

Lecture Presentation

Chapter 16

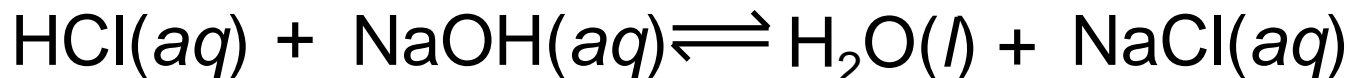
Applications of Aqueous Equilibria

16.1, 16.2, 16.3, 16.7,
16.10, 16.14, 16.16 c,
16.20, 16.22, 16.26

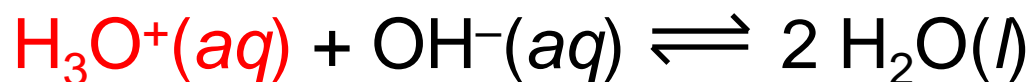
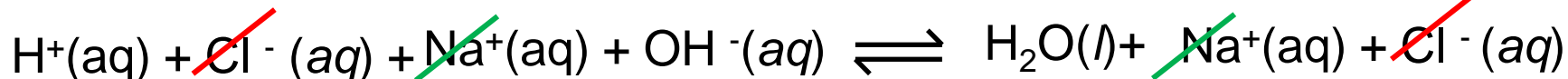
John E. McMurry
Robert C. Fay

Neutralization Reactions

Strong Acid–Strong Base



Assume complete dissociation bc strong acid / strong base
strong acid/base & salts dissociates 100%
weak acids / bases do NOT



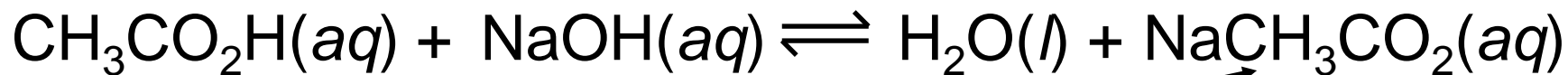
(net ionic equation)

After neutralization: **pH = 7**

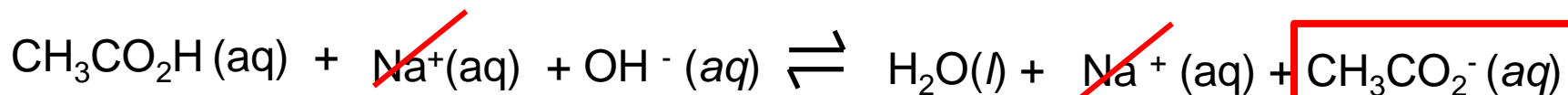
$$K = 1/K_w = 1/(1.0 \times 10^{-14}) = 1.0 \times 10^{14}$$

Neutralization Reactions

Weak Acid–Strong Base

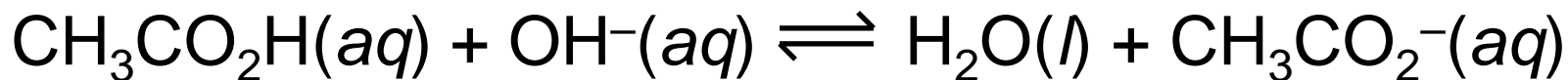


Assume complete dissociation:



strong acid/base & salts dissociates 100%
weak acids / bases do NOT

conjugate base of weak acid
is weak base

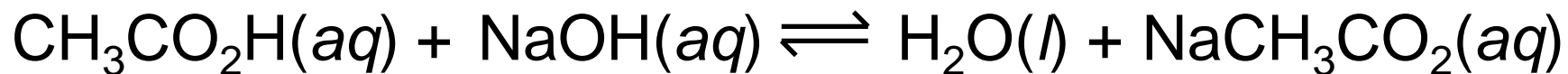


(net ionic equation)

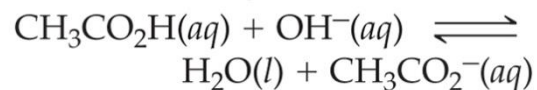
after neutralization: **pH > 7**
(basic)

Neutralization Reactions

Weak Acid–Strong Base



Weak acid–strong base neutralization

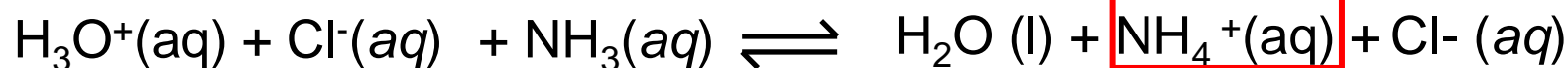


Neutralization Reactions

Strong Acid–Weak Base

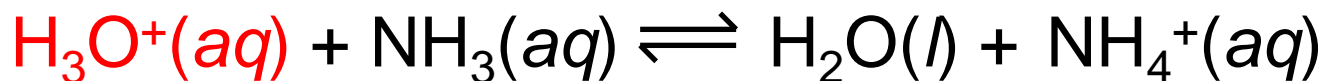


Assume complete dissociation:



strong acid/base & salts dissociates 100%
weak acids / bases do NOT

conjugate acid of weak base
is weak acid

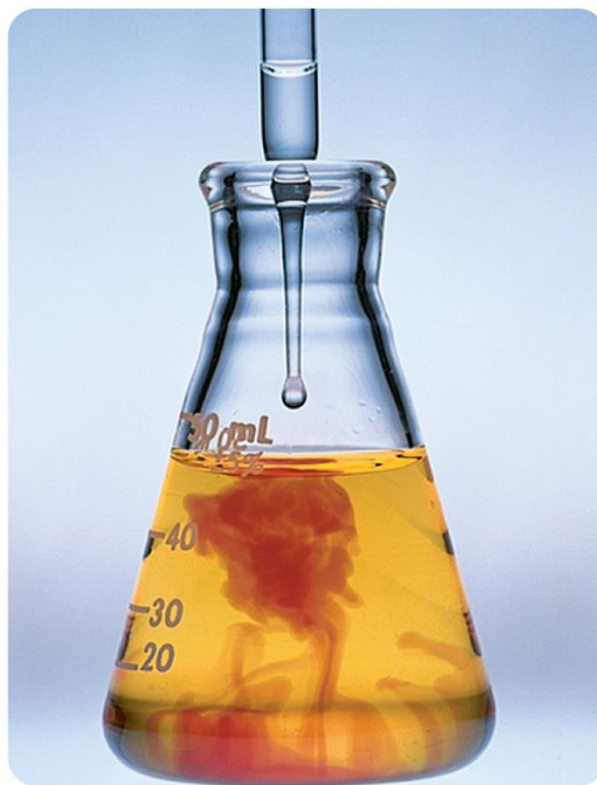


(net ionic equation)

After neutralization: **pH < 7**
(acidic)

Neutralization Reactions

Strong Acid–Weak Base



Strong acid–weak base neutralization



Neutralization Reactions

Weak Acid–Weak Base



assume complete dissociation of salt



(net ionic equation)

After neutralization: **pH = ?**

strong acid/base & salts dissociates 100%
weak acids / bases do NOT

HW 16.1: Neutralization Reactions

(a) Give the neutralization reaction for the following acid / base reaction (HF is weak acid, KOH is strong base)



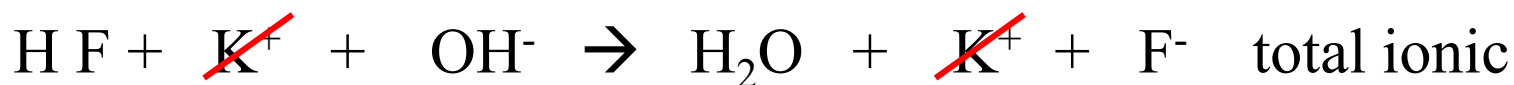
(Can use either 2 or 3 blanks for answer – 2 blanks if you do not show the salt dissociated, 3 blanks if you show the salt dissociated)

(b) Is the product of the neutralization.

[(acidic) (basic) (neutral)] (choose one)

HW 16.1: Neutralization Reactions

(a) Give the neutralization reaction for the following acid / base reaction



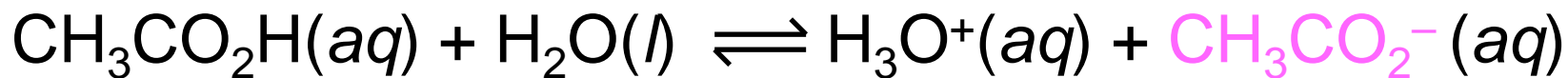
(b) Is the product of the neutralization.

[(acidic) (**basic**) (neutral)] (choose one)

K⁺ is neutral, F⁻ is weak base so weak base

The Common-Ion Effect

Common-Ion Effect: The *shift in the position of an equilibrium* on addition of a substance that provides an ion in **common** with one of the **ions** already involved in the equilibrium



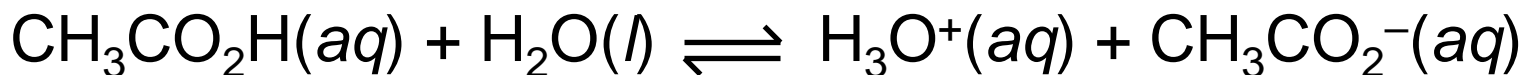
Add Na CH_3CO_2



The Common-Ion Effect

Calculate the pH of a solution that is prepared by dissolved 0.10 mol of acetic acid and 0.10 mol sodium acetate in enough water to make 1.00 L of solution.

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



0.10		≈ 0	0.10
$-x$		$+x$	$+x$
$0.10 - x$		x	$0.10 + x$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CH}_3\text{CO}_2^-]}{[\text{CH}_3\text{CO}_2\text{H}]}$$

The Common-Ion Effect

$$1.8 \times 10^{-5} = \frac{(x)(0.10 + x)}{(0.10 - x)} \approx \frac{x(0.10)}{0.10}$$

$$x \ll 0.10$$

$$x = [\text{H}_3\text{O}^+] = 1.8 \times 10^{-5} \text{ M}$$

$$\text{pH} = -\log([\text{H}_3\text{O}^+]) = -\log(1.8 \times 10^{-5}) = 4.74$$

The Common-Ion Effect

With common ion

$\text{CH}_3\text{CO}_2\text{H}(aq) - \text{CH}_3\text{CO}_2\text{Na}(aq)$
 $[\text{H}_3\text{O}^+] = 1.8 \times 10^{-5} \text{ M}$
pH = 4.74



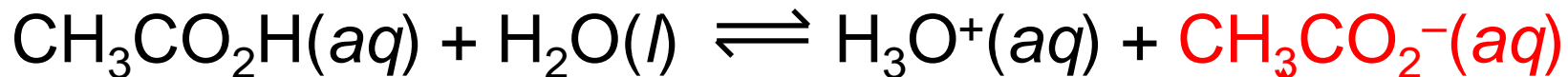
Without common ion

$\text{CH}_3\text{CO}_2\text{H}(aq)$
 $[\text{H}_3\text{O}^+] = 1.3 \times 10^{-3} \text{ M}$
pH = 2.89

The difference in pH is revealed by the color of the indicator methyl orange, which changes from yellow to red in the pH range 4.4–3.2.

The Common-Ion Effect

Le Châtelier's Principle

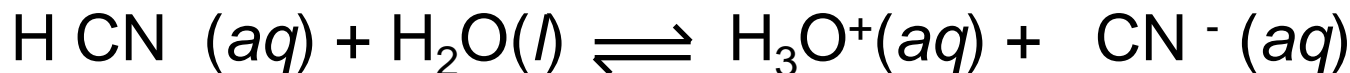


The addition of acetate ion to a solution of acetic acid suppresses the dissociation of the acid. The equilibrium shifts to the left.

