

1 of 4 Chem 103 Practice problems #3

Practice problems from week 3:

Dear students,

Here are some of the problems I presented to you in lecture. I've put the solutions at the very end so you can try the problems first and look at the solutions. Be aware that midterm #1 is coming up soon (next week, Friday) and so don't procrastinate.

(1) Suppose a solution contains 0.100 M A^- and 0.200 M HA, what is its pH? (given: $K_a = 1.8 \times 10^{-5}$)

(2) A solution contains 0.100 M HA and 0.100 M NaA. What is the pH? (given: $K_a = 1.8 \times 10^{-5}$)

(3) If 20.0 mL of 0.30 M HCl is added to 50.0 mL of a buffer containing 0.50 M acetic acid (HA, $K_a = 1.8 \times 10^{-5}$) and 0.50M sodium acetate.

(4) Suppose that we add 20.0 mLs of 0.150 M NaOH to 50.0 mL solution of 0.200 M nitrous acid (HNO_2 , $K_a = 4.5 \times 10^{-4}$), what is the resulting pH?

(5) Consider the titration of 25.0 mL of 0.100 M acetic acid (HA, $K_a = 1.8 \times 10^{-5}$) with 0.100M NaOH.

(a) what is $V_e = ?$ b) determine pH at:

i) region I, $V_{NaOH} = 0$

ii) region II, $V_{NaOH} = 20.0\text{mL}$

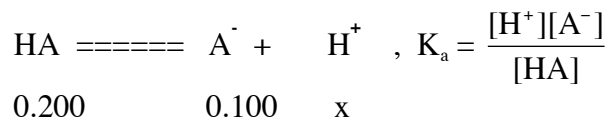
iii) region III, $V_{NaOH} = 25.0\text{ mL}$

iv) region IV, $V_{NaOH} = 30.0\text{ mL}$

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Solutions:

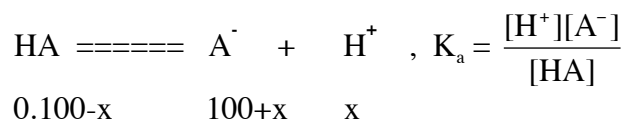
(1) Use the K_a equilibrium as your starting point.



$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{x(0.100)}{(0.200)} = 1.8 \times 10^{-5} \quad x = 3.6 \times 10^{-5} \text{ M} = [\text{H}^+]$$

$$\text{pH} = -\log(3.6 \times 10^{-5}) = 4.44$$

(2) Solution:



$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = \frac{x(0.100)}{(0.100)} = 1.8 \times 10^{-5} \quad (\text{note: } .100+x \approx .100 \text{ etc})$$

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

$$\text{pH} = -\log(1.8 \times 10^{-5}) = 4.74 \quad (\text{same as } \text{p}K_a)$$

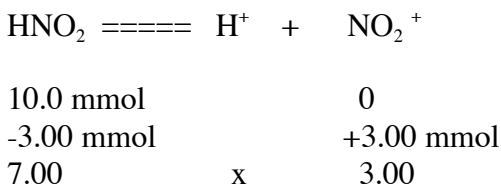
(3) Solution: Do ICE approach (here it is best to follow the moles)

	$\text{HA} \rightleftharpoons \text{A}^- + \text{H}^+ , K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$	
I	25mmol 25mmol 6mmol	
C	+ 6 mmol -6 mmol -6 mmol	$[\text{H}^+] = \frac{[\text{HA}][K_a]}{[\text{A}^-]}$
E	31 mmol+x 19 mmol+x x	$[\text{H}^+] = \frac{[31/V_t][1.8 \times 10^{-5}]}{[19/V_t]} = \frac{[31][1.8 \times 10^{-5}]}{[19]}$
	$[\text{H}^+] = 2.9_4 \times 10^{-5} \Rightarrow \text{pH} = 4.53$	

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(4) Solution: (note: here a weak acid solution becomes a buffer solution upon addition of strong base).

Start with the equilibrium equation:



$$\text{so } x = [\text{H}^+] = \frac{[\text{HNO}_2][K_a]}{[\text{NO}_2^-]} = \frac{3.00(4.5 \times 10^{-4})}{7.00}$$

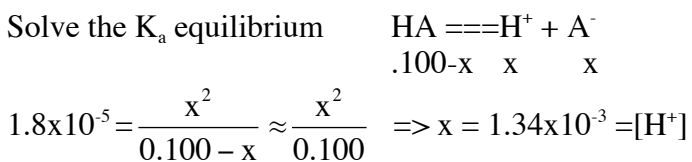
$$x = 1.93 \times 10^{-4} = [\text{H}^+] \Rightarrow \text{pH} = -\log(1.93 \times 10^{-4}) = 3.715$$

(5) (a) at equivalence: $M_{\text{HA}}^{\circ} V_{\text{HA}}^{\circ} = M_{\text{NaOH}}^{\circ} V_e$

$$V_e = \frac{0.100\text{M}(25.0\text{mL})}{0.100\text{M}} = 25.0 \text{ mL}$$

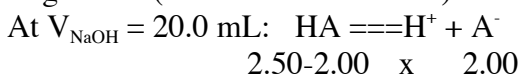
(b) pH calculations at various points:

Region I. No base has been added:



$$\text{pH} = -\log(1.34 \times 10^{-3}) = 2.87$$

Region II. (here follow the moles)



$$x = (0.50)(1.8 \times 10^{-5}) / 2.00 = 4.5 \times 10^{-6} = [\text{H}^+]$$

$$\text{pH} = -\log(4.5 \times 10^{-6}) = 5.347$$

region III (here use K_b equilibrium)

$V_{\text{NaOH}} = V_e$ and all HA is converted to A^-
Same as a pure solution of the weak base, A^- :

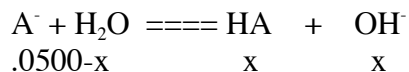
We need to things first:

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$$[A^-] = \frac{\text{moles } A^-}{\text{total volume}} = \frac{M_{HA}^{\circ} V_{HA}^{\circ}}{V_{HA}^{\circ} + V_e} = \frac{0.100M(25.0mL)}{(25.0 + 25.0)mL} = 0.0500 M A^-$$

$$\text{and, } K_b : K_b = \frac{1.00 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.55 \times 10^{-10}$$

K_b equilibrium:



$$\frac{x^2}{0.0500 - x} = 5.55 \times 10^{-10} \Rightarrow x = 5.27 \times 10^{-6} = [OH^-]$$

$$pOH = -\log(5.27 \times 10^{-6}) = 5.278 ;$$

$$pH = 14.00 - 5.278 = 8.72$$

Region IV: past the equivalence point
Treat it as a strong base solution:

$$pH = 14.00 - pOH = 14.00 - (-\log [NaOH]_{\text{excess}})$$

solve for the $[NaOH]_{\text{excess}}$:

$$[NaOH]_{\text{excess}} = \frac{\text{total mol NaOH} - \text{mol NaOH reacted}}{\text{total volume}}$$

$$= \frac{M_{NaOH}^{\circ} V_{NaOH} - M_{NaOH}^{\circ} V_e}{V_{HA}^{\circ} + V_{NaOH}} = \frac{(0.100M)(30.0mL) - (0.100M)(25.0mL)}{(25.0mL + 30.0mL)}$$

$$= 9.09 \times 10^{-3} M \Rightarrow pOH = 2.041 \Rightarrow pH = 14.00 - 2.041 = 11.96$$

(6)

(7)