

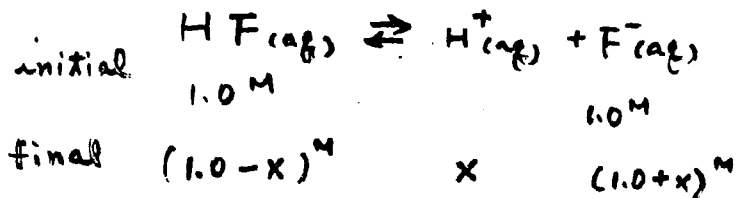
- I. Acid-Base Equilibria
- II. Solubility Equilibria
- III. Complex Ion Equilibria

§ 15.1 Solutions of Acids or Bases Containing a Common Ion

Ex 15.1

Calculate  $[H^+]$  and the percent dissociation of HF in a solution containing 1.0 M HF ( $K_a = 7.2 \times 10^{-4}$ ) and 1.0 M NaF.

sol: major species  $H^+$ ,  $F^-$ ,  $Na^+$ , and  $H_2O$



$$K_a = \frac{(1.0 + x)x}{1.0 - x} = 7.2 \times 10^{-4}$$

$$\approx \frac{1.0x}{1.0} \quad x = 7.2 \times 10^{-4}$$

$$\therefore \text{percent dissociation} = \frac{7.2 \times 10^{-4}}{1.0} \times 100\% = 0.072\% \quad \#$$

A buffered solution: one that resists a change in its pH when either  $OH^-$  or  $H^+$  are added.

Ex (v) Blood !!

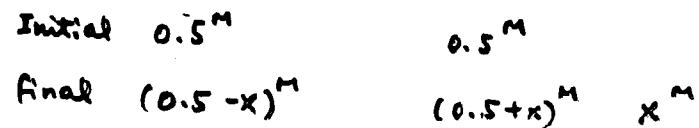
(Take-home exercise! Why blood is a buffer?)

- (1) weak acid and its salt
- (2) weak base and its salt

Ex 15.2

A buffered solution contains 0.5 M acetic acid ( $CH_3COOH$ )  $K_a = 1.8 \times 10^{-5}$  and 0.5 M  $NaCH_3COO$ . Calculate the pH of this solution.

sol:



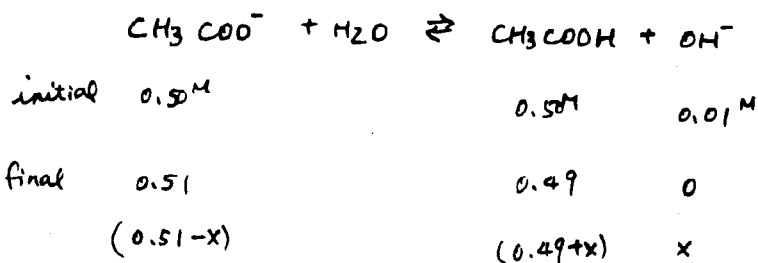
$$K_a = \frac{(0.5 + x)x}{(0.5 - x)} \approx x = 1.8 \times 10^{-5}$$

$$\therefore x = 1.8 \times 10^{-5}$$

$$pH = -\log 1.8 \times 10^{-5} = 4.74 \quad \#$$

Ex 15.5: Buffer. 1L soln 0.01 mol solid NaOH. please calculate pH change

Sol:



$$K_b = \frac{K_w}{K_a} = \frac{10^{-14}}{1.8 \times 10^{-5}} = \frac{x(0.49+x)}{(0.51-x)} \approx x \frac{0.49}{0.51}$$

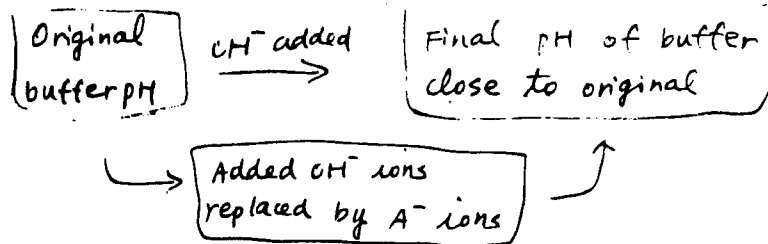
$$[\text{OH}^-] = 5.5 \times \frac{0.51}{0.49} \times 10^{-10} = 5.6 \times 10^{-10}$$

$$[\text{H}^+] = \frac{10^{-14}}{5.6 \times 10^{-10}} = 1.7 \times 10^{-5}$$

$$\therefore \text{pH} = -\log [\text{H}^+] \approx 4.76$$

only 0.02 pH difference.

Buffering: How does it work?



$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

$$[\text{H}^+] = K_a \frac{[\text{HA}]}{[\text{A}^-]}$$

$\therefore$  pH is determined by the ratio of  $\frac{[\text{HA}]}{[\text{A}^-]}$  !!

$\therefore$  当  $\text{OH}^-$  is added, 少量  $\text{HA} \rightarrow \text{A}^-$ .  
 $\frac{[\text{HA}]}{[\text{A}^-]} \approx \text{constant}$

$\therefore$  pH 不变 !!

Original  
 $\frac{[\text{HA}]}{[\text{A}^-]}$

Final  
 $\frac{[\text{HA}]}{[\text{A}^-]}$  close to original

$\downarrow$  The  $\text{OH}^-$  added changes  $\text{HA}$  to  $\text{A}^-$ , but  $[\text{HA}]$  and  $[\text{A}^-]$  are large compared to the  $[\text{OH}^-]$  added

$$[H^+] = K_a \frac{[HA]}{[A^-]}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$-\log [H^+] = -\log \left( K_a \frac{[HA]}{[A^-]} \right)$$

$$= -\log K_a - \log \frac{[HA]}{[A^-]}$$

$$= -\log K_a + \log \frac{[A^-]}{[HA]}$$

$$pH = pK_a + \log \left( \frac{[base]}{[acid]} \right)$$

Henderson - Hasselbalch equation

For a particular buffering system (acid-conjugate base pair)

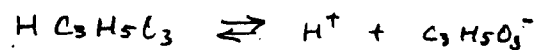
all solutions that have the same ratio

$\frac{[A^-]}{[HA]}$  will have the same pH !!

