

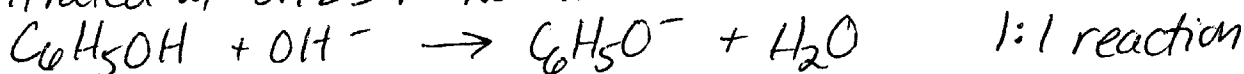
# Answers to "More Acid/Base, Buffer Problems"

Chem IB

$$1. (0.515 \text{ g } C_6H_5OH) \left( \frac{1 \text{ mol}}{94.108 \text{ g}} \right) = 5.472 \times 10^{-3} \text{ mol phenol}$$

$$\frac{5.472 \times 10^{-3} \text{ mol}}{.100 \text{ L}} = .05472 \text{ M phenol}$$

titrated w/ 0.123 M NaOH

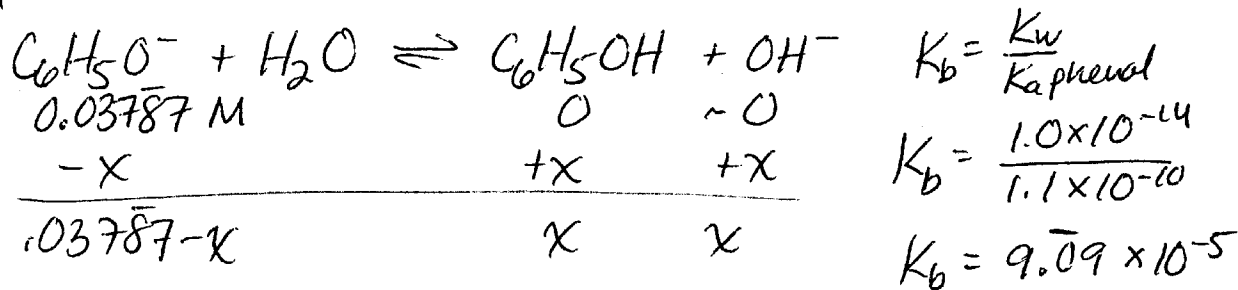


$$5.472 \times 10^{-3} \text{ mol phenol reacts with } 5.472 \times 10^{-3} \text{ mol NaOH} \\ (5.472 \times 10^{-3} \text{ mol NaOH}) \left( \frac{\text{L NaOH}}{.123 \text{ mol}} \right) = 0.04449 \text{ L NaOH} \\ \text{needed to reach the equivalence pt}$$

$$\text{@ equivalence point } V_{\text{total}} = 100. \text{ mL} + 44.5 \text{ mL} \\ = 144.5 \text{ mL}$$

@ equivalence point, all  $C_6H_5OH$  has been converted to  $C_6H_5O^-$ , a weak base. This will hydrolyze.

$$[C_6H_5O^-] = \frac{5.472 \times 10^{-3} \text{ mol}}{.1445 \text{ L}} \\ = 0.03787 \text{ M}$$



$$K_b = \frac{x^2}{.03787 - x}$$

$$\frac{[C_6H_5O^-]}{K_b} = 417 \text{ not ok to neglect } x.$$

neglect  $x$

$$x = \sqrt{(9.09 \times 10^{-5})(.03787)}$$

$$x = 1.86 \times 10^{-3}$$

2nd approx:

$$x = \sqrt{(9.09 \times 10^{-5})(.03787 - 1.86 \times 10^{-3})}$$

$$x = 1.81 \times 10^{-3}$$

#1 continued

3rd approx (just to be sure)  $x = \sqrt{(9.09 \times 10^{-5})(.03787 - 1.81 \times 10^{-3})}$  $x = 1.81 \times 10^{-3} \text{ M}$  represents  $[\text{OH}^-]$  and  $[\text{C}_6\text{H}_5\text{OH}]$  at equilibrium. $\text{pOH} = 2.74$      $\text{pH} = 11.26$      $[\text{H}_3\text{O}^+] = 5.5 \times 10^{-12} \text{ M}$ 

$$[\text{Na}^+] = \frac{\text{initial moles Na}^+}{\text{total volume of solution}} = \frac{5.472 \times 10^{-3} \text{ mol Na}^+}{.1495 \text{ L}}$$

$$= 0.0379 \text{ M Na}^+$$

$$[\text{C}_6\text{H}_5\text{O}^-] = .03787 - 1.81 \times 10^{-3} \text{ M} = 0.0361 \text{ M}$$

