Answer on Question#74456 – Chemistry – General chemistry

Question: How long would it take to electroplate all the Au^{3+} in 0.260 L of 0.220 M $[Au(CN)_4]Cl$ solution with a current of 2.50 A?

Solution:

$$M = \frac{mol}{L}; A = \frac{C}{s}$$

1. Find moles of Au^{3+}

$$n(Au^{3+}) = n([Au(CN)_4]Cl) = V \times C_M = 0.260\ L \times 0.220 \frac{mol}{L} = 0.0572\ mol$$

2. Each Au^{3+} ion requires three electron to become a gold atom:

Au^{3+} (aq) + $3e^-$ → Au (s)

$$n(e^-) = 3 \times 0.057\ mol = 0.1716\ mol$$

3. We should calculate the quantity of charge carried by these electrons:

$$Q = \frac{F}{n(e^-)} = \frac{96485 \frac{C}{mol}}{0.1716\ mol} = 562267\ \text{C}$$

4. Find the time:

$$t = \frac{Q}{I} = \frac{562267\ C}{2.50\ \frac{C}{s}} = 2.25 \times 10^5 s$$

Answer: $2.25 \times 10^5 s$