Ex 15.4

calculate pH of a solution containing 0.75 M lactic acid ( $K_a = 1.4 \times 10^{-4}$ ) and 0.25 M sodium lactate.  $HC_3H_5O_3$

sol: major species:  $HC_3H_5O_3$ ,  $Na^+$ ,  $C_3H_5O_3^-$ , and  $H_2O$



$$K_a = \frac{[H^+][C_3H_5O_3^-]}{[HC_3H_5O_3]} = 1.4 \times 10^{-4}$$

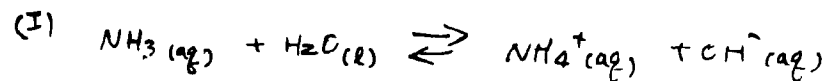
$$\therefore [H^+] = 1.4 \times 10^{-4} \times \frac{[HC_3H_5O_3]}{[C_3H_5O_3^-]} = 4.2 \times 10^{-4} M$$

Ex. 15.5

0.25 M  $NH_3$  ( $K_b = 1.8 \times 10^{-5}$ ) & 0.40 M  $NH_4Cl$ .

calculate the pH of the solution.

sol: major species:  $NH_3$ ,  $NH_4^+$ ,  $Cl^-$  and  $H_2O$



$$K_b = 1.8 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

initial 0.25 M

0.40 M

0.25 - x

(0.40 + x)

x

$$K_b = 1.8 \times 10^{-5} = \frac{(0.40 + x)x}{(0.25 - x)} \approx \frac{0.40}{0.25} x$$

$$x = 1.1 \times 10^{-5}$$

$$\therefore pH = -\log(1.1 \times 10^{-5}) = 4.95$$

$$pH = 9.05$$

(II) If use Henderson - Hasselbalch equation

$$pH = pK_a + \log \left( \frac{[base]}{[acid]} \right)$$

$$= p \left( \frac{K_w}{K_b} \right) + \log \left( \frac{[NH_3]}{[NH_4^+]} \right)$$

$$= 9.25 + \log \left( \frac{0.25}{0.40} \right)$$

$$= 9.05$$

15.6

Buffering capacity: the amount of protons or hydroxide ions the buffer can absorb without a significant change in pH

The capacity of a buffered solution is determined by the magnitudes of  $[HA]$  and  $[A^-]$ .

( $[HA]$  &  $[A^-]$  高, buffering capacity 高)

The pH of a buffered solution is determined by the ratio  $\frac{[A^-]}{[HA]}$

(Henderson-Hasselbalch equation)

$$pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

Ex 15.7

Calculate the change in pH that occurs when 0.010 mol gaseous HCl is added to 1.0 L of each of following solutions: acetic acid  $K_a = 1.8 \times 10^{-5}$

Solution A:  $5.00 M$   $HCH_3COO$  &  $5.00 M$   $NaCH_3COO$

Solution B:  $0.050 M$   $HCH_3COO$  &  $0.050 M$   $NaCH_3COO$

sol: (1) pH of soln A:  $pK_a + \log\left(\frac{[A^-]}{[HA]}\right) = pK_a + \log 1 = 4.74$   
& soln B

(2) main part:  $HCH_3COO$ ,  $CH_3COO^-$ ,  $H^+$ ,  $OH^-$



soln A  
initial  $5.00 M$                        $5.00 M$        $0.01 M$

final  $5.01$                                $4.99$        $0$

$5.01 - x$                                $4.99 + x$        $x$

$$K_a = \frac{(4.99+x) \cdot x}{5.01-x} \approx \frac{4.99}{5.01} \cdot x = 1.8 \times 10^{-5}$$

$$x = (1.8 \times 10^{-5}) \cdot \frac{5.01}{4.99} = 1.81 \times 10^{-5} \quad \text{or } pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right)$$

$$pH = -\log(1.81 \times 10^{-5}) \approx 4.74 \quad \# \quad = 4.74 + \log\left(\frac{4.99}{5.01}\right) \approx 4.74$$

soln B  
initial  $0.05$                                $0.05$        $0.01$

final  $0.06$                                $0.04$        $0$

$0.06 - x$                                $0.04 + x$        $x$

$$K_a = \frac{(0.04+x) \cdot x}{(0.06-x)} = 1.8 \times 10^{-5}$$

$$x = 2.7 \times 10^{-5}$$

$$pH = -\log(2.7 \times 10^{-5}) = 4.56 \quad \#$$

solution A, which contains larger quantities of HA &  $A^-$ , has higher buffering capacity!

\* pKa of the weak acid to be used as the buffer should be as close as possible to the desired pH.

$$pH = pKa + \log \left( \frac{[A^-]}{[HA]} \right)$$

↓ when  $\frac{[A^-]}{[HA]} = 1$  (most effective buffer)

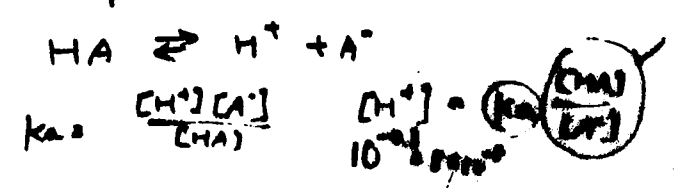
$$pH = pKa \quad \text{when} \quad \frac{[A^-]}{[HA]} = 1 \quad \frac{1}{1} = 1 \quad \frac{1}{0.1} = 10$$

Ex 15.8 preparing a buffer

A chemist needs a solution buffered at pH 4.30 please choose one from the following acids

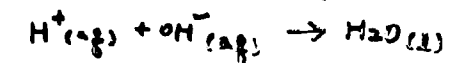
- a. chloroacetic acid ( $pKa = 1.35 \times 10^{-3}$ )
- b. propanoic acid ( $Ka = 1.3 \times 10^{-5}$ )
- c. benzoic acid ( $Ka = 6.4 \times 10^{-5}$ )
- d. hypochlorous acid ( $Ka = 3.5 \times 10^{-8}$ )

calculate  $\frac{[HA]}{[A^-]}$  required for each system to yield a pH of 4.30

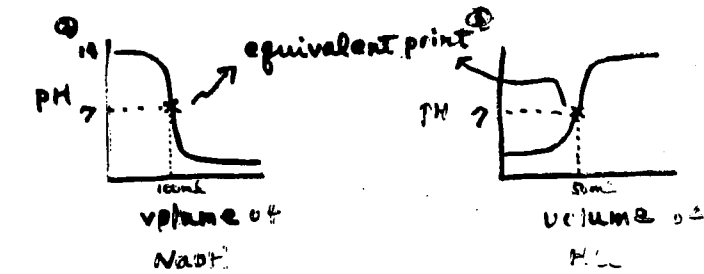


The progress of an ~~appropriate~~ acid-base titration is often monitored by plotting the pH of the solution being analyzed as a function of the amount of titrant added. Such a plot is called pH curve or titrant curve.