So:	$[\text{Na}^+] = 0.0379 \text{ M}$	$[\text{C}_6\text{H}_5\text{O}^-] = 0.0361 \text{ M}$
	$[\text{H}_3\text{O}^+] = 5.5 \times 10^{-12} \text{ M}$	$[\text{C}_6\text{H}_5\text{OH}] = 1.8 \times 10^{-3} \text{ M}$
	$[\text{OH}^-] = 1.8 \times 10^{-3} \text{ M}$	$\text{pH} = 11.26$

2. 0.150 M KOH + 0.230 M  $\text{H}_3\text{PO}_4$  (100 mL)need a buffer,  $\text{pH} = 2.50$  $\text{H}_3\text{PO}_4$   $K_a = 6.9 \times 10^{-3}$      $\text{p}K_a = 2.16$ if  $\text{pH} = 2.50$ ,  $[\text{H}_3\text{O}^+] = 3.16 \times 10^{-3} \text{ M}$ 

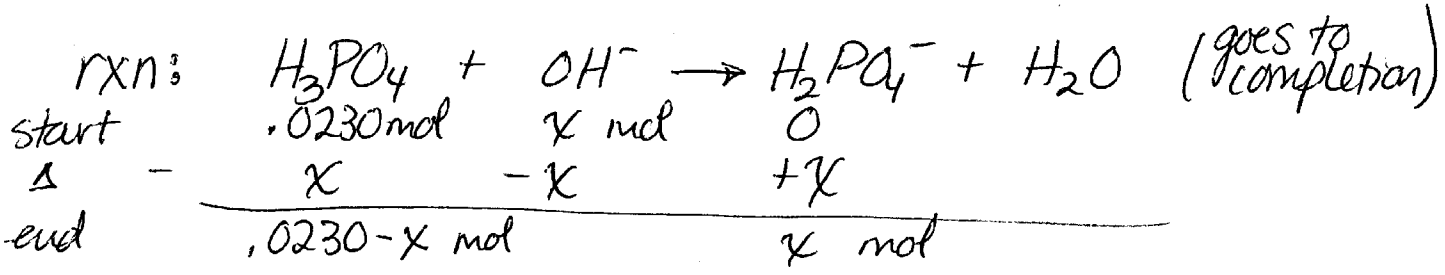
$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

$$\frac{K_a}{[\text{H}^+]} = \frac{6.9 \times 10^{-3}}{3.16 \times 10^{-3}} = 2.18 = \frac{[\text{A}^-]}{[\text{HA}]}$$

2.18 mol  $\text{H}_2\text{PO}_4^-$  : 1 mol  $\text{H}_3\text{PO}_4$  in the buffer.

$$\text{start with: } (.100 \text{ L}) \left( \frac{.230 \text{ mol H}_3\text{PO}_4}{1 \text{ L}} \right) = 0.0230 \text{ mol H}_3\text{PO}_4$$

#2 continued



@ end-buffer.  $\frac{\text{H}_2\text{PO}_4^-}{\text{H}_3\text{PO}_4} = \frac{2.18}{1} = \frac{x}{.0230 - x}$

$$2.18(.0230 - x) = x$$

$$.0502 - 2.18x = x$$

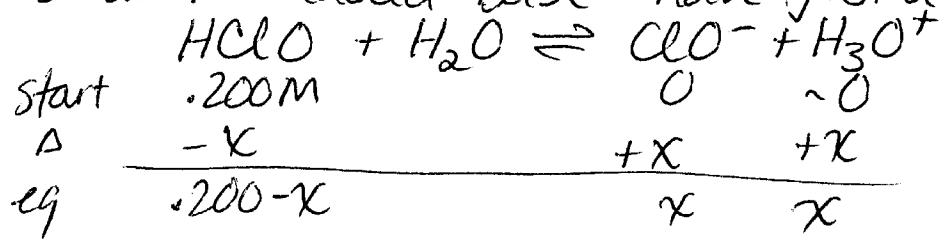
	+ 2.18x	+ 2.18x
<hr/>		
.0502		= 3.18x

x = 0.0158  
represents # moles OH<sup>-</sup> added

$(0.0158 \text{ mol OH}^-) \left( \frac{1 \text{ L}}{.150 \text{ mol KOH}} \right) = .105 \text{ L KOH to add}$

add 105 mL KOH  
(1.1 x 10<sup>2</sup> mL KOH)

