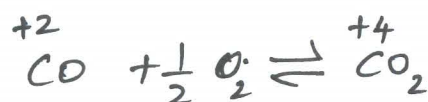


Relation of E° and the Equilibrium Constant:

13-24
~~13-24~~ The free energy change for the reaction
 $\text{CO} + \frac{1}{2}\text{O}_2 \rightleftharpoons \text{CO}_2$ is $\Delta G^\circ = -257 \text{ kJ per mole}$
 of CO at 298 K.

(a) Find E° for the reaction.



Ans

$$E^\circ = \frac{-\Delta G^\circ}{nF}$$

$$= \frac{-(-257 \times 10^3 \text{ J/mol})}{(2) (9.6485 \times 10^4 \text{ C/mol})}$$

$$E^\circ = 1.33 \text{ V}$$

Where ΔG° = Free energy change.

E° = Standard electrode potential.

n = number of electrons.

F = is the Faraday constant.

(28)

(b) Find the equilibrium constant for the reaction.

Ans:

$$K = 10^{nE^0/0.05916}$$

$$K = 10^{(2)(1.33)/0.05916}$$

$$= 10^{44.96}$$

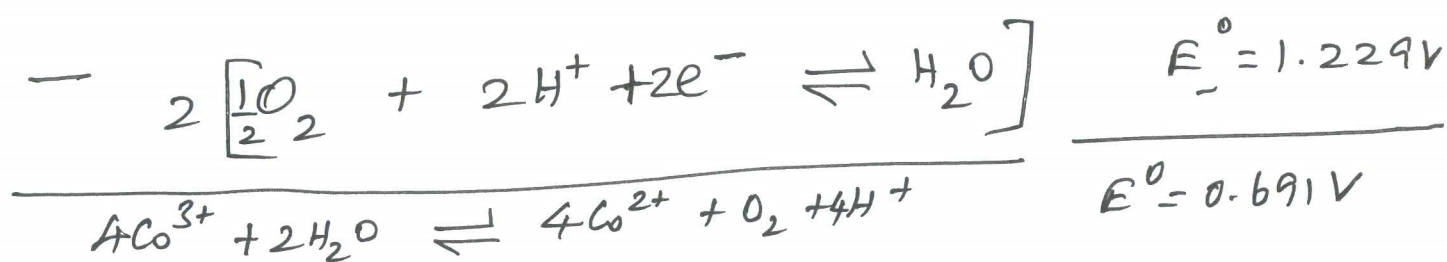
$K = 1 \times 10^{45}$

$$\therefore n = 2$$

$$E^0 = 1.33 \text{ V}$$

13-25

~~13-25~~ Calculate E° , ΔG° , and k for each of the following reactions.