$\text{CH}_3\text{CO}_2^-(aq)$
Added common ion

HW 16.2: The Common-Ion Effect

What is the concentration of all species for the reaction of HCN ? $K_a = 4.9 \times 10^{-10}$ The solution has 0.025 M HCN and 0.010 M Na CN

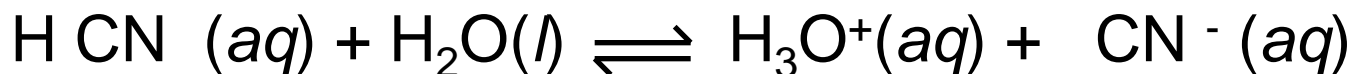


0.025		≈ 0	0.010
-x		+x	+x
0.025 - x		x	0.010 + x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]}$$

HW 16.2: The Common-Ion Effect

What is the concentration of all species for the reaction of HCN ? $K_a = 4.9 \times 10^{-10}$ The solution has 0.025 M HCN and 0.010 M Na CN [on quiz 8 stuff from GC I (a) # moles from grams (b) mL & Molarity – A+ portion of Q8]



0.025		0	0.010
-x		+x	+x
0.025-x (Assume x is small)		x	0.010+x (assume x is small)

$$K_a = \frac{[x][0.010+x]}{[0.025-x]} = \frac{[x][0.010]}{[0.025]}$$

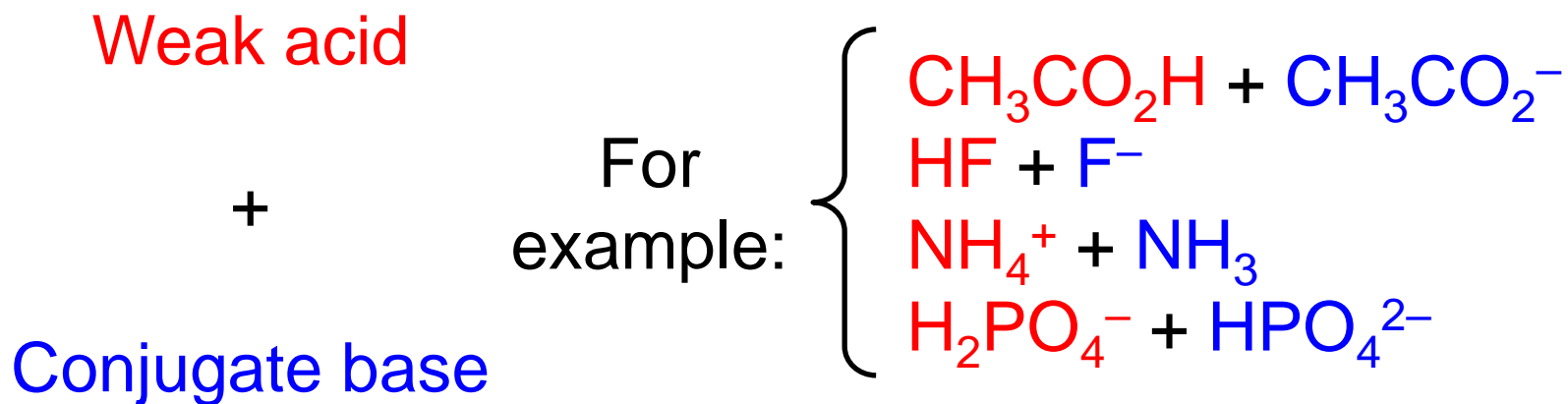
$$x = 1.2 \times 10^{-9} = \text{H}_3\text{O}^+$$

$$\text{HCN} = 0.025$$

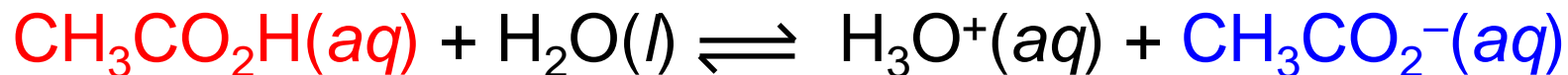
$$\text{CN}^- = 0.010$$

Buffer Solutions

Buffer Solution: A solution that contains a **weak acid** and its **conjugate base** and resists drastic changes in pH



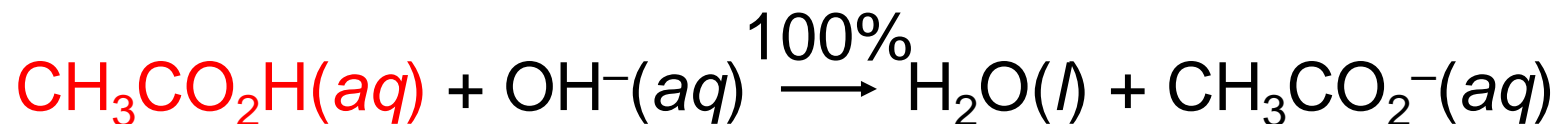
Buffer Solutions



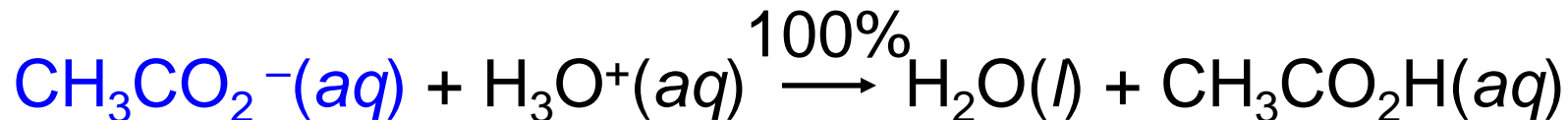
Weak acid

Conjugate base
(NaCH_3CO_2)

Addition of OH^{1-} to a buffer: **change strong base to weak base**



Addition of H_3O^{1+} to a buffer: **change strong acid to weak acid**



Buffer Solutions

+ HCl to NONBUFFER

(a) 1.00 L of 1.8×10^{-5} M HCl
(pH = 4.74)



(b) The solution from **(a)**
turns yellow (pH > 5.4)
after addition of only a few
drops of 0.10 M NaOH.



+ HCl to BUFFER

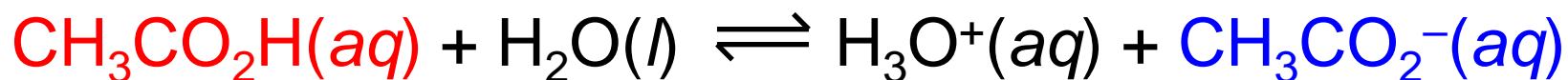
(c) 1.00 L of a 0.10 M acetic
acid–0.10 M sodium
acetate buffer solution
(pH = 4.74)



(d) The solution from **(c)** is
still red (pH < 5.4) after
addition of 100 mL of
0.10 M NaOH.

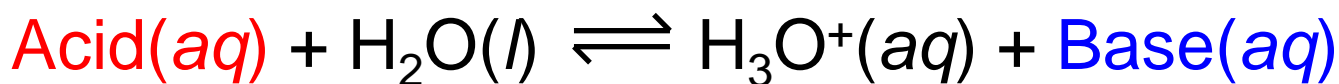


The Henderson–Hasselbalch Equation (Derivation of Henderson-Hasselbalch)



Weak acid

Conjugate base



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{Base}]}{[\text{Acid}]}$$

$$[\text{H}_3\text{O}^+] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$$

The Henderson–Hasselbalch Equation

$$[\text{H}_3\text{O}^+] = K_a \frac{[\text{Acid}]}{[\text{Base}]}$$

$$-\log([\text{H}_3\text{O}^+]) = -\log(K_a) - \log \frac{[\text{Acid}]}{[\text{Base}]}$$

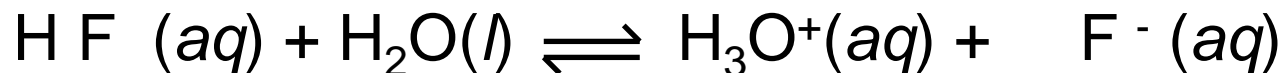
$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{Base}]}{[\text{Acid}]} \right)$$

Henderson-Hasselbalch
Equation

End 4/6 Monday
A sect, 4/7
Tuesday C sect

HW 16.3: Buffer Solutions

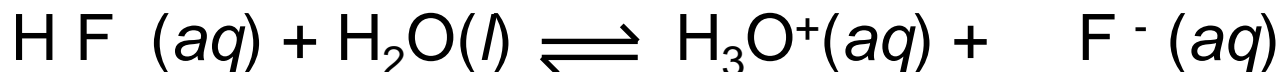
What is the pH of a buffer solution with 0.25 M HF and 0.50 M NaF. $K_a = 3.5 \times 10^{-4}$. (using ICE table or Henderson Hasselbalch)



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]}$$

HW 16.3: Buffer Solutions

What is the pH of a buffer solution with 0.25 M HF and 0.50 M NaF. $K_a = 3.5 \times 10^{-4}$. (using ICE table or Henderson Hasselbalch)
[on quiz 8 stuff from GC I (a) # moles from grams (b) mL & Molarity – A+ portion of Q8]



0.25		≈ 0	0.50
-x		+x	+x
0.250 - x		x	0.50 + x

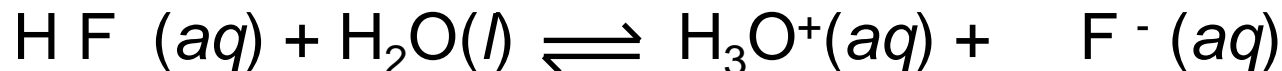
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(0.50+x)}{0.250-x} \sim \frac{(x)(0.50)}{0.250} = 3.5 \times 10^{-4}$$

$$x = \frac{(3.5 \times 10^{-4})(0.250)}{(0.50)} = 1.75 \times 10^{-4}$$

$$\text{pH} = -\log(1.75 \times 10^{-4}) = 3.757$$

HW 16.3: Buffer Solutions

What is the pH of a buffer solution with **0.25 M HF** and **0.50 M NaF**. $K_a = 3.5 \times 10^{-4}$. What is the pH after the addition of 0.002 mol HNO_3 ?



Using Henderson – Hasselbalch:

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log (3.5 \times 10^{-4}) + \log \frac{0.50}{0.25}$$

$$\text{pH} = 3.456 + 0.30 = 3.757$$

HW 16.4: Buffer Solutions

Buffer Solution: A solution that contains a **weak acid** and its **conjugate base** and resists drastic changes in pH

Given the following reagents, fill the blank with the other molecule needed to make a buffer solution and label each as (A) for acid or (B) for base.

