IN-Class Group QUIZ spring 2017 Name: ____________ ____________ ____________

1. Consider titrating 100.0 mL of 0.200 M acetic acid ($K_a = 1.8 \times 10^{-5}$) with 0.100 M NaOH. Calculate the pH of the resulting solution at the following points of the titration:

A. 0.0 mL of NaOH have been added 2.72 pure weak acid

B. 50.0 mL of NaOH have been added 4.26 Buffer zone

C. half-way point pH = pKA 4.74

D. equivalence point 8.78

E. 250.0 mL of NaOH have been added 12.15

F. Which of the following would be the best indicator to use for this titration? Justify your answer.

Methyl red $K_a = 1.0 \times 10^{-5}$ pKa = 5

Thymol blue $K_a = 1.3 \times 10^{-9}$ pKa = 8.89

Alizarin yellow $K_a = 6.3 \times 10^{-12}$ pKa = 11.20

Thymol Blue since its pKa is closest to the Equiv. pt

2. If 25.0 mL of 0.10 M $NH_3(aq)$ (K_b for NH_3 is 1.8×10^{-5} at 25 °C) and 60.0 mL of 0.20 M $NH_4Cl(aq)$ are mixed, determine:

A. the pH of the resulting solution 8.58

B. the pH of the resulting solution after 10.0 mL of 0.20 M HCl(aq) is added 5.04 if .25M HCl $\boxed{\text{7.81 for .2M}}$

3. Exactly 0.400 L of 0.50 M Pb^{2+} & 1.60 L of 2.5×10^{-8} M Cl^- are mixed together to form 2.00L. Calculate Q and predict if a ppt will occur. What if 2.5×10^{-2} Cl^- was used? $K_{sp} = 1.17 \times 10^{-5}$

A: $Q = 4 \times 10^{-17}$ NO

B: $Q = 4 \times 10^{-5}$ yes ppt

4. What is the molar solubility of lead(II) chloride in 1.0 L of solution that contains 2.0×10^{-2} mol of HCl? 0.00857M $\boxed{\text{approx = .02M}}$

5. Consider zinc hydroxide, $Zn(OH)_2$, where $K_{sp} = 3 \times 10^{-17}$.

A. Determine the solubility of zinc hydroxide in pure water. Molar $S = 1.96 \times 10^{-6}$ M

B. How does the solubility of zinc hydroxide in pure water compare with that in a solution buffered at pH 6.00? Quantitatively demonstrate the difference (if any) in solubility. Is zinc hydroxide more or less soluble at pH 6.00? pOH = 8

3×10^{-5} M Approx [redacted] M Because the solubility has increased from pure H_2O, the solubility has ~~increased~~ more soluble at pH 6

C. If enough base is added, the OH^- ligand can coordinately bind with the Zn^{+2} ion to form the soluble zincate ion, $[Zn(OH)_4]^{-2}$. The formation constant, K_f, of the full complex ion $[Zn(OH)_4]^{-2}$ can be calculated from the following successive equilibrium expressions shown:

$Zn^{2+}(aq) + OH^- \leftrightarrow ZnOH^+(aq)$ $K_1 = 2.5 \times 10^4$

$ZnOH^+(aq) + OH^-(aq) \leftrightarrow Zn(OH)_2(s)$ $K_2 = 8.0 \times 10^6$

$Zn(OH)_2(s) + OH^-(aq) \leftrightarrow Zn(OH)_3^-(aq)$ $K_3 = 70$

$Zn(OH)_3^-(aq) + OH^-(aq) \leftrightarrow Zn(OH)_4^{2-}(aq)$ $K_4 = 33$

$K_T = K_1 + K_2 + K_3 + K_4$

Determine the value of K_f for the zincate ion. $\boxed{4.62 \times 10^{14}}$ K_T

$Cr(OH)_3$ $\boxed{K_{sp} = 6.3 \times 10^{-31}}$

6. Calculate the free ion concentration of Cr^{3+} when 0.01 moles of chromium(III) nitrate is dissolved in 2.00 liters of a pH 10 buffer. Cr^{3+} $\boxed{.00497M}$

$[H^+] = 1 \times 10^{-10}$ $[OH^-] = 1 \times 10^{-4}$ M

7. Calculate the pH required to precipitate out ZnS from a solution mixture containing 0.010 M Zn^{2+} and 0.01M Cu^{2+}. Will CuS precipitate out under these conditions? $ZnS = 2 \times 10^{-25}$ $CuS = 1.27 \times 10^{-36}$

8. Will a precipitate of silver carbonate form ($K_{sp} = 8.46 \times 10^{-12}$) when 100.0 mL of 1.00×10^{-4} M $AgNO_3(aq)$ and 200.0 mL of 3.00×10^{-3} M $Na_2CO_3(aq)$ are mixed? What will be the remaining concentration of ions present in solution?

$Q_{sp} < K_{sp}$ no ppt

$\frac{.00001 \text{ mol } Ag^+}{.3} = .000033M\ Ag^+$ $Q = [Ag^+]^2[CO_3^{2-}]$ $(1 \times 10^{-4} \frac{mol}{L})(2L) = \frac{.0002m}{2L}$

$.002M\ CO_3^{2-}$ 2.22×10^{-12}

$K = 8.5 \times 10^{-12}$

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