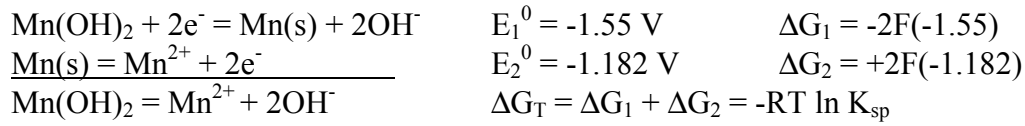


1] Method 1



$$\Delta G_T = \Delta G_1 + \Delta G_2$$

$$= -2F(-1.55) + 2F(-1.182) = -RT \ln K_{sp}$$

$$-2/0.0257 (1.55-1.182) = \ln K_{sp}$$

$$K_{sp} = 3.7\text{e-}13$$

Method 2

$$E_{\text{cell}} = E^0(\text{cath}) - E^0(\text{anod}) = -1.55 - (-1.182) \text{ volts} = -0.368 \text{ volts}$$

$$\Delta G = -2F(-0.368) = -RT \ln K_{sp}$$

$$K_{sp} = \exp[-2/0.0257 * 0.368] = 3.7\text{e-}13$$

2] First consider that

$$\text{O-O} + \text{e}^- = \text{R-O} \quad E = E_1^0 - 0.0592 \log \frac{[\text{R-O}]}{[\text{O-O}]}$$

$$\text{O-R} + \text{e}^- = \text{R-R} \quad E = E_2^0 - 0.0592 \log \frac{[\text{R-R}]}{[\text{O-R}]}$$

$$\text{For } X_1 = \frac{[\text{O-O}]}{[\text{R-O}]} = \frac{[\text{O-O}]}{[\text{O}_L - \text{R}_R] + [\text{R}_R - \text{O}_R]} \quad \text{multiple by } \frac{1}{\frac{[\text{O-R}]}{1}} \text{ we get } \frac{1}{[\text{O-R}]}$$

For clarity we will drop the L and R designation.

$$\frac{[\text{O-O}]}{[\text{O-R}]} = \frac{x_2}{1 + \frac{x_3}{x_4}} = \frac{x_2 x_4}{x_3 + x_4} = X_1 \quad (\#1)$$

remember that we defined:

$$x_1 = [\text{O}_L - \text{O}_R] / [\text{R}_L - \text{O}_R]$$

$$\begin{aligned}x_2 &= [\text{O}_L - \text{O}_R] / [\text{O}_L - \text{R}_R] \\x_3 &= [\text{R}_L - \text{O}_R] / [\text{R}_L - \text{R}_R] \\x_4 &= [\text{O}_L - \text{R}_R] / [\text{R}_L - \text{R}_R]\end{aligned}$$

Now consider X_2

$$X_2 = \frac{[R - O]}{[R - R]} = \frac{[R - O] + [O - R]}{[R - R]} = \frac{[R - O]}{[R - R]} + \frac{[O - R]}{[R - R]} = x_3 + x_4$$

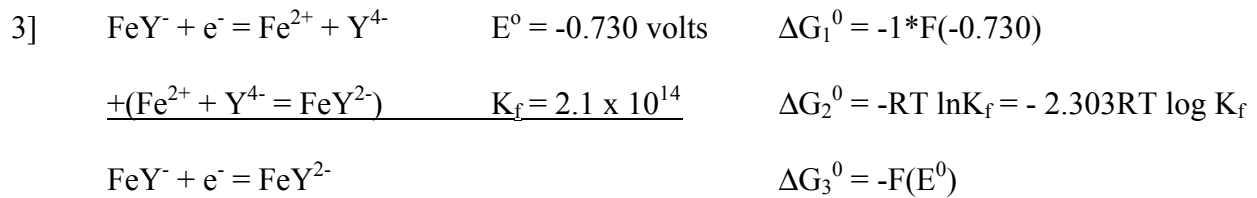
Note that at equilibrium $x_1 = x_2 = x_3 = x_4 = x$

$$X_1 = \frac{x^2}{2x} = \frac{x}{2} \quad X_2 = 2x$$

$$E = E_1^0 - 0.0592 \log \frac{1}{X_2} = E = E_2^0 - 0.0592 \log \frac{1}{X_2}$$

