Quantitative electrolysis of aqueous copper(II) sulfate – worksheet

Questions after observing the demonstration

1) Calculate the number of moles of copper removed from the anode (weight lost (g)/ A_r[Cu]).

2) Calculate the charge which has been passed through the electrolyte in Coulombs (current (A) x time (s)).

3) Deduce from **1** and **2** the charge required to remove one mole of copper.

4) Copper ions are doubly positively charged, so two moles of electrons are involved in removing one mole of copper ($Cu (s) \rightarrow Cu^{2+} (aq) + 2e^-$), so deduce the quantity of electricity equivalent to one mole of electrons (one-half of the value just calculated in **3**). This quantity of electricity is known as the Faraday constant.

5) The official value for the Faraday is 96500 Coulombs. How well does this agree with the value you have just calculated? What are the major sources of error in this experiment?

6) Write the equation for the reaction occurring at the cathode.

7) Predict how the concentration of the copper(II) sulfate solution changes, if at all, as the electrolysis takes place. Explain your answer.

8) Why is it better to use the loss in weight of the anode as opposed to the gain in weight of the cathode?