

## Answer on the question #63394, Chemistry / Physical Chemistry

### Question:

1 dm<sup>3</sup> of a solution of 2.0 M CuSO<sub>4</sub> is electrolysed using platinum electrodes by passing (5) 4.50 A current for 9000 s. Calculate 1) the mass of Cu deposited, and 2) the amount of Cu<sup>2+</sup> in the solution at the end of electrolysis.

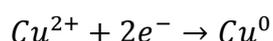
### Solution:

Faraday's law of electrolysis tells that the mass of the substance, liberated at the electrode in grams is proportional to the quantity of electricity  $Q$  passed through the solution:

$$m = \left(\frac{Q}{F}\right) \left(\frac{M}{z}\right)$$

where  $M$  is the molar mass of the substance,  $z$  is the charge of the ion and  $F$  is Faraday's constant  $F = 96485 \text{ C mol}^{-1}$ .

During the electrolysis in the solution of CuSO<sub>4</sub>, the following reaction takes place:



Thus, the charge of ion is +2. Molar mass of Cu is 63.546 g mol<sup>-1</sup>. Quantity of electricity, of simply charge is the product of current and time:

$$Q = It = 4.50(\text{A}) \cdot 9000(\text{s}) = 4.05 \cdot 10^4(\text{C})$$

Getting the mass of copper deposited on the electrode:

$$m = \left(\frac{4.05 \cdot 10^4(\text{C})}{96485 (\text{C mol}^{-1})}\right) \left(\frac{63.546 (\text{g mol}^{-1})}{2}\right) = 13.34 (\text{g})$$

Also, we get the amount of copper reduced and left in the solution:

$$n_{\text{red}} = \frac{m}{M} = \frac{13.34(\text{g})}{63.546 (\text{g mol}^{-1})} = 0.210 (\text{mol})$$

$$n_{\text{all}} = cV = 2(\text{M}) \cdot 1(\text{L}) = 2 (\text{mol})$$

$$n_{\text{left}} = n_{\text{all}} - n_{\text{red}} = 2 - 0.210 = 1.790 (\text{mol})$$

And the concentration of copper is finally

$$c = \frac{n_{\text{left}}}{V} = \frac{1.790(\text{mol})}{1(\text{L})} = 1.790 (\text{mol L}^{-1})$$

**Answer:** 1) 13.34g, 2) 1.790 mol L<sup>-1</sup>

Remark: The solubility of CuSO<sub>4</sub> in water at 30°C is 1.502 mol L<sup>-1</sup>. This raises a question about the reality of the exercise case.