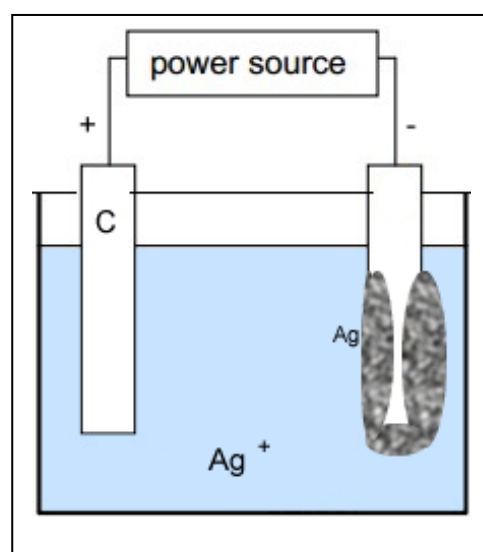
$$\Rightarrow 1.575 \times 96500 = 151988\text{C}$$

*Step 4 Calculate the time in seconds*

$$\Rightarrow Q = It$$

$$\Rightarrow Q/I = t$$

$$\Rightarrow 151988/20.0 = 7599.4 \text{ seconds} = 7599.4 / 60 = 127 \text{ minutes}$$



Example 3. How many Faradays of charge are required to pass through the electrolytic cell to produce 7.10 grams of  $\text{Cl}_2$  gas in a Down's cell using molten  $\text{NaCl}$ ?

*Step 1 Calculate the mol of  $\text{Cl}_2$  gas*

$$\Rightarrow 7.10/71.0 = 0.100 \text{ mol of } \text{Cl}_2$$

*Step 2 Calculate the mol of electrons needed. Write the half equation for the formation of  $\text{Cl}_2$  at the anode.*



*From the stoichiometry we see that for every mol of  $\text{Cl}_2$  formed 2 mol of electrons are also produced.*

*$\Rightarrow 0.200$  mol of electrons must flow through the cell.*

*Step 3 Hence 0.200 Faradays of charge is required.*

Example 4. A steady current of 50.0 Amps flowed through an electrolytic cell for 30.0 minutes to deposit 21.45 grams of an unknown metal whose ion has a 1+ charge. Identify this metal.

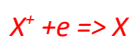
*Step 1 Calculate the mol of electrons that passed through the cell.*

$$\Rightarrow Q = It = 50.0 \times 30.0 \times 60 = 90000 \text{ C}$$

$$\Rightarrow n_{\text{electrons}} = 90000 / 96500 = 0.9326$$

*Step 2 Find the mol of metal deposited.*

*Since the charge on the metal ion is 1+ we can write the following balanced half equation.*



*Hence for every mol of electrons we produce one mol of metal.*

*$\Rightarrow 0.9326$  mol of metal*

*Step 3 Find the atomic mass of the metal and hence identify it.*

$$\text{atomic mass} = \text{mass} / \text{mol} = 21.45 / 0.9326 = 23$$

$\Rightarrow \text{Na}$

Example 5 Two cells containing  $\text{Cu}^{2+}$  and  $\text{Cr}^{3+}$  ions respectively are connected in series. The same current runs through both cells. The setup is pictured on the right.

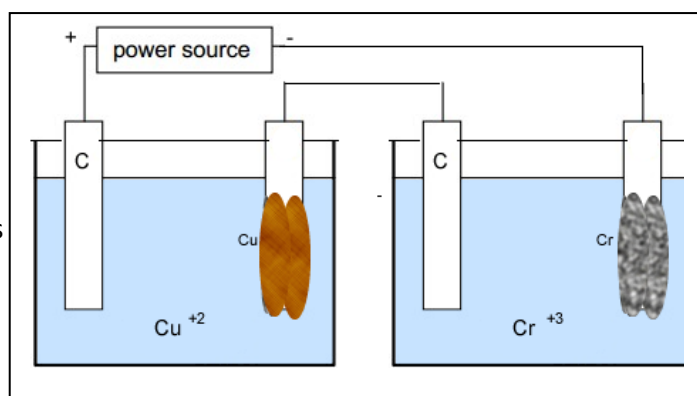
A steady current of 20.0 Amps runs for 12.0 minutes. What mass of each element, in grams, is deposited at the cathode of each cell?

