Electrochemistry

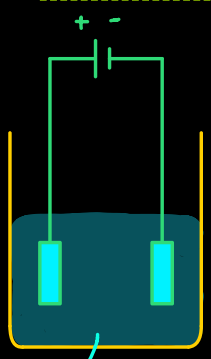
- Electrolytical cell
  - > Electrical energy to Chemical energy
  - > Based on non spontaneous reactions
  - > E cell is -ve
  - > Anode is +ve, Cathode is -ve

Reduction always at cathode

oxidation always at anode

- Electrochemical cell
  - > Chemical energy to Electrical energy
  - > Based on spontaneous reactions
  - > E cell is +ve
  - > Anode is -ve, cathode is +ve

Electrolysis of Molten NaCl



Reduction (cathode)

at -ve electrode: $Na^+ + e^- \rightarrow Na$

at +ve electrode: $2Cl^- \rightarrow Cl_2 + 2e^-$

Oxidation (Anode)

Electrolysis of NaCl(aq)

- -ve electrode : one of these
  $Na^+ + e^- \rightleftharpoons Na$ (-2.71 V) ①
  $2H^+ + 2e^- \rightleftharpoons H_2$ (0.00 V) ②
  Choose based on the one that's more likely to go forward, so ②

- +ve electrode : one of these
  $Cl_2 + 2e^- \rightleftharpoons 2Cl^-$ (1.36V) ③
  $O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$ (1.23V) ④
  Choose based on the one that's more likely to go backward, so ④

∴ At cathode (-ve electrode) ⇒ $2H^+ + 2e^- \rightarrow H_2$ ×2

At Anode (+ve electrode) ⇒ $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

So, overall equation : $2H_2O \rightarrow 2H_2 + O_2$

Faraday's constant (F)

> F is basically charge per mol of electrons

$F = L \cdot e$ (L: Avogadro's constant) ∴ $F = 6.02 \times 10^{23} \times 1.6 \times 10^{-19} = 96320$

but actual value = $96500\ C\ mol^{-1}$

> Use $Q = It$ to find total charge
> $\frac{\text{Total charge}}{96500}$ = moles of electrons passed
> check ratio of electrons to Element deposited
> find mass by $n \times M_r$

$CuSO_4 \longrightarrow Cu^{2+} + SO_4^{2-}$

$Cu^{2+} + 2e^- \longrightarrow Cu$

$Q = It$
$= 1.5 \times 3 \times 60^2$
$= 16200\ C$

$x = \frac{16200}{1.6 \times 10^{-19}}$

$= 1.0125 \times 10^{23}$

$n = \frac{5.09}{63.5} = 0.080157$ ← moles of Cu

Cu : e^-
1 : 2

$\therefore$ moles of $e^- = 2 \times 0.080157$
$= 0.16031$

1.0125×10^{23} electrons $= 0.16031$ mol
x $= 1$ mol

$x = 6.3158 \times 10^{23}$
$\approx 6.32 \times 10^{23}$