Lesson 3 –electrolysis-Faraday's Laws

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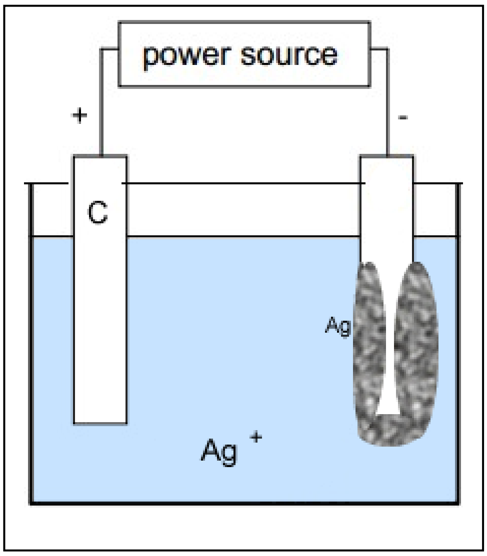
Faraday's first law states that the mass of metal deposited at the electrode of an electrolytic cell is directly proportional to the electrical charge passed through the cell.
Charge (coulombs) = current (amps) X time (seconds)
=> Q = I X t

The charge of one mol of electrons is 96,500 Q. An amount of charge of 96,500Q is known as a Faraday, hence 1F = 96,500Q 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 $Ag^{+}(aq) + e \Rightarrow Ag(s)$. A steady current of 20.0 Amps was applied for 30.0 minutes. What mass of silver was deposited?



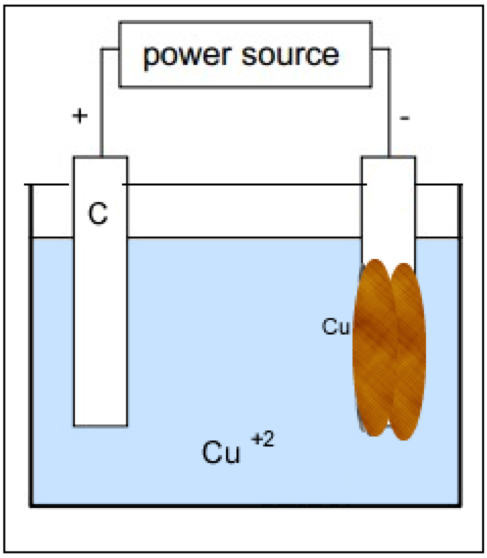
Step 1 Calculate the charge in Q
=> Q = I X t = 20.00 X 30.0 X 60 = 36000 Q

Step 2 Calculate the mol of electrons
=> 36000 / 96500 = 3.73×10^{-1} *mol*

Step 3 Calculate the mol of silver deposited
$Ag^{+}(aq) + e \Rightarrow Ag(s)$.
According to the stoichiometry for every one mol of electrons delivered one mol of silver metal is deposited.
=> 3.73×10^{-1} *mol of silver is deposited.*

Step 4 Calculate the mass of silver metal.
=> mass = $n_{Silver} \times 107.9 = 3.73 \times 10^{-1} \times 107.9 = 40.3$ *grams*

Example 2 At the cathode of an electrolytic cell 50.0 grams copper metal needs to be deposited according to the half equation $Cu^{+2}(aq) + 2e \Rightarrow Cu(s)$. 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.
=> $n_{copper} = 50.0 / 63.5 = 0.7874$
Step 2 Calculate the mol of electrons needed.
$Cu^{+2}(aq) + 2e \Rightarrow Cu(s)$.
For every mol of copper deposited two mol of electrons are needed.
=> 2 X 0.7874 = 1.575
Step 3 Calculate the charge that 1.575 mol of electrons represents.
=> 1.575 X 96500 = 151988Q
Step 4 Calculate the time in seconds
=> Q = It
=> Q/I = t
=> 151988/20.0 = 7599.4 seconds = 7599.4 / 60 = 127 minutes

Example 3. How many Faradays of charge are required to pass through the electrolytic cell to produce 7.10 grams of Cl_2 gas in a Down's cell using molten NaCl?