1. $\text{CH}_3\text{CO}_2\text{H}$ () _____ ()

2. CN^- () _____ ()

3. NH_3 () _____ ()

HW 16.4: Buffer Solutions

Buffer Solution: A solution that contains a **weak acid** and its **conjugate base** and resists drastic changes in pH

Given the following reagents, fill the blank with the other molecule needed to make a buffer solution and label each as (A) for acid or (B) for base.



pH Titration Curves

Titration: adding measured volume (known concentration) to react with a known volume of the solution of another substance (the standard solution) with unknown concentration (adding acid to base or adding base to acid)

Equivalence Point: The point with stoichiometric equivalence of the two reagents. (moles H^+ = moles OH^-)

pH Titration Curves

(strong base vs. strong acid)

(a) A pH titration in which 0.100 M NaOH is added slowly from a buret to an HCl solution of unknown concentration. The pH of the solution is measured with a pH meter and is recorded as a function of the volume of NaOH added.

buret - base

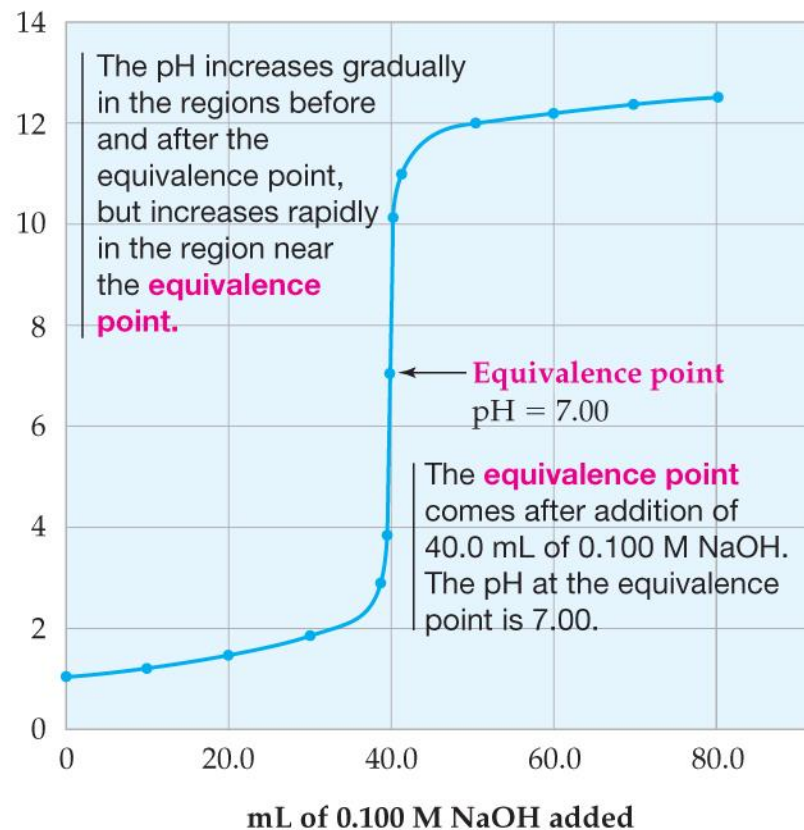


Acid of unknown concentration

(b) The pH titration curve for titration of 40.0 mL of 0.100 M HCl with 0.100 M NaOH.

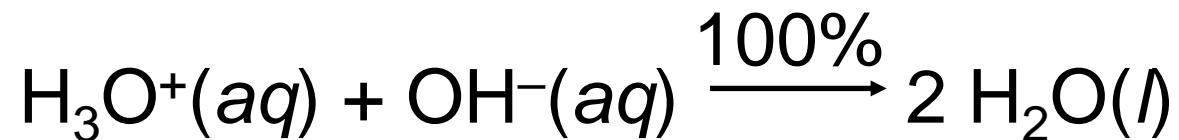
fast pH change near equivalence pt.

pH



volume base added

Strong Acid–Strong Base Titrations



Strong Acid–Strong Base Titrations

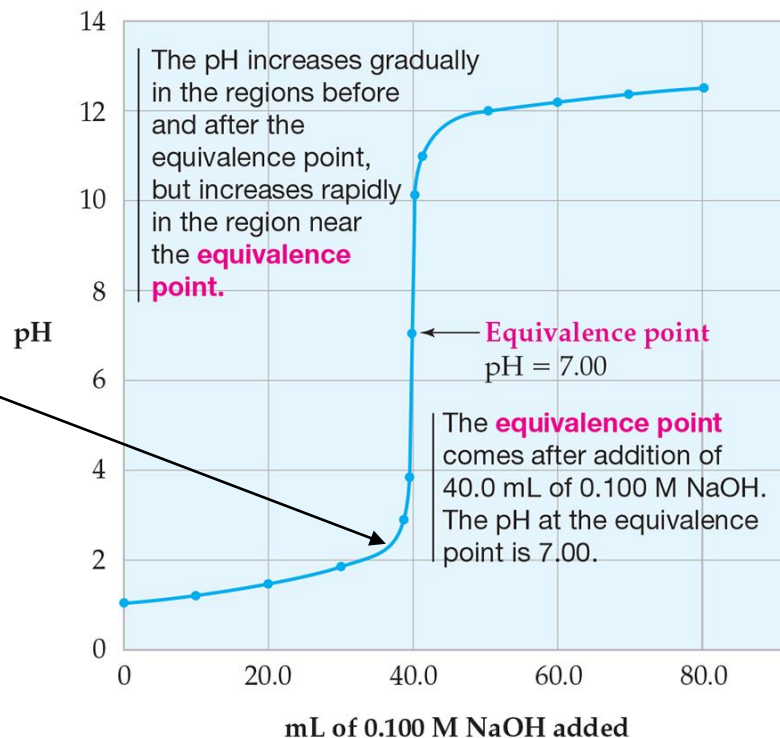
1. Before Addition of Any NaOH

0.100 M HCl

0.100 M HCl x 40.0 mL with
0.100 M NaOH

Equivalence point is 40.0 mL of
added NaOH

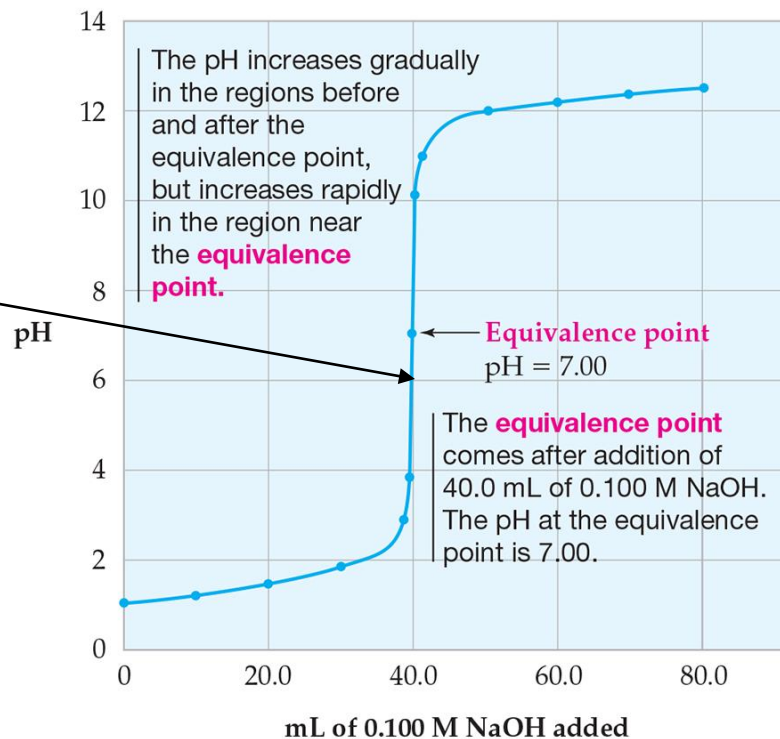
pH=7.00 for strong acid/strong
base titration



Strong Acid–Strong Base Titrations

2. Before the Equivalence Point

Excess H_3O^+

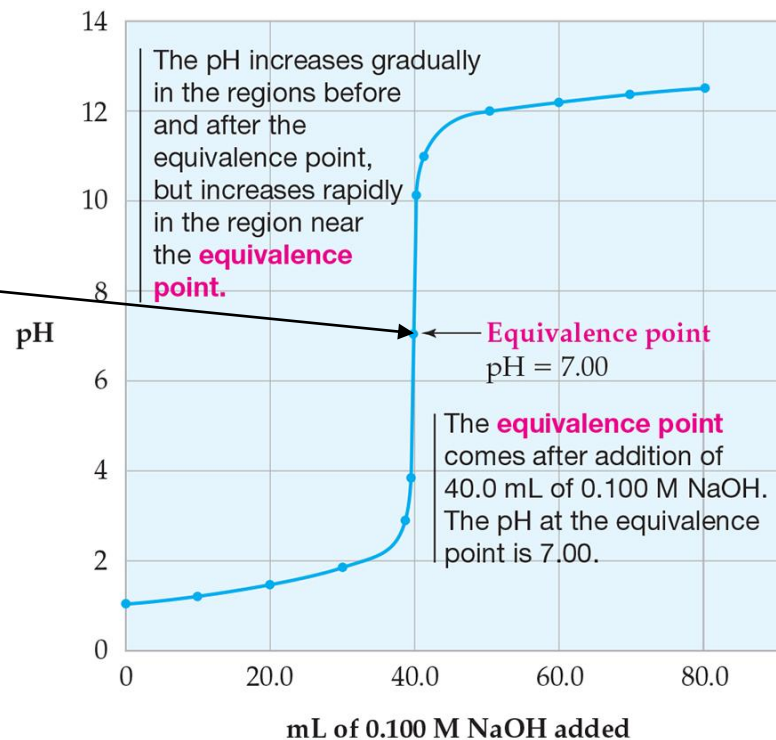


Strong Acid–Strong Base Titrations

3. At the Equivalence Point

The quantity of H_3O^+ is equal to the quantity of added OH^- .

