

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 $pH = pK_a$ 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 | $K_a = 1.0 \times 10^{-5}$ | $pK_a = 5$ | <i>Thymol blue since its pKa is closest to the equiv. pt</i> |
| Thymol blue | $K_a = 1.3 \times 10^{-9}$ | $pK_a = 8.89$ | |
| Alizarin yellow | $K_a = 6.3 \times 10^{-12}$ | $pK_a = 11.20$ | |

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 $0.25M$ HCl 7.81 for $0.2M$

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

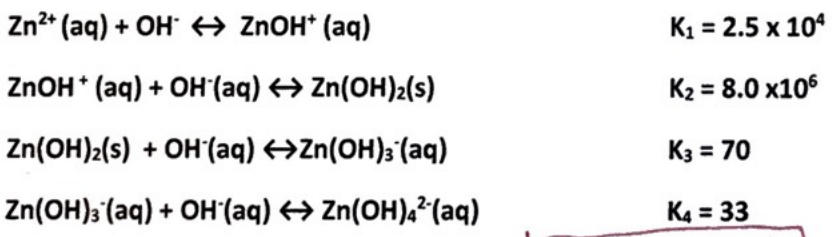
4. What is the molar solubility of lead(II) chloride in 1.0 L of solution that contains 2.0×10^{-2} mol of HCl? $Q = 4 \times 10^{-17}$ NO *yes ppt* $0.00857M$ $approx = 0.02M$

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

- A. Determine the *molar* solubility of zinc hydroxide in pure water. $S = 1.96 \times 10^{-6} M$
- 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? $pOH = 8$

6.00? $3 \times 10^{-5} M$ *Approx* ~~0.02M~~ *Because the solubility has increased from pure H2O, 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:



$K_T = K_1 + K_2 + K_3 + K_4$

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

4.162×10^{14} K_T

Determine the value of K_f for the zincate ion.

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. $[Cr^{3+}] = 1.00497M$

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 $1.00 \times 10^{-4} M AgNO_3(aq)$ and 200.0 mL of $3.00 \times 10^{-3} M Na_2CO_3(aq)$ are mixed? What will be the remaining concentration of ions present in solution?

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

$Q = [Ag^+]^2 [CO_3^{2-}]$
 2.22×10^{-12}
 $K = 8.5 \times 10^{-12}$

$\frac{0.0001 \text{ mol } Ag^+}{0.3} = 0.00033M Ag^+$
 $\frac{0.006 \text{ mol } CO_3^{2-}}{0.3} = 0.02M CO_3^{2-}$

$(1 \times 10^{-4} \frac{mol}{L}) (2L) = \frac{0.0002 \text{ mol}}{2L}$

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 - Methyl red $K_a = 1.0 \times 10^{-5}$ $pK_a = 5$
 - Thymol blue** $K_a = 1.3 \times 10^{-9}$ $pK_a = 8.89$
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Thymol blue since its pKa is closest to the equiv. pt

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$Q = 4 \times 10^{-17}$ NO *yes ppt*
 $Q = 4 \times 10^{-5}$ YES *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? $0.0085M$
 $approx = 0.02M$

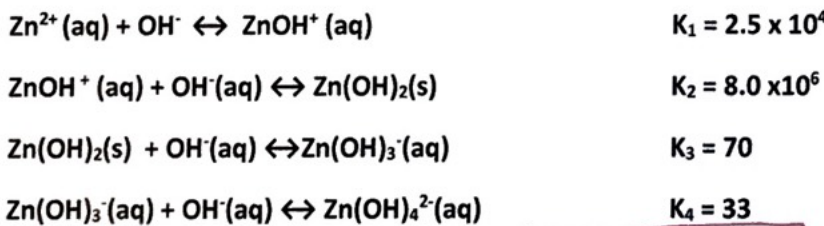
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B. How does the solubility of zinc hydroxide in pure water compare with that in a solution buffered at pH 6.00? Quantitatively demonstrate the difference (if any) in solubility. Is zinc hydroxide more or less soluble at pH 6.00?

$3 \times 10^{-5} M$ *approx* ~~3.33 x 10^-6 M~~ *Because the solubility has increased from pure H2O, 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:



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$K_f = 6.3 \times 10^{-31}$

4.62×10^{14} K_T

Determine the value of K_f for the zincate ion.

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.

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

$1.00497 M$

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? $ZnS = 2 \times 10^{-25}$ $CuS = 1.27 \times 10^{-36}$

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$\frac{0.00001 \text{ mol } Ag^+}{0.3} = 3.33 \times 10^{-5} M Ag^+$

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

$Q = [Ag^+]^2 [CO_3^{2-}]$
 2.22×10^{-12}
 $K = 8.5 \times 10^{-12}$

$(1 \times 10^{-4} \frac{mol}{L}) (2L) = \frac{0.0002 \text{ mol}}{2L}$