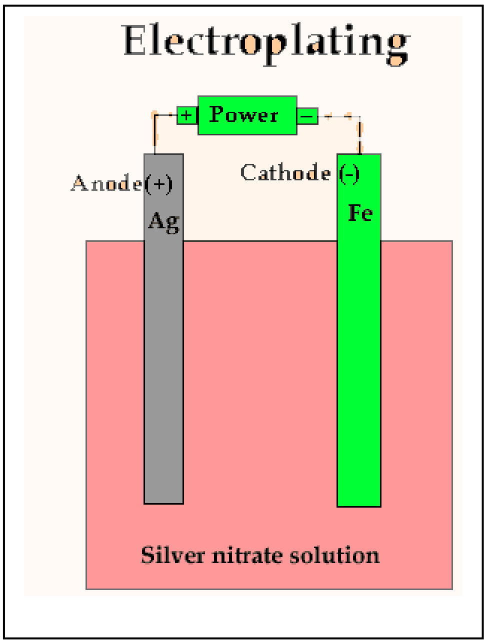
Lesson 4 electrolysis- electroplating

Electroplating involves the deposition of a layer of metal on top of another metal. Often the silverware you see on dinner tables is cheap iron coated with a thin layer of silver. This is done by a process called electroplating in an electrolytic cell. A normal electrolytic cell is used for electroplating with the following conditions.

- The metal to be coated is placed at the negative terminal (cathode)
- The metal we wish to coat with is placed at the positive terminal (anode)
- The electrolyte is a solution of positive metal ions of the metal we wish to coat with.
- The concentration of the metal ion in the electrolyte stays constant during the electrolysis process.



As is shown on the right, If we wish to coat iron metal with a layer of silver then the electrolyte used should contain silver ions and the anode should be silver metal, while the iron is connected to the negative terminal.
Electrons are pumped out of the silver metal in effect corroding it and releasing silver ions into the electrolyte solution.

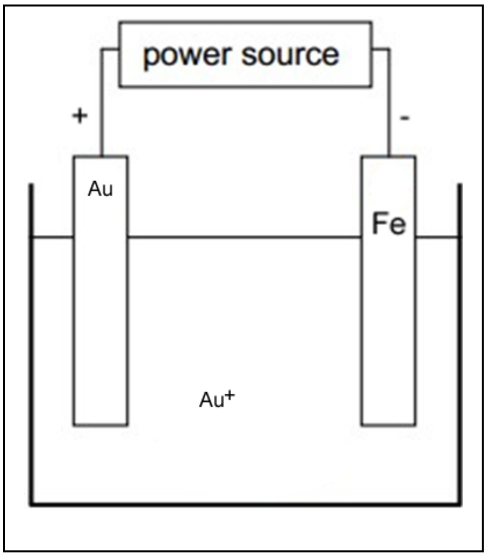
$$Ag_{(s)} \Rightarrow Ag^{+}_{(aq)} + e$$

The positive ions migrate over to the negative electrode (the metal we wish to coat) and regain their electrons to form silver metal once again.

$$Ag^{+}_{(aq)} + e \Rightarrow Ag_{(s)}$$

The ions that migrate are not always the ones that have just being released from the silver metal but

Example 1 A medal is gold plated in an electrolytic cell. The following information is obtained
Mass of medal before gold plating = 23.2 g
Current = 0.900 A
Mass of anode before gold plating = 30.0 g
Time the current was turned on for = 23.0 minutes



a) Calculate, the mass of gold deposited on the medal.
Step 1 Calculate the mol of electrons delivered in 23.0 minutes
=> Q = It = 0.900 X 23.0 X 60 = 1242 C
=> $n_{electrons}$ = 1242 / 96500 = 0.1287
Step 2 write the balanced half equation for the reaction occurring at the cathode. Au(aq) + e => Au(s)
Step 3 Calculate the mol of Au deposite
=> 0.1287
Step 4 Calculate the mass of Au
=> 0.1287 X 197 = 2.54g
b) Write the balanced chemical equation taking place at the anode. *$Au(s) \Rightarrow Au^{+}(aq) + e$*
c) Calculate the final mass of the anode. *The mass of gold deposited at the cathode is equal to the mass lost at the gold anode, 2.54 grams.*

d) Calculate the time, in minutes, required to deposit 4.52 grams of gold if a current of 1.10A is used.

