

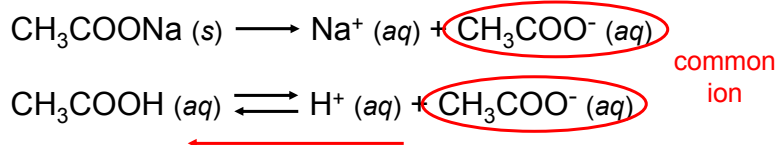
Acid-Base Equilibria and Solubility Equilibria

Chapter 16



The Common Ion Effect

A shift in equilibrium caused by the addition of a compound having an ion in common with the dissolved substance.



The common ion effect can be used to produce a **BUFFER SOLUTION** = a solution of a weak acid or base and its conjugate, e.g. CH_3COOH and CH_3COONa

By controlling the ratio of weak acid/base to its conjugate, we can shift the equilibrium to whatever $[\text{H}^+]$ and therefore pH we want.

Example 16.2 - Which of the following solutions are buffer solutions? (a) $\text{KH}_2\text{PO}_4/\text{H}_3\text{PO}_4$, (b) $\text{NaClO}_4/\text{HClO}_4$.

Buffer Solutions and the Henderson-Hasselbach Equation

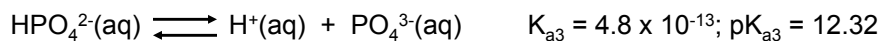
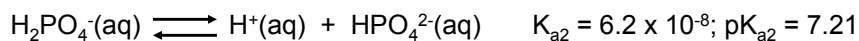
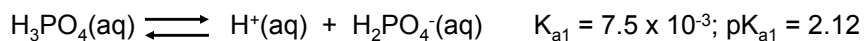
How to calculate the ratio of weak acid/base to its conjugate for a given pH (and vica-versa)

Example 16.1 - (a) calculate the pH of a solution containing 0.20 M CH_3COOH and 0.30 M CH_3COONa . (b) what would the pH of a 0.20 M CH_3COOH solution be if no salt were present?

A Buffer in Action		
	<u>Weak Acid & Salt</u> e.g. $\text{CH}_3\text{COO}^- / \text{CH}_3\text{COOH}$	<u>Weak Base & Salt</u> e.g. $\text{NH}_4^+ / \text{NH}_3$
If add H^+		
If add OH^-		

Preparing a Buffer Solution With a Specific pH

Example 16.4 - Describe how you would prepare a phosphate buffer with a pH of about 7.40.



Titration

In a **titration** a solution of accurately known concentration (titrant) is added gradually to another solution of unknown concentration (analyte) until the chemical reaction between the two solutions is complete.

Equivalence point – the point at which the reaction is complete

Indicator – substance that changes color at (or near) the equivalence point

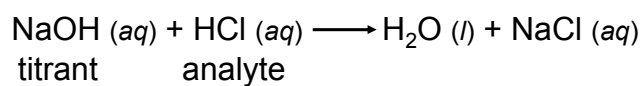


Slowly add base
to unknown acid
UNTIL

The indicator
changes color
(pink)



Strong Acid-Strong Base Titrations



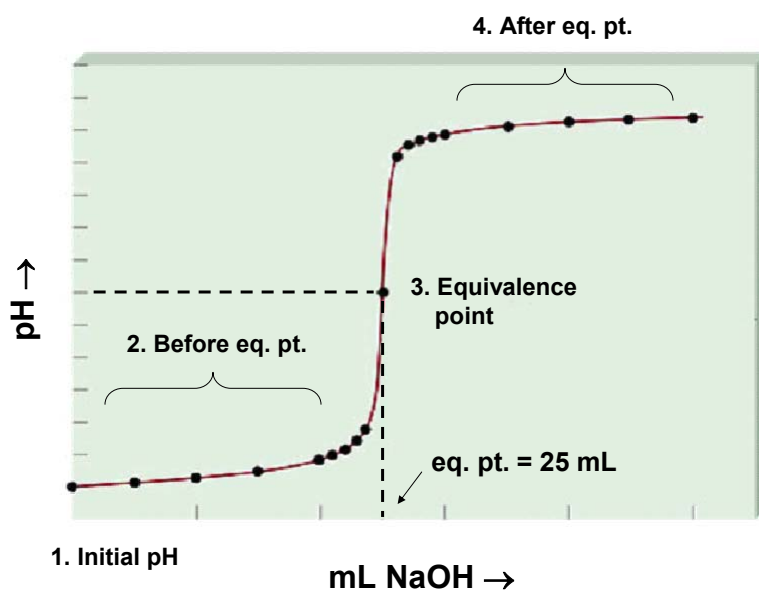
net ionic equation =

0.10 M NaOH added to 25 mL of 0.10 M HCl



How much NaOH has to be added to reach the equivalence point?