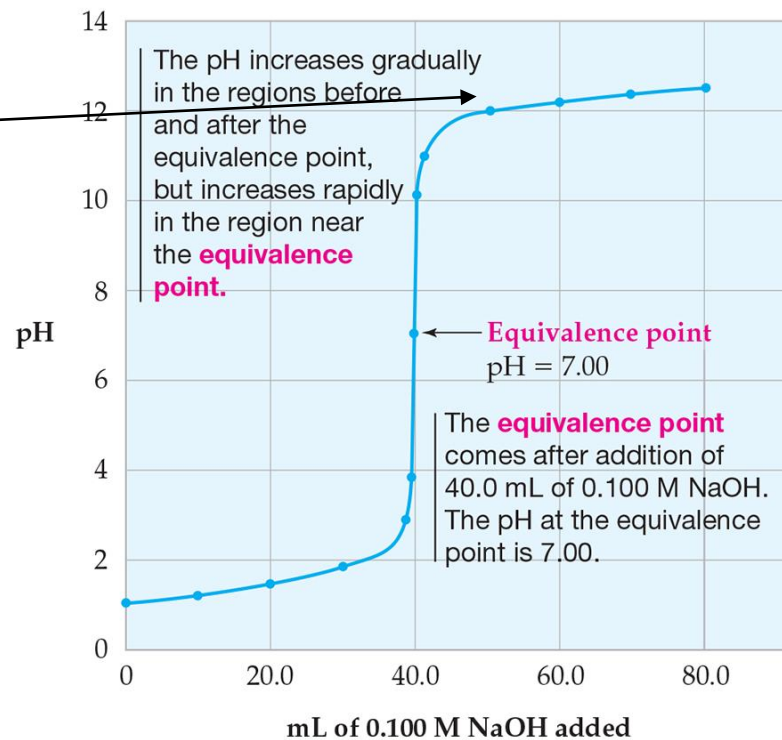
$\text{pH} = 7$



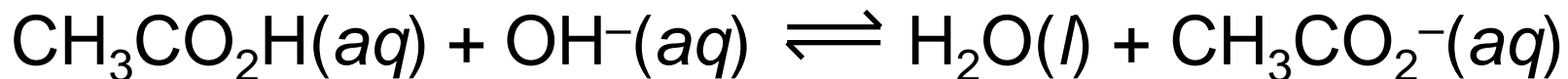
Strong Acid–Strong Base Titrations

4. Beyond the Equivalence Point

Excess OH^-



Weak Acid–Strong Base Titrations

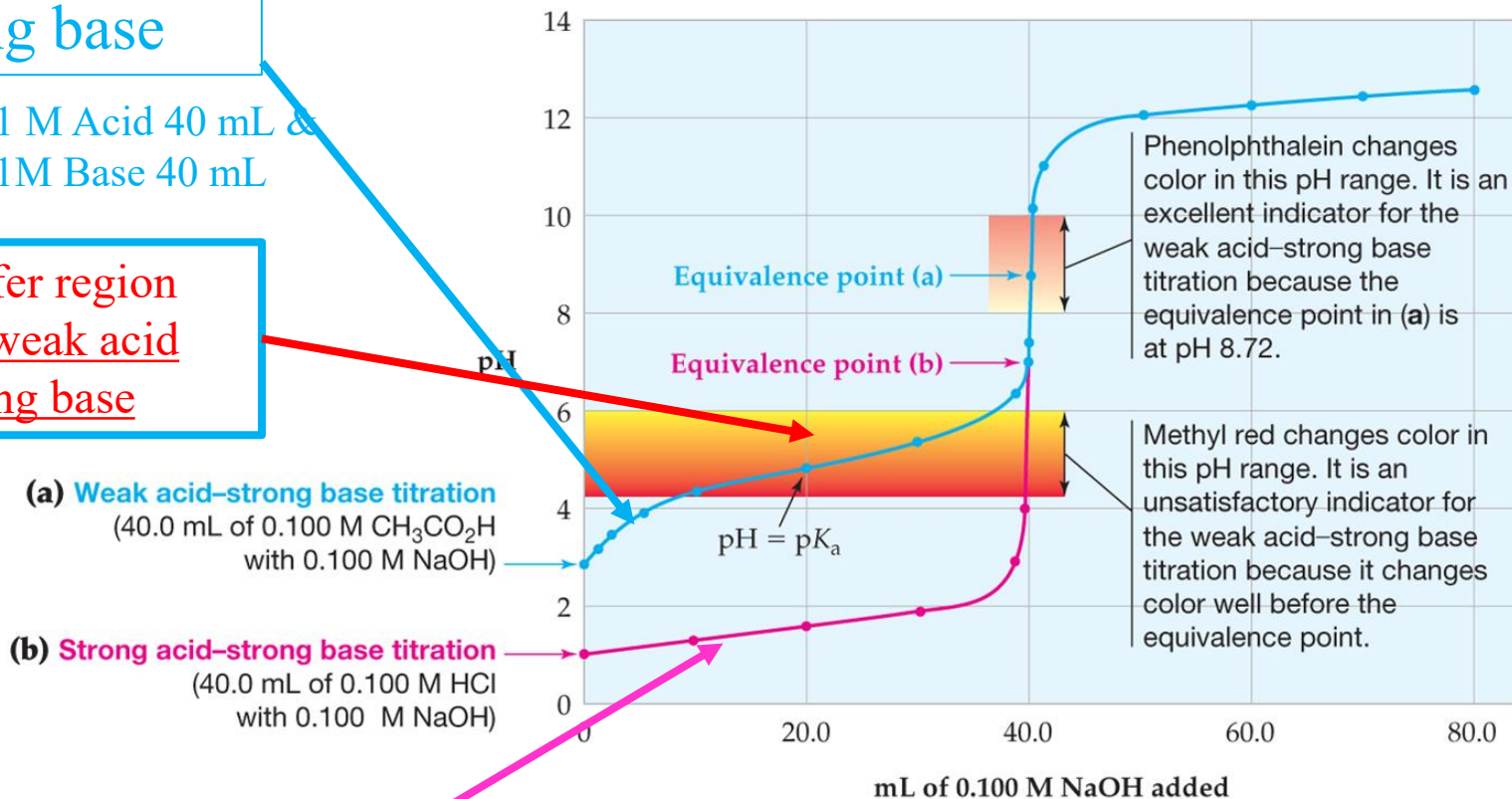


pH = basic at equivalence pt for weak acid/strong base titration

weak acid vs
strong base

0.1 M Acid 40 mL &
0.1 M Base 40 mL

Buffer region
for weak acid
strong base



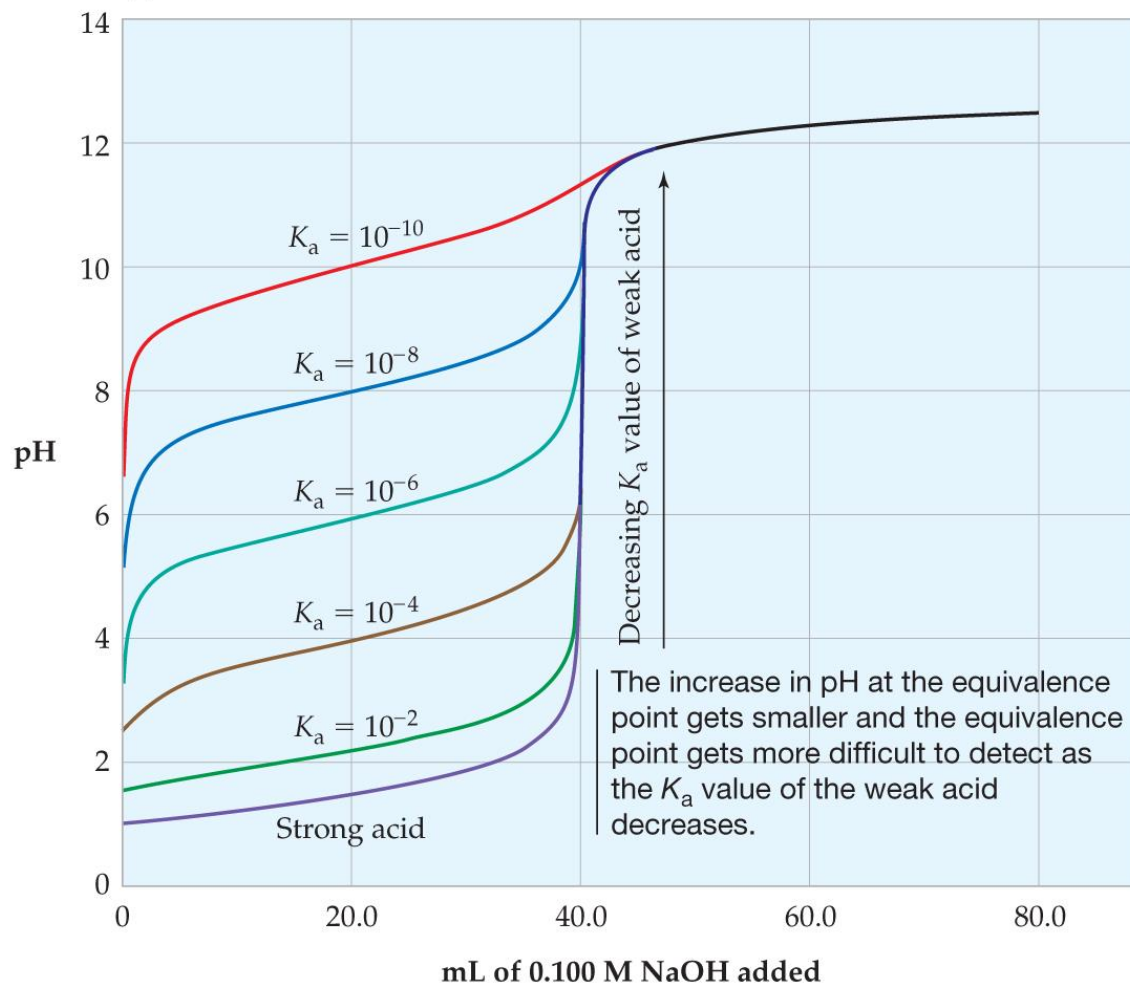
strong acid vs
strong base

0.1 M acid 40 mL &
0.1 M base 40 mL

Either phenolphthalein or methyl red can be used for the strong acid–strong base titration because the curve rises very steeply in the region of the equivalence point in (b) at pH 7.00.

Weak Acid–Strong Base Titrations

The curves shown are for titration of 40.0 mL of 0.100 M solutions of various weak acids with 0.100 M NaOH. In each case, the equivalence point comes after addition of 40.0 mL of the base.

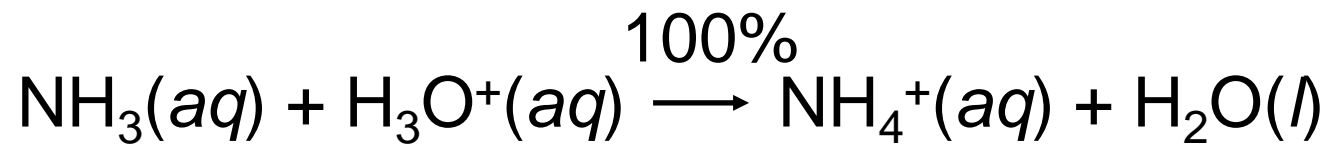


Bottom curve is strong acid vs strong base
0.1 M 40 mL 0.1 M 40 mL

Rest of plots are for various other weak acids of 40 mL of 0.1 M. All other acids have equivalence point at 40 mL NaOH.

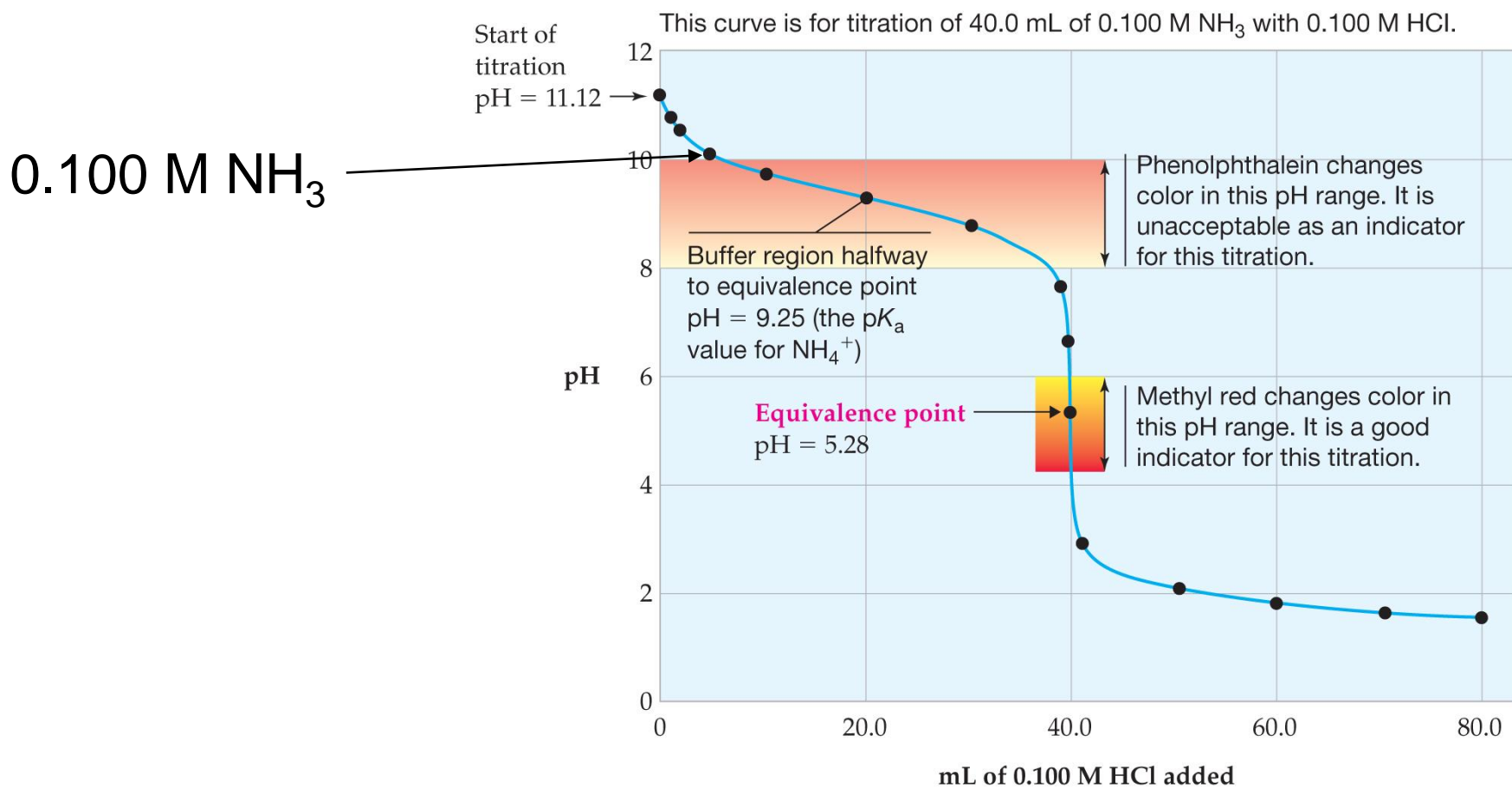
Weak acid curve is less sharp and well defined.