$$\therefore \Delta G^\circ = -4FE^\circ$$

$$= -4(9.6485 \times 10^4)(0.691)$$

$$= -266684.54$$

$$= -2.7 \times 10^5 \text{ J}$$

$$\therefore k = 10^{4E^\circ / 0.05916}$$

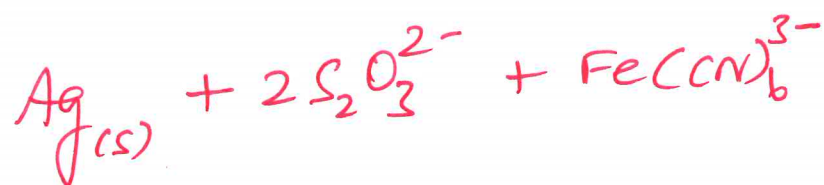
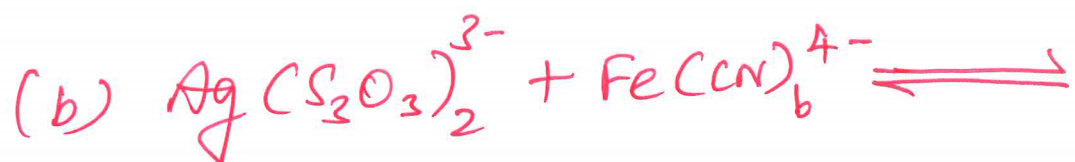
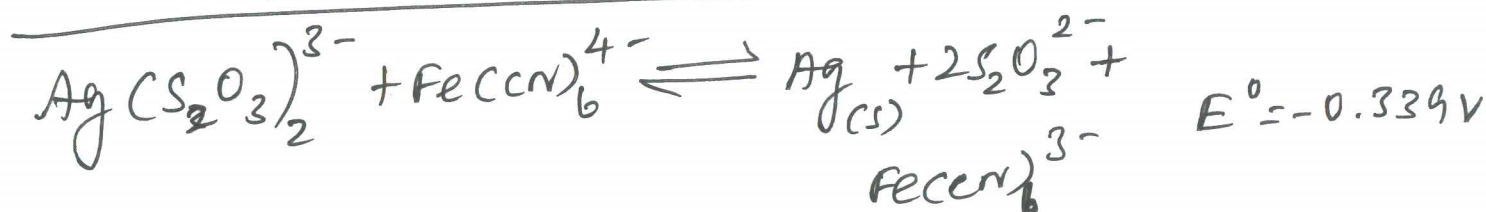
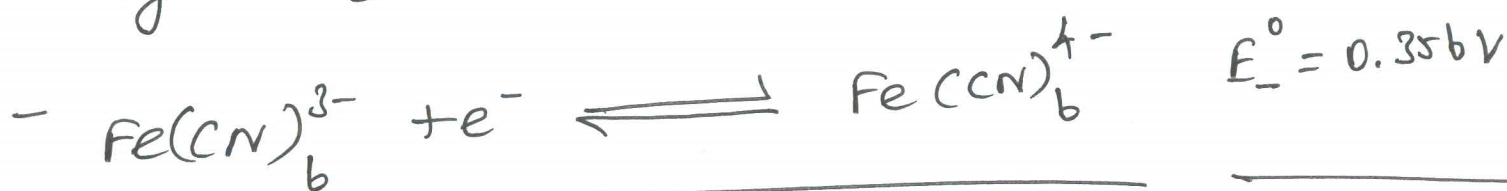
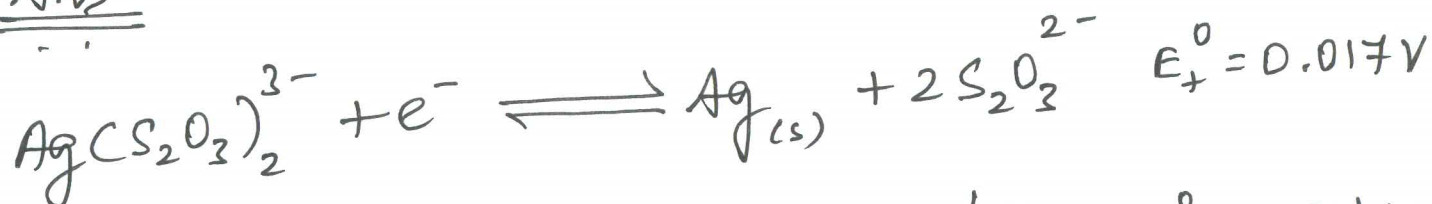
$$= 10^{4(0.691) / 0.05916}$$

$$= 10^{46.7}$$

$$\boxed{k = 10^{47}}$$

$$\begin{aligned} n &= 4 \\ F &= 9.6485 \times 10^4 \\ E^\circ &= 0.691\text{V} \end{aligned}$$

(40) (29)

Ans

$$\Delta G^0 = -nFE^0$$

$$= -1(9.6485 \times 10^4)(-0.339)$$

$$\Delta G^0 = 32.7 \text{ kJ}$$

$$K = 10^{nE^0/0.05916}$$

$$= 10^{1(-0.339)/0.05916}$$

$$= 10^{-5.73} = 1.9 \times 10^{-6}$$

$$\therefore K = 1.9 \times 10^{-6}$$

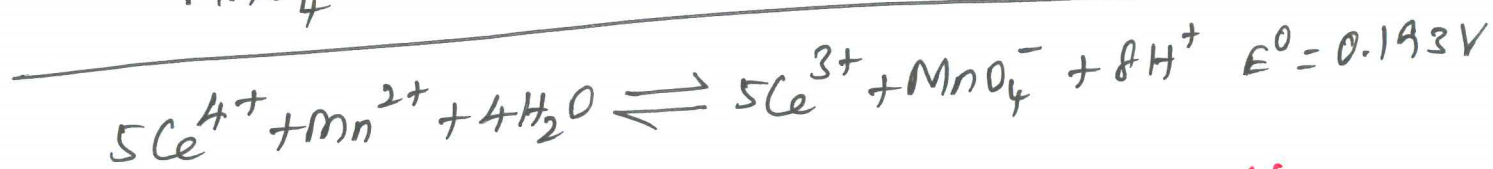
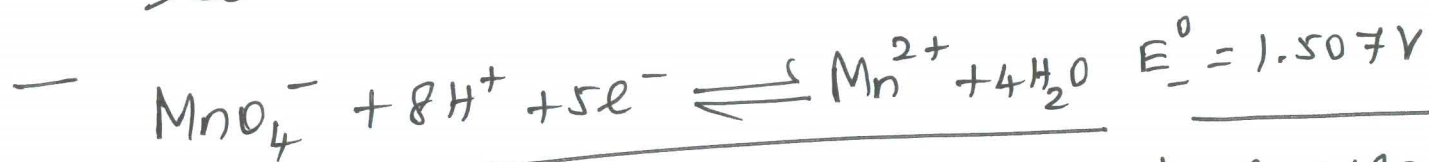
13-26