$$E = E_1^0 + 0.0592 \log \frac{x}{2} = E_2^0 + 0.0592 \log 2x$$

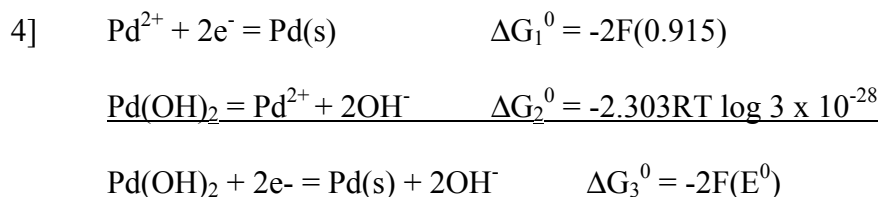
$$\frac{E_1^0 - E_2^0}{0.0592} = \log 4 \quad \Delta E^0 = E_1^0 - E_2^0 = 0.0356$$



$$\Delta G_3^0 = \Delta G_1^0 + \Delta G_2^0$$

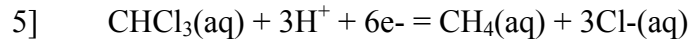
$$-F(E^0) = -1 \cdot F(-0.730) - 2.303RT \log K_f$$

$$E^0 = (-0.730) + 0.0592 \log 2.1 \times 10^{14} = -0.118 \text{ V}$$



$$\Delta G_3^0 = \Delta G_1^0 + \Delta G_3^0 = -2F(E^0) = -2F(0.915) - 2.303RT \log 3 \times 10^{-28}$$

$$E^0 = 0.100 \text{ Volts}$$



$$\Delta G^0 \text{ for } \text{CH}_4(\text{aq}) \quad H = 67.4 \text{ kPa/mol (1000)} = 6.74\text{e}4 \text{ J/mol}$$

$$\Delta G^0(\text{CH}_4(\text{aq})) = -50.84 \text{ kJ/mol} + 8.314 \text{ J/K-mol } 298\text{K} \ln 6.74\text{e}4 \text{ J/mol} = -23.3 \text{ kJ/mol}$$

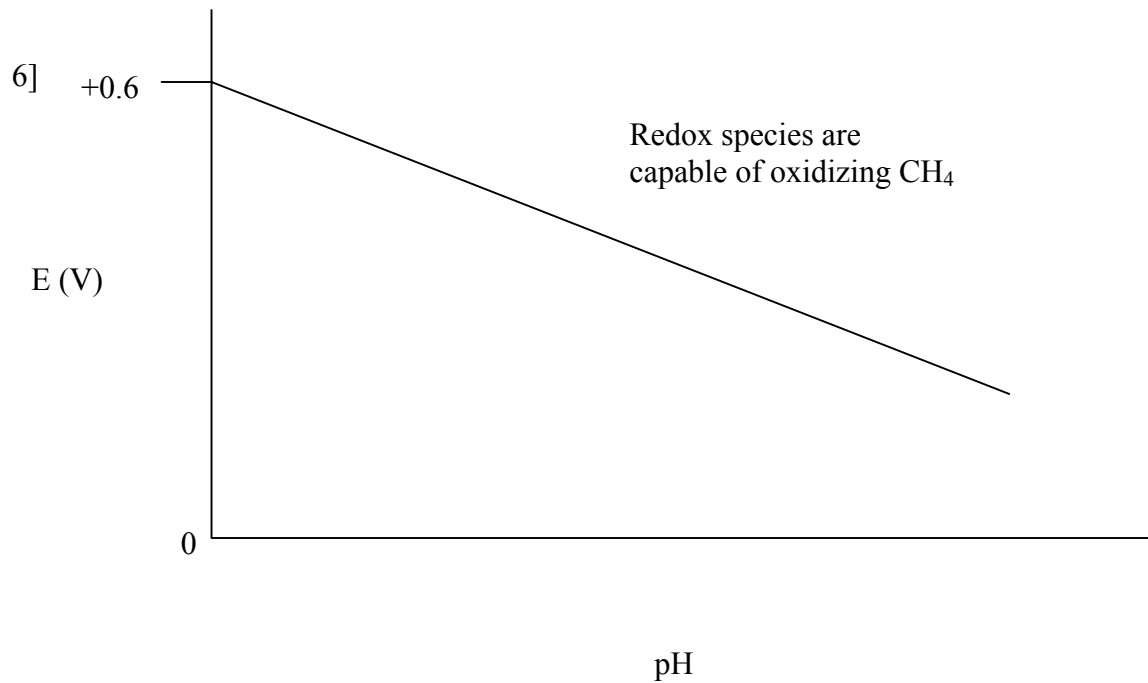
$$\Delta G^0 \text{ for } \text{CHCl}_3(\text{aq}) \quad H(\text{CHCl}_3) = 0.38 \text{ kPa m}^3/\text{mol (1000)} = 380 \text{ J/mol}$$

