LOHS AP Chemistry - Titration Calculations Chp 17 – Addit Aspects of Aqueous Equil Schrempp

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×

a. 40.0 mL of 0.0900 M HNO_3

$$\left(40.0\,\text{mL } HNO_3\right)\left(\frac{0.0900\,\text{mol } HNO_3}{1000\,\text{mL}}\right)\left(\frac{1\,\text{mol NaOH}}{1\,\text{mol } HNO_3}\right)\left(\frac{1000\,\text{mL}}{0.0850\,\text{mol NaOH}}\right) = 42.4\,\text{mL NaOH}$$

b. 35.0 mL of 0.0720 M HBr

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

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

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

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

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

moles $H^+ = M_{HBr} \times L_{HBr} = (.200\,M)(0.020\,L) = 4.00\times10^{-3}$ mol

Moles $OH^- = M_{NaOH} \times L_{NaOH}$

a. 15.0 mL

b. 19.9 mL

c. 20.0 mL

	mL_{HBr}	mL_{NaOH}	Total Vol	Mol H^+	mol OH^-	Excess ION (Molarity = moles/vol)	pH
a.	20.0	15.0	35.0	4.0×10^{-3}	3.0×10^{-3}	0.0286 M (H^+)	1.544
b.	20.0	19.9	39.9	"	3.98×10^{-3}	5×10^{-4} [H^+]	3.3
c.	20.0	20.0	40.0	"	4.00×10^{-3}	1.0×10^{-7} [H^+]	7.0
d.	20.0	20.1	40.1	"	4.02×10^{-3}	5.0×10^{-4} (H^+]	10.7
e.	20.0	35.0	55.0	"	7.00×10^{-3}	0.0545 (OH^-)	12.737

d. 20.1 mL

e. 35.0 mL

(3)

E13.7 (a) Weak base problem: $K_b = 1.8\times10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]}$ $\longrightarrow$ $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

at equilibrium, $[OH^-] = x$, $[NH_3] = (0.030 - x)$; $[NH_4^+] = x$

$1.8\times10^{-5} = \frac{x^2}{(0.030-x)} \simeq \frac{x^2}{0.030}$; $x = [OH^-] = 7.348\times10^{-4}$ $\boxed{pH = 14 - 3.13 = 10.87}$

(b-f) Calculate mol NH_3 and mol NH_4^+ after the acid-base Rxn.

$0.030\ M\ NH_3 \times 0.0300\,L = 9.0\times10^{-4}$ mol NH_3 present initially

(b.)

	NH_3	+ HCl	$\longrightarrow$ NH_4^+	+ Cl^-
before Rxn	9.0×10^{-4} mol	$(0.025\,M \times 0.01\,L) = 2.5\times10^{-4}$ mol	0 mol	-
After Rxn	6.5×10^{-4} mol	0 mol	2.5×10^{-4} mol	—

using the acid dissociation equilibrium for NH_4^+ (so that we calculate $[H^+]$ directly, (rather than via pOH)

$NH_4^+ \rightleftharpoons H^+ + NH_3$

$K_a = \frac{[H^+][NH_3]}{NH_4} = \frac{K_w}{K_b \text{ for } NH_3} = \frac{1.0\times10^{-14}}{1.8\times10^{-5}} = 5.56\times10^{-10}$

$[NH_3] = \frac{6.5\times10^{-4}\text{ mol}}{0.040\,L} = 0.01625\,M$ $\quad [NH_4^+] = \frac{2.5\times10^{-4}\text{ mol}}{0.040\,L} = 6.25\times10^{-3}\,M$

$[H^+] = \frac{5.56\times10^{-10}[NH_4^+]}{[NH_3]} = \frac{5.56\times10^{-10}(6.25\times10^{-3})}{(0.01625)} = 2.14\times10^{-10}$; $\boxed{pH = 9.67}$

(we will assume $[H^+]$ is small compared to $[NH_3]$ and $[NH_4^+]$)

(c.)

