

1. a) Calculate the pH of a buffer solution prepared by adding 2.58 g of NH_4Cl to 125.0 mL of 0.225 M NH_3 . For NH_3 , $K_b = 1.8 \times 10^{-5}$ (5 pts)

$$M \text{ NH}_4\text{Cl} = \left(\frac{2.58 \text{ g}}{0.125 \text{ L}} \right) \left(\frac{1 \text{ mol NH}_4\text{Cl}}{53.5 \text{ g NH}_4\text{Cl}} \right) = \frac{0.386 \text{ moles NH}_4\text{Cl}}{L}$$

NH_4Cl is a soluble salt, so it breaks up into ions in solution: $\text{NH}_4\text{Cl} \rightarrow \text{NH}_4^+ + \text{Cl}^-$

$[\text{NH}_4^+] = [\text{NH}_4\text{Cl}] = 0.386 \text{ M}$ $\text{NH}_4^+ \leftarrow$ **NH_4^+ is a weak acid; the conjugate acid of NH_3**

$[\text{NH}_3] = 0.225 \text{ M}$ $\text{NH}_3 \leftarrow$ **NH_3 is a weak base**



	$[\text{NH}_3] \text{ M}$	$[\text{H}_2\text{O}]$	$[\text{NH}_4^+] \text{ M}$	$[\text{OH}^-] \text{ M}$
I	0.225		0.386	0
C	-x		+x	+x
E	0.225 - x		0.386 + x	x

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} \quad 1.8 \times 10^{-5} = \frac{(0.386)x}{(0.225)} \quad x = [\text{OH}^-] = 1.05 \times 10^{-5} \text{ M}$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log 1.05 \times 10^{-5} = 4.98 \quad \text{pH} = 14 - 4.98 = \mathbf{9.02}$$

Note: If you use $\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+$, you will need to convert K_b to K_a , and find $x = [\text{H}^+]$.

- b) Calculate the pH after adding 25.0 mL of 0.215 M HBr to the 125.0 mL buffer solution in part a. (6 pts)

$$\text{strong acid: HBr} \rightarrow \text{H}^+ + \text{Br}^-; \quad [\text{H}^+] = [\text{HBr}] = \frac{0.215 \text{ M} \times 25.0 \text{ ml}}{150.0 \text{ ml}} = 0.0358 \text{ M}$$

$$[\text{NH}_4^+] = \frac{0.386 \text{ M} \times 125.0 \text{ ml}}{150.0 \text{ ml}} = 0.322 \text{ M}; \quad [\text{NH}_3] = \frac{0.225 \text{ M} \times 125.0 \text{ ml}}{150.0 \text{ ml}} = 0.188 \text{ M}$$

H^+ will react with the weak base, NH_3 : $[\text{NH}_3] \downarrow$ and $[\text{NH}_4^+] \uparrow$



	$[\text{H}^+] \text{ M}$	$[\text{NH}_3] \text{ M}$	$[\text{NH}_4^+] \text{ M}$
I	0.0358	0.188	0.322
C	-0.0358	-0.0358	+0.0358
After neutralization	0	0.152	0.358

reestablish eq: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ (or $\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+$)

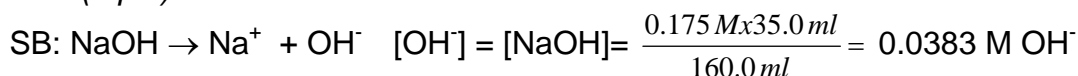
	$[\text{NH}_3] \text{ M}$	$[\text{H}_2\text{O}]$	$[\text{NH}_4^+] \text{ M}$	$[\text{OH}^-] \text{ M}$
I	0.152		0.358	0
C	-x		+x	+x
E	0.152 - x		0.358 + x	x

$$1.8 \times 10^{-5} = \frac{(0.358)x}{(0.152)} \quad x = [\text{OH}^-] = 7.64 \times 10^{-6} \text{ M}$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log 7.64 \times 10^{-6} = 5.12 \quad \text{pH} = 14 - 5.12 = \mathbf{8.88}$$

***note pH decreased a little when SA added**

c) Calculate the pH after adding 35.0 mL of 0.175 M NaOH to the 125.0 mL buffer solution in part a. (6 pts)



$$[\text{NH}_4^+] = \frac{0.386 \text{ M} \times 125.0 \text{ ml}}{160.0 \text{ ml}} = 0.302 \text{ M}; \quad [\text{NH}_3] = \frac{0.225 \text{ M} \times 125 \text{ ml}}{160.0 \text{ ml}} = 0.176 \text{ M}$$

OH^- will react with the weak acid, NH_4^+ : $[\text{NH}_4^+] \downarrow$ and $[\text{NH}_3] \uparrow$

acid-base neutralization: $\text{OH}^- + \text{NH}_4^+ \rightarrow \text{NH}_3 + \text{H}_2\text{O}$

	$[\text{OH}^-] \text{ M}$	$[\text{NH}_4^+] \text{ M}$	$[\text{NH}_3] \text{ M}$	
I	0.0383	0.302 M	0.176	
C	-0.0383	-0.0383	+0.0383	
After neutralization	0	0.264	0.214	

reestablish eq: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$ (or $\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+$)

