6. Electrochemical Cell - Quantitative (V, 9)

Voltage along the wire in an electrochemical cell depends on the half reactions.

a) Cell Potential

i) cell potential (E_{cell}) is the maximum voltage of an electrochemical cell

ii) standard cell potential ($E°_{cell}$) is the maximum voltage when:

concentration of all ions is 1.0 M,
temperature is 25°C and,
pressure is 1 atm (101.325 kPa)

} "standard state"

iii) purpose of $E°_{cell}$ is to:

① *calculate voltage of a particular electrochemical cell*
② *predict whether a particular redox reaction will occur (spontaneous)*

iv) $E°_{cell}$ is calculated from the difference in voltage between the half reactions

b) Table of "Standard Reduction Potentials"

i.e. "Voltages for Half Reactions written as a Reduction"

i) Hydrogen half reaction is set at an arbitrary voltage of 0.00 V

$H_{2\,(g)} \rightarrow 2H^{+} + 2\,e^{-}$ $\quad E° = 0.00$ V

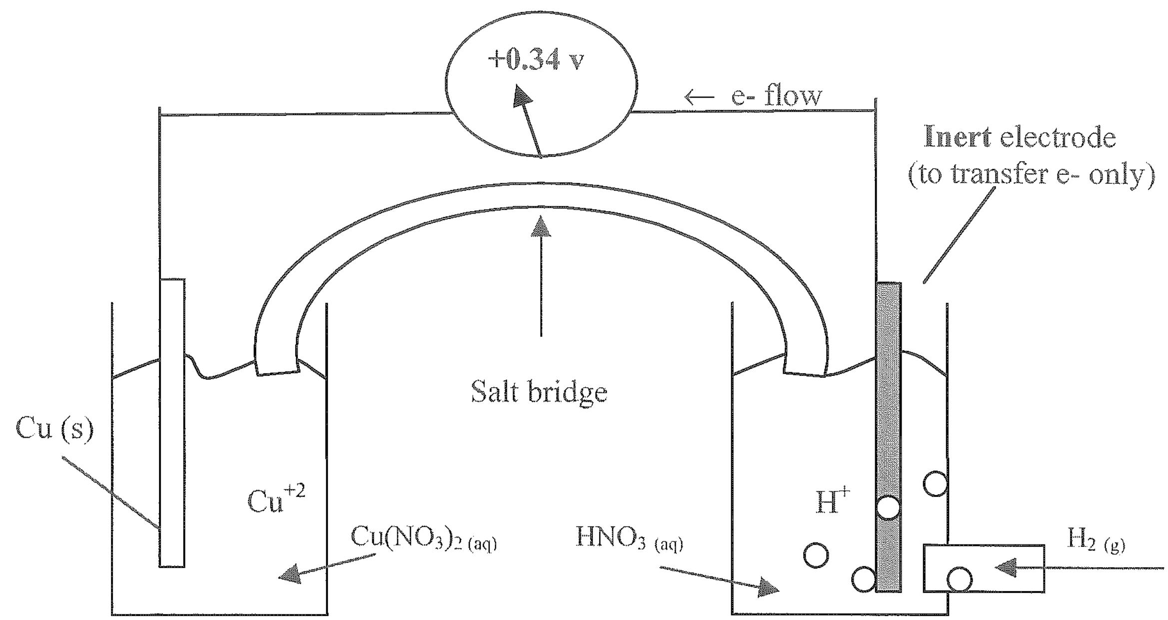
ii) Example: $Zn^{+2} + 2e^{-} \rightarrow Zn_{(s)}$ $\quad$ **$E° = -0.76$ V**

This reduction produces 0.76 V *less* than the hydrogen half reaction.

iii) Example: $Cu^{+2} + 2e^{-} \rightarrow Cu_{(s)}$ $\quad$ **$E° = +0.34$ V**

This reduction produces 0.34 V *more* than the hydrogen half reaction.

iv) Other half reaction voltages are relative compared to the hydrogen half cell.



Reduction:
$Cu^{+2} + 2e- \rightarrow Cu_{(s)}$

"Cathode"

Oxidation:
$H_{2\,(g)} \rightarrow 2H^{+} + 2e-$

"Anode"

v) Each reduction half reaction can be written as an oxidation.

$Cu^{+2} + 2e- \rightarrow Cu_{(s)}$ $\quad E° = +0.34$ V

$Cu_{(s)} \rightarrow Cu^{+2} + 2e-$ $\quad$ **E° = - 0.34 V** – reduction rxn in reverse!

c) Calculating the E°cell for an Electrochemical Cell

i) Example: ***$Cu^{+2} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{+2}$***

Cathode reaction: $Cu^{+2} + 2e- \rightarrow Cu_{(s)}$ $\quad E° = +0.34$ V

Anode reaction: $Zn_{(s)} \rightarrow Zn^{+2} + 2e-$ $\quad E° = +0.76$ V (reverse rxn = reverse sign!!)

E°cell = +1.10 V

ii) Example: $Fe^{+2} + Mg_{(s)} \rightarrow Fe_{(s)} + Mg^{+2}$

Cathode reaction: $Fe^{+2} + 2e^- \rightarrow Fe_{(s)}$ E° = -0.45 V

Anode reaction: $Mg_{(s)} \rightarrow Mg^{+2} + 2e^-$ E° = +2.37 V (reverse rxn = reverse sign!!)

E°cell = +1.92 V

iii) Example: $2Ag^{+} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{+2}$

Cathode reaction: $2Ag^{+} + 2e^- \rightarrow 2Ag_{(s)}$ E° = +0.80 V (do not multiple E° by 2)

Anode reaction: $Cu_{(s)} \rightarrow Cu^{+2} + 2e^-$ E° = -0.34 V (reverse rxn = reverse sign!!)

E°cell = +0.46 V

iv) Example: $2MnO_4^- + 4H_2O + 3Ni_{(s)} \rightarrow 2MnO_{2\,(s)} + 8OH^- + 3Ni^{+2}$

Cathode reaction: $2(\ MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_{2\,(s)} + 4OH^-\ )$ E° = +0.60 V

Anode reaction: $3(\ Ni_{(s)} \rightarrow Ni^{+2} + 2e^-\ )$ E° = +0.26 V (reverse rxn = reverse sign!!)

E°cell = +0.86 V

Do Questions: #36-38, 40-41 page 224-225

Note: Hebden uses formula E°cell = E°red - E°ox
Do NOT reverse sign of oxidation reaction if using this formula.