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, $V_{NaOH} = 20.0$mL
iii) region III, $V_{NaOH} = 25.0$ mL
iv) region IV, $V_{NaOH} = 30.0$ mL

Solutions:

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

$$HA \;=\!=\!=\!=\!=\!=\; A^- + H^+ \quad , \; K_a = \frac{[H^+][A^-]}{[HA]}$$

0.200   0.100   x

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{x(0.100)}{(0.200)} = = 1.8x10^{-5} \qquad x = 3.6x10^{-5}\ M = [H^+]$$

$pH = -\log(3.6x10^{-5}) = 4.44$

(2) Solution:

$$HA \;=\!=\!=\!=\!=\!=\; A^- + H^+ \quad , \; K_a = \frac{[H^+][A^-]}{[HA]}$$

0.100-x   100+x   x

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{x(0.100)}{(0.100)} = 1.8x10^{-5} \quad \text{(note: .100+x≈.100 etc)}$$

$x = 1.8\ x10^{-5}\ M = [H^+] \ = K_a$

$pH = -\log(1.8x10^{-5}) = 4.74$ (same as pK_a)

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

$$HA \;=\!=\!=\!=\!=\!=\; A^- + H^+ \quad , \; K_a = \frac{[H^+][A^-]}{[HA]}$$

	HA	A^-	H^+
I	25mmol	25mmol	6mmol
C	+ 6 mmol	-6 mmol	-6 mmol
E	31 mmol +x	19 mmol+x	x

$$[H^+] = \frac{[HA][K_a]}{[A^-]}$$

$$[H^+] = \frac{[31/V_t][1.8x10^{-5}]}{[19/V_t]} = \frac{[31][1.8x10^{-5}]}{[19]}$$

$[H^+] = 2.9_4x10^{-5}$ => **pH = 4.53**

(50.0mL x 0.50 M HA)

(50.0mL x 0.50 M HA)

(20.0mL x 0.30M HCl)

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

Start with the equilibrium equation:

HNO_2 ===== H^+ + $NO_2^{\ +}$

10.0 mmol		0
-3.00 mmol		+3.00 mmol
7.00	x	3.00

$$\text{so } x = [H^+] = \frac{[HNO_2][K_a]}{[NO_2^-]} = \frac{3.00(4.5x10^{-4})}{7.00}$$

$x = 1.9_3x\ 10^{-4} = [H^+]$ => $pH = -\log(1.9_3x10^{-4}) = 3.715$

(5) (a) at equivalence: $M^o_{HA}V^o_{HA} = M^o_{NaOH}V_e$

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

(b) pH calculations at various points:

Region I. No base has been added:

Solve the K_a equilibrium HA ===H^+ + A^-

.100-x x x

$$1.8x10^{-5} = \frac{x^2}{0.100 - x} \approx \frac{x^2}{0.100} \quad \Rightarrow x = 1.34x10^{-3} = [H^+]$$

$pH = -\log(1.34x10^{-3}) = 2.87$

Region II. (here follow the moles)

At V_{NaOH} = 20.0 mL: HA ===H^+ + A^-

2.50-2.00 x 2.00

$x = (0.50)(1.8x10^{-5}) / 2.00 = 4.5x10^{-6} = [H^+]$

$pH = -\log(4.5x10^{-6}) = 5.347$

region III (here use K_b equilibrium)

$V_{NaOH} = V_e$ and all HA is converted to A^-

Same as a pure solution of the weak base, A^-:

We need to things first:

$$[A^-] = \frac{\text{moles } A^-}{\text{total volume}} = \frac{M^o_{HA}V^o_{HA}}{V^o_{HA} + V_e} =$$

$$\frac{0.100\text{M}(25.0)\text{mL}}{(25.0+25.0)\text{mL}} = 0.0500 \text{ M } A^-$$

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

K_b equilibrium:

$A^- + H_2O$ ==== HA + OH^-

.0500-x x x

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

pOH = -log(5.27x10^{-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 $[NaOH]_{excess}$)

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

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

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

= 9.09x10^{-3} M => pOH = 2.041 => pH = 14.00-2.041 = 11.96

(6)

(7)