(41)

~~100220~~. A solution contains 0.100 M Ce^{3+} , $1.00 \times 10^{-4} \text{ M Ce}^{4+}$, $1.00 \times 10^{-4} \text{ M Mn}^{2+}$, 0.100 M MnO_4^- and 1.00 M HClO_4 .

(a) write a balanced net reaction that can occur between the species in this solution.

Ans:



(b) Calculate ΔG^0 and K for the reaction.

Ans:

$$\Delta G^0 = -nFE^0$$

$$= -5(9.6485 \times 10^4) (0.193) = -93108.025$$

$$\boxed{\Delta G^0 = -93.1 \text{ kJ}}$$

$$K = 10^{5E^0/0.05916}$$

$$= 10^{5(0.193)/0.05916} = 10^{16.31} = 2 \times 10^{16}$$

$$\therefore \boxed{K = 2 \times 10^{16}}$$

(c) Calculate E for the conditions given. (42)

Ans:

$$E = E_+ - E_-$$

$$= \left\{ 1.70 - \frac{0.05916}{5} \log \frac{[\text{Ce}^{3+}]^5}{[\text{Ce}^{4+}]^5} \right\} -$$

$$\left\{ 1.507 - \frac{0.05916}{5} \log \frac{[\text{Mn}^{2+}]}{[\text{MnO}_4^-][\text{H}^+]^8} \right\}$$

$$= \left\{ 1.70 - \frac{0.05916}{5} \log \frac{[0.100]^5}{[1.00 \times 10^{-4}]^5} \right\} -$$

$$\left\{ 1.507 - \frac{0.05916}{5} \log \frac{[1.00 \times 10^{-4}]}{[0.100][1.00]} \right\}$$

$$E = -0.020 \text{ V.}$$

(d) Calculate ΔG for the conditions given.

Ans:

$$\Delta G = -5FE$$

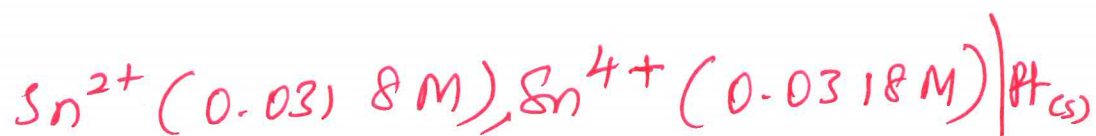
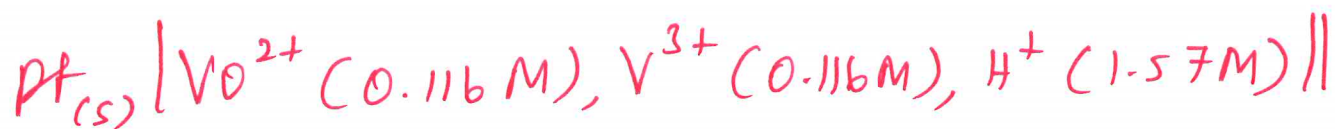
$$= -5(9.6485 \times 10^4)(-0.020)$$

