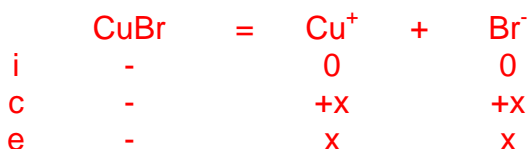


Problem Set 5 –Equilibrium Review

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.



a

$$K_{sp} = [Cu^+][Br^-] = (x)(x) = 5 \times 10^{-9}$$

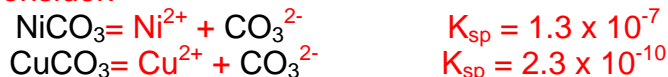
$$x = (K_{sp})^{1/2} = \mathbf{7.1 \times 10^{-5} M}$$

b

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

2. A solution contains 0.25 M Ni(NO₃)₂ and 0.25 M Cu(NO₃)₂. Can the metal ions be separated by slowly adding Na₂CO₃? 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 Na₂CO₃. (Ignore activities.)

Two equilibria to consider:



1. What [CO₃²⁻] is needed to lower each ion's concentration to 1% of its initial value?

Target [] = 0.01(0.25M) = 2.5 x 10⁻³ M

$$\text{Ni}^{2+}: [\text{CO}_3^{2-}] = \frac{K_{sp}}{[\text{Ni}^{2+}]} = \frac{1.3 \times 10^{-7}}{2.5 \times 10^{-3} \text{ M}} = \mathbf{5.2 \times 10^{-5} \text{ M CO}_3^{2-}}$$

$$\text{Cu}^{2+}: [\text{CO}_3^{2-}] = \frac{K_{sp}}{[\text{Cu}^{2+}]} = \frac{2.3 \times 10^{-10}}{2.5 \times 10^{-3} \text{ M}} = \mathbf{9.2 \times 10^{-8} \text{ M CO}_3^{2-}}$$

So, Cu²⁺ will precipitate first.

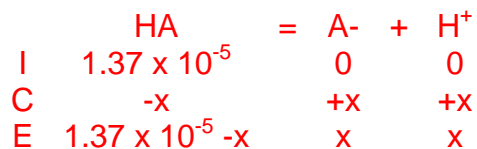
2. Will Ni²⁺ precipitate if [CO₃²⁻] = 9.2 x 10⁻⁸ M?

$$Q = [\text{Ni}^{2+}][\text{CO}_4^{2-}] = (0.25 \text{ M})(9.2 \times 10^{-8} \text{ M}) = 2.3 \times 10^{-8}$$

Since Q < K_{sp} for NiCO₃, Ni²⁺ will not precipitate before [Cu²⁺] = 2.5 x 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.)

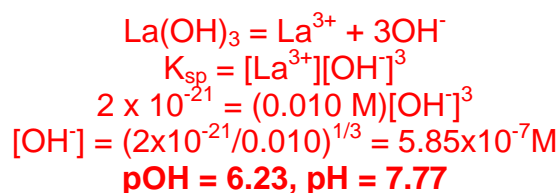


For iodoacetic acid, $pK_a = 3.175$, $K_a = 6.68 \times 10^{-4}$

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

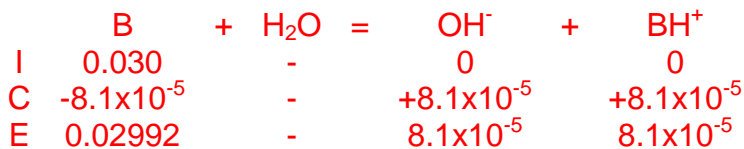
After some algebra, we find $x^2 + K_a x - 1.37 \times 10^{-5} K_a = 0$ and with the quadratic equation, we find $x = [H^+] = 1.34 \times 10^{-5}$ M and **pH = 4.87, pOH = 14 - pH = 9.13**

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?



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, $(0.27/100) \times 0.030 \text{ M} = 8.1 \times 10^{-5} \text{ M } BH^+$ is produced



$$K_b = \frac{[OH^-][BH^+]}{[B]} = \frac{(8.1 \times 10^{-5})(8.1 \times 10^{-5})}{0.02992} = \mathbf{2.19 \times 10^{-7}}$$