Strong Acid-Strong Base Titration Curve



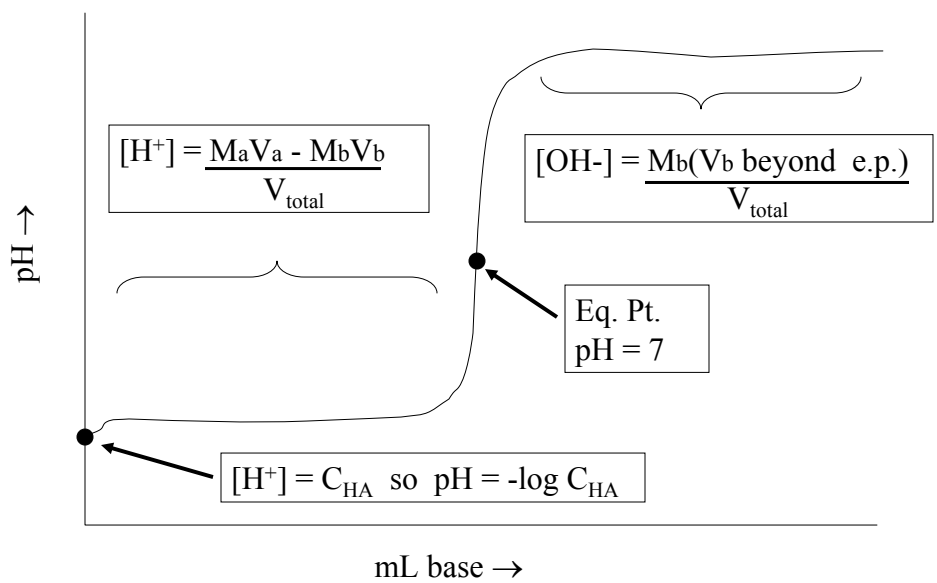
1. Initial pH

2. Before eq. pt.

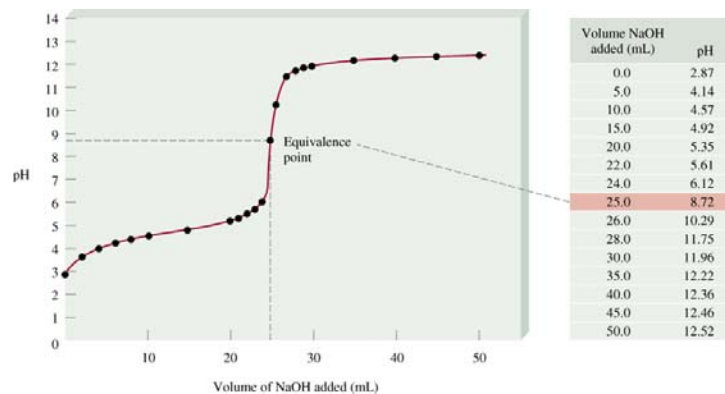
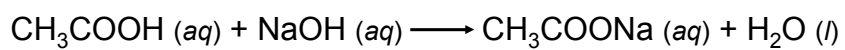
3. Equivalence point

4. After eq. pt.

Strong Acid - Strong Base Titration (both monoprotic)
(analyte) (titrant)



