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AP Chemistry

Name BRACKEN KEY
Period ____

Acid – Base Titration Problems Interactive Practice
Strong & Strong, Strong & Weak, Equiv Point

1. Calculate the pH of the resulting solution that is formed when 200. mL of 1.00 M NaOH is added to 150. mL of 1.00 M $HC_2H_3O_2$.

$HC_2H_3O_2 + OH^- \rightarrow C_2H_3O_2^- + H_2O$

$(1.00)(.150)$ $(1.00)(.200)$

.150 moles $HC_2H_3O_2$ | .200 moles OH^-

EXCESS STRONG!

.200
−.150
.050 moles leftover OH^-

$[OH^-] = \frac{.050}{.350} = .14\ M$

pOH = .85

pH = 13.15

2. Calculate the pH of the solution that is formed by adding 150. mL of 0.10 M NaOH to 200.0 mL of 0.10 M of HF.

$HF + OH^- \rightarrow H_2O + F^-$

$(.10)(.200)$ $(.10)(.150)$

.020 moles HF | .015 moles OH^-

.015 moles F^- produced

TOTAL VOLUME = .350 Liters

.020
−.015
.005 moles HF leftover

$HF + H_2O \rightleftharpoons H_3O^+ + F^-$

$\frac{.005}{.350} = .014\ M$ | 0 | $\frac{.015}{.350} = .043\ M$

$-x$ | $+x$ | $+x$

$.014 - x$ | x | $.043 + x$

$K_a = 6.8 \times 10^{-4} = \frac{[x][.043+x]}{[.014-x]}$ $\quad x = 2.2 \times 10^{-4}\ M = [H_3O^+]$

pH = 3.66

3. Calculate the pH at the **equivalence point** in a titration reaction involving 50.0 mL of 1.00 M HCl with 1.500 M NH_3.

$NH_3 + H^+ \longrightarrow NH_4^+$

$(1.00)(.050)$

.050 moles NH_3 = .050 moles H^+

.050 moles NH_4^+ produced

$.050 = (1.500)(V)$

V = .033 Liters NH_3 needed

TOTAL VOLUME = .083 Liters

	NH_4^+	+ H_2O ⇌	H_3O^+	+	NH_3
$\frac{.050}{.083}$ = .60 M			0		0
	−x		+x		+x
	.60 − x		x		x

$K_a = 5.6 \times 10^{-10} = \frac{x^2}{[.60 - x]}$

$x = 1.8 \times 10^{-5} M = [H_3O^+]$

pH = 4.74

4. What is the pH of the resulting solution formed by adding 350 mL of 0.100 M HCl to a beaker containing 220 mL of 0.20 M NaOH?

$H^+ + OH^- \longrightarrow H_2O$

$(.100)(.35)$ $(.20)(.22)$

.035 moles H^+

.044 moles OH^-

$.044 - .035 =$ .009 moles OH^- leftover

$\frac{.009}{.570} = .016\ M = [OH^-]$

pOH = 1.80

pH = 12.20

5. Calculate the pH at the **equivalence point** in a titration reaction involving 300.0 mL of 0.10 M HF with 0.20 M KOH.

$HF + OH^- \longrightarrow H_2O + F^-$

$(.10)(.300) = .030$ moles HF $= .030$ moles OH^-

$.030 = (.20)(V)$

$V = .15$ Liters KOH needed

$.030$ moles F^- produced

TOTAL VOLUME $= .45$ Liters

$F^- + H_2O \rightleftharpoons HF + OH^-$

$\frac{.030}{.45} = .067$ M		0	0
$-x$		$+x$	$+x$
$.067 - x$		x	x

$K_b = 1.5 \times 10^{-11} = \frac{x^2}{[.067 - x]}$

$x = 1.0 \times 10^{-6}$ M $= [OH^-]$

$pOH = 6.00$

$pH = 8.00$

6. What is the pH of the solution formed if 250. mL of 1.00 M HBr and 400.0 mL of 1.00 M NH_3 are mixed together?

$NH_3 + H^+ \longrightarrow NH_4^+$

$(1.00)(.400) = .400$ moles NH_3

$(1.00)(.250) = .250$ moles H^+

$.250$ moles NH_4^+ produced

$.400 - .250 = .150$ moles NH_3 Leftover

TOTAL VOLUME $= .650$ Liters

$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$

$\frac{.150}{.650} = .23$ M		$\frac{.250}{.650} = .38$ M	0
$-x$		$+x$	$+x$
$.23 - x$		$.38 + x$	x

$K_b = 1.8 \times 10^{-5} = \frac{[.38 + x][x]}{[.23 - x]}$

$x = [OH^-] = 1.1 \times 10^{-5}$ M

$pOH = 4.96$

$pH = 9.04$

7. What will be the pH at the **equivalence point** in a titration that involves 540 mL of 0.100 M HCN and 0.200 M KOH?

$HCN + OH^- \longrightarrow H_2O + CN^-$

$(.100)(.54)$

.054 moles HCN = .054 moles OH^-

.054 moles CN^- produced

$.054 = (.200)(V)$

V = .27 Liters KOH needed

TOTAL VOLUME = .81 Liters

$CN^- + H_2O \rightleftharpoons HCN + OH^-$

	CN^-		HCN	OH^-
	.054/.81 = .067 M		0	0
	−x		+x	+x
	.067 − x		x	x

$K_b = 2.0 \times 10^{-5} = \dfrac{x^2}{[.067 - x]}$

$x = [OH^-] = .0011\ M$

$pOH = 2.94$

pH = 11.06

8. What is the final pH of the solution if 300. mL of 1.20 M NaOH and 320. mL of 1.00 M $HC_2H_3O_2$ are mixed?

$HC_2H_3O_2 + OH^- \longrightarrow H_2O + C_2H_3O_2^-$

$(1.00)(.320)$ .320 moles $HC_2H_3O_2$

$(1.20)(.300)$ .360 moles OH^-

EXCESS STRONG!

TOTAL VOLUME = .620 Liters

.360
−.320
.040 moles OH^- Leftover

$[OH^-] = \dfrac{.040}{.620} = .065\ M$

$-\log[.065] = pOH = 1.19$

pH = 12.81