

Review Sheet #3 Friday, December 12, 2003

1] Given that $\alpha_{y^{4-}} = 3.8e-9$ at pH 4.00 & $\alpha_{y^{4-}} = 1.9e-18$ at pH 1.00 what is the conditional formation constant for FeY^- at those pH's. $\log K_f = 25.1$ (2 minutes)

Remember that

$$K_f = [FeY^-] / [Fe^{3+}][Y^{4-}] \qquad [Y^{4-}] = \alpha_{y^{4-}}[EDTA]$$

$$K_f = [FeY^-] / [Fe^{3+}]\alpha_{y^{4-}}[EDTA]$$

$$K_f' = \alpha_{y^{4-}}K_f = [FeY^-] / [Fe^{3+}][EDTA]$$

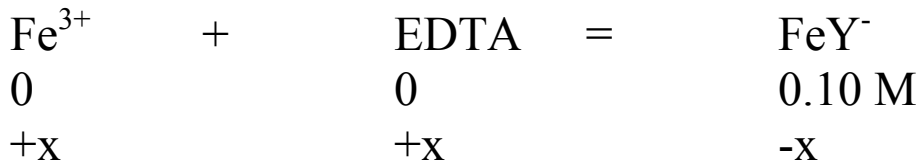


$$\text{At pH 4.00} \qquad K_f' = \alpha_{y^{4-}}K_f = 3.8e-9 * 1.3e25 = 4.9e16$$

$$\text{At pH 1.00} \qquad K_f' = 1.9e-18 * 1.3e25 = 2.5e7$$

2] Calculate the concentrations of free Fe^{3+} in a 0.10 M FeY^- solution at pH 4.00 and 1.00. (3 minutes)

2] answer



$$0.10 - x / x^2 = K_f'$$

$$x = 1.4\text{e-}9 \text{ @ pH } 4.00$$

$$x = 6.3\text{e-}5 \text{ @ pH } 1.00$$

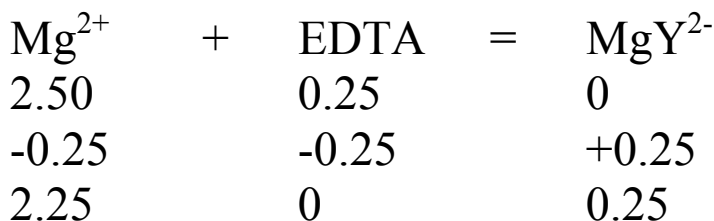
EDTA titrations

3] Calculate the concentration of free Mg^{2+} in a solution of 50.0 mL of 0.0500 M Mg^{2+} when 5.00 mL of 0.0500 M EDTA is added at pH 10.00. (3 minutes)



$$\text{Initial } \text{Mg}^{2+} = 0.0500 \text{ M} * 50.0 \text{ mL} = 2.50 \text{ mmol}$$

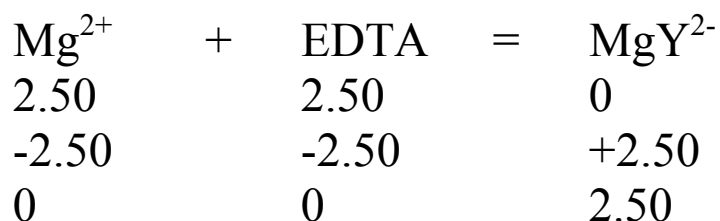
$$\text{Added EDTA} = 0.0500 * 5.00 \text{ mL} = 0.25 \text{ mmol}$$



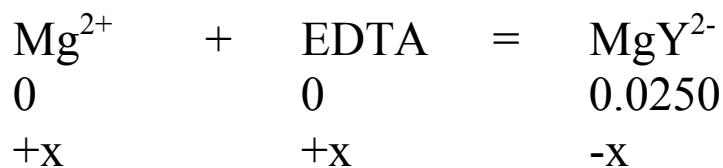
$$[\text{Mg}^{2+}] = 2.25 \text{ mmol} / 55.00 \text{ mL} = 0.0409 \quad \text{pMg} = 1.39$$

4] When 50.0 mL of 0.0500 M EDTA is added. (3 minutes)

added EDTA = 0.0500 M * 50.0 mL = 2.50 mmol



$[Mg^{2+}] = 2.50 \text{ mmol} / 100 \text{ mL} = 0.0250 \text{ M}$

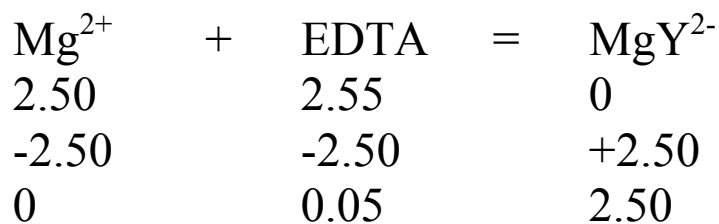


$$0.0250 - x / x^2 = 2.2e8$$

$$x = 1.07e-5 \quad pMg = 4.97$$

5] When 51.00 of 0.0500 M EDTA is added. (3 minutes)

