

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 $[\text{Au}(\text{CN})_4]\text{Cl}$ solution with a current of 2.50 A?

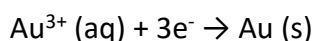
Solution:

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

1. Find moles of Au^{3+}

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

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



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

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

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

4. Find the time:

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

Answer: $2.25 \times 10^5 \text{ s}$