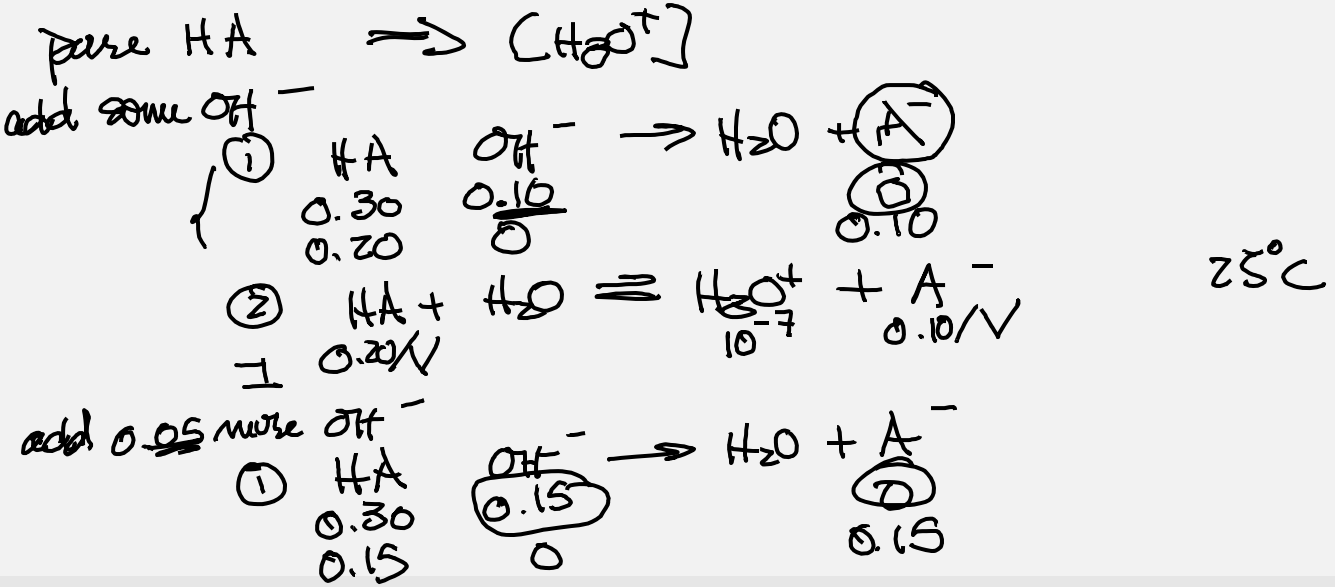
Lecture 20 CH131 Summer 1

Wednesday, June 26, 2019

- Complete practice problems



Electrochemistry practice problem: 17:35

$K = 2.8 \times 10^{31}$, so orange

35. By using the half-cell potentials in Appendix E, calculate the equilibrium constant at 25°C for the reaction in problem 33. Dichromate ion ($Cr_2O_7^{2-}$) is orange, and Cr^{3+} is light green in aqueous solution. If 2.00 L of 1.00 M $HClO_2$ solution is added to 2.00 L of 0.50 M $Cr(NO_3)_3$ solution, what color will the resulting solution have?

$$\log(K) = \frac{n_e \, 0.31\,V}{0.0592\,V} = \frac{6 \times 0.31}{0.0592}$$

$$= 31.31$$

$$K = 10^{31}$$

$$= 2.8 \times \ldots 31$$

$$3HClO_2(aq) + 2Cr^{3+}(aq) + 12H_2O(l) \rightarrow \text{green} \rightarrow \text{orange}$$

$$3HClO(aq) + Cr_2O_7^{2-}(aq) + 8H_3O^+(aq)$$

$$6e^- + 6H^+ + 3HClO_2 \rightarrow 3HClO + 3H_2O$$

$$7H_2O + 2Cr^{3+} \rightarrow Cr_2O_7^{2-} + 14H^+ + 6e^-$$

LR

$3HClO_2$ 2.00 + $2Cr^{3+}$ 1.00 → $Cr_2O_7^{2-}$

$$1.00\,Cr^{3+} \times \frac{1\,Cr_2O_7^{2-}}{2\,Cr^{3+}} = 0.50\,Cr_2O_7^{2-}$$

$$2.00\,HClO_2 \times \frac{1\,Cr_2O_7^{2-}}{3\,HClO_2} = 0.66\,Cr_2O_7^{2-}$$

Electrochemistry practice problem: 17.39

$[H_3O^+] = 2.9 \times 10^{-3}$

$K_a \approx [H_3O^+]$

39. A galvanic cell consists of a Pt $|H_3O^+(1.00\ M)|H_2(g)$ cathode connected to a Pt $|H_3O^+(aq)|H_2(g)$ anode in which the concentration of H_3O^+ is unknown but is kept constant by the action of a buffer consisting of a weak acid, HA(0.10 M), mixed with its conjugate base, A^-(0.10 M). The measured cell potential is $E_{cell} = 0.150$ V at 25°C, with a hydrogen pressure of 1.00 atm at both electrodes. Calculate the pH in the buffer solution, and from it determine the K_a of the weak acid.

$H_2(g) \rightarrow 2H^+ + 2e^-$ anode

$2H^+(1.00) \rightarrow 2H^+(x)$ cathode

$Q = \frac{(x)^2}{(1.00)^2} = x^2$

$E = E^\circ - \frac{0.05912\,V}{2}\log(Q) = 0.150\,V$

$E^\circ = 0$

$\log(Q) = \frac{-2 * 0.150\,V}{0.05912\,V} = \log(x^2) = 2\log(x)$

$\log(x) = \frac{0.150}{0.05912} = 2.54$

$x = 2.9 \times 10^{-3}$

	$HA + H_2O \rightleftharpoons H_3O^+ + A^-$		
I	0.10	10^{-7}	0.10
E	0.10 − x	$10^{-7} + x \approx x$	0.10 + x ≈ 0.10

$K_a = \frac{x(0.10)}{0.10} = x$

Electrochemistry practice problem: 17.41

cathode = reduction

anode = oxidation

25°C

41. A galvanic cell is constructed in which the overall reaction is

$$Br_2(\ell) + H_2(g) + 2\,H_2O(\ell) \longrightarrow 2\,Br^-(aq) + 2\,H_3O^+(aq)$$

(a) Calculate E° for this cell.

(b) Silver ions are added until AgBr precipitates at the cathode and $[Ag^+]$ reaches 0.060 M. The cell potential is then measured to be 1.710 V at pH = 0 and = 1.0 atm. Calculate $[Br^-]$ under these conditions.

(c) Calculate the solubility product constant K_{sp} for AgBr.

$E^\circ = E^\circ(Br_2|Br^-) - E^\circ(H^+|H_2)$

$= 1.065V - 0$

$= 1.065V$

$E = E^\circ - \frac{0.05912}{2}\log(Q)$

$Q = \frac{(Br^-)^2(H^+)^2}{P_{H_2}} = (Br^-)^2$

$P_{H_2} = 1$ atm

$(H^+) = 10^{-pH} = 1$

$1.710V = 1.065V - \frac{0.05912V}{2}\log[Br^-]^2$

$\left(\frac{1.710 - 1.065}{0.05912}\right) = -\log(Br^-)$

$[Br^-] = 1.2 \times 10^{-11}$

$AgBr(s) \rightleftharpoons Ag^+(aq) + Br^-(aq)$

0.060 $\quad$ 1.2×10^{-11}

$K_{sp} = [Ag^+][Br^-]$

$= 0.060 \times 1.2\times10^{-11}$

$= 7.2 \times 10^{-13}$