Step 1 Calculate the mol of Cl_2 gas
=> 7.10/71.0 = 0.100 mol of Cl_2

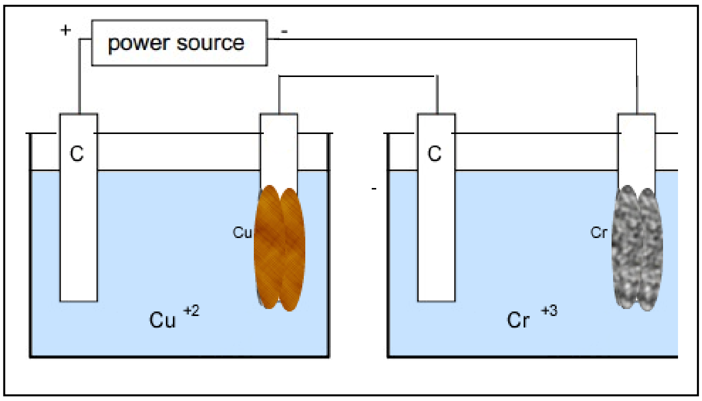
Step 2 Calculate the mol of electrons needed. Write the half equation for the formation of Cl_2 at the anode.
$2Cl^-(l)$ => $Cl_2(g)$ + 2e
From the stoichiometry we see that for every mol of Cl_2 formed 2 mol of electrons are also produced.
=> 0.200 mol of electrons must flow through the cell.
Step 3 Hence 0.200 Faradays are 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.
=> Q = It = 50.0 X 30.0 X 60 = 90000 C
=> $n_{electrons}$ = 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.
X^+ +e => X
Hence for every mol of electrons we produce one mol of metal.
=> 0.9326 mol of metal
Step 3 Find the atomic mass of the metal and hence identify it.
atomic mass = mass / mol = 21.45 / 0.9326 = 23

⇨ *Na*

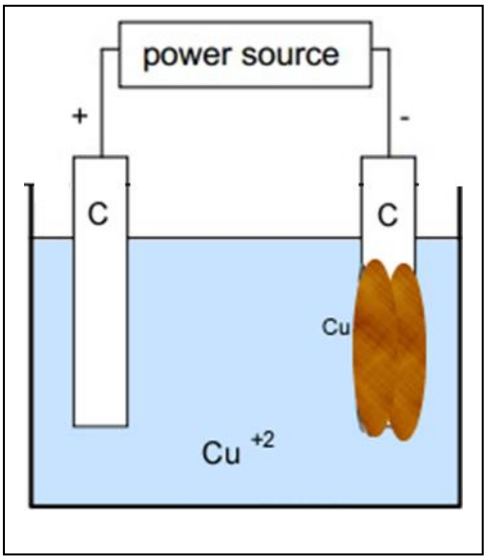
Example 5 Two cells containing Cu^{2+} and Cr^{3+} ions respectively are connected in series. The same current runs through both cells. The setup is pictured on the right.



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
=> Q = It = 20.0 X 12.0 X 60 = 14400
=> $n_{electrons}$ =14400/96500 = 0.1492
Step 2 Calculate the mol of copper and Chromium that will be deposited.
Write balanced half equations for the reduction of each metal.
$Cu^{2+}(aq)$ + 2e => Cu(s)
$Cr^{3+}(aq)$ + 3e => Cr(s)
n_{copper} = 0.1492 / 2 = 0.0746
$n_{chromium}$ = 0.1492 / 3 = 0.0497
Step 3 Calculate the mass of each metal in grams.
Mass of copper = 0.0746 X 63.5 = 4.74 grams, Mass of Cr = 0.0497 X 52.0 = 2.58 grams

Example 6 A charge of 0.250 Faraday was passed through an electrolytic cell with 0.100M solution of $CuSO_4$, as shown on the right.