	NH_3	+ HCl	$\longrightarrow$ NH_4^+	+ Cl^-
before Rxn:	9.0×10^{-4} mol	$(0.025M \times 0.020L)$ 5.0×10^{-4} mol	0 mol	—
After Rxn:	4.0×10^{-4} mol	0 mol	5.0×10^{-4} mol	—

$[NH_3] = \frac{4.0\times10^{-4}\text{ mol}}{0.050\,L} = 0.0080\,M$ $\quad [NH_4^+] = \frac{5.0\times10^{-4}\text{ mol}}{0.0500\,L} = 0.010\,M$

$[H^+] = \frac{5.56\times10^{-10}(0.010)}{(0.0080)} = 6.94\times10^{-10}$; $\boxed{pH = 9.16}$

(d.)

	NH_3	+ HCl	$\longrightarrow$ NH_4^+	+ Cl^-
before Rxn	9.0×10^{-4} mol	$(0.025M \times 0.0350L)$ 8.75×10^{-4} mol	0 mol	—
After Rxn	0.25×10^{-4} mol	0 mol	8.75×10^{-4} mol	—

$[NH_3] = \frac{0.25\times10^{-4}\text{ mol}}{0.0650\,L} = 3.846\times10^{-4}\,M$ $\quad [NH_4^+] = \frac{8.75\times10^{-4}\text{ mol}}{0.0650\,L} = 0.01346\,M$

$[H^+] = \frac{5.56\times10^{-10}(0.01346)}{3.846\times10^{-4}} = 1.946\times10^{-8}$; $\boxed{pH = 7.7}$

e)

	NH_3	+	HCl	$\rightleftharpoons$	NH_4^+	+	Cl^-
before Rxn	9.0×10^{-4} mol		$(0.025M \times .0376L)$ 9.4×10^{-4} mol		0 mol		—
After Rxn	0 mol		0 mol		9.0×10^{-4} mol		—

at the equivalence point $[H^+] = [NH_3] = x$

$$[NH_4^+] = \frac{9.0\times10^{-4}\,M}{0.0660\,L} = 0.01364\,M$$

$$5.56\times10^{-10} = \frac{x^2}{0.01364}\ ;\quad x = [H^+] = 2.754\times10^{-6} \quad \boxed{pH = 5.56}$$

f)

	NH_3	+	HCL	$\rightleftharpoons$	NH_4^+	+	Cl^-
before Rxn	9.0×10^{-4} mol		$(0.025M)(0.0370L)$ 9.25×10^{-4} mol		0 mol		—
After Rxn	0 mol		0.25×10^{-4} mol		9.0×10^{-4} mol		

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

$$[H^+] = \frac{0.25\times10^{-4}\,mol}{0.0670\,L} = 3.731\times10^{-4}\,M \quad \boxed{pH = 3.4}$$

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

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

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

The concentration of the salt at the equivalence point is:

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

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

b) 0.0444 M $Na^+C_3H_5O_3^-$; $C_3H_5O_3^- + H_2O \rightleftharpoons HC_3H_5O_3 + OH^-$

$$K_b = \frac{[HC_3H_5O_3][OH^-]}{[C_3H_5O_3^-]} = \frac{K_w}{K_a} = \frac{1.0\times10^{-14}}{1.4\times10^{-4}} = 7.14\times10^{-11}$$

$[HC_3H_5O_3] = [OH^-]$; $[C_3H_5O_3^-] \simeq 0.0444$

$[OH^-]^2 \simeq 0.0444(7.14\times10^{-11})$; $[OH^-] = 1.78\times10^{-6}$ pOH = 5.75 $\boxed{pH = 8.25}$

c) $CrO_4^{2-} + H_2O \rightleftharpoons HCrO_4^- + OH^-$

$$K_b = \frac{[HCrO_4^-][OH^-]}{[CrO_4^{2-}]} = \frac{K_w}{K_a} = 3.33\times10^{-8}$$

$[OH^-]^2 \simeq 0.0444(3.33\times10^{-8})$; $[OH^-] = 3.845\times10^{-5}$ $\boxed{pH = 9.58}$