Today, April 6th: Sections 18.12-18.14
Sunday: problem club with Ali
Monday: start ch 19

Review Questions

18-2 #4

$E = 1.56\,V$ $\quad \Delta G = -nFE$ $\quad \Delta G^\circ = -nFE^\circ$ $\quad F = 96500\ \frac{J}{elect \cdot V}$

$Zn + 2MnO_2 + 2NH_4^+ \longrightarrow 2NH_3 + Mn_2O_3 + Zn^{2+} + H_2O$

2e⁻ ox per Zn; n = least common multiple

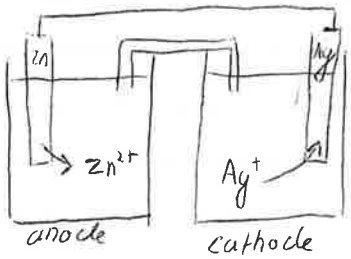
1e⁻ reduction per Mn; 2e⁻ total

$\Delta G = -nFE$

$\Delta G = (-2\ electrons)\left(96500\ \frac{J}{e^- \cdot V}\right)(1.56\,V)$ (have to divide by 1000 to get to kJ)

$\Delta G = -301\ kJ$

18-1 2f.



$Zn|Zn^{2+}||Ag^+|Ag$

$Zn + 2Ag \rightarrow Zn^{2+} + 2Ag$

1e⁻ reduction, 2 tot; 2e⁻ oxidation, 2 tot

Ag $M_I = 2.00g$ $\quad n_{Ag^+} = 0.250\,mol$

Zn $M_I = 3.00g$ $\quad n_{Zn^{2+}} = 0.25\,mol$ $\quad M_f = 1.2g$

$\Delta m_{Zn} = 1.80g$

$\Delta n_{Zn} = \frac{1.80g \mid mol}{65.39g} = 2.75 \times 10^{-2}\,mol$

$\frac{0.0275\,mol\ Zn \mid 2\,mol\ Ag}{1\,mol\ Zn} = 0.0551\,mol\ Ag$

$\frac{.0551\ \overset{mol}{Ag} \mid 107.87g\ Ag}{mol\ Ag} = 5.94g + 2.00g = \boxed{7.94g}$

$n_{Ag^+} = 0.025\,mol$

$-\ 0.0551\,mol$ reacted

$n_{Ag^+} = \frac{0.195\,mol}{.25\,L} = \boxed{0.780\,M}$

$\Delta G = \Delta G^\circ + RT\ln Q$

(-) $Q < K_c$ shift right

(+) $Q > K_c$ shift left

* at equilibrium $\Delta G = 0$

$0 = \Delta G^\circ + RT\ln K$

$\Delta G^\circ = -RT\ln K$

Nernst:

$E = E^\circ - \frac{0.0592}{n}\log Q$

* at equilibrium $E = 0$

$E^\circ = \frac{0.0592}{n}\log K$

Q + K

gases: Q_P K_P

soln: Q_c K_c

Mixed

$2H^+ + 2e^- \rightarrow H_2 \quad E^\circ_{red} = 0.00V$

$Zn \rightarrow Zn^{2+} + 2e^- \quad E^\circ_{ox} = 0.26V$

$2H^+ + Zn \rightarrow H_2 + Zn^{2+} \quad E^\circ_{rxn} = +0.26V$

$Q = \frac{P_{H_2} \times [Zn^{2+}]}{[H^+]^2}$

$Q = \frac{1.0\,atm \times 1.0M}{(4\times10^{-3}M)^2}$

$Q = 6.25\times10^4$

Calc ΔG° for the reaction

$\Delta G^\circ = -nFE^\circ$

$\Delta G^\circ = -(2e^-)\left(96.5\frac{kJ}{mole^-\cdot V}\right)(.26\,V)$

sorry black pen ran out :(

Calc E if $P_{H_2} = 1.00\,atm$, $[Zn^{2+}] = 1.00M$, $[H^+] = 4.0\times10^{-3}M$

$E = E^\circ - \frac{0.0592}{n}\log Q$

$E = 0.26V - \frac{0.0592}{2}\log(6.25\times10^4)$

$E = 0.118V$ not as spontaneous than under standard conditions

* don't multiply E values *

$2Ag^+ + 2e^- \rightarrow 2Ag \quad E^\circ_{red} = +0.80V$

~~$Cu^{2+} + 2e^- \rightarrow Cu \quad E^\circ_{red} = +0.34V$~~

$Cu \rightarrow Cu^{2+} + 2e^- \quad E^\circ_{ox} = -0.34V$

$2Ag^+ + Cu \rightarrow Cu^{2+} + 2Ag \quad E^\circ_{rxn} = +0.46$

LCM = 2 = n

Suppose $E = 0.19V$. Calc pH (calc H^+) - from $2H^+ + Zn \rightarrow H_2 + Zn^{2+}$

$[Zn^{2+}] = 1.00M \quad P_{H_2} = 1.0atm$

$$E = E^\circ - \frac{0.0592}{2} \log \frac{(1.0atm)(1M\ Zn)}{[H^+]^2}$$

$$0.19 = 0.26 - \frac{0.0592}{2} \log \frac{(1atm)(1M\ Zn)}{[H^+]^2}$$

$H^+ = 6.57 \times 10^{-2} M$

$pH = -\log(6.57 \times 10^{-2} M)$

$pH = 1.18$

Electrolytic cells - nonspontaneous cells ($E < 0$, $\Delta G > 0$)

charge = current x time

(coulomb) = (amps) x (s)

1 mol e⁻ = 96500 coulombs

How many moles e⁻ are delivered from a 4.52A system in 30min?

charge = (4.52A)(30x60s)

charge = 8140C

$$\frac{8140C}{96500C} \left| \frac{1 mol\ e^-}{} \right. = 0.0843\ mol\ e^-$$

$Cu^{+1} + e^- \rightarrow Cu \quad n_{Cu^+} = \frac{0.0843 mol\ e^-}{1 mol\ e^-} \left| \frac{1 mol\ Cu^+}{} \right. = 0.0843 mol\ Cu^+ = 0.0843 mol\ Cu$

$Cu^{2+} + 2e^- \rightarrow Cu \quad n_{Cu} = \frac{0.0843 mole^-}{2 mole^-} \left| \frac{1 mol\ Cu}{} \right. = 0.0421 mol\ Cu$

$Cr^{3+} + 3e^- \rightarrow Cu \quad n_{Cr} = \frac{0.0843 mole^-}{3 mole^-} \left| \frac{1 mol\ Cr}{} \right. =$