Weak Base–Strong Acid Titrations



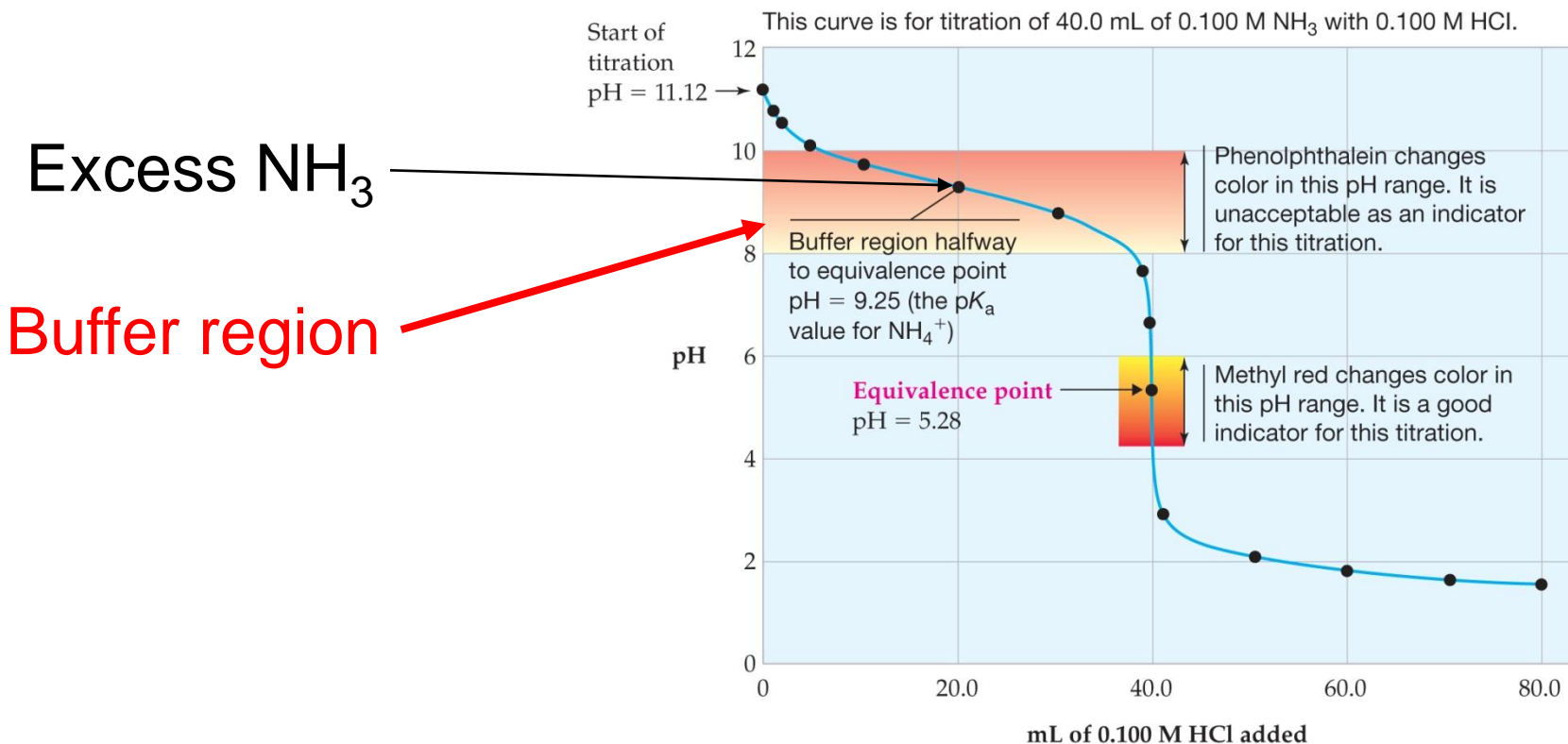
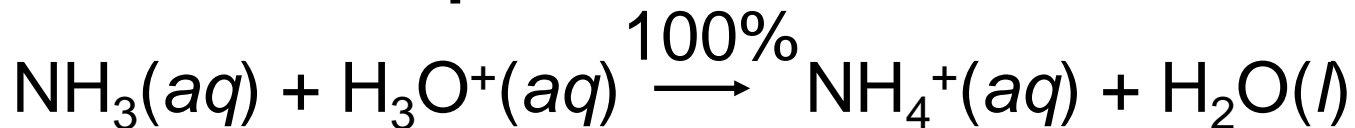
Weak Base–Strong Acid Titrations