Weak Acid-Strong Base Titration Curve

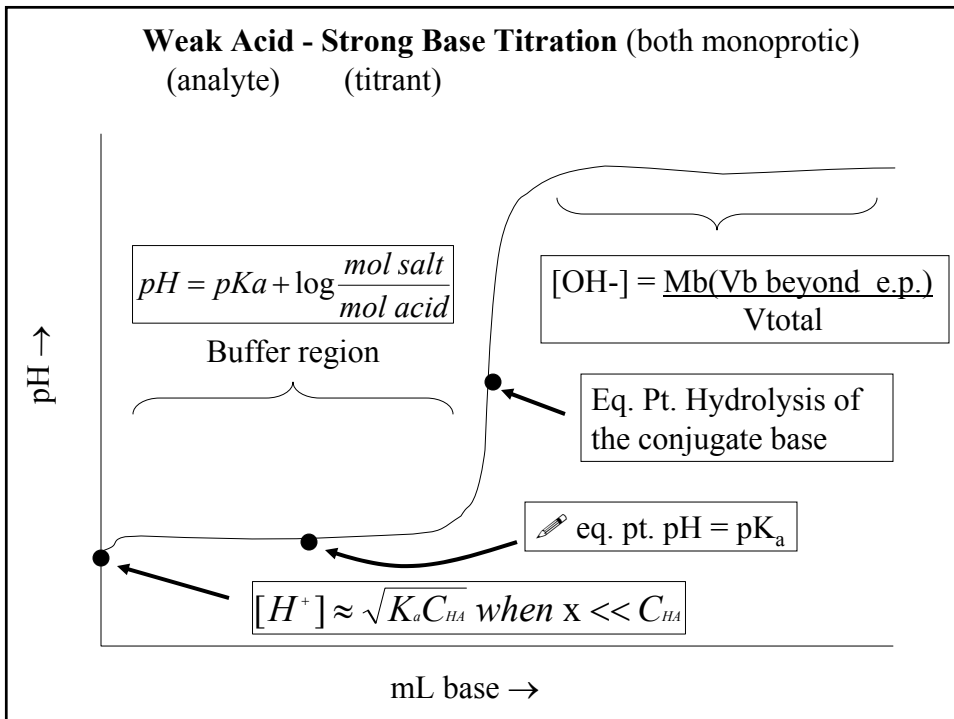


1. Initial pH

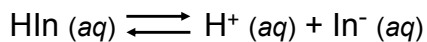
2. Before eq. pt.

3. Equivalence point

4. After eq. pt.



Acid-Base Indicators



$\frac{[\text{HIn}]}{[\text{In}^-]} \geq 10$ Color of acid (HIn) predominates

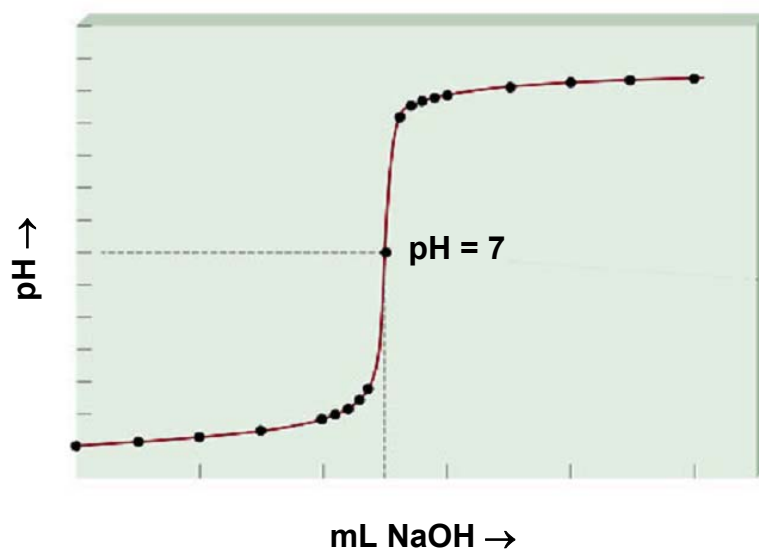
$\frac{[\text{HIn}]}{[\text{In}^-]} \leq 10$ Color of conjugate base (In⁻) predominates

Table 16.1 Some Common Acid-Base Indicators

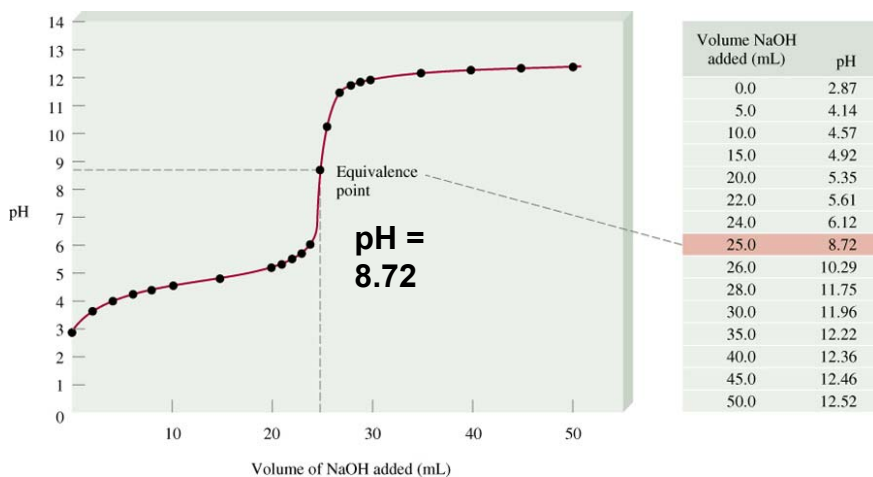
Indicator	C o l o r		pH Range*
	In Acid	In Base	
Thymol blue	Red	Yellow	1.2–2.8
Bromophenol blue	Yellow	Bluish purple	3.0–4.6
Methyl orange	Orange	Yellow	3.1–4.4
Methyl red	Red	Yellow	4.2–6.3
Chlorophenol blue	Yellow	Red	4.8–6.4
Bromothymol blue	Yellow	Blue	6.0–7.6
Cresol red	Yellow	Red	7.2–8.8
Phenolphthalein	Colorless	Reddish pink	8.3–10.0

* The pH range is defined as the range over which the indicator changes from the acid color to the base color.