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



Step 1 Calculate the mol of electrons.
=> $n_{electrons} = 0.250$

Step 2 Write the balanced half equation for the reduction of Cu^{2+}.
$Cu^{2+}(aq) + 2e \Rightarrow Cu(s)$

Step 3 Calculate the mol of Cu formed on the cathode.
=> $n_{copper} = 0.250 / 2 = 0.125$

Step 4 Calculate the mass of copper.
=> 0.125 X 63.5 = 7.94 grams

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
=> Q = It
=> Q = 25.0 X 12 X 60 X 60 = 1080000C
Step 2 Calculate the mol of electrons delivered.
=> Q/96500
=>1080000/96500 = 11.12 = 11 (2 sig figs)

Step 3 Write the balanced half equation for the reduction of Cu^{2+}.
$Cu^{2+}(aq) + 2e \Rightarrow Cu(s)$

Step 4 Calculate the mol of Cu formed
=> $n_{copper} = 11 / 2 = 5.5$
Step 5 Calculate the mass of copper.
=> 5.5 X 63.5 = 350 grams (2 sig figs)

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 $CuSO_4$ 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
=> Q =It
=> Q = 20.0 X 8.00 X 60 X 60 = 576000C
Step 2 Find the mol of electrons delivered.
=> Q/96500
=> 576000/96500 = 5.97
Step 3 Write the balanced half equation for the reduction of Cu^{2+}.
$Cu^{2+}(aq) + 2e \Rightarrow Cu(s)$

Step 4 Calculate the mol of Cu formed

=> 5.97 / 2

=> 2.98 mol

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

=> mol of Cu^{2+} present initially = C X V = 1.10 X 10.00 = 11.0 mol

=> mol of Cu^{2+} used = 2.98

=> mol of Cu^{2+} remaining = 11.0 – 2.98 = 8.0

Step 6 calculate the concentration.

=> 8.0 / 10.00 = 0.80 M

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.

Step 1 Write the reduction reaction for Li^+

=> $Li^+(l)$ + e => Li(l)

Step 2 Calculate the amount of charge delivered. Be careful to convert hours into seconds.

=> Q = It = 6.80 X 36.0 X 60 X 60 = 881280C

Step 3 Find the mol of electrons

=> 881280/96500 = 9.132

Step 4 Find the mol of Li produced

=> according to the stoichiometry the mol ratio of electrons to lithium deposited is 1:1

=> 9.132

Step 5 Calculate the mass of Li

=> mass = n X 6.9 = 9.132 X 6.9 = 63 grams

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 2.5×10^{-6} A, what is the amount of magnesium metal, in mol, reacting each second. Give the answer to the right number of significant figures.

Step 1 Write the balanced half equation for the reduction of magnesium

=> $M^{2+}(aq)$ + 2e => Mg(s)

Step 2 Calculate the charge delivered every second

=> Q = 2.5 X 10^{-6} X 1 = 2.5 X 10^{-6} C(the number 1 is a constant and as such can have as many siginificant figures as required.)

Step 3 Calculate the mol of electrons delivered every second.

=> $n_{electrons}$ = 2.5 X 10^{-6} / 96500 = 2.6 X 10^{-11} mol

Step 4 calculate the mol of magnesium reacting

=> ½ X 2.6 X 10^{-11} mol = 1.3 X 10^{-11}mol (2 sig figs)

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

Step 1 write a balanced half equation for the reduction of Zn^{2+}

Step 2 calculate the mol of Zn => 0.850 / 65.4 = 0.012997

Step 3 calculate the mol of electrons 0.01339 X 2 = 0.02599

Step 4 calculate the charge 0.02677 X 96500 = 2508 C

Stpe 5 calculate the current Q/t =I => 2508 / (35.0 X 60) = 1.19 A (3 sig figs)