$$= 9648.5 \text{ J}$$

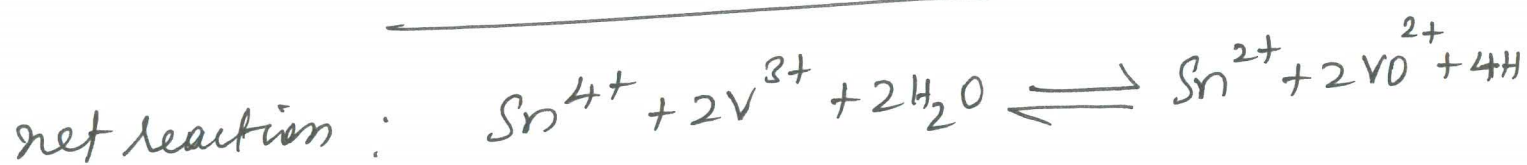
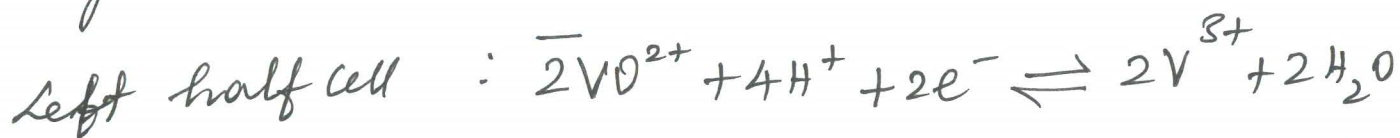
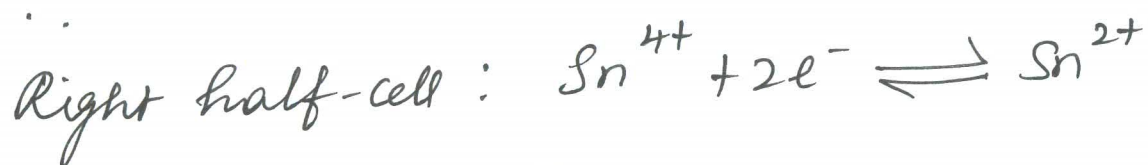
$$\Delta G = +10 \text{ KJ}$$

13-27

~~43~~ 43. For the following cell, E (not E°) = -0.289V .
 Write the net cell reaction and calculate its equilibrium constant. Do not use E° values from Appendix H to answer this question.



Soln



$$E = E^\circ - \frac{0.05916}{2} \log \frac{[\text{VO}^{2+}]^2 [\text{H}^+]^4 [\text{Sn}^{2+}]}{[\text{V}^{3+}]^2 [\text{Sn}^{4+}]}$$

$$-0.289 = E^\circ - \log \frac{(0.116)^2 (1.57)^4 (0.0318)}{(0.116)^2 (0.0318)}$$

$$\therefore E^\circ = -0.266\text{V}$$

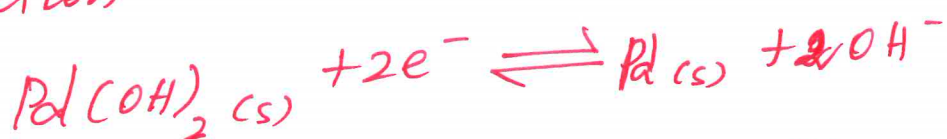
$$K = 10^{nE^{\circ}/0.05916}$$

$$= 10^{2(-0.266)/0.05916}$$

$$= 10^{-8.99}$$

$$K = 1.0 \times 10^{-9}$$

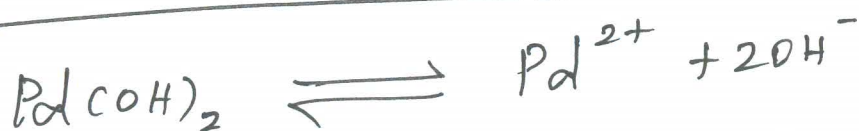
13-28
~~13-29~~. Calculate the standard potential for the half reaction



given that K_{sp} for $\text{Pd}(\text{OH})_2$ is 3×10^{-28} and given that the standard potential for the reaction



Ans:



$$E = E_{+}^{\circ} - 0.915$$

(45)

$$\text{Ant } K_{sp} = 3 \times 10^{-28}$$

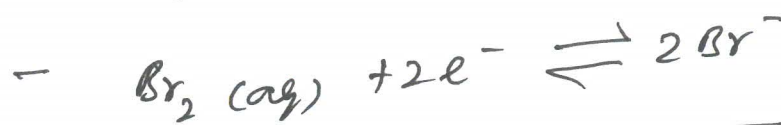
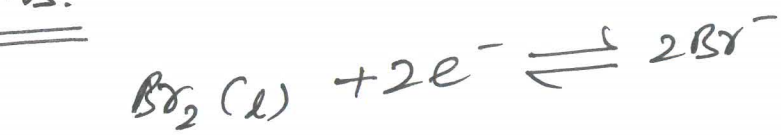
$$\therefore E^{\circ} = \frac{0.05916}{2} \log K_{sp} = -0.814$$

$$\therefore -0.814 = E_{+}^{\circ} - 0.915$$

$$\therefore E_{+}^{\circ} = 0.101$$

13-29. From the standard potentials for reduction of $\text{Br}_2(\text{aq})$ and $\text{Br}_2(\text{l})$ in Appendix H, calculate the solubility of Br_2 in water at 25°C . Express your answer as g/L.