Choose an Indicator for Our SA-SB Titration Curve



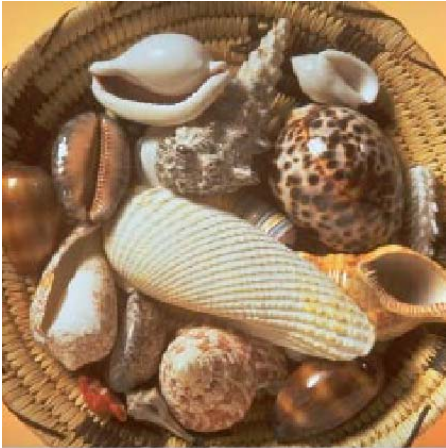
Choose an Indicator for Our WA-SB Titration Curve

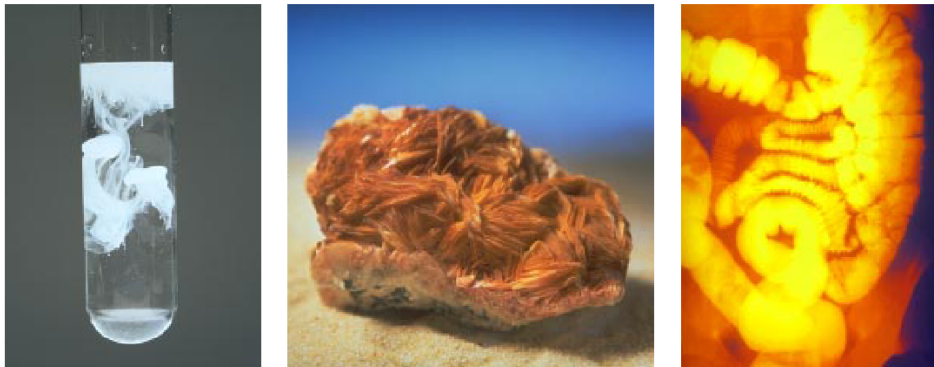


Solubility Equilibria

Calculating the concentration of sparingly soluble salts in solution (compounds that violated the solubility rules do ionize to a very limited extent)







The Solubility Product K_{sp}

Another equilibrium constant that allows calculation of the amount of a compound that will dissolve in water

e.g. $\text{AgCl}(s)$



Note: for now we aren't going to consider hydrolysis of the ions that form, or the more complicated charged species that can be produced sometimes:

e.g.

Calculating K_{sp} from Molar Solubilities (s)

Molar solubility (mol/L) is the number of moles of solute dissolved in 1 L of a saturated solution.

Example 16.8 - the molar solubility of calcium sulfate is found experimentally to be 0.67 g/L. Calculate the value of K_{sp} for calcium sulfate.

Calculating Molar Solubilities (s) from K_{sp}

Example 16.9 - using the data in Table 16.2, calculate the solubility of copper(II) hydroxide in g/L.

Calculating solubility in general -

Comparing Solubilities by Comparing K_{sp} 's

AgCl(s) =	$K_{sp} = 1.8 \times 10^{-10}$
AgBr(s) =	$K_{sp} = 3.3 \times 10^{-13}$
AgI(s) =	$K_{sp} = 1.5 \times 10^{-16}$

BaF ₂ =	$K_{sp} = 1.7 \times 10^{-6}$
CaF ₂ =	$K_{sp} = 3.9 \times 10^{-11}$



Do not compare solubilities for compounds that break up into different numbers of ions:

$$S_{AgCl} =$$

$$S_{CaF_2} =$$

Predicting if a Precipitate will Form

Compare Q_c to K_{sp} :

$Q_c < K_{sp}$ (unsaturated solution)

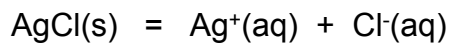
$Q_c = K_{sp}$ (saturated solution)

$Q_c > K_{sp}$ (supersaturated solution)

Example 16.10 - exactly 200 mL of 0.0040 M BaCl_2 are added to exactly 600 mL of 0.0080 M K_2SO_4 . Will a precipitate form?

The Common Ion Effect and Solubility

The presence of a common ion **decreases** the solubility of a salt, e.g. add Ag⁺ (as AgNO₃)



Note: K_{sp} doesn't change, just concentrations

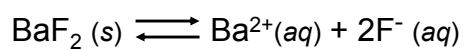
Example 16.12 - Calculate the solubility of silver chloride (in g/L) in a 6.5 x 10⁻³ M silver nitrate solution.

pH and Solubility

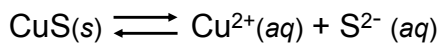
A. The salt dissociates into OH⁻ ions



B. The salt dissociates into a basic ion



Another example of a salt dissociating into a basic ion:



Qualitative Analysis of Cations