I. Strong Acid - Strong Base Titration



Ex. 9 50.0 ml of 0.200 M  $HNO_3$ , titrated w/ 0.100 M NaOH. (see page 739)  
 100.0 ml of 0.50 M NaOH w/ 1.0 M HCl

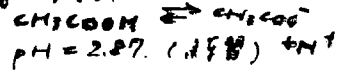


II. Weak acid - strong base titration

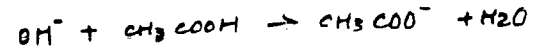
(see page 740)  
 Step 1 a stoichiometry problem. The reaction of hydroxide ( $OH^-$ ) w/ the weak acid is assumed to run to completion, the conc. of the acid remaining and the conjugate base formed are determined

Ex. 50.0 ml of 0.10 M acetic acid

w/ 0.10 M NaOH (see page 744)



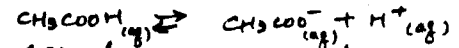
- A. No NaOH has been added
- B. 10.0 ml of 0.10 M NaOH has been added
- (a) the stoichiometry problem



before 10.0 ml x 0.10 M = 1.0 mmole  
50.0 ml x 0.10 M = 5.0 mmole

after 0 mmole 4.0 mmole 1.0 mmol

(b) The equilibrium problem



$$K_a = \frac{[\text{H}^+][\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = \frac{(1.0+x)x}{(4.0-x)}$$

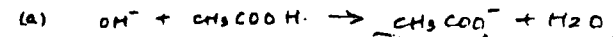
4.0 mmole (4-x)  
1.0 mmole (1.0+x)  
x

$$= 1.8 \times 10^{-5}$$

$$x \approx 7.24 \times 10^{-5}$$

$$\text{pH} = 4.14$$

(c) 25.0 ml (total) of 0.10 M NaOH has been added.



before 2.5 mmol 5.0 mmol 0

after 0 mmol 2.5 mmol 2.5 mmol

half way of the equivalence point



before 2.5 mmol 2.5 mmol

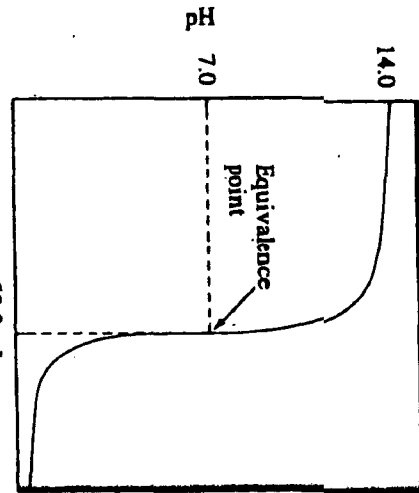
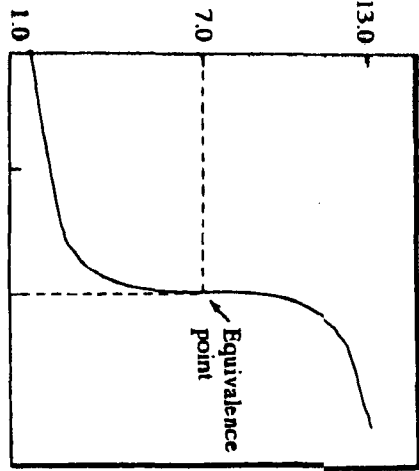
after  $\frac{(2.5-x)}{50+25}$  M

$\frac{(2.5+x)}{(50+25)}$  M

$$K_a = \frac{(2.5+x)x}{(2.5-x)}$$

$$\approx x$$

$$\therefore [\text{H}^+] = K_a$$



$$\text{pH} = -\log 0.2 = 0.69$$

$$0.2 \text{ M HNO}_3 \rightarrow 0.2 \text{ M H}^+$$

Figure 15.1 and 15.2 The pH curves for the titration of a strong acid and a strong base



initial  $\frac{5.0 \text{ mmol}}{(50.0 + 50.0) \text{ mL}}$   
 $= 0.050 \text{ M}$

after  $0.050 - x$                        $x$                        $x$

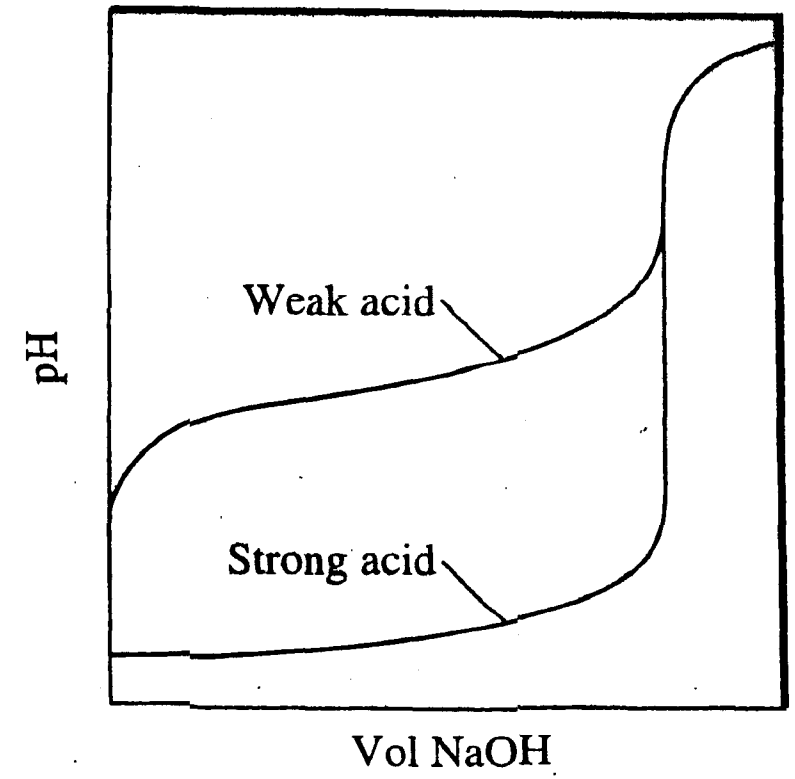
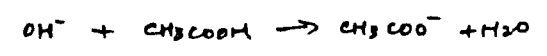
$$K_b = \frac{x^2}{(0.050 - x)} = \frac{K_w}{K_a} \approx \frac{10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$x \approx 5.3 \times 10^{-6} = [\text{OH}^-]$$

