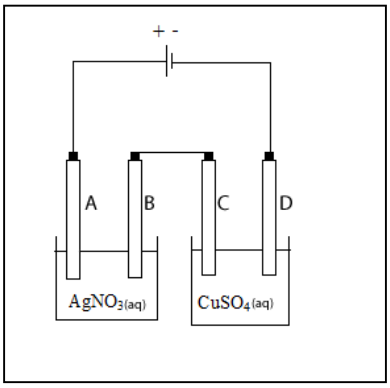
Friday quiz
electrolysis – worksheet 12

In attempt to produce silver and copper metals the electrolytic cells are setup in series as shown on the right.



i. Give the products formed (if any) at each electrode.

A. ____*O_2 and H^+ ions*

B._____*Ag(s)*

C.______ *O_2 and H^+ ions*

D._____*Cu(s)*

ii. Assuming 100% efficiency calculate the mol of silver metal and copper metals deposited if an electric charge of 2.50 faradays is passed through the circuit.

A charge of 2.5 faradays is equivalent to 2.5 mol of electrons.

=> $Ag^+(aq) + e \rightarrow Ag(s)$

=> 2.5 mol of silver is produced at electrode B

=> $Cu^{2+}(aq) + 2e \rightarrow Cu(s)$

=> 1.25 mol of Cu is deposited on electrode D

iii. What will be the change in mass at each electrode? Explain your answer and calculate the mass change.

Electrode B = 2.5 mol of silver will be deposited hence the mass will increase by

=> 2.5 X 107.9 = 270 grams

Electrode C no change in mass

Electrode D = 1.25 mol of copper is deposited hence th mass wil increase by

=> 1.25 X 63.5 = 79.4 grams

Electrode E no change in mass.

iv. Indicate the direction of electron flow.

From electrode D to A

v. If a current of 1.25 A runs through the circuit calculate the time, in hours, required to deposit 0.136 Kg of copper.

Step 1 calculate the mol of copper.

=> 136 / 63.5 = 2.14

Step 2 calculate the mol of electrons necessary to deposit 2.14 mol of copper

=> $Cu^{+2}(aq) + 2e \rightarrow Cu(s)$

=> 4.28 mol of electrons

Step 3 calculate the charge of 4.28 mol of electrons

=> 4.28 X 96500 = 413020 C

Step 4 calculate the time

=> Q = It

=> 413020 = 1.25 X t

=> 413020 / 1.25 = 330416 s = 91.8 hours