Ans:



$$E_{+}^{\circ} = 1.078 \text{ V}$$

$$E_{-}^{\circ} = 1.098 \text{ V}$$

$$E^{\circ} = -0.020 \text{ V}$$

(46)

At equilibrium, $E=0$.

Therefore $0 = -0.020 - \frac{0.05916}{2} \log \frac{[Br_2(aq)]}{[Br_2(l)]}$

$$0 = -0.020 - 0.02958 \log k$$

$$\therefore 0.020 = -0.02958 \log k$$

$$\therefore k = \frac{[Br_2(aq)]}{[Br_2(l)]}$$

$$\log k = \frac{0.020}{-0.02958} = -0.6761$$

$$\therefore k = 10^{-0.6761} = 0.211 \text{ M.}$$

$$k = 0.211 \text{ M}$$

That is, the solubility of Br_2 in water

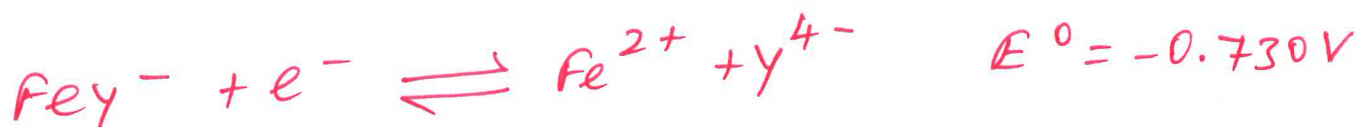
$$\text{is } 0.211 \text{ M} = 34 \text{ g/L}$$

13-30

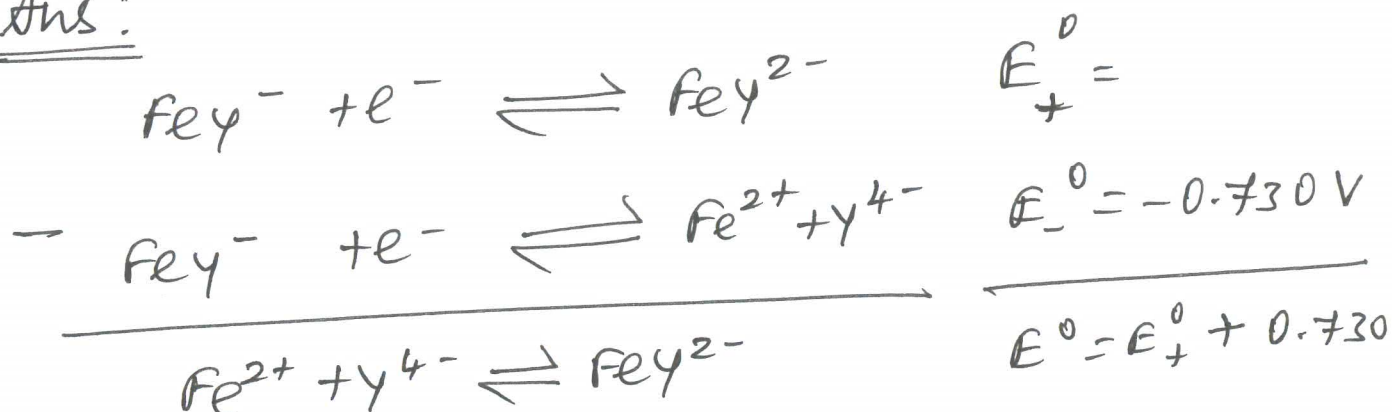
~~13-30~~. Given the following information

(where Y is EDTA).

A7



Calculate the standard potential for the reaction

Ans:

$$\text{But } E^0 = \frac{0.05916}{1} \log [k_f \text{ for FeY}^{2-}]$$

$$= 0.05916 \times \log (2.1 \times 10^{14})$$

$$\boxed{E^0 = 0.847 \text{ V}}$$

48

$$\therefore E^0 = E_+^0 - E_-^0$$

$$0.847 = E_+^0 - (-0.730)$$

$$0.847 = E_+^0 + 0.730$$

$$\therefore E_+^0 = 0.847 - 0.730$$

$$E_+^0 = 0.117 \text{ V}$$