

## Faraday's Laws-Stoichiometry

The amount of product of an electrolysis reaction depends on

- i) mass
- ii) time
- iii) electrical current

Definitions -  $1 \text{ Coulomb} = 1 \text{ Ampere} / \text{sec.}$   
(charge) (current)  
 $e^-$

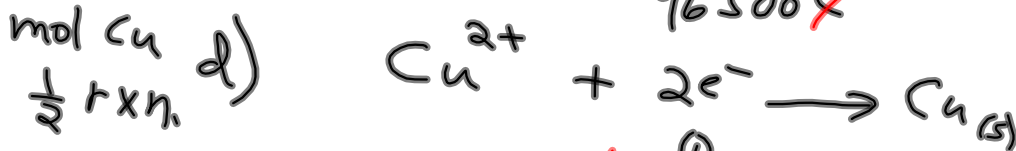
Faraday's constant  $1 \text{ mole } e^- = 96500 \text{ C}$

Example - What mass of  $\text{Cu}$  is deposited on a cathode dipped in  $\text{Cu}^{2+}\text{SO}_4(\text{aq})$  by a 5.0 A for 30 min.

sec. a)  $30 \text{ min.} \times \frac{60 \text{ s}}{1 \text{ min.}} = 1800 \text{ sec.}$

c b)  $\text{Coulombs} = 5.0 \text{ A} \times 1800 \text{ s} = 9000 \text{ C}$

moles  $e^-$  d)  $9000 \text{ C} \times \frac{1 \text{ mole } e^-}{96500 \text{ C}} = 0.093 \text{ mol } e^-$



e)  $0.093 \text{ mole } e^- \times \frac{1 \text{ mol Cu}}{2 \text{ mole } e^-} = 0.047 \text{ mol Cu}$

Mass Cu f)  $0.047 \text{ mol Cu} \times \frac{63.55 \text{ g Cu}}{1 \text{ mol Cu}} = 3.0 \text{ g Cu}$

## Faraday's Laws-Time

How much time is needed to deposit 15.8 g Ag using a 25.0 A current?



b)  $15.8 \cancel{\text{g Ag}} \times \frac{1 \text{ mole Ag}}{107.87 \cancel{\text{g Ag}}} = 0.146 \text{ mol Ag}$

b)  $0.146 \cancel{\text{ mol Ag}} \times \frac{1 \text{ mol } e^-}{1 \cancel{\text{ mol Ag}^+}} = 0.146 \text{ mol } e^-$

c)  $0.146 \cancel{\text{ mol } e^-} \times \frac{96500 \text{ C}}{1 \cancel{\text{ mol } e^-}} = 14135 \text{ C}$   
also

d)  $\frac{14135 \cancel{\text{ A-s}}}{25.0 \cancel{\text{ A}}} = 565 \text{ sec.}$       14135 A-sec.

Questions -

P. 793

# 22, 23, 21 ← mass

↑      ↑  
Time    current