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 current, *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:

Charge (coulomb) = current (in ampere) x 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 A x 6.00 min x (60 s / 1 min)
Q = 1.26×10^3 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

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

$Au^{3+}_{(aq)} + 3e^- \rightarrow Au_{(s)}$ — 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.

Moles of electrons, n_e = Q/F ie. n = charge/ charge per mole of electrons

- Since ***Q = It*** then ***n_e = (I t)/F***

  Ie. n_e = current(A) x time/ 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 x 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.

$Zn^{2+}_{(aq)} + 2e- \rightarrow Zn_{(s)}$

I = 750 mA = 0.750 A
t = 3.25 h x (60 min/1 h) x (60 s / 1 min) = 1.17×10^4 s
F = 96 500 C/mol

Step 1: **n_e = (I t) / F** = (0.750 A x 1.17×10^4 s) / 96 500 C/mol
= 0.0909 mol e^- (this is the total moles of electrons)

Step 2: ***G*** ***R***

$Zn^{2+}_{(aq)} + 2e- \rightarrow Zn_{(s)}$

n_R = n_G x R/G = 0.0909 mol e- x 1/2
$n_{Zn(s)}$ = 0.0455 mol of Zn produced

Step 3: **m = nM**
m = 0.0455 mol x 65.39 g/mol
m = 2.98 g

Example 2
How many grams of gold will be deposited when 5000 C pass through a solution by the following equation.

$AuCl_4^-{}_{(aq)} + 3e- \rightarrow Au_{(s)} + 4Cl^-_{(aq)}$

Q = 5000 C
F = 96 500 C/mol

Step 1: **n_e = Q/F**

$n = \frac{5000\ C}{96\ 500 \times C/mol}$

n = 0.0518 moles of electrons

Step 2: ***G*** ***R***

$AuCl_4^-{}_{(aq)} + 3e- \rightarrow Au_{(s)} + 4Cl^-_{(aq)}$

n_R = n_G x R/G = 0.0518 mol e- x 1/3
$n_{Au(s)}$ = 0.0173 mol of $Au_{(s)}$ will be produced

Step 3:
m = nM
m = 0.0173 mol x 196.97 g/mol
m = 3.41 g

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.

$Mg^{2+}_{(aq)} + 2e- \rightarrow Mg_{(s)}$

I = 60 A F = 96 500 C/mol
t = 4.00 h

Step 1: $\mathbf{n_e = It/F}$

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

n = 8.95 mol of e-

Step 2: ***G*** ***R***

$Mg^{2+}_{(aq)} + 2e- \rightarrow Mg_{(s)}$

$\mathbf{n_R = n_G \times R/G}$ = 8.95 mol e- x 1/2
$n_{Mg(s)}$ = 4.48 mol of Mg produced

Step 3: m = nM
m = 4.48 mol x 24.32 g/mol
m = 109 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.

$Au(CN)_2^-{}_{(aq)} + e- \rightarrow Au_{(s)} + 2CN^-_{(aq)}$

I = 3.50 A
m = 129 g
F = 96 500 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*

n = m/M
n = 129/ 196.97 g/mol
n =0.655 mol of gold

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

$Au(CN)_2^-{}_{(aq)} + e- \rightarrow Au_{(s)} + 2CN^-_{(aq)}$

$\mathbf{n_R = n_G \times R/G}$ = 0.655 mol e- x 1/1
n = 0.655 mol e-

Step 3: *Determine the time*

$t = n_eF/I$

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

$t = 1.81 \times 10^5$ s
t = ***5.02 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 $CuSO_{4(aq)}$ 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.

   $Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$

   $Cr^{3+}_{(aq)} + 3e^- \rightarrow Cr_{(s)}$

   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