$$\text{pH} = 14 + \log [\text{OH}^-] = 8.72$$

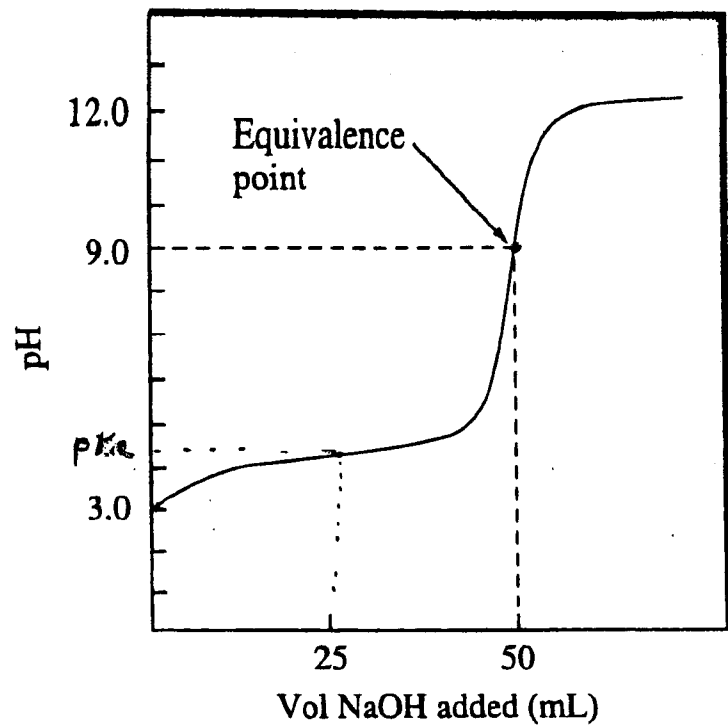
pH > 7 when the equivalent point reached!

(F) 过量



**Comparison of strong and weak acid titration curves**

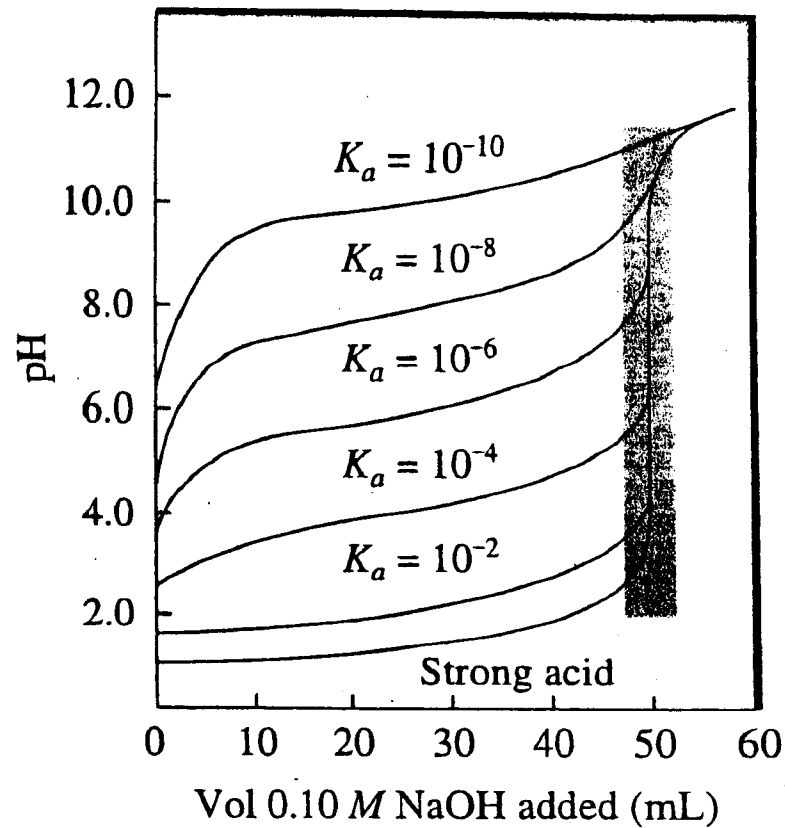
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**Figure 15.3**  
**The titration curve for a weak acid**



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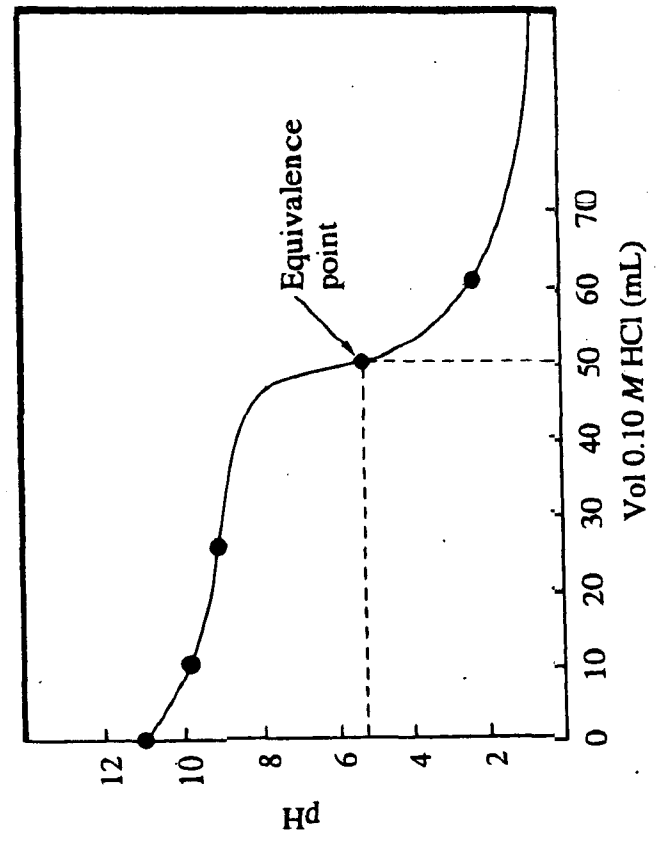
**Figure 15.4**  
**The shape of the titration curve as a function of  $K_a$**



