

Name: KEY Per \_\_\_\_\_ Date: \_\_\_\_\_

17.27

1. How many milliliters of 0.0850 M NaOH are required to titrate each of the following solutions to the equivalence point: *At equivalence point, moles of base added equals moles acid initially present. solve stoichiometrically, remembering mol = M x*

a. 40.0 mL of 0.0900 M HNO<sub>3</sub>

$$(40.0 \text{ mL HNO}_3) \left( \frac{0.0900 \text{ mol HNO}_3}{1000 \text{ mL}} \right) \left( \frac{1 \text{ mol NaOH}}{1 \text{ mol HNO}_3} \right) \left( \frac{1000 \text{ mL}}{0.0850 \text{ mol NaOH}} \right) = 42.4 \text{ mL NaOH}$$

b. 35.0 mL of 0.0720 M HBr

$$(35.0 \text{ mL HBr}) \left( \frac{0.0720 \text{ M HBr}}{1000 \text{ mL}} \right) \left( \frac{1 \text{ mol NaOH}}{1 \text{ mol HBr}} \right) \left( \frac{1000 \text{ mL}}{0.0850 \text{ mol NaOH}} \right) = 29.6 \text{ mL NaOH}$$

c. 50.0 mL of a solution that contains 1.85 g of HCl per liter

$$\left( \frac{1.85 \text{ g HCl}}{1 \text{ L soln}} \right) \left( \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \right) = 0.05074 = 0.0507 \text{ M HCl}$$

$$(50 \text{ mL HCl}) \left( \frac{0.05074 \text{ mol HCl}}{1000 \text{ mL}} \right) \left( \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} \right) \left( \frac{1000 \text{ mL soln}}{0.0850 \text{ mol NaOH}} \right) = 29.8 \text{ mL NaOH}$$

17.29

2. A 20.0 mL sample of 0.200 M HBr solution is titrated with 0.200 M NaOH solution. Calculate the pH of the solution after the following volumes of base have been added: *strong acid - strong base titration*

a. 15.0 mL

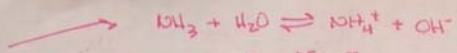
$$\begin{aligned} \text{moles H}^+ &= M_{\text{HBr}} \times L_{\text{HBr}} = (0.200 \text{ M}) (0.020 \text{ L}) = 4.00 \times 10^{-3} \text{ mol} \\ \text{moles OH}^- &= M_{\text{NaOH}} \times L_{\text{NaOH}} \end{aligned}$$

	mL HBr	mL NaOH	Total Vol	Mol H <sup>+</sup>	Mol OH <sup>-</sup>	Excess Ion	pH
b. 19.9 mL	a. 20.0	15.0	35.0	4.0 × 10 <sup>-3</sup>	3.0 × 10 <sup>-3</sup>	0.0286 M (H <sup>+</sup> )	1.544
	b. 20.0	19.9	39.9	"	3.98 × 10 <sup>-3</sup>	5 × 10 <sup>-4</sup> [H <sup>+</sup> ]	3.3
	c. 20.0	20.0	40.0	"	4.00 × 10 <sup>-3</sup>	1.0 × 10 <sup>-7</sup> [H <sup>+</sup> ]	7.0
	d. 20.0	20.1	40.1	"	4.02 × 10 <sup>-3</sup>	5.0 × 10 <sup>-4</sup> [H <sup>+</sup> ]	10.7
c. 20.0 mL	e. 20.0	35.0	55.0	"	7.00 × 10 <sup>-3</sup>	0.0545 (OH <sup>-</sup> )	12.737

d. 20.1 mL

e. 35.0 mL

3

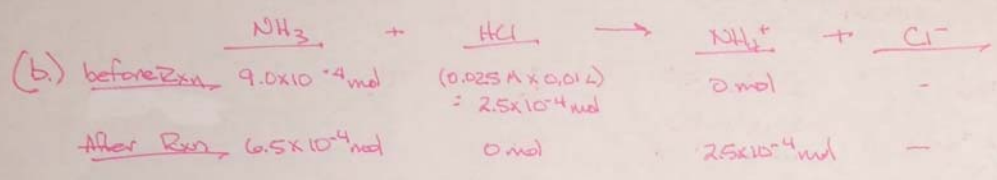


(a) weak base problem:  $K_b = 1.8 \times 10^{-5} = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$

