Today: chapter 18 part 3
Friday: class canceled
Sunday: problem club with Al: canceled
Monday: start nuclear

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balance in basic solution:

3e⁻ reduction / Mn 42e⁻ total

acidic: $8H^+ + 14MnO_4^- + 3I_2 \rightarrow 14MnO_2 + 6IO_4^- + 4H_2O$

(+7 -2 on MnO_4^-; 0 on I_2; +4 -2 on MnO_2; +7 -2 on IO_4^-)

7e⁻ oxidation / I
14e⁻ oxidation / 2I

42 e⁻ total

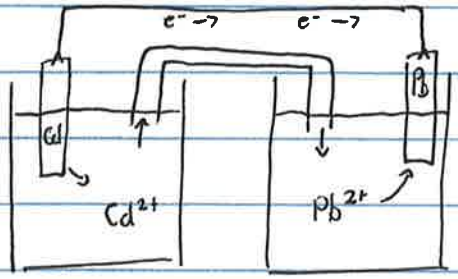
add OH^- to both sides to cancel H^+

basic: $8OH^- + 8H^+ + 14MnO_4^- + 3I_2 \rightarrow 14MnO_2 + 6IO_4^- + 4H_2O + 8OH^-$

$4H_2O + 14MnO_4^- + 3I_2 \rightarrow 14MnO_2 + 6IO_4^- + 4H_2O + 8OH^-$

$4H_2O + 14MnO_4^- + 3I_2 \rightarrow 14MnO_2 + 6IO_4^- + 8OH^-$

$Pb|Pb^{2+}$ and $Cd|Cd^{2+}$ => galvanic cell $\Delta G^\circ < 0$ $E^\circ > 0$



anode: $Cd \rightarrow Cd^{2+} + 2e^-$

cathode: $Pb^{2+} + 2e^- \rightarrow Pb$

$Cd|Cd^{2+}(1M)||Pb^{2+}(1M)|Pb$

$Cd \rightarrow Cd^{2+} + 2e^-$ $E^\circ = +0.40V$

$Pb^{2+} + 2e^- \rightarrow Pb$ $E^\circ = -0.13V$

$Cd + Pb^{2+} \rightarrow Cd^{2+} + Pb$ $E^\circ = 0.27V$

$K_c = \frac{[Cd^{2+}]}{[Pb^{2+}]}$

volt = J/coul

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

$\Delta G^\circ = -(2\ mol\ e^-)(96500\ coul/mole^-)(+0.27\ J/coul)$

$\Delta G^\circ = -52110\ J = \boxed{-52.11\ kJ}$

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

$+nFE^\circ = RT\ln K_c$

$$E^\circ = \frac{RT\ln K_c}{nF} \qquad \ln K_c = \frac{nFE^\circ}{RT}$$

$$\ln K_c = \frac{2\text{ mol } e^-}{\text{mol } e^-} \left| \frac{96500\text{ coul}}{\text{coul}} \right| \frac{0.27\text{ J}}{8.314\text{ J}} \left| \frac{\text{K}}{298\text{ K}} \right| = 21.03$$

$$K_c = e^{21.03} = 1.36 \times 10^9$$

non-standard conditions (for example $[A] \neq 1.0M$)

$\Delta G = \Delta G^\circ + RT\ln Q_c$ ← MICE

$-nFE = -nFE^\circ + RT\ln Q_c$

$$E = E^\circ - \frac{RT}{nF}\ln Q_c$$ ← Nernst equation

$$E = E^\circ - \frac{0.0572}{n}\log Q_c$$

$Pb^{2+} + Cd \rightarrow Pb + Cd^{2+}$ $\quad K_c = \frac{[Cd^{2+}]}{[Pb^{2+}]}$

Calculate E for $Cd|Cd^{2+}(1M)\,||\,Pb^{2+}(2M)|Pb$.

$$E = (0.27V) - \frac{0.0572}{2\text{ mol } e^-}\log\left(\frac{(1M)}{(2M)}\right)$$

$$E = 0.279\text{ V}$$

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

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

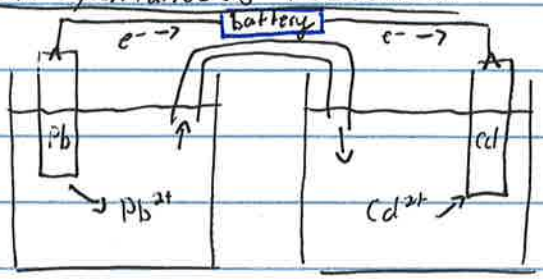
$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag \quad E^\circ = 0.46V \quad Q_c = \frac{[Cu^{2+}]}{[Ag^+]^2}$

Calculate E for $Cu | Cu^{2+}(1.0M) || Ag^{+}(0.04M) | Ag$

$$E = (0.46) - \frac{0.0572}{2\,mol\,e^-} \log\left(\frac{(1M)}{(0.04)^2}\right)$$

$$E = 0.38V$$

non-spontaneous reactions (electrolytic cells)



$E^\circ = -.27V$

Calculate e^- transferred (pumped by battery)

charge = current x time

(coul) = (amps) x (sec)

4.0 amp supply, operates for 30 min. Calculate charge delivered

charge = (4.0 amp)(30 min x 60 s/min)

charge = 7200 coul

$$n_{e^-} = \frac{7200\,coul \mid 1\,mol\,e^-}{\underbrace{96500\,coul}_{F}} = \boxed{0.0746\,mol\,e^-}$$

$2e^- + Cd^{2+} \rightarrow Cd$

$$n_{Cd^{2+}} = \frac{0.0746\,mole^- \mid 1\,mol\,Cd^{2+}}{2\,mol\,e^-} = 0.0373\,mol\,Cd^{2+}\ \text{undergoing reduction}$$

Cd(s) increases by 0.0373 mol

Pb(s) decreases by 0.0373 mol

Pb^{2+} increases by 0.0373 mol