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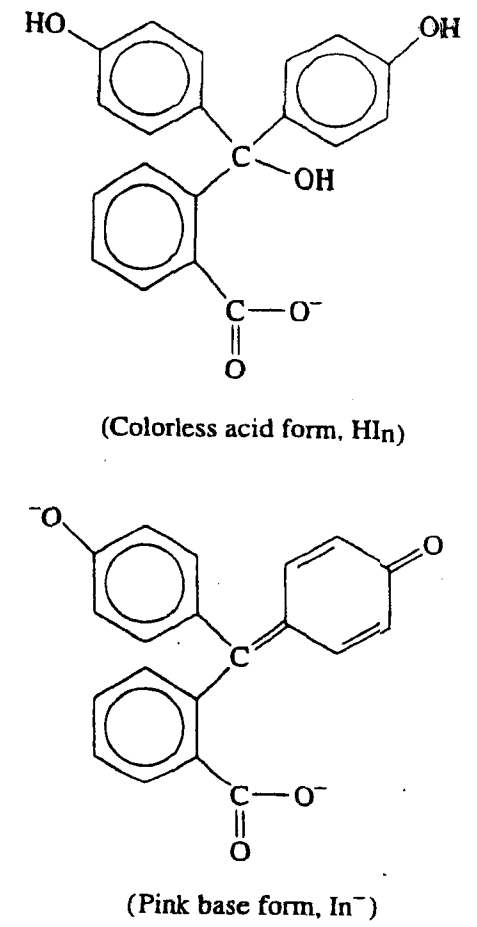
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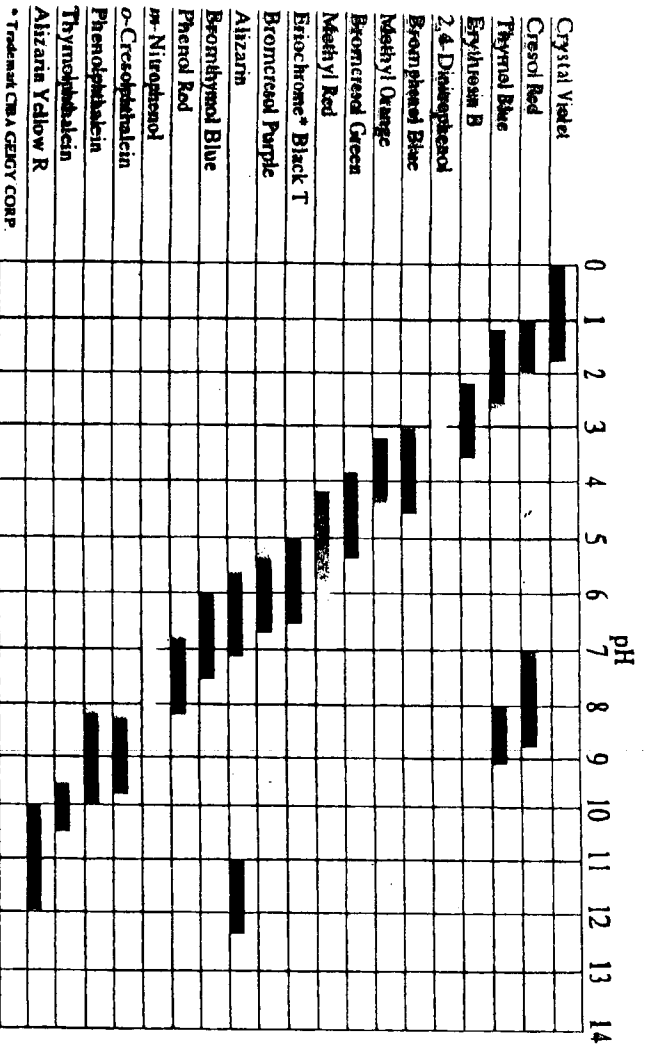


**Figure 15.5**  
**The pH curve for the titration of NH<sub>3</sub> with HCl**  
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**Figure 15.6**  
**The acid and base forms of phenolphthalein**  
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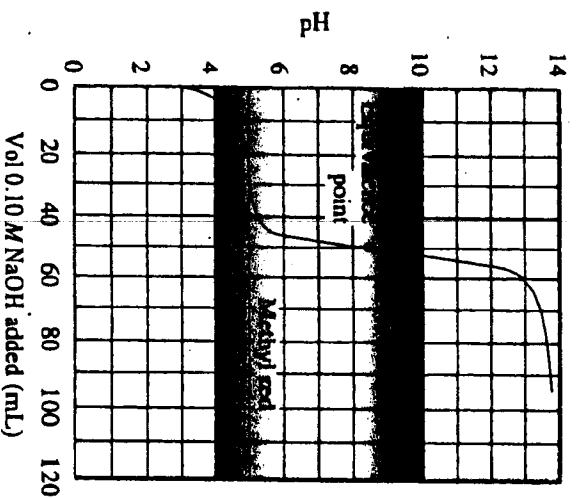
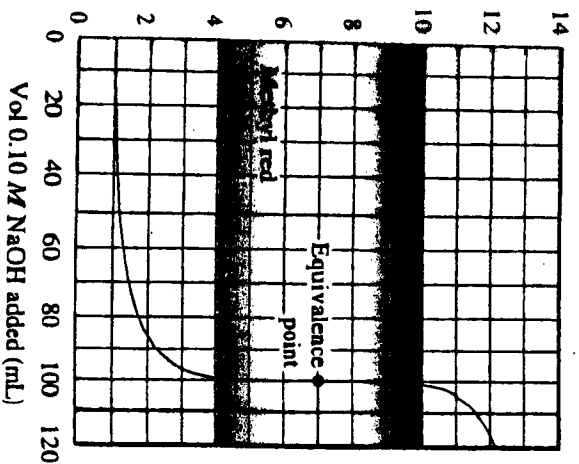


The pH ranges shown are approximate. Specific transition ranges depend on the indicator solvent chosen.