Step 1 Calculate the mol of gold

=> 4.52 /197 = 0.02294

Step 2 calculate the mol of electrons needed

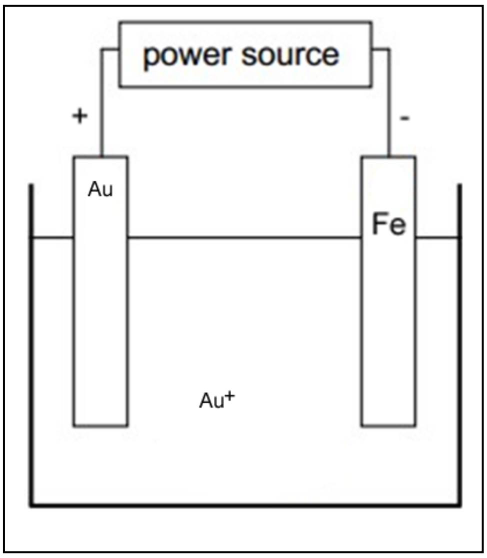
=> $n_{electrons}$ = 0.02294

Step 3 calculate the charge delivered.

=> 0.02294 X 96500 = 2213 C

Step 4 Calculate the time needed

=> t = 2214 / 1.10 = 2013 s = 33.5 min.



Example 2 An electroplating process uses a solution of chromium(III) sulfate, $Cr_2(SO_4)_3$, to deposit a thin layer of chromium on the surface of an object. A current of 6.00 A is maintained. How long does it take, in seconds, to deposit 0.0202 mol chromium onto the surface?

Step 1 Write the balanced chemical equation for the reaction that takes place at the cathode

=> Cr^{3+}(aq) + 3e => Cr(s)

Step 2 Calculate the mol of electrons needed.

=> $n_{electrons}$ = 3 X 0.0202 = 0.0606

Step 3 calculate the charge delivered.

=> Q = n X 96500 = 0.0606 X 96500 = 5848 C

Step 4 Calculate the time taken in seconds

=> Q/I =t

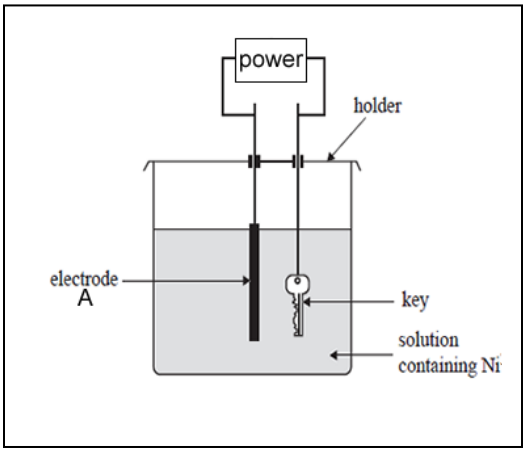
=> 5848 / 6.00 = 975 s

Example 3. An iron key of mass 15.34 grams is to be coated with a layer of nickel. The electrolytic cell, shown on the right was set up. A current of 0.990 A was applied for 20.00 minutes. The key was then removed, dried and reweighed.

a) What is the polarity of electrode "A"? *positive.*

b) What should electrode A be composed of? *Nickel metal*

c) What was the final mass of the key, given to the right number of significant figures?



Step 1 Write a balanced half equation for the reaction occurring at the cathode => Ni^{3+} + 3e => Ni(s)

Step 2 Calculate the mol of electrons delivered

=> calculate charge --- Q = It = 0.990A X 20.00 X 60 = 1188

=> $n_{electrons}$ = 1188 / 96500 = 0.01231 mol

Step 3 Calculate the mol of Ni deposited.

=> 0.01231 / 3 = 0.00410

Step 4 Calculate the mass of Ni deposited and hence the final mass of the key.

