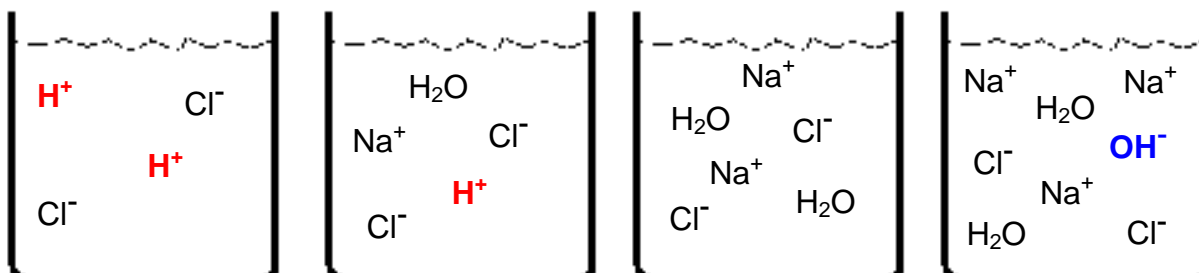


SA-SB TITRATION PROBLEMS KEY

1. In the beakers below:
 - A. Draw 2 moles HCl. **SA solution (SA completely ionizes)**
 - B. Draw what happens when 1 mole NaOH is added. **Excess SA (moles SA > moles SB)**
 - C. Draw what happens when 2 moles NaOH are added. **Eq pt (moles SA = moles SB); For SA-SB have neutral salt solution so pH = 7**
 - D. Draw what happens when 3 moles NaOH are added. **Excess SB (moles SB > moles SA)**



Note: for titration problems, you will need to calculate moles or mmoles of acid and base.

$$M = \frac{\text{mol}}{L} = \frac{\text{mmol}}{\text{mL}} \Rightarrow \text{thus, moles} = L \times M \text{ or mmoles} = \text{mL} \times M$$

General steps for SA-SB titration:

1. Write A-B neutralization reaction.
 2. Calculate mmol of acid and base present.
 3. Subtract smaller amount (LR) to find mmol of excess acid or base remaining.
 4. Divide mmol excess acid by total V (mL) to calculate $[\text{H}_3\text{O}^+]$ (or mmol excess base by the total mL to find $[\text{OH}^-]$)
 5. $\text{pH} = -\log [\text{H}_3\text{O}^+]$ (or find pOH if $[\text{OH}^-]$ is present and convert to pH)
2. Consider titrating 20.00 mL of 0.200 M **HCl** with 0.100 M **NaOH**.

SA **SB**

What volume of NaOH must be added to reach the equivalence point?



at EQ pt: moles acid = moles of base

$$\text{mmoles HCl} = 20.00 \text{ mL} \times \left(\frac{0.200 \text{ mmol HCl}}{\text{mL}} \right) = 4.00 \text{ mmol HCl} = 4.00 \text{ mmol NaOH}$$

$$\text{mL NaOH} = 4.00 \text{ mmol} \left(\frac{\text{mL}}{0.100 \text{ mmol}} \right) = 40.0 \text{ mL NaOH} = V \text{ NaOH at eq pt}$$

Calculate the pH of the solution after the following volumes of NaOH have been added:

- a) 0 ml of NaOH added – only **SA** is present initially:

Since SA completely ionizes: $[\text{HCl}] = [\text{H}_3\text{O}^+] = 0.200 \text{ M HCl}$ (3 sf)

$$\text{pH} = -\log [\text{H}_3\text{O}^+] \quad \text{pH} = -\log (0.200) \quad \text{pH} = 0.700 \checkmark \text{low pH for SA}$$

- b) 5.00 mL NaOH - When SB is added to SA, a neutralization reaction occurs:



$$\text{mmol HCl} = 20.00 \text{ mL} \times 0.200 \text{ M} = 4.00 \text{ mmol HCl}$$

$$\text{mmol NaOH} = 5.00 \text{ mL} \times 0.100 \text{ M} = 0.500 \text{ mmol NaOH (Limiting Reactant!)}$$

mmol excess acid = 4.00 mmol – 0.500 mmol = 3.50 mmol HCl

For SA: $[H_3O^+] = [HCl]$ $[H_3O^+] = \frac{3.50\text{mmoles}}{25.00\text{mL}} = 0.140\text{ M}$

$pH = -\log [H_3O^+]$ $pH = -\log 0.140$ **pH = 0.854** ✓ excess SA, pH < 7

c) 40.00 mL NaOH mmol NaOH = 40.00 mL × 0.100 M = 4.00 mmol NaOH

from part b: mmol HCl = 4.00 mmol HCl

moles acid = moles base ⇒ pH = 7 at Eq pt for SA-SB titration!