at equilibrium,  $[\text{OH}^-] = x$ ,  $[\text{NH}_3] = (0.030 - x)$ ,  $[\text{NH}_4^+] = x$

$$1.8 \times 10^{-5} = \frac{x^2}{(0.030 - x)} \approx \frac{x^2}{0.030}; \quad x = [\text{OH}^-] = 7.348 \times 10^{-4} \quad \boxed{\text{pH} = 14 - 3.13 = 10.87}$$

(b-f) Calculate mol  $\text{NH}_3$  and mol  $\text{NH}_4^+$  after the acid-base Rxn.  
 $0.030 \text{ M NH}_3 \times 0.0300 \text{ L} = 9.0 \times 10^{-4} \text{ mol NH}_3$  present initially



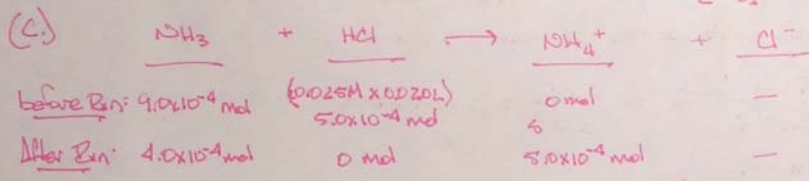
using the acid dissociation equilibrium for  $\text{NH}_4^+$  (so that we calculate  $[\text{H}^+]$  directly),  
 $\text{NH}_4^+ \rightleftharpoons \text{H}^+ + \text{NH}_3$  (easier than via pOH)

$$K_a = \frac{[\text{H}^+][\text{NH}_3]}{[\text{NH}_4^+]} = \frac{K_w}{K_b \text{ for NH}_3} = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.56 \times 10^{-10}$$

$$[\text{NH}_3] = \frac{6.5 \times 10^{-4} \text{ mol}}{0.040 \text{ L}} = 0.01625 \text{ M} \quad [\text{NH}_4^+] = \frac{2.5 \times 10^{-4} \text{ mol}}{0.040 \text{ L}} = 6.25 \times 10^{-3} \text{ M}$$

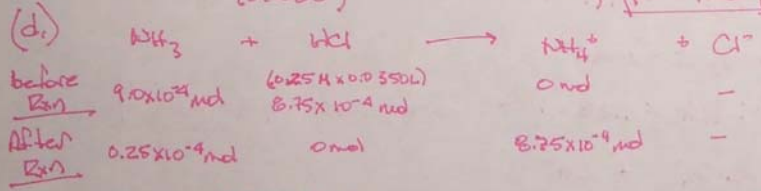
$$[\text{H}^+] = \frac{5.56 \times 10^{-10} [\text{NH}_4^+]}{[\text{NH}_3]} = \frac{5.56 \times 10^{-10} (6.25 \times 10^{-3})}{(0.01625)} = 2.14 \times 10^{-10} \quad \boxed{\text{pH} = 9.67}$$

(we will assume  $[\text{H}^+]$  is small compared to  $[\text{NH}_3]$  and  $[\text{NH}_4^+]$ )



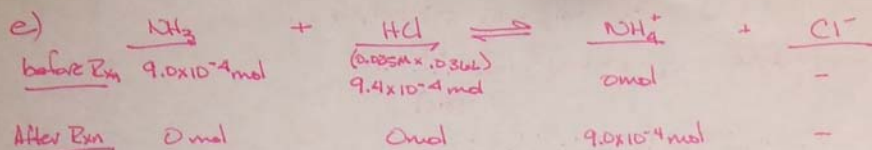
$$[\text{NH}_3] = \frac{4.0 \times 10^{-4} \text{ mol}}{0.050 \text{ L}} = 0.0080 \text{ M} \quad [\text{NH}_4^+] = \frac{5.0 \times 10^{-4} \text{ mol}}{0.0500 \text{ L}} = 0.010 \text{ M}$$

$$[\text{H}^+] = \frac{5.56 \times 10^{-10} (0.010)}{(0.0080)} = 6.94 \times 10^{-10} \quad \boxed{\text{pH} = 9.16}$$



$$[\text{NH}_3] = \frac{0.25 \times 10^{-4} \text{ mol}}{0.0650 \text{ L}} = 3.846 \times 10^{-4} \text{ M} \quad [\text{NH}_4^+] = \frac{8.75 \times 10^{-4} \text{ mol}}{0.0650 \text{ L}} = 0.01346 \text{ M}$$