$$\Delta G^0(\text{CHCl}_3(\text{aq})) = -70.5 \text{ kJ/mol} + 8.314 \text{ J/K-mol } 298\text{K} \ln 380 \text{ J/mol} = -55.8 \text{ kJ/mol}$$

$$\Delta G_{\text{rxn}}^0 = -23.3 \text{ kJ/mol} + 3(-131.2 \text{ kJ/mol}) - (-55.8 \text{ kJ/mol}) = -361.1 \text{ kJ/mol}$$

$$E^0 = -361.1\text{e}3 \text{ J/mol}/6 (96484 \text{ coul}) = 0.623 \text{ volts}$$

$$E = 0.623 - 0.0592/2 \text{ pH}$$



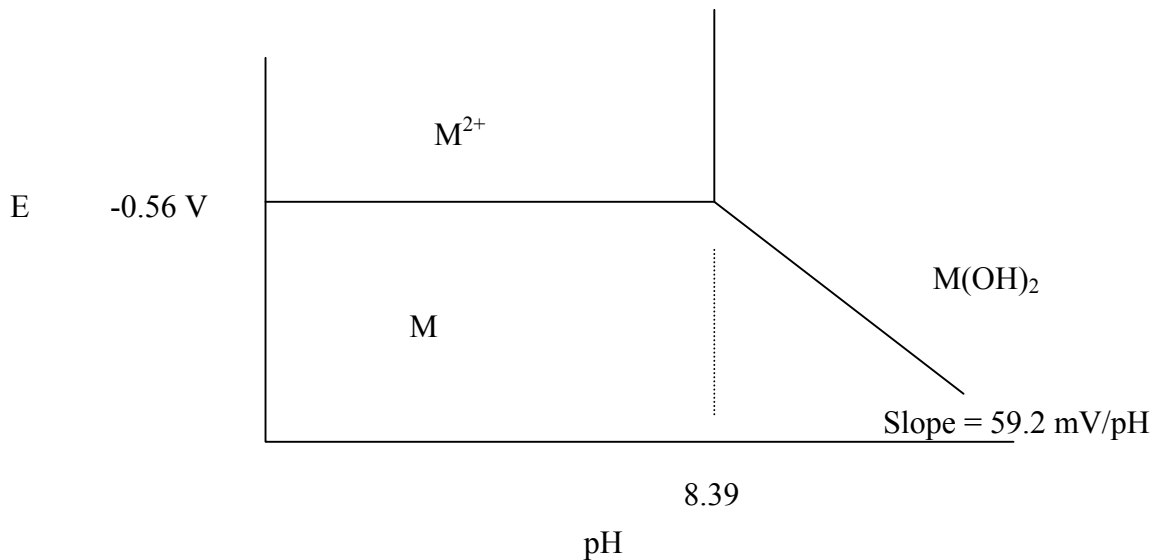
7] $E = 0.623 - 0.0592/2 \text{ pH}$

$\text{pH} = 7$

$$E = 0.623 - 0.0592/2 * 7 = 0.416 \text{ V}$$

8] $K_{sp} = [M^{2+}][OH^-]^2 = 6.0 \times 10^{-12}$ let $[Fe^{2+}] = 1M$

$pOH = 5.61$
 $pH = 8.39 = pK_a$



9] $E_{cathode} = E^0 - 0.0592/2 \log [Sn^{2+}]/[Sn^{4+}] = 0.154 - 0.0296 \log 0.150/0.250 = 0.161$ volts

$E_{anode} = E^0 - 0.0592/2 \log [SO_4^{2-}] = -0.350 - 0.0296 \log 0.200 = -0.329$ V

$E_{cell} = 0.161 - (-0.329) = 0.490$ V this is a spontaneous reaction.

10 Start with $Pb^{2+} + 2e^- = Pb(s)$

The Nernst eqn is $E = E^0_{red} + 0.0592/2 \log [Pb^{2+}]$

Consider that $Pb^{2+} + x [OH^-] = Pb(OH)_x$ $K_{sp} = [Pb^{2+}][OH^-]^x$

Solve for $[Pb^{2+}]$ and sub into Nernst eqn above

$E = E^0_{red} + 0.0592/2 \ln (K_{sp}/[OH^-]^x) = Const + x 0.0592/2 \log [OH^-]$

Plot of E vs. $-\ln [OH^-]$

Slope ≈ 0.17 volts/2.1 = $x 0.0592/2$ $x \approx 3$