Chemistry 350 | Spring 2016
20 points | Due Monday, February 29

Problem Set 5 –Equilibrium Review

Complete all problems on separate paper. Show all work for credit.

1. Use the solubility product to calculate the solubility of CuBr in water expressed as (a) moles per liter and (b) grams per 100 mL.

	CuBr	=	Cu^+	+	Br^-
i	-		0		0
c	-		+x		+x
e	-		x		x

a

$$K_{sp} = [Cu^+][Br^-] = (x)(x) = 5 \times 10^{-9}$$
$$x = (K_{sp})^{1/2} = \mathbf{7.1\times10^{-5}\ M}$$

b

$$\frac{7.1\text{x}10\text{-}5 \text{ mol Cu}}{1 \text{ L}} \times \frac{143.45 \text{ g Cu}}{1 \text{ mol Cu}} = \frac{0.010 \text{ g}}{\text{L}} = \frac{\mathbf{0.0010\ g}}{\mathbf{100\ mL}}$$

2. A solution contains 0.25 M $Ni(NO_3)_2$ and 0.25 M $Cu(NO_3)_2$. Can the metal ions be separated by slowly adding Na_2CO_3? Assume that for successful separation, 99% of the metal ion must be precipitated before the other metal ion begins to precipitate, and assume that no volume change occurs upon addition of Na_2CO_3. (Ignore activities.)

Two equilibria to consider:

$NiCO_3 = Ni^{2+} + CO_3^{2-}$ $\quad K_{sp} = 1.3 \times 10^{-7}$

$CuCO_3 = Cu^{2+} + CO_3^{2-}$ $\quad K_{sp} = 2.3 \times 10^{-10}$

1. What $[CO_3^{2-}]$ is needed to lower each ion's concentration to 1% of its initial value?

Target [] = 0.01(0.25M) = 2.5×10^{-3} M

$$Ni^{2+}: [CO_3^{2-}] = \frac{K_{sp}}{[Ni^{2+}]} = \frac{1.3 \times 10^{-7}}{2.5 \times 10^{-3}\text{ M}} = \mathbf{5.2 \times 10^{-5}\ M\ CO_3^{2-}}$$

$$Cu^{2+}: [CO_3^{2-}] = \frac{K_{sp}}{[Cu^{2+}]} = \frac{2.3 \times 10^{-10}}{2.5 \times 10^{-3}\text{ M}} = \mathbf{9.2 \times 10^{-8}\ M\ CO_3^{2-}}$$

So, Cu^{2+} will precipitate first.

2. Will Ni^{2+} precipitate if $[CO_3^{2-}] = 9.2 \times 10^{-8}$ M?

$Q = [Ni^{2+}][CO_4^{2-}] = (0.25\text{ M})(9.2 \times 10^{-8}\text{ M}) = 2.3 \times 10^{-8}$

Since $Q<K_{sp}$ for $NiCO_3$, Ni^{2+} will not precipitate before $[Cu^{2+}] = 2.5 \times 10^{-3}$ M.
Separation is feasible.

3. Find the pH and pOH of a 1.37×10^{-5} M solution of iodoacetic acid. (Ignore activities.)

	HA	=	A-	+	H^+
I	1.37×10^{-5}		0		0
C	-x		+x		+x
E	1.37×10^{-5} -x		x		x

For iodoacetic acid, $pK_a = 3.175$, $K_a = 6.68 \times 10^{-4}$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{x^2}{1.37 \times 10^{-5} - x} = 6.68\times10^{-4}$$

After some algebra, we find $x^2+K_a x-1.37 \times 10^{-5}K_a = 0$ and with the quadratic equation, we find $x = [H^+] = 1.34\times10^{-5}$ M and **pH = 4.87, pOH = 14-pH = 9.13**

4. An acidic solution containing 0.010 M La^{3+} is treated with NaOH until $La(OH)_3$ precipitates. At what pH does this occur?

$$La(OH)_3 = La^{3+} + 3OH^-$$
$$K_{sp} = [La^{3+}][OH^-]^3$$
$$2 \times 10^{-21} = (0.010\ M)[OH^-]^3$$
$$[OH^-] = (2\times10^{-21}/0.010)^{1/3} = 5.85\times10^{-7} M$$

pOH = 6.23, pH = 7.77

5. In a 0.030 M solution of the weak base, B, 0.27% of B underwent hydrolysis to make BH^+. Find K_b for the base.

If 0.27% undergoes reaction, (0.27/100)*0.030M = 8.1×10^{-5} M BH^+ is produced

	B	+	H_2O	=	OH^-	+	BH^+
I	0.030		-		0		0
C	-8.1×10^{-5}		-		$+8.1\times10^{-5}$		$+8.1\times10^{-5}$
E	0.02992		-		8.1×10^{-5}		8.1×10^{-5}

$$K_b = \frac{[OH^-][BH^+]}{[B]} = \frac{(8.1\times10^{-5})(8.1\times10^{-5})}{0.02992} = \mathbf{2.19\times10^{-7}}$$