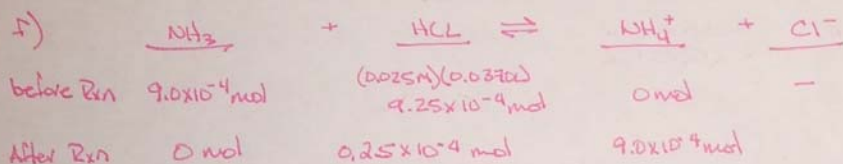
$$[\text{H}^+] = \frac{5.56 \times 10^{-10} (0.01346)}{3.846 \times 10^{-4}} = 1.946 \times 10^{-8} \quad \boxed{\text{pH} = 7.7}$$



at the equivalence point  $[\text{H}^+] = [\text{NH}_3] = x$

$$[\text{NH}_4^+] = \frac{9.0 \times 10^{-4} \text{ M}}{0.0166 \text{ L}} = 0.01364 \text{ M}$$

$$5.56 \times 10^{-10} = \frac{x^2}{0.01364}; \quad x = [\text{H}^+] = 2.754 \times 10^{-6} \quad \boxed{\text{pH} = 5.56}$$



Past the equivalence point,  $[\text{H}^+]$  from the excess HCl determines the pH

$$[\text{H}^+] = \frac{0.25 \times 10^{-4} \text{ mol}}{0.067 \text{ L}} = 3.731 \times 10^{-4} \text{ M} \quad \boxed{\text{pH} = 3.4}$$

4) The volume of NaOH solution required in all cases is

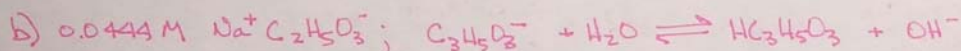
$$V_{\text{base}} = \frac{V_{\text{acid}} \times M_{\text{acid}}}{M_{\text{base}}} = \frac{(0.100) V_{\text{acid}}}{(0.080)} = 1.25 V_{\text{acid}}$$

The total volume at the equivalence point is  $V_{\text{base}} + V_{\text{acid}} = 2.25 V_{\text{acid}}$

The concentration of the salt at the equivalence point is:

$$\frac{M_{\text{acid}} V_{\text{acid}}}{2.25 V_{\text{acid}}} = \frac{0.100}{2.25} = 0.0444 \text{ M}$$

a) 0.0444 M NaBr,  $\boxed{\text{pH} = 7.00}$



$$K_b = \frac{[\text{HC}_2\text{H}_5\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{O}_2^-]} = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{1.4 \times 10^{-4}} = 7.14 \times 10^{-11}$$

$$[\text{HC}_2\text{H}_5\text{O}_2] = [\text{OH}^-]; \quad [\text{C}_2\text{H}_5\text{O}_2^-] \approx 0.0444$$

$$[\text{OH}^-]^2 \approx 0.0444 (7.14 \times 10^{-11}); \quad [\text{OH}^-] = 1.78 \times 10^{-6} \quad \text{pOH} = 5.75 \quad \boxed{\text{pH} = 8.25}$$



$$K_b = \frac{[\text{HCrO}_4^-][\text{OH}^-]}{[\text{CrO}_4^{2-}]} = \frac{K_w}{K_a} = 3.33 \times 10^{-8}$$

$$[\text{OH}^-]^2 \approx 0.0444 (3.33 \times 10^{-8}); \quad [\text{OH}^-] = 3.845 \times 10^{-5} \quad \boxed{\text{pH} = 9.58}$$