1. Before Addition of Any HCl



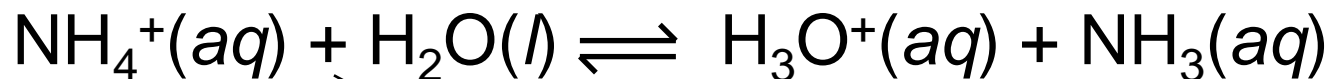
Weak Base–Strong Acid Titrations

2. Before the Equivalence Point

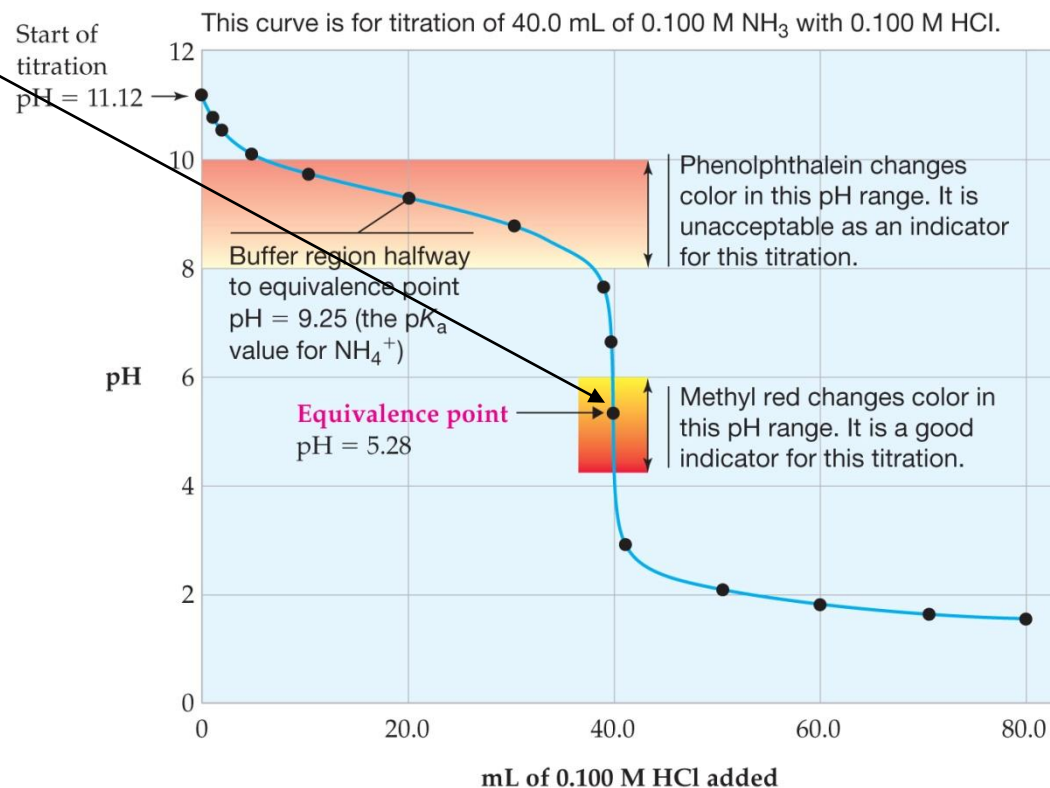


Weak Base–Strong Acid Titrations

3. At the Equivalence Point



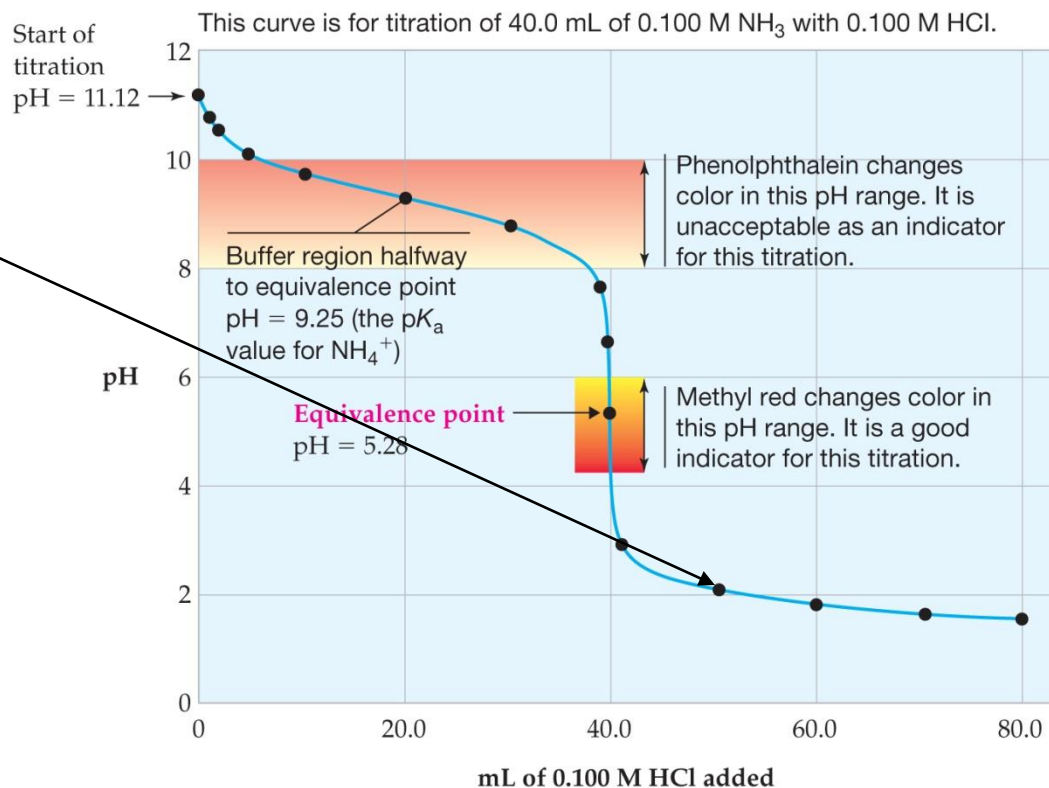
pH < 7



Weak Base–Strong Acid Titrations

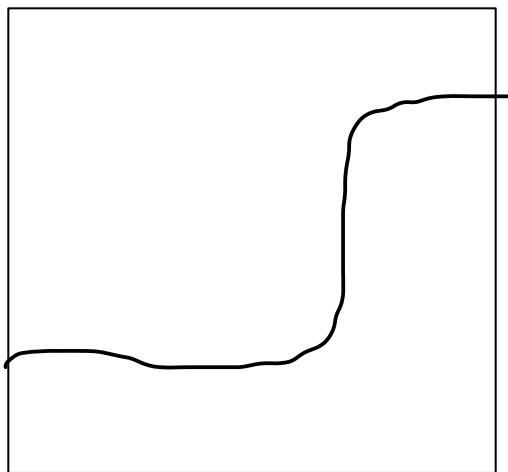
4. Beyond the Equivalence Point

Excess H_3O^+

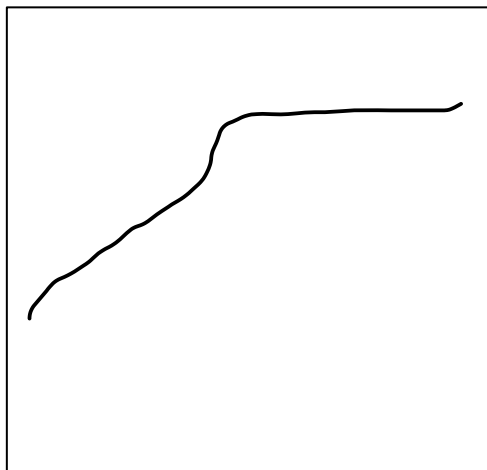


HW 16.4: pH Titration Curves

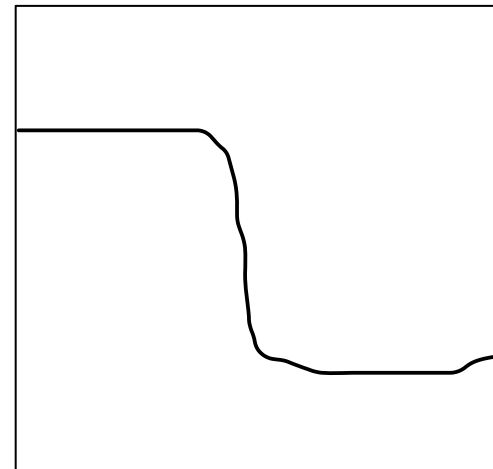
Label the following titration curve as (A) strong acid added to strong base (B) strong base added to strong acid (C) strong base added to weak acid (choose one letter under each #) (only weak acid and weak base being titrated (stuff in beaker) has buffer, strong acid/strong base being titrated has NO BUFFER)



1.(A)(B)(C)



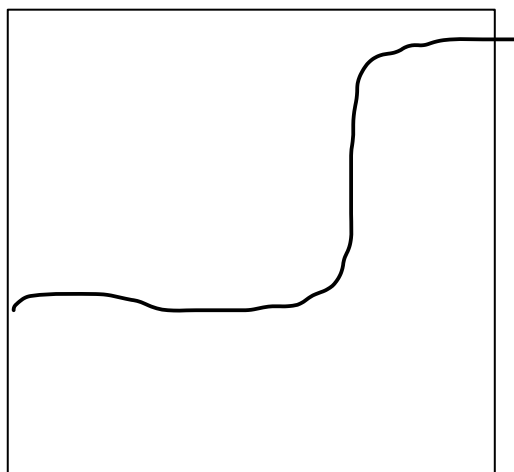
2.(A)(B)(C)



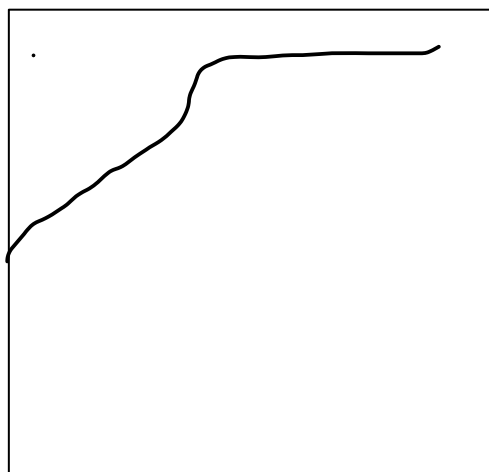
3.(A)(B)(C)

HW 16.4: pH Titration Curves

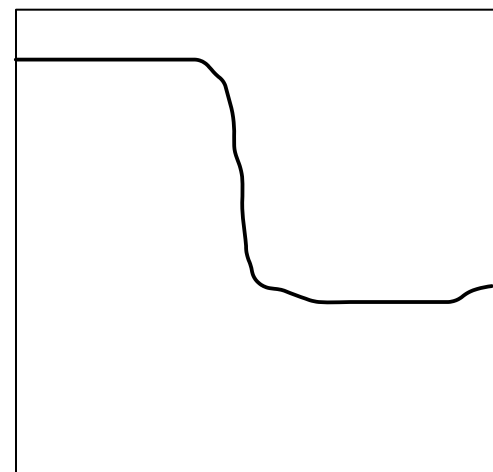
Label the following titration curve as (A) strong acid added to strong base (B) strong base added to strong acid (C) strong base added to weak acid (choose one letter under each #)



1.(A)(B)(C)



2.(A)(B)(C)



3.(A)(B)(C)

HW 16.5: pH Titration Curves

If you do a titration of 100.0 mL of 0.50 M HCl with 0.25 M NaOH, what is the volume of base at the equivalence point?

$$(M_{\text{acid}} V_{\text{acid}} = M_{\text{base}} V_{\text{base}})$$

End 4/8/20 Wed
A & C section

HW 16.5: pH Titration Curves

If you do a titration of 100.0 mL of 0.50 M HCl with 0.25 M NaOH, what is the volume of base at the equivalence point?

$$(M_{\text{acid}} V_{\text{acid}} = M_{\text{base}} V_{\text{base}})$$

$$V_{\text{acid}} = 100.0 \text{ mL HCl}$$

$$M_{\text{acid}} = 0.50 \text{ M HCl}$$

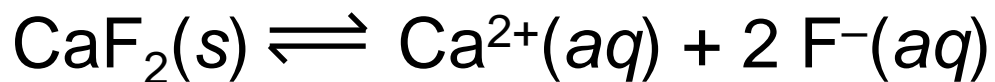
$$M_{\text{base}} = 0.25 \text{ M}$$

$$(0.50 \text{ M HCl})(100.0 \text{ mL HCl}) = (0.25 \text{ M NaOH})(x \text{ mL NaOH})$$

$$x \text{ mL NaOH} = 200.0 \text{ mL NaOH}$$

End Quiz 8 & Test 4

Solubility Equilibria – quantitative look at precipitation (we did qualitative precipitation in General Chemistry I Lecture)



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

Equilibrium constant in
heterogeneous solution.



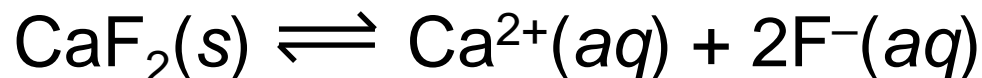
Solubility Equilibria



$$K_{sp} = [M^{n+}]^m [X^{y-}]^x$$

Measuring K_{sp} and Calculating Solubility from K_{sp}

If the concentrations of $\text{Ca}^{2+}(\text{aq})$ and $\text{F}^{-}(\text{aq})$ in a saturated solution of calcium fluoride are known, K_{sp} may be calculated.



$$[\text{Ca}^{2+}] = 2.0 \times 10^{-4} \text{ M} \quad [\text{F}^{-}] = 4.1 \times 10^{-4} \text{ M}$$

$$K_{sp} = [\text{Ca}^{2+}][\text{F}^{-}]^2 = (2.0 \times 10^{-4})(4.1 \times 10^{-4})^2 = \boxed{3.4 \times 10^{-11}}$$

(at 25 °C)

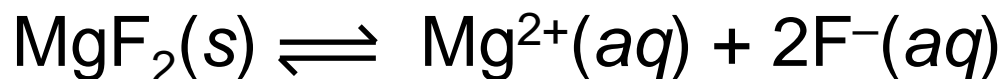
Measuring K_{sp} and Calculating Solubility from K_{sp}

TABLE 16.3 K_{sp} Values for Some Ionic Compounds at 25 °C

Name	Formula	K_{sp}
Aluminum hydroxide	$\text{Al}(\text{OH})_3$	1.9×10^{-33}
Barium carbonate	BaCO_3	2.6×10^{-9}
Calcium carbonate	CaCO_3	5.0×10^{-9}
Calcium fluoride	CaF_2	3.5×10^{-11}
Lead(II) chloride	PbCl_2	1.2×10^{-5}
Lead(II) chromate	PbCrO_4	2.8×10^{-13}
Silver chloride	AgCl	1.8×10^{-10}
Silver sulfate	Ag_2SO_4	1.2×10^{-5}

Measuring K_{sp} and Calculating Solubility from K_{sp}

Calculate the molar solubility of MgF_2 in water at 25 °C.



	x	$2x$
--	-----	------

$$K_{sp} = 7.4 \times 10^{-11} = [Mg^{2+}][F^-]^2 = (x)(2x)^2$$

$$4x^3 = 7.4 \times 10^{-11}$$

$$x = [Mg^{2+}] = \text{Molar solubility} = \boxed{2.6 \times 10^{-4} \text{ M}}$$

HW 16.6: Measuring K_{sp} and Calculating Solubility from K_{sp}

Calculate the molar solubility of Ag Cl in water at 25 °C.



--	--	--

$$K_{sp} = 1.8 \times 10^{-10} = [\text{Ag}^+][\text{Cl}^-]$$

HW 16.6: Measuring K_{sp} and Calculating Solubility from K_{sp}

Calculate the molar solubility of Ag Cl in water at 25 °C.



	x	x
--	-----	-----

$$K_{sp} = 1.8 \times 10^{-10} = [\text{Ag}^+][\text{Cl}^-] = (x)(x)$$

$$x^2 = 1.8 \times 10^{-10}$$

$$x = [\text{Ag}^+] = \text{Molar solubility} = \boxed{1.3 \times 10^{-5} \text{ M}}$$

Factors That Affect Solubility (Qualitative for Final Exam)

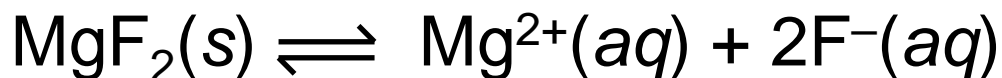
End 3/14 T
C section

1. common ion effect – Le Chatelier
2. Effect of pH – common ion effect with H_3O^+
3. Effect of complex ions – $\text{Ag}(\text{NH}_3)_2$ soluble (AgCl with added ammonia is soluble)
4. Amphoterism - Al^{+3} soluble in acid and base

Factors That Affect Solubility

Solubility and the Common-Ion Effect

Calculate the molar solubility of MgF_2 in 0.10 M NaF at 25 °C.



	x	$0.10 + 2x$
--	-----	-------------

$$K_{\text{sp}} = 7.4 \times 10^{-11} = [\text{Mg}^{2+}][\text{F}^{-}]^2 = (x)(0.10 + 2x)^2$$

$$7.4 \times 10^{-11} = (x)(0.10 + 2x)^2 \approx (x)(0.10)^2$$

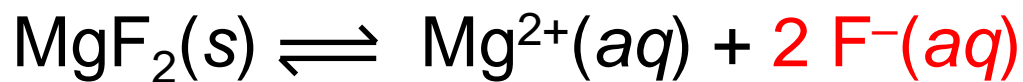
$$x = [\text{Mg}^{2+}] = \text{Molar solubility} = \frac{7.4 \times 10^{-11}}{(0.10)^2} = \boxed{7.4 \times 10^{-9} \text{ M}}$$

$x \ll$ than 0.10 (bc K_{sp} is very small)

so $0.10 + 2x \sim 0.10$

Factors That Affect Solubility

Solubility and the **Common-Ion** Effect



Molar solubility: $2.6 \times 10^{-4} \text{ M}$
(without common ion)