$$\text{added EDTA} = 0.0500 \text{ M} * 51.0 \text{ mL} = 2.55 \text{ mmol}$$



$$[\text{MgY}^{2-}] = 2.50 \text{ mmol} / 101 \text{ mL} = 2.47\text{e-}2 \text{ M}$$

$$[\text{EDTA}] = 0.05 \text{ mmol} / 101 \text{ mL} = 4.95\text{e-}4$$

$$K_f' = [\text{MgY}^{2-}] / [\text{Mg}^{2+}][\text{EDTA}] = 2.47\text{e-}2 \text{ M} / [\text{Mg}^{2+}] * 4.95\text{e-}4$$

$$K_f' = 2.2\text{e}8$$

$$[\text{Mg}^{2+}] = 2.3\text{e-}7$$

$$\text{pMg} = 6.64$$

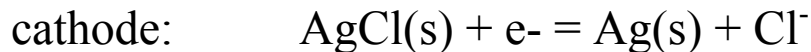
6) Calculate the E_{cell} for the following:

$\text{Cd(s)}/0.010 \text{ M Cd(NO}_3)_2(\text{aq})//0.50 \text{ M KCl(aq)}/\text{AgCl(s)}/\text{Ag(s)}$



(8 minutes)

6] answer



$$E = E^0 - 0.0592 \log [\text{Cl}^-] = 0.222 - 0.0592 \log (0.50) = 0.240 \text{ V}$$



$$E = E^0 - 0.0592/2 \log 1/[\text{Cd}^{2+}]$$

$$= -0.402 - 0.0592/2 \log 1/(0.0100) = -0.461 \text{ V}$$

$$E_{\text{cell}} = E_{\text{cath}} - E_{\text{anod}} = 0.240 - (-0.461) = 0.701 \text{ V}$$

Potentiometric titrations

7] A 20.00 mL solution of 0.1004 M KI was titrated with 0.0845 M AgNO₃. Calculate the voltage when V_{Ag⁺} = 15.00 mL. Assume the NHE.

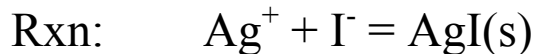
Draw cell silver wire + NHE



$$\mathbf{K_{sp}(\text{AgI}) = 8.3e-17}$$

(6 minutes)

7] answer



$$\text{Initial I}^- = 20.00 \text{ mL} * 0.1004 \text{ M} = 2.008 \text{ mmol}$$

$$\text{Added Ag}^+ = 15.00 \text{ mL} * 0.0845 \text{ M} = 1.27 \text{ mmol}$$

Ag^+	+	I^-	=	AgI(s)
1.27		2.008		--
-1.27		-1.27		--
0		0.740		--

$$[\text{I}^-] = 0.740 \text{ mmol} / 35.00 \text{ mL} = 0.0211 \text{ M}$$

$$E = E^0 - 0.0592 \log 1/[\text{Ag}^+]$$

$$K_{\text{sp}} = 8.3\text{e-}17 = [\text{Ag}^+][\text{I}^-]$$

$$[\text{Ag}^+] = 8.3\text{e-}17 / 0.0211 = 3.9\text{e-}15 \text{ M}$$

$$E = 0.799 - 0.0592 \log \{1 / 3.9\text{e-}15\} = - 0.0538 \text{ V}$$

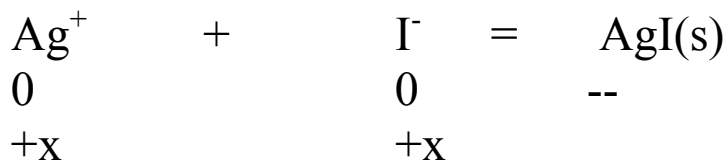
8] E_{cell} when $V_{\text{Ag}^+} = 23.76 \text{ mL}$

(5 minutes)

8] answer

$$\text{added Ag}^+ = 23.76 \text{ mL} * 0.0845 \text{ M} = 2.008 \text{ mmol}$$

This is the eq. pt.



$$x^2 = K_{sp} = 8.3e-17$$

$$x = 9.1e-9$$

$$E = E^0 - 0.0592 \log 1/[\text{Ag}^+]$$

$$E = 0.799 - 0.0592 \log \{1 / 9.1e-9\} = 0.323 \text{ V}$$

9] E_{cell} when $V_{\text{Ag}^+} = 25.00 \text{ mL}$ (3 minutes)

we are beyond the eq. pt.

$$\text{excess Ag}^+ = (25.00 - 23.76) \text{ mL} * 0.0845 \text{ M} = 0.105 \text{ mmol}$$

$$[\text{Ag}^+] = 0.105 \text{ mmol} / 45.00 \text{ mL} = 2.33e-3 \text{ M}$$

$$E = 0.799 - 0.0592 \log \{1 / 2.33e-3\} = 0.643 \text{ V}$$

10] Change each potential from the NHE scale to the SCE scale.

$$E_{\text{scale}} = 0.241 \text{ V}$$

$$\text{For \# 7} \quad E = -0.0538 \text{ V} - 0.241 \text{ V} = -0.294 \text{ V}$$

$$\text{For \# 8} \quad E = 0.323 - 0.241 = 0.082 \text{ V}$$

$$\text{For \# 9} \quad E = 0.643 - 0.241 = 0.402 \text{ V}$$