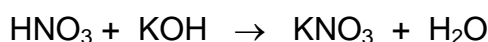
	$[\text{NH}_3] \text{ M}$	$[\text{H}_2\text{O}]$	$[\text{NH}_4^+] \text{ M}$	$[\text{OH}^-] \text{ M}$
I	0.214		0.264	0
C	-x		+x	+x
E	$0.214 - x$		$0.264 + x$	x

$$1.8 \times 10^{-5} = \frac{(0.264)x}{(0.214)} \quad x = [\text{OH}^-] = 1.46 \times 10^{-5} \text{ M}$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log 1.46 \times 10^{-5} = 4.84 \quad \text{pH} = 14 - 4.84 = \mathbf{9.16}$$

* note pH increased a little when SB added

2. Calculate the pH when 275.0 mL of 0.114 M HNO_3 is titrated with 245.0 mL of 0.315 M KOH. (4 pts)



$$\text{moles HNO}_3 = 0.275 \text{ L} \times 0.114 \text{ M HNO}_3 = 0.0314 \text{ moles H}^+$$

$$\text{moles KOH} = 0.245 \text{ L} \times 0.315 \text{ M KOH} = 0.0772 \text{ moles OH}^-$$

$$\text{moles OH}^- \text{ remaining} = 0.0772 \text{ moles} - 0.0314 \text{ moles} = 0.0458 \text{ moles}$$

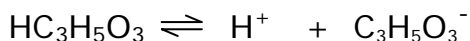
$$\text{M OH}^- = \frac{4.58 \times 10^{-2} \text{ moles}}{0.5200 \text{ L}} = 8.81 \times 10^{-2} \text{ M OH}^-$$

$$\text{pOH} = -\log 8.81 \times 10^{-2} = 1.055 \quad \text{pH} = 14 - 1.055 = \mathbf{12.945}$$

3. A 12.5 mL sample of 0.400 M HC₃H₅O₃, lactic acid, is titrated with 0.200 M KOH. For lactic acid, K_a = 1.4 × 10⁻⁴.

a) What is the initial pH of the acid solution? (3 pts)

For a weak acid [HA] ≠ [H⁺] because the weak acid is only partially ionized!



	[HC ₃ H ₅ O ₃]	[H ⁺]	[C ₃ H ₅ O ₃ ⁻]
I	0.400	0	0
C	-x	+x	+x
E	0.400 - x	x	x

$$K_a = \frac{[\text{H}^+][\text{C}_3\text{H}_5\text{O}_3^-]}{[\text{HC}_3\text{H}_5\text{O}_3]} \quad 1.4 \times 10^{-4} = \frac{x^2}{0.400}$$

$$[\text{H}^+] = x = \sqrt{0.400(1.4 \times 10^{-4})} = 7.48 \times 10^{-3}$$

$$\text{pH} = -\log 7.48 \times 10^{-3} = 2.13$$

b) What is the pH after the addition of 25.0 mL of 0.200 M KOH? (6 pts)

$$\text{moles HC}_3\text{H}_5\text{O}_3 = 12.5 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.400 \text{ moles HC}_3\text{H}_5\text{O}_3}{\text{L}} \right) = 5.00 \times 10^{-3} \text{ moles HC}_3\text{H}_5\text{O}_3$$

$$\text{moles KOH} = 25.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.200 \text{ moles KOH}}{\text{L}} \right) = 5.00 \times 10^{-3} \text{ moles KOH}$$

acid-base neutralization: $\text{HC}_3\text{H}_5\text{O}_3 + \text{OH}^- \rightarrow \text{C}_3\text{H}_5\text{O}_3^- + \text{H}_2\text{O}$

I	0.00500	0.00500	0
C	-0.00500	-0.00500	+0.00500
moles after neut.	0	0	0.00500

At EQ Pt: moles acid = moles base; conjugate base reacts with H₂O

$$[\text{C}_3\text{H}_5\text{O}_3^-] = \frac{5.00 \times 10^{-3} \text{ moles}}{0.0375 \text{ L}} = 0.133 \text{ M}$$

weak base hydrolysis: $\text{C}_3\text{H}_5\text{O}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_3\text{H}_5\text{O}_3 + \text{OH}^-$

I	0.133		0	0
C	-x		x	x
E	0.133 - x		x	x

$$K_b \text{ for } \text{C}_3\text{H}_5\text{O}_3^- = \frac{1 \times 10^{-14}}{1.4 \times 10^{-4}} = 7.1 \times 10^{-11}$$

$$K_b = \frac{[\text{HC}_3\text{H}_5\text{O}_3][\text{OH}^-]}{[\text{C}_3\text{H}_5\text{O}_3^-]} \quad 7.1 \times 10^{-11} = \frac{x^2}{0.133}$$

$$x = [\text{OH}^-] = \sqrt{0.133(7.1 \times 10^{-11})} = 3.07 \times 10^{-6}$$

$$\text{pOH} = -\log 3.07 \times 10^{-6} = 5.51 \quad \text{pH} = 14 - 5.51 = 8.49 \text{ at equiv pt}$$

⇒ **pH > 7 BECAUSE THE CONJUGATE BASE REACTS WITH WATER TO FORM OH⁻!**