Figure 15.8

**The pH ranges of various common indicators**

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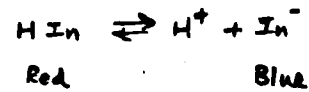
Figures 15.9 and 15.10

**Choosing an indicator for the titration of a strong acid and a weak acid**

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acid-base indicator: which marks the end point of a titration by changing color

Sample Ex: a hypothetical indicator,  $\text{HIn}$   
a weak acid with  $K_a = 1.0 \times 10^{-8}$



$$K_a = \frac{[\text{H}^+][\text{In}^-]}{[\text{HIn}]}$$

$$\frac{K_a}{[\text{H}^+]} = \frac{[\text{In}^-]}{[\text{HIn}]}$$

when  $\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{1}{10}$  the color will start to change

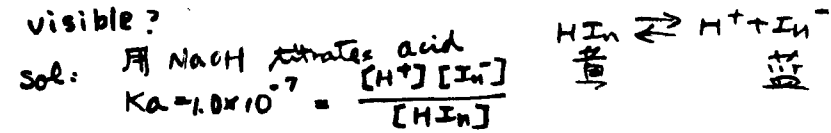
$$\log K_a - \log [\text{H}^+] = \log \frac{[\text{In}^-]}{[\text{HIn}]}$$

$$\begin{aligned} \text{pH} &= \text{p}K_a - 1 \quad \text{[red] to [blue]} \\ &= 8 - 1 = 7 \end{aligned}$$

Ex. 15.11

15.14

Bromthymol blue, an indicator w/ a  $K_a = 1.0 \times 10^{-7}$  is yellow in its  $\text{HIn}$  form, and blue in its  $\text{In}^-$  form. Suppose we put a few drops of this indicator in a strongly acidic soln. If the soln is then titrated w/  $\text{NaOH}$ , at what pH will the indicator color change first be visible?



assume that the color change is visible when

$$\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{1}{10}$$

Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log \left( \frac{[\text{In}^-]}{[\text{HIn}]} \right)$$

$$\text{pH} = \text{p}K_a + \log \left( \frac{1}{10} \right) = \text{p}K_a - 1 = 7 - 1 = 6$$

when a basic soln is titrated, the indicator  $\text{HIn}$  will initially exist as  $\text{In}^-$  form

when  $\frac{[\text{In}^-]}{[\text{HIn}]} = \frac{10}{1}$  color starts to change

$$\begin{aligned} \therefore \text{pH} &= \text{p}K_a + \log 10 = \text{p}K_a + 1 = 7 + 1 = 8 \\ \therefore \text{pH} &= \text{p}K_a \pm 1 \quad \text{or} \quad \text{p}K_a = \text{pH} \pm 1 \end{aligned}$$

indicator end point (where the color changes) and the titration equivalence point to be as close as possible.

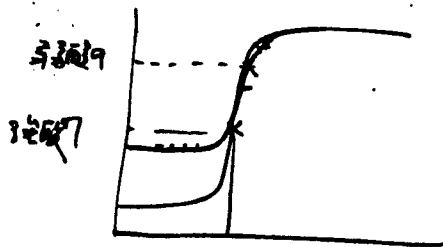
### I Strong acid - strong base titration

Indicator color changes will be sharp, w/ the addition of a single drop  
 Choose indicators w/ end points as far apart as pH 5 and pH 9.

### II. Weak acid is titrated

↓  
 choose pKa very close to the equivalence point of the indicator.

Ex. see Fig 15.9 & 15.10



## 15.6 Solubility Equilibria

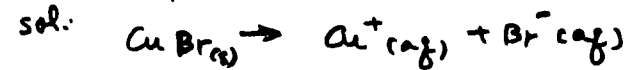


$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2$$

$K_{sp}$ : solubility product constant

or solubility product (unit is normally omitted!)

Ex: 15.12  
 Copper (I) bromide has a measured solubility of  $2.0 \times 10^{-4}$  mol/L at 25°C. Calculate its  $K_{sp}$  value.



$$\begin{aligned} K_{sp} &= [\text{Cu}^+][\text{Br}^-] \\ &= (2.0 \times 10^{-4})(2.0 \times 10^{-4}) \\ &= 4.0 \times 10^{-8} \end{aligned}$$

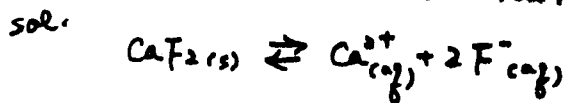
Ex 15.14.

The  $K_{sp}$  value for copper (II) iodate,  $\text{Cu}(\text{IO}_3)_2$  is  $1.4 \times 10^{-7}$  at 25°C. Calculate its solubility!



$$\begin{aligned} K_{sp} &= [\text{Cu}^{2+}][\text{IO}_3^-]^2 \\ &= x \cdot (2x)^2 \\ &= 4x^3 = 1.4 \times 10^{-7} \\ x &= 3.3 \times 10^{-3} \text{ mol/L} \end{aligned}$$

Calculate the solubility of solid  $\text{CaF}_2$  ( $K_{sp} = 4.0 \times 10^{-11}$ ) in a  $0.025 \text{ M NaF}$  solution.



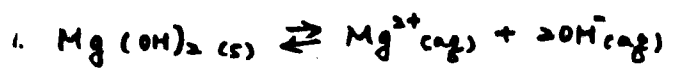
initial		0.025
final	x	0.025+2x

$$K_{sp} = x(0.025+2x)^2 = 4.0 \times 10^{-11}$$

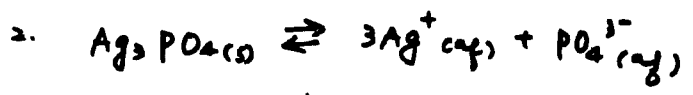
$$\approx x(0.025)^2$$

$$\therefore x \approx 6.4 \times 10^{-8} \text{ mol/L}$$

pH & solubility

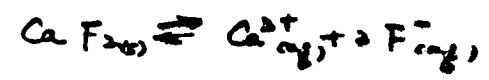


When add  $\text{OH}^{-}$  ion, solubility  $\downarrow$



When add  $\text{H}^{+}$   $+\text{H}^{+}$   
 solubility  $\uparrow$   $\downarrow$   
 $\text{HPO}_4^{2-}$

915. Precipitation and Qualitative Analysis 15-16



$Q = [\text{Ca}^{2+}]_0 [\text{F}^{-}]_0^2$

When  $Q > K_{sp}$  ppt forms

$Q < K_{sp}$  ppt will not form

See Ex 15.16

Ex. I. 100.0 ml of  $0.0500 \text{ M Pb}(\text{NO}_3)_2$

& 200.0 ml of  $0.100 \text{ M NaI}$ .

