Cell Potential

Electromotive force - potential difference = Volts
charge

Work = EMF
↓
q

Cell Potential

EMF = potential difference = Volt = work Joule
charge Coulomb

System (us) → surroundings work = $-qE$

$q = nF \rightarrow 96,485\ C/mole\ e^-$
(n = #moles electrons transferred)

$Ag^+ + 1e^- \rightarrow Ag$

work = $-qE$ E = cell potential
work = $-(nF)E$

$W = -(1.33\ mole\ e^- \times 96,485\ \frac{Coloumb}{mole^-})\ 2.10\ \frac{Joules}{Coloumb}$
$= -269,000\ J = -269\ kJ$

ΔG = Gibb's Free Energy = the amount of energy available to work

$W = -nFE$
$\Delta G = -nFE$
ΔG = negative is spontaneous
ΔG = positive not spontaneous

E = positive $\rightarrow$ ΔG = neg. = spontaneously
E = negative $\rightarrow$ ΔG = pos = not spontaneous
E = 0 $\rightarrow$ ΔG = zero = equilibrium
dead battery

$\Delta G = -nF(-E)$
$\Delta G = +$

$\Delta G = -nF(0)$
$\Delta G = 0$

$Cu^{+2}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{+2}(aq)$

$Cu^{+2} + 2e^- \rightarrow Cu$ $E = .337V$ red
~~$Fe^{+2} + 2e^- \rightarrow Fe$ $E = -.44V$ ox~~
$Fe \rightarrow Fe^{+2} + 2e$ $E = .44\ V$

$Cu^{+2} + Fe \rightarrow Cu + Fe^{+2}$ $E = .777\ V$
n = 2 moles electrons
$\Delta G = -nFE$ F = 96,485 ℰ/mole

$\Delta G = -(2 \times 96{,}485)(.777V) = -149{,}938$ ℰ
-150 kJ work done on the surroundings

$NO_3^- + 4H^+ + 3e^- \rightarrow NO + H_2O$ $= .96V$ red
$Au^{+3} + 3e^- \rightarrow Au^0$ $= \pm 1.50V$ red
$2.46\ V$

$HNO_3 + H^+ + NO_3^- + Au \rightarrow Au^{+3} + NO + H_2O$

$E = .96V$ red
E $-1.50\ V$ ox
$E = -.54\ V$

$\Delta G = -nFE$
$-(3 \times 96{,}485 \times 2.46$
= ~~neg spontaneous~~
$= -(3 \times 96{,}485)(-.54V)$
$\Delta G = +$ = not sponf.

Cell Potential

1M 1M (M) (M)

$2Al(s) + 3Mn^{+2}(aq) \rightarrow 2Al^{+3}(aq) + 3Mn(s)$

1.3 M M(s) electrons flow

a. $[Al^{+3}] = 2.0\ M$ ← Left E_{cell} will decrease
$[Mn^{+2}] = 1.0\ M$

b. $[Mn^{+2}] = 3.0$ → Right E_{cell} will increase

Exercise 7

cathode red: $2(VO_2^+ + 2H^+ + e^- \rightarrow VO_2^+ + H_2O)$ $E = 1.00\ V$

anode ox: ~~$Zn + 2e^- \rightarrow Zn$~~ $Zn \rightarrow Zn^{+2} + 2e^-$ $E = -.76\ V$ ✓

$E = +.76$

$E^\circ_{cell} = 1.76\ V$

E_{cell} Q

$2VO_2^+ (aq) + 4H^+ (aq) + Zn (s) \rightarrow 2VO_2^+ (aq) + 2H_2O (l) + Zn^{+2} (aq)$

$$Q = \frac{[VO_2^+]^2 [Zn^{+2}]}{[VO_2^+]^2 [H^+]^4}$$

$$Q = \frac{[2.0]^2 [.1]}{[.01]^2 [.50]^4} = 4. \times 10^{-5}$$

$$E_{cell} = E^\circ_{cell} - \frac{.0592}{n} \log Q$$

$$= 1.76 - \frac{.0592}{2} \log (4 \times 10^{-5})$$

$$= 1.76 - (.0296 \times 4.398) = \boxed{1.89\ V}$$

Exercise 8

$S_4O_6^{-2} + Cr^{+2} \rightarrow Cr^{+3} + S_2O_3^{-2}$

$S_4O_6^{-2} + 2e^- \rightarrow 2S_2O_3^{-2}$ $E = .17V$ red

~~$Cr^{+3} + 1e^- \rightarrow Cr^{+3}$~~ ~~$E = -.50V$~~ ox

$Cr^{+2} \rightarrow Cr^{+3} + 1e$ $E = .50V$

$E^\circ = .67V$

$E = \frac{RT}{nF} \ln K$

$R = 8.315\ J/mol \cdot K$

$E = .67V$

$n = 2$

$F = 96.485\ J/V$

$T = 298K$

$.67 = \frac{(8.315 \times 298)}{(2 \times 96,485)} \ln K$

$129290 = 2478 \ln K$

$\ln K = 52.18$

$K = e^{52.18}$

$K = 4.59 \times 10^{22}$ very far to the right