✓ **All the acid and base has been neutralized and all that remains is Na⁺, Cl⁻ and water. All neutral substances!**

d) 50.00 mL NaOH mmol NaOH = 50.00 mL × 0.100 M = 5.00 mmol NaOH

from part b: mmol HCl = 4.00 mmol HCl (**Limiting Reactant!**)

mmol excess base = 5.00 mmol - 4.00 mmol = 1.00 mmol NaOH

For SB: $[OH^-] = [NaOH]$ $[OH^-] = \frac{1.00\text{mmoles}}{70.00\text{mL}} = 0.01429\text{ M}$

$pOH = -\log (0.01429) = 1.845$ **pH = 12.155** ✓ excess SB, pH > 7

WA-SB TITRATION PROBLEMS KEY

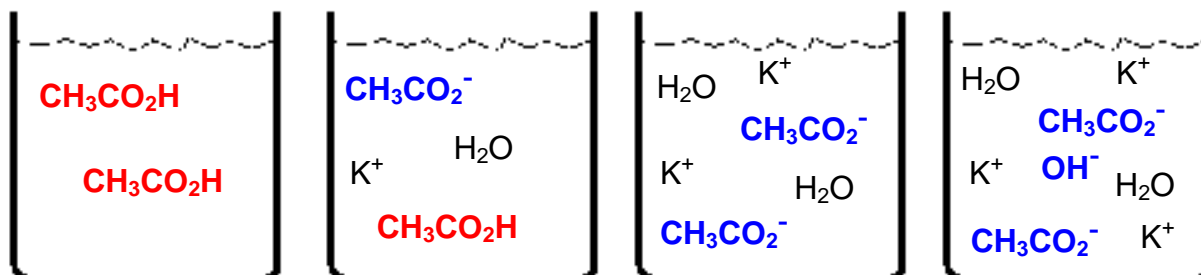
1. In the beakers below.

A. Draw 2 moles of acetic acid. **WA** (For 2 moles of WA – assume it has not ionized)

B. Draw what happens when 1 mole KOH is added. (Buffer: WA, conj base)

C. Draw what happens when 2 moles KOH are added. **Eq pt (WB hydrolysis)**

D. Draw what happens when 3 moles KOH are added. **Excess SB remains**



2. Consider the titration of 50.00 mL of 0.100 M CH_3CO_2H with 0.150 M KOH . For acetic acid, $K_a = 1.8 \times 10^{-5}$

For WA-SB titration problems, you must consider what substances are present and write the appropriate hydrolysis or neutralization reaction!

What volume of NaOH must be added to reach the equivalence point?

neutralization: $CH_3CO_2H + NaOH \rightarrow NaCH_3CO_2 + H_2O$

at EQ pt: moles acid = moles of base

$\text{mmoles} = 50.00\text{ mL} \times \left(\frac{0.100\text{mmol } CH_3CO_2H}{\text{mL}} \right) = 5.00\text{ mmol } CH_3CO_2H = 5.00\text{ mmol } KOH$

$\text{mL } KOH = 5.00\text{ mmol} \left(\frac{\text{mL}}{0.150\text{ mmol}} \right) = 33.33\text{ mL } NaOH = V\text{ NaOH at eq pt}$

Calculate the pH of the solution after the following volumes of NaOH have been added:

a) 0 mL KOH – **we just have a WA solution!**

Strategy: write WA hydrolysis, ICE & K_a , $x = [H_3O^+]$

	$CH_3CO_2H + H_2O(l) \rightleftharpoons H_3O^+ + CH_3CO_2^-$			
I (M)	0.100		0	0
C (M)	-x		+x	+x
E (M)	0.100 - x		x	x

$$K_a = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H]} \quad \text{Assume } x \text{ is small!} \quad 1.8 \times 10^{-5} = \frac{x^2}{0.100}$$

$$x = [H_3O^+] = \sqrt{(1.8 \times 10^{-5}) \times (0.100)} = 1.34 \times 10^{-3} \text{ M}$$

$$pH = -\log [H_3O^+] \quad pH = -\log 1.34 \times 10^{-3} \quad \text{pH} = 2.87 \quad \checkmark \text{relatively low pH for WA}$$

Other regions of WA curve:

1. Write WA-SB neutralization reaction.
2. Calculate mmol of acid and base present.
3. Plug mol into change table to determine what remains after neutralization:
 - ⇒ If moles HA > moles SB, both HA and A^- are present so it's a buffer problem.
 - ⇒ If moles HA = moles SB, only A^- is left so its WB problem (eq pt).
 - ⇒ If moles SB > moles WA, excess OH^- is left so its SB problem.

b) 10.00 mL KOH mmol KOH = mmol OH^- = 10.00 mL \times 0.150 M = 1.50 mmol KOH

mmol CH_3CO_2H = 50.00 mL \times 0.100 M = 5.00 mmol CH_3CO_2H

neutralization: $CH_3CO_2H + OH^- \rightarrow CH_3CO_2^- + H_2O$

Neutralization rxn	$CH_3CO_2H + OH^- \rightarrow CH_3CO_2^- + H_2O(l)$		
initial (mmol)	5.00	1.50	0
Change (mmol)	-1.50	-1.50	+1.50
Final (mmol)	3.50	0	1.50

CH_3CO_2H (WA) and $CH_3CO_2^-$ (WB) remain = buffer solution!