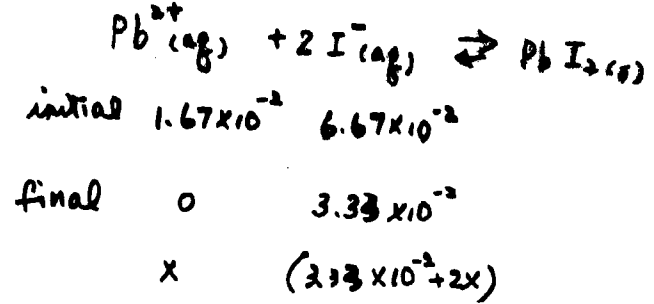
$\text{PbI}_2$  ( $K_{sp} = 1.4 \times 10^{-8}$ ) will it ppt?

$[\text{Pb}^{2+}]_0 = \frac{(100.0 \text{ ml} \cdot 0.0500 \text{ M})}{300.0 \text{ mL}} = 1.67 \times 10^{-2} \text{ M}$

$[\text{I}^{-}]_0 = \frac{(200.0 \text{ ml} \cdot 0.100 \text{ M})}{300.0 \text{ mL}} = 6.67 \times 10^{-2} \text{ M}$

$Q = [\text{Pb}^{2+}]_0 [\text{I}^{-}]_0^2 = (1.67 \times 10^{-2}) (6.67 \times 10^{-2})^2$   
 $= 7.43 \times 10^{-5} > K_{sp} \therefore$  will ppt.

II. After ppt. what are the concentrations equilibrium of  $[\text{I}^{-}]$  &  $[\text{Pb}^{2+}]$ ?



$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$$

$$= x(3.33 \times 10^{-2} + 2x)^2 = 1.4 \times 10^{-8}$$

$$\approx x(3.33 \times 10^{-2})^2 = 1.4 \times 10^{-8}$$

$$x = 1.3 \times 10^{-5}$$

$$[\text{Pb}^{2+}] = 1.3 \times 10^{-5} \text{ M}$$

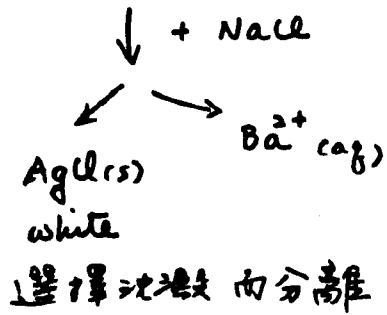
$$[\text{I}^{-}] = (3.33 \times 10^{-2} + 2x)$$

$$= 3.33 \times 10^{-2} \text{ M}$$

15-14

Selective Precipitation  
 Mixtures of metal ions in aqueous solution are often separated by selective precipitation.

Ex.  $\text{Ba}^{2+}(\text{aq})$  and  $\text{Ag}^{+}(\text{aq})$  ions mixture



Ex 15.18  
 A solution contains:  
 $[\text{Cu}^{2+}] = 1.0 \times 10^{-4} \text{ M}$   
 $[\text{Pb}^{2+}] = 2.0 \times 10^{-3} \text{ M}$

$\text{I}^{-}$  is added into the solution, which will ppt first  $\text{CuI}$  or  $\text{PbI}_2$ ?

$$K_{sp} = [\text{Cu}^{2+}][\text{I}^{-}] \quad K_{sp} = 5.3 \times 10^{-12}$$

$$5.3 \times 10^{-12} = 1.0 \times 10^{-4} \text{ M} \cdot [\text{I}^{-}] \quad [\text{I}^{-}] = 5.3 \times 10^{-8} \text{ M}$$

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2 \quad K_{sp} = 1.4 \times 10^{-8}$$

$$= 2.0 \times 10^{-3} [\text{I}^{-}]^2 \quad [\text{I}^{-}] = \sqrt{\frac{1.4 \times 10^{-8}}{2.0 \times 10^{-3}}} = 2.6 \times 10^{-3} \text{ M}$$

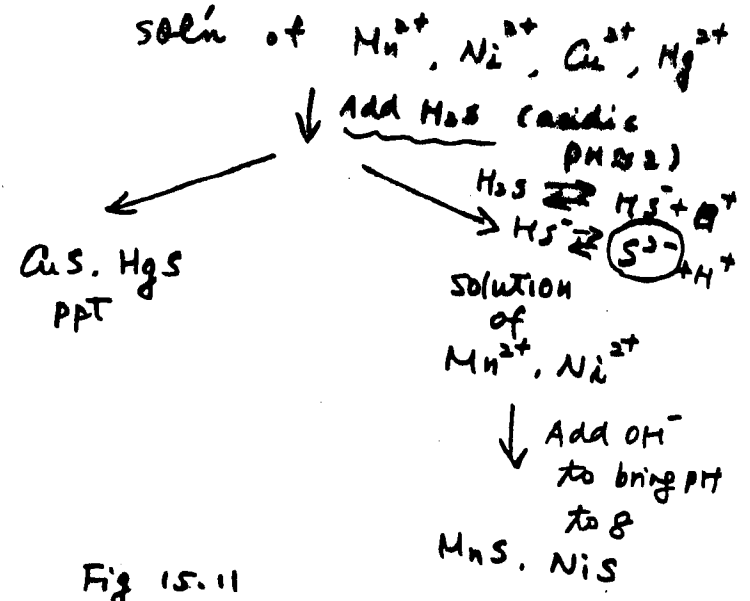


Fig 15.11

Fig 15.12

see page 771-772

### Complex Ion Equilibria

#### § 15.8

\* A complex ion is a charged species consisting of a metal ion surrounded by ligands.

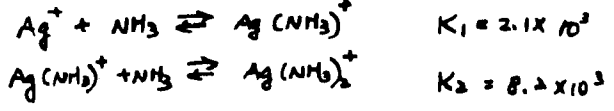
\* A ligand is simply a Lewis base - a molecule or ion having a lone electron pair that can be donated to an empty orbital on the metal ion to form a covalent bond.