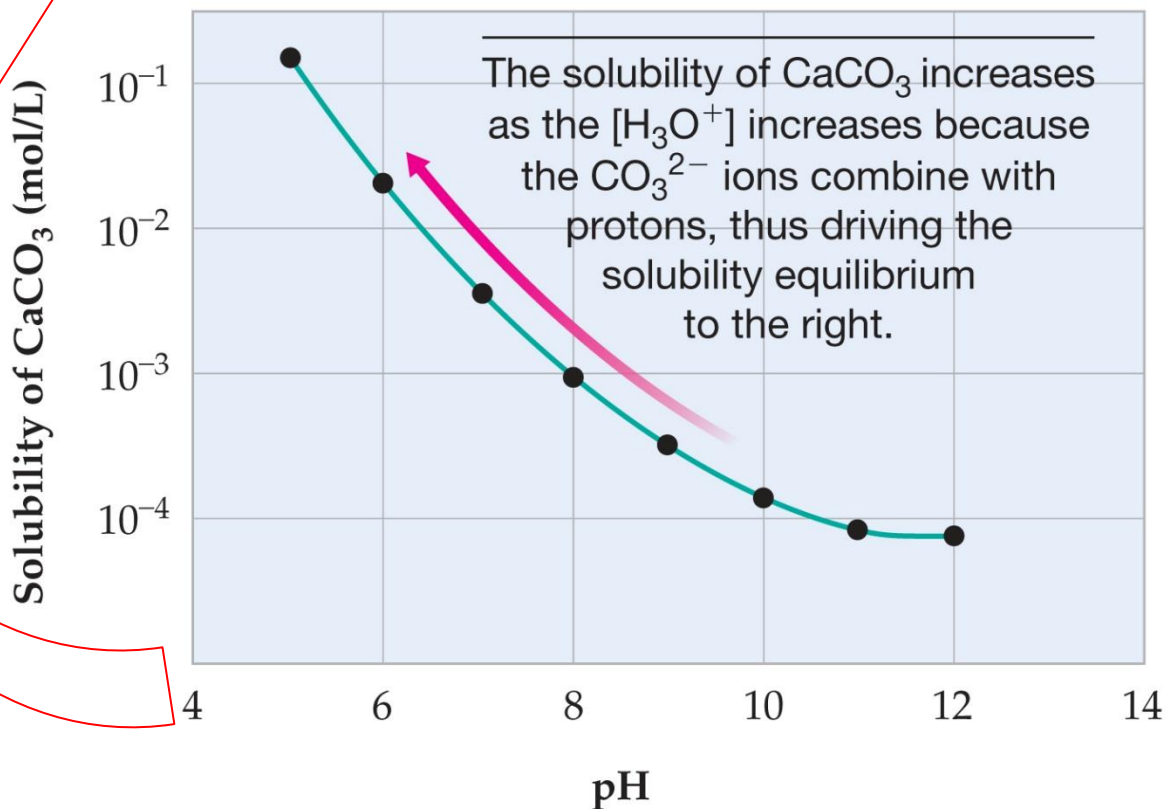
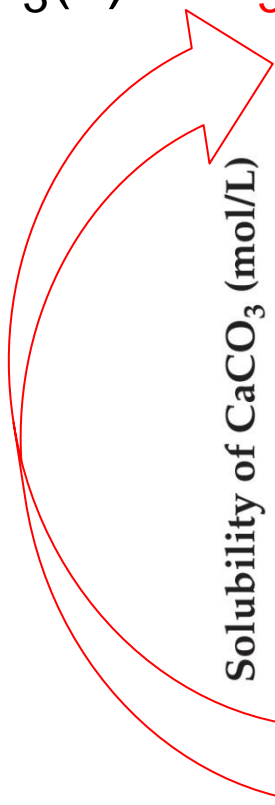
Molar solubility in 0.10 M NaF: $7.4 \times 10^{-9} \text{ M}$
(with common ion effect)

Why does the solubility decrease in the presence of a common ion?

Le Châtelier's Principle

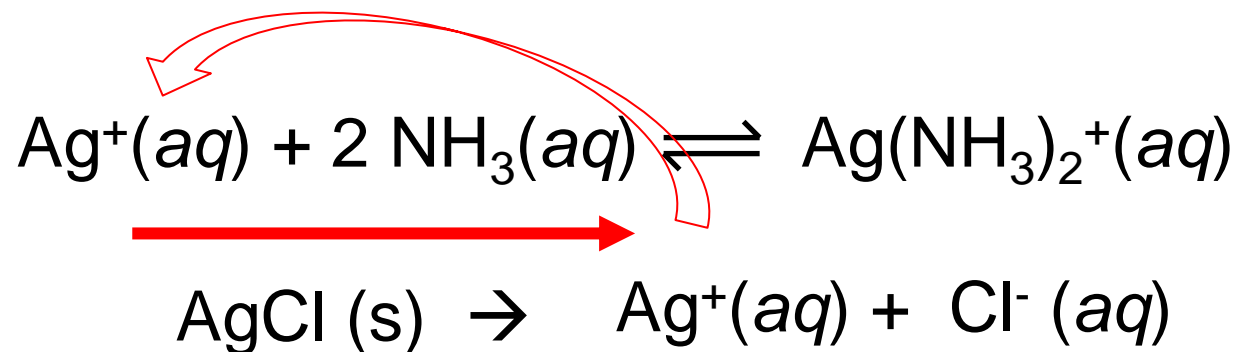
Factors That Affect Solubility

Solubility and the pH of the Solution – Le Chatelier (Common Ion of H_3O^+)



Factors That Affect Solubility

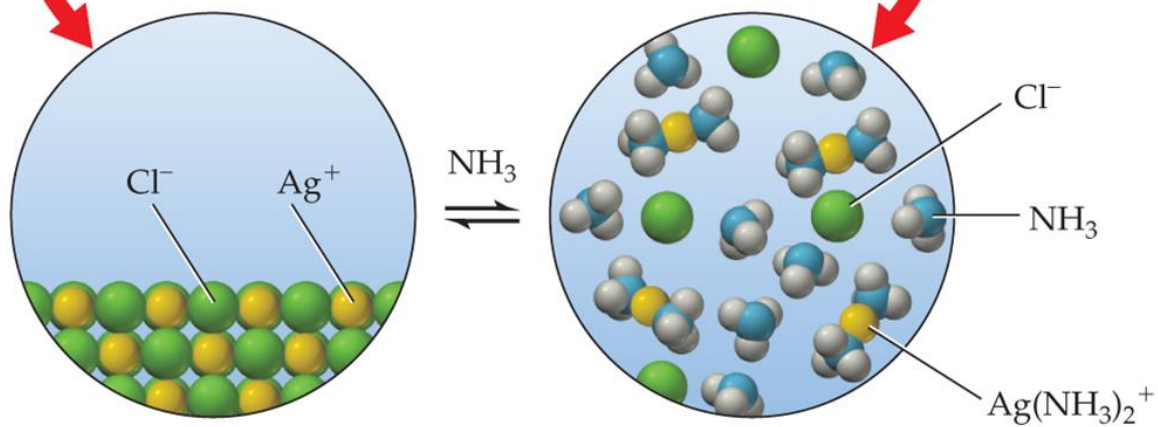
Solubility and the Formation of Complex Ions (common ion le Chatelier's)



Silver chloride is quite insoluble in water . . .

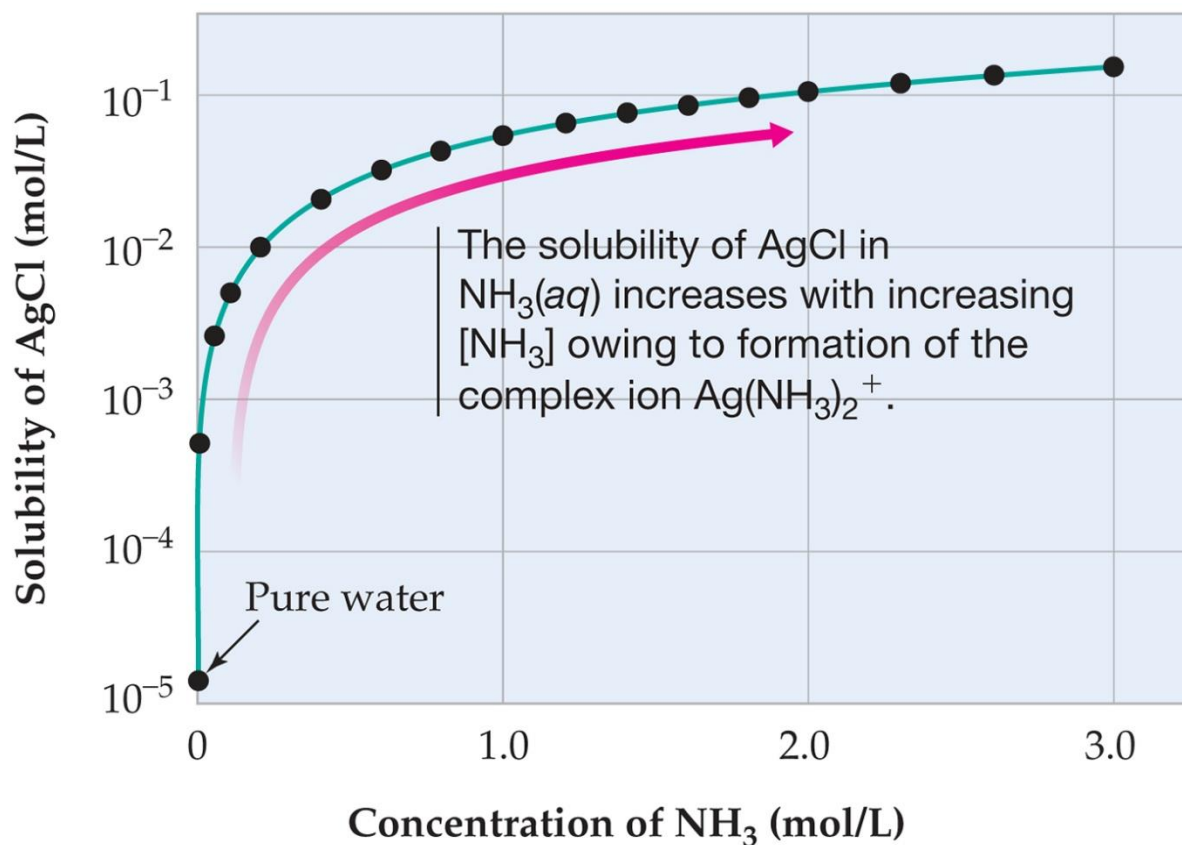
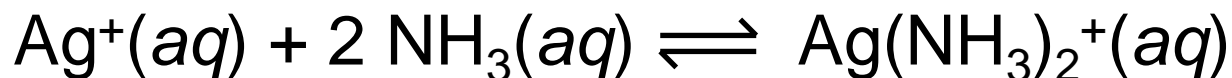


. . . but dissolves on addition of an excess of aqueous ammonia.



Factors That Affect Solubility

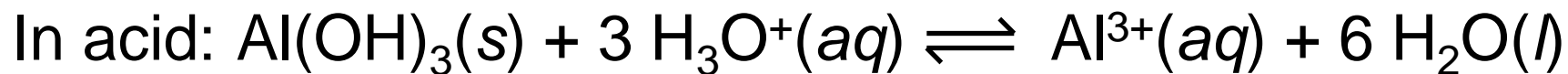
Solubility and the Formation of Complex Ions



Factors That Affect Solubility

Solubility and Amphoterism (both acidic & basic)

Aluminum hydroxide is soluble *both* in strongly acidic and in strongly basic solutions.