=> mass of Ni = 0.00410 X 58.7 = 0.241 grams , hence final mass of key is

=> 15.34 + 0.241 = 15.58 grams

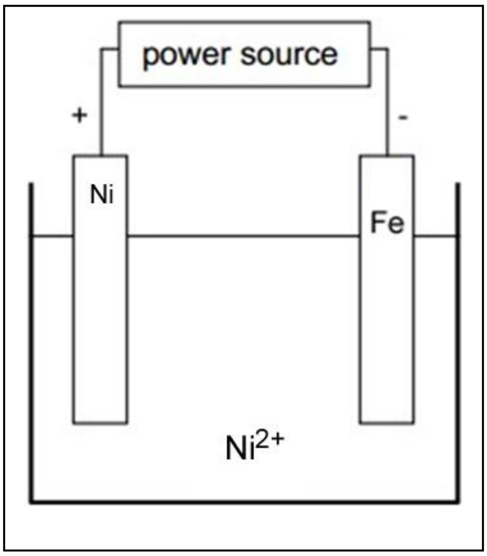
1) A medal is nickel plated in an electrolytic cell. The following information is obtained

Mass of medal before nickel plating = 19.2 g

Current = 0.870 A

Mass of anode before nickel plating = 20.0 g

Time the current was turned on for = 13.0 minutes



a) Calculate, the mass of nickel deposited on the medal.

Step 1 Calculate the mol of electrons delivered

=> Q = It = 0.870 X 13.0 X 60 = 679 C

=> $n_{electrons}$ = 679 / 96500 = 7.03 X 10^{-3}

Step 2 write the balanced half equation for the reaction occurring at the cathode. Ni^{2+}(aq) + 2e => Ni(s)

Step 3 Calculate the mol of Ni deposite

=> 7.03 X 10^{-3} / 2 = 3.52 X 10^{-3}

Step 4 Calculate the mass of Ni

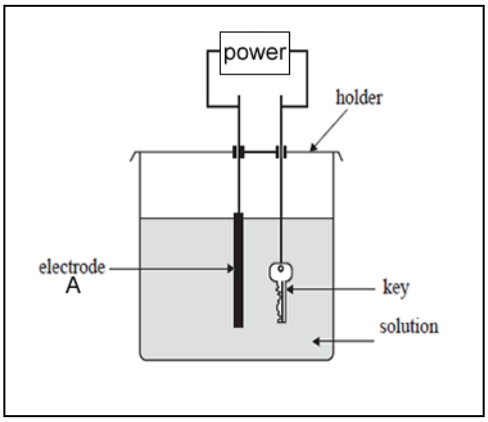
=> 3.52 X 10^{-3} X 58.7 = 0.207g

b) Write the balanced chemical equation taking place at the anode. *Ni(s) => Ni^{2+} (aq) + 2e*

c) Calculate the final mass of the anode. *The mass of nickel deposited at the cathode is equal to the mass lost at the nickel anode, 0.207 grams.*

=> final mass of anode = 20.0 – 0.207 = 19.8 grams

2) A key is to be covered with a fine layer of silver metal in an electrolytic cell as shown on the right.



a) What is electrode A made from? *Silver metal*

b) What is the polarity of electrode A? *positive*

c) Write the equation to the half reaction that occurs at the anode. *Ag(s) => e + Ag^{+} (aq)*

d) What is an appropriate electrolyte? *$AgNO_3$(aq)*

e) What is the mass of silver deposited on the key if a current of 1.04 A was delivered for 18.0 minutes?

Step 1 Write a balanced half equation for the reaction occurring at the cathode => Ag^{+} (aq) + e => Ag(s)

Step 2 Calculate the mol of electrons delivered

=> calculate charge --- Q = It = 1.04 A X 18.00 X 60 = 1123

=> $n_{electrons}$ = 1123 / 96500 = 0.01164 mol

Step 3 Calculate the mol of Ag deposited.=> 0.01164

Step 4 Calculate the mass of Ag deposited

=> mass of Ag = 0.01164 X 107.9 = 1.26 grams

f) Another custom made key is to have a layer of silver 0.203 mm deep coated on a key with surface area 12.92 cm^2. How long, in minutes, should the key be left in the electrolytic cell if a current of 0.981 A is used and the density of silver is 10.48 g/cm^3?

Step 1 Calculate the volume of silver. Be careful to convert units (0.203 mm = 0.0203 cm)

=> volume of Ag to be deposited = 0.0203 X 12.92 = 0.262 cm^3.

Step 2 Calculate the mass of Ag

=> mass = density X volume = 10.48 g/cm^3 X 0.262 = 2.75 grams

Step 3 Calculate the mol of silver.

=> 2.75 / 107.9 = 0.0255 mol

Step 4 Calculate the mol of electrons needed => 0.0255

Step 5 Calculate the charge delivered

=> 0.0255 X 96500 = 2461 C

Step 6 Calculate the time

=> t = Q/I = 2461 / 0.981 = 2508 s = 41.8 minutes