*Step 1 Calculate the mol of electrons that flow through the circuit*

*Step 2 Calculate the mol of copper and chromium that will be deposited.*

*Write balanced half equations for the reduction of each metal.*

*Step 3 Calculate the mass of each metal in grams.*



Example 6 A charge of 0.250 Faraday was passed through an electrolytic cell with 0.100M solution of  $\text{CuSO}_4$ , as shown on the right.

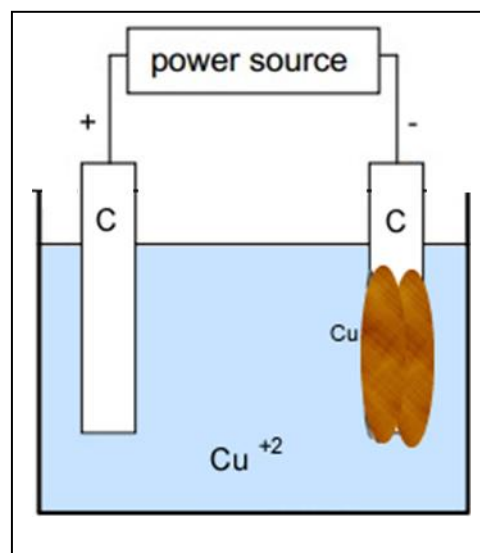
a) Calculate the mass of copper deposited on the cathode.

*Step 1 Calculate the mol of electrons.*

*Step 2 Write the balanced half equation for the reduction of  $\text{Cu}^{2+}$ .*

*Step 3 Calculate the mol of CU formed on the cathode.*

*Step 4 Calculate the mass of copper.*



b) The same electrolytic cell was used by a student to produce copper metal. A current of 25.0 Amps was passed through the cell for 12 hours. What mass of copper was deposited?

*Step 1 Find the charge delivered*

$$\Rightarrow Q = It$$

*Step 2 Calculate the mol of electrons delivered.*

$$\Rightarrow Q/96500$$

*Step 3 Write the balanced half equation for the reduction of  $\text{Cu}^{2+}$ .*

*Step 4 Calculate the mol of Cu formed*

*Step 5 Calculate the mass of copper.*

c) Another student used the electrolytic cell shown on the right to deposit copper at the cathode. This cell used a 10.00 litre 1.10 M  $\text{CuSO}_4$  solution. A current of 20.0 Amps was passed through the cell for 8.00 hours. What was the final concentration of the electrolyte after the current was turned off?

*Step 1 find the charge delivered*

$$\Rightarrow Q = It$$

*Step 2 Find the mol of electrons delivered.*

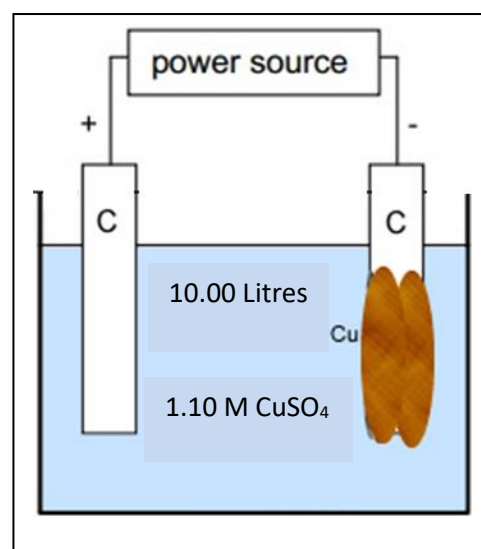
$$\Rightarrow Q/96500$$

*Step 3 Write the balanced half equation for the reduction of  $\text{Cu}^{2+}$ .*

*Step 4 Calculate the mol of Cu formed*

*Step 5 Subtract the mol of Cu formed from the mol of  $\text{Cu}^{2+}$  originally in the solution.*

*Step 6 calculate the concentration.*



Example 7 Lithium metal is produced by the electrolysis of molten lithium chloride, LiCl. Calculate the mass of lithium metal produced in 36.0 hours using a current of 6.80 amps to the right number of significant figures.

*Example 8* Corrosion of an iron pipe can be prevented by connecting it to a magnesium bar buried in the ground. The magnesium corrodes in preference to the iron. If the average current flowing between the two metals is  $2.5 \times 10^{-6}$  A, what is the amount of magnesium metal, in mol, reacting each second. Give the answer to the right number of significant figures.

Example 9 A mass of 0.850 g of zinc is produced in 35.0 minutes. Calculate the electric current, in amps, supplied to the cell during the electrolysis. Express your answer to an appropriate number of significant figures.