Factors That Affect Solubility

Solubility and Amphoterism

Aluminum hydroxide, a gelatinous white precipitate, forms on addition of aqueous NaOH to $\text{Al}^{3+}(\text{aq})$.



The precipitate dissolves on addition of excess aqueous NaOH, yielding the colorless $\text{Al}(\text{OH})_4^-$ ion. The precipitate also dissolves in aqueous HCl, yielding the colorless Al^{3+} ion.

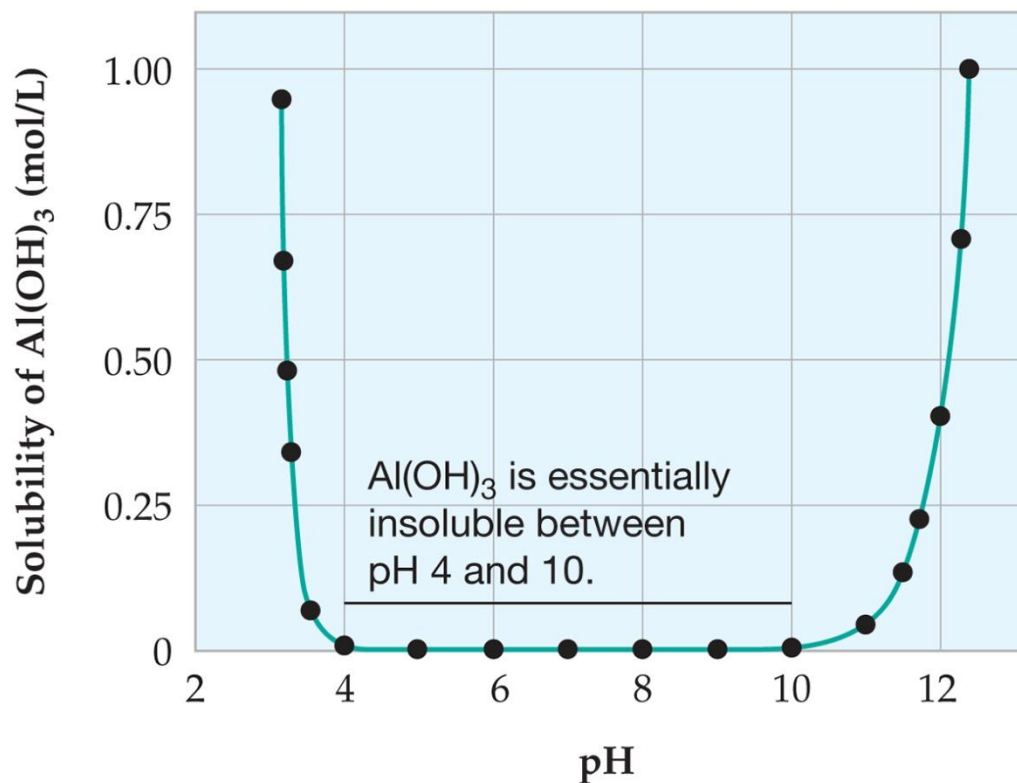


Factors That Affect Solubility

Solubility and Amphoterism

Al(OH)₃ dissolves in strongly acidic solutions.

Al(OH)₃ dissolves in strongly basic solutions.



ANSWERS TO HOMEWORK IN CLASS

HW 16.1: Neutralization Reactions

(a) Give the neutralization reaction for the following acid / base reaction



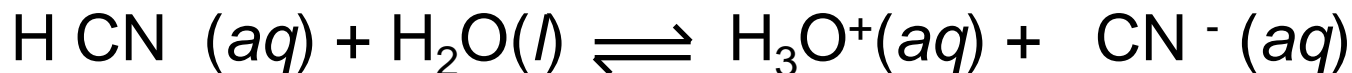
(b) Is the product of the neutralization.

[(acidic) **(basic)** (neutral)] (choose one)

K is neutral, F is weak base so weak base

HW 16.2: The Common-Ion Effect

What is the concentration of all species for the reaction of HCN ? $K_a = 4.9 \times 10^{-10}$ The solution has 0.025 M HCN and 0.010 M Na CN

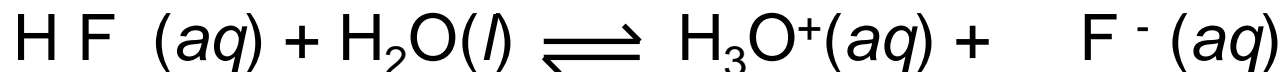


0.025		≈ 0	0.010
-x		+x	+x
0.025 - x		x	0.010 + x

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]}$$

HW 16.3: Buffer Solutions

What is the pH of a buffer solution with 0.25 M HF and 0.50 M NaF. $K_a = 3.5 \times 10^{-4}$. (using ICE table or Henderson Hasselbalch)



0.25		≈ 0	0.50
-x		+x	+x
0.250 - x		x	0.50 + x

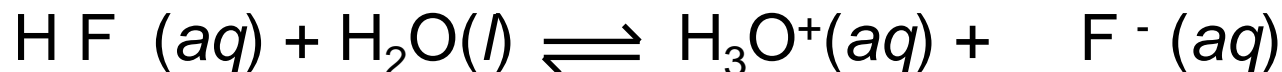
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(0.50+x)}{0.250-x} \sim \frac{(x)(0.50)}{0.250} = 3.5 \times 10^{-4}$$

$$x = \frac{(3.5 \times 10^{-4})(0.250)}{(0.50)} = 1.75 \times 10^{-4}$$

$$\text{pH} = -\log(1.75 \times 10^{-4}) = 3.757$$

HW 16.2: Buffer Solutions

What is the pH of a buffer solution with 0.25 M HF and 0.50 M NaF. $K_a = 3.5 \times 10^{-4}$. What is the pH after the addition of 0.002 mol HNO_3 ?



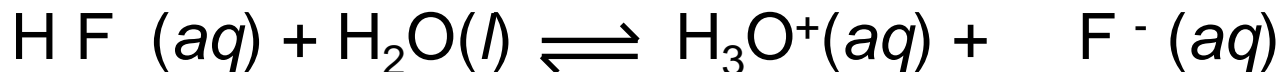
Using Henderson – Hasselbalch:

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log (3.5 \times 10^{-4}) + \log \frac{0.50}{0.25}$$

$$\text{pH} = 3.456 + 0.30 = 3.757$$

HW 16.3: Buffer Solutions

What is the pH of a buffer solution with 0.25 M HF and 0.50 M NaF. $K_a = 3.5 \times 10^{-4}$. (using ICE table or Henderson Hasselbalch)



0.25		≈ 0	0.50
-x		+x	+x
0.250 - x		x	0.50 + x

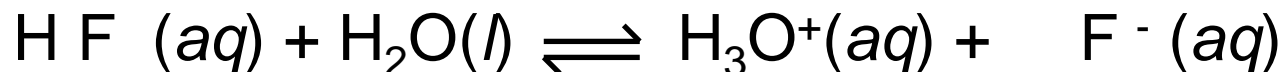
$$K_a = \frac{[\text{H}_3\text{O}^+][\text{F}^-]}{[\text{HF}]} = \frac{(x)(0.50+x)}{0.250-x} \sim \frac{(x)(0.50)}{0.250} = 3.5 \times 10^{-4}$$

$$x = \frac{(3.5 \times 10^{-4})(0.250)}{(0.50)} = 1.75 \times 10^{-4}$$

$$\text{pH} = -\log(1.75 \times 10^{-4}) = 3.757$$

HW 16.3: Buffer Solutions

What is the pH of a buffer solution with 0.25 M HF and 0.50 M NaF. $K_a = 3.5 \times 10^{-4}$. What is the pH after the addition of 0.002 mol HNO_3 ?



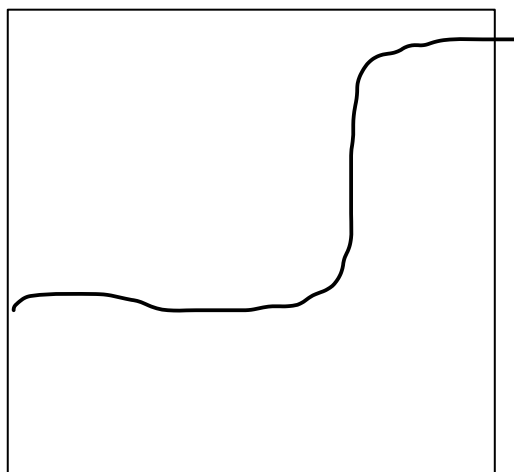
Using Henderson – Hasselbalch:

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log (3.5 \times 10^{-4}) + \log \frac{0.50}{0.25}$$

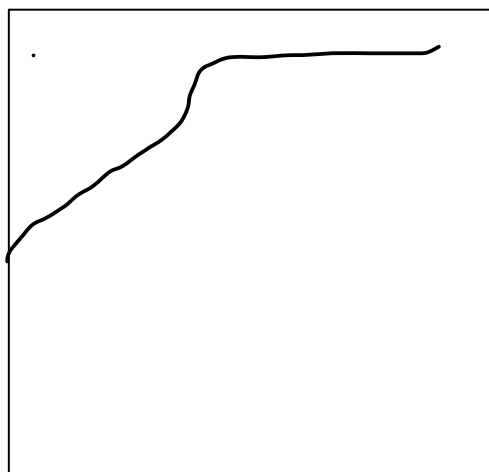
$$\text{pH} = 3.456 + 0.30 = 3.757$$

HW 16.4: pH Titration Curves

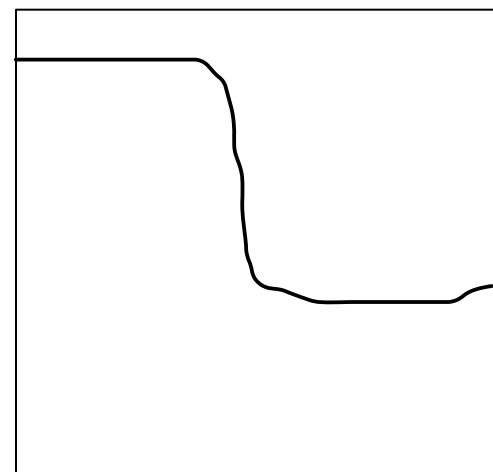
Label the following titration curve as (A) strong acid added to strong base (B) strong base added to strong acid (C) strong base added to weak acid (choose one letter under each #)



1.(A)(B)(C)



2.(A)(B)(C)



3.(A)(B)(C)

HW 16.5: pH Titration Curves

If you do a titration of 100.0 mL of 0.50 M HCl with 0.25 M NaOH, what is the volume of base at the equivalence point?

$$(M_{\text{acid}} V_{\text{acid}} = M_{\text{base}} V_{\text{base}})$$

$$V_{\text{acid}} = 100.0 \text{ mL HCl}$$

$$M_{\text{acid}} = 0.50 \text{ M HCl}$$

$$M_{\text{base}} = 0.25 \text{ M}$$

$$(0.50 \text{ M HCl})(100.0 \text{ mL HCl}) = (0.25 \text{ M NaOH})(x \text{ mL NaOH})$$

$$x \text{ mL NaOH} = 200.0 \text{ mL NaOH}$$

HW 16.6: Measuring K_{sp} and Calculating Solubility from K_{sp}

Calculate the molar solubility of Ag Cl in water at 25 °C.



	x	x
--	-----	-----

$$K_{sp} = 1.8 \times 10^{-10} = [\text{Ag}^+][\text{Cl}^-] = (x)(x)$$

$$x^2 = 1.8 \times 10^{-10}$$

$$x = [\text{Ag}^+] = \text{Molar solubility} = \boxed{1.3 \times 10^{-5} \text{ M}}$$