Ex.  $H_2O$ ,  $NH_3$ ,  $Cl^-$ , and  $CN^-$

\* The number of ligands attached to a metal ion is called coordination number

Ex. coordination number = 6	$Co(H_2O)_6^{2+}$	$Ni(NH_3)_6^{2+}$
" = 4	$CoCl_4^{2-}$	$Cu(NH_3)_4^{2+}$
" = 2	$Ag(NH_3)_2^+$	

EX. A solution prepared by mixing 100.0ml of 2.0M  $NH_3$  with 100.0ml of  $1.0 \times 10^{-3} M AgNO_3$ . Please calculate the concentrations of  $Ag^+$ ,  $NH_3$ ,  $Ag(NH_3)^+$ , and  $Ag(NH_3)_2^+$



See page 779 Figure 16.13

Solution of  $Ag^+$ ,  $Hg_2^{2+}$ ,  $Pb^{2+}$

↓ Add cold  $HCl(aq)$

$AgCl(s)$ ,  $Hg_2Cl_2(s)$ ,  $PbCl_2(s)$   
ppt

↓ heat

Solution of  $Pb^{2+}$

↓ Add  $CrO_4^{2-}$   
 $PbCrO_4(s)$  ppt  
yellow

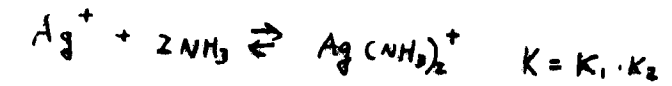
$AgCl(s)$ ,  $Hg_2Cl_2(s)$

↓ Add  $NH_3(aq)$

Solution of  $Ag(NH_3)_2^+$ ,  $Cl^-$

↓ Add  $H^+$   
 $AgCl(s)$   
white

ppt of  $Hg_2(s)$  black  
+  $HgNH_2Cl(s)$   
white



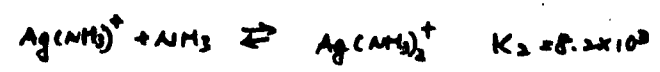
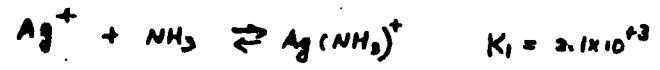
$$\frac{1.0 \times 10^{-4} M \cdot 1.0 M^2}{200.0 mL} = \frac{2.0 \times 10^{-4} M}{200.0 mL}$$

initial  $5.0 \times 10^{-4} M$        $1.0 M$

final  $\sim 0$        $(1.0 - 5.0 \times 10^{-4}) M$        $5.0 \times 10^{-4} M$

$$[NH_3] \approx 1.0 M$$

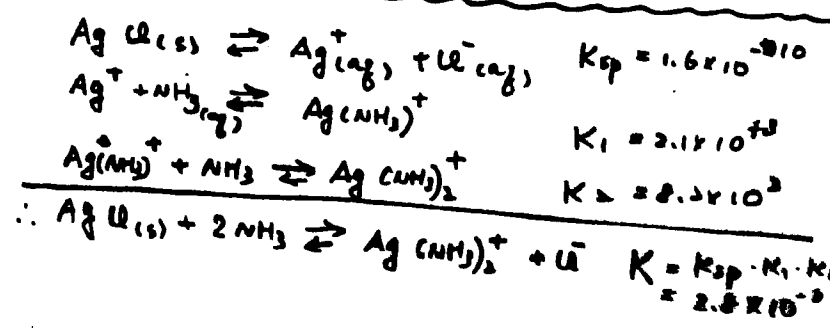
$$[Ag(NH_3)_2^+] \approx 5.0 \times 10^{-4} M$$



$$K_2 = \frac{[Ag(NH_3)_2^+]}{[Ag(NH_3)^+][NH_3]} = \frac{5.0 \times 10^{-4} M}{x \cdot 1 M} = 8.2 \times 10^3$$

$$K_1 = \frac{[Ag(NH_3)^+]}{[Ag^+][NH_3]} = \frac{6.1 \times 10^{-8} M}{y \cdot 1.0 M} = 2.1 \times 10^3$$

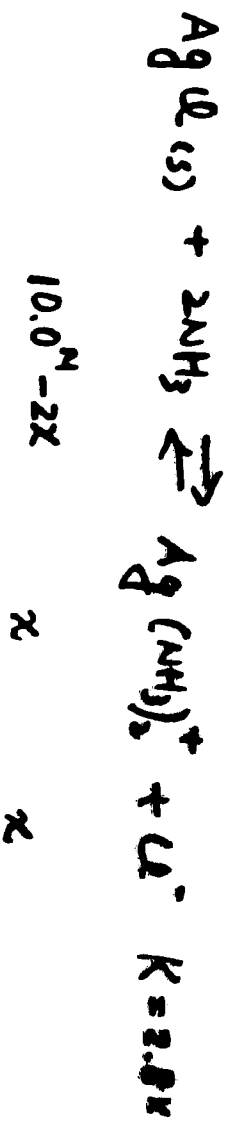
$$y = \frac{6.1 \times 10^{-8} M}{2.1 \times 10^3} = 2.9 \times 10^{-11} M = [Ag^+]$$





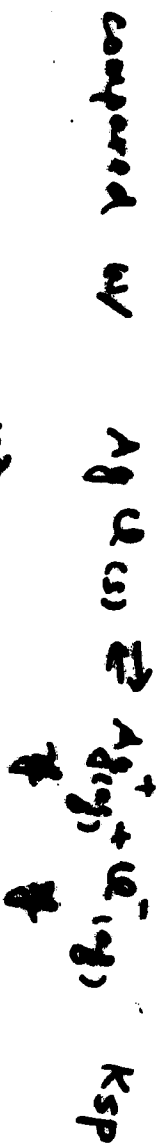
If we add  $10.0\text{ M NH}_3$ , calculate the solubility of  $\text{AgCl}$ :

Soln:



$$K = \frac{x \cdot x}{(10.0 - 2x)^2} = 2.8 \times 10^{-3}$$

$$\frac{x}{10.0 - 2x} = \sqrt{2.8 \times 10^{-3}} \quad \therefore x = 0.48\text{ M}$$



$$x \gg y \quad y^2 = K_{sp} \quad y = \sqrt{K_{sp}} = 1.3 \times 10^{-5}$$