***Can use HH equation – have WA and Conjugate Base!**

$$pH = -\log K_a + \log \left(\frac{[A^-]}{[HA]} \right) \quad pH = -\log 1.8 \times 10^{-5} + \log (1.50/3.50)$$

$$pH = 4.74 + -.368 \quad \text{pH} = 4.380$$

pH ↑ after SB added; now have buffer soln containing acetic acid and acetate ions.

c) the equivalence point volume of KOH = 33.33 mL WA-SB eq pt is WB problem!

from part b, mmol CH_3CO_2H = 5.00 mmol CH_3CO_2H

mmol KOH = mmol OH^- = 33.33 mL \times 0.150 M = 5.00 mmol KOH

neutralization: $CH_3CO_2H + OH^- \rightarrow CH_3CO_2^- + H_2O$

Neutralization rxn	$CH_3CO_2H + OH^- \rightarrow CH_3CO_2^- + H_2O(l)$		
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initial (mmol)	5.00	5.00	0	
Change (mmol)	-5.00	-5.00	+5.00	
Final (mmol)	0	0	5.00	

At eq pt: all WA & SB have reacted, only CH_3CO_2^- (WB) remains!

⇒ Set up WB hydrolysis rxn, solve K_b for $x = [\text{OH}^-]$

Convert mmol (or mol) acetate to Molarity – must use M in K_b expression!

$$M = \frac{\text{mmol}}{\text{mL acid} + \text{mL base}} \quad [\text{CH}_3\text{CO}_2^-] = \frac{5.00 \text{ mmol}}{50.00 + 33.33 \text{ mL}} = 0.0600 \text{ M}$$

WB hydrolysis: $\text{CH}_3\text{CO}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{CO}_2\text{H} + \text{OH}^-$

	$\text{CH}_3\text{CO}_2^- + \text{H}_2\text{O}$	\rightleftharpoons	$\text{CH}_3\text{CO}_2\text{H} + \text{OH}^-$
I (M)	0.0600		0
C (M)	-x		+x
E (M)	0.0600 - x		x

Need K_b instead of K_a for WB: K_b for $\text{CH}_3\text{CO}_2^- = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$

$$K_b = \frac{[\text{CH}_3\text{CO}_2\text{H}][\text{OH}^-]}{[\text{CH}_3\text{CO}_2^-]} \quad 5.56 \times 10^{-10} = \frac{x^2}{0.0600}$$

$$x = [\text{OH}^-] = \sqrt{0.0600 (5.56 \times 10^{-10})} = 5.78 \times 10^{-6} \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log 5.78 \times 10^{-6} = 5.24 \quad \text{pH} = 14 - 5.24 = \mathbf{8.76}$$

⇒ ✓ At the equivalence point for a WA/SB titration, the $\text{pH} > 7$ due to the OH^- produced by the conjugate base hydrolysis reaction.

d) 50.00 mL of KOH

from part b, mmoles $\text{CH}_3\text{CO}_2\text{H} = 5.00 \text{ mmol CH}_3\text{CO}_2\text{H}$

mmoles KOH = mmol $\text{OH}^- = 50.00 \text{ mL} \times 0.150 \text{ M} = 7.50 \text{ mmol OH}^-$

neutralization: $\text{CH}_3\text{CO}_2\text{H} + \text{OH}^- \rightarrow \text{CH}_3\text{CO}_2^- + \text{H}_2\text{O}$

Neutralization rxn	$\text{CH}_3\text{CO}_2\text{H} + \text{OH}^-$	\rightarrow	$\text{CH}_3\text{CO}_2^- + \text{H}_2\text{O}(l)$
initial (mmol)	5.00	7.50	0
Change (mmol)	-5.00	-5.00	+5.00
Final (mmol)	0	2.50	5.00

moles excess SB remaining = 2.50 mmoles KOH

$$M \text{ OH}^- = M \text{ KOH} = \frac{2.50 \text{ mmol}}{100 \text{ mL}} = 0.0250 \text{ M OH}^-$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log 0.0250 = 1.602 \quad \text{pH} = 14 - 1.602 = \mathbf{12.398}$$

✓ high pH since SB not all neutralized

*Excess NaOH remains - this is the primary source of OH^- . We can neglect the hydrolysis of the conjugate base because this contributes a relatively small amount of OH^- compared to the amount that comes directly from the excess NaOH.