3. a. No added base - have just a weak acid in water.



$K_a = 3.5 \times 10^{-8} = \frac{x^2}{.200 - x}$  ← neglect x

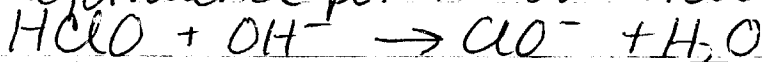
$x = \sqrt{.200(3.5 \times 10^{-8})} = 8.37 \times 10^{-5} \text{ M H}^+$

pH = 4.08

#3 continued

p. 4

3b. @ equivalence point: all HClO is converted to ClO<sup>-</sup>



ClO<sup>-</sup> is a weak base, and will hydrolyze  
we need to know its concentration in the  
final solution.

$$(.0500 \text{ L HClO}) \left( \frac{.200 \text{ mol HClO}}{1 \text{ L}} \right) = 0.0100 \text{ mol HClO}$$

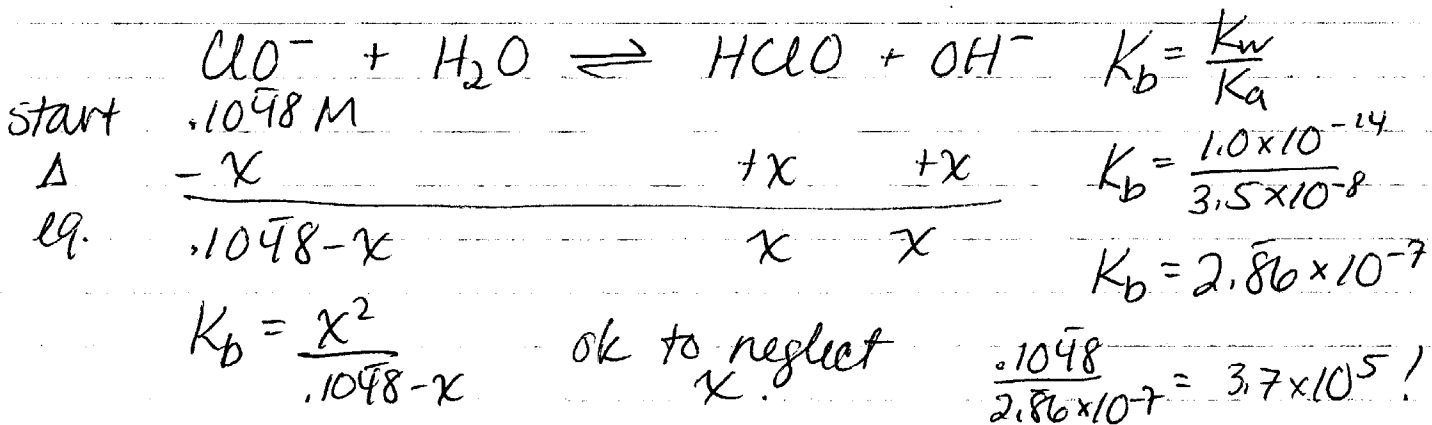
so, we need to add 0.0100 mol NaOH to reach  
the equivalence point. (the reaction is 1:1)

$$(0.0100 \text{ mol NaOH}) \left( \frac{1 \text{ L}}{.220 \text{ mol NaOH}} \right) = .04545 \text{ L NaOH}$$

$$\text{so } V_{\text{total}} = 50.0 \text{ mL} + 45.45 \text{ mL} \quad \text{to add}$$
$$= 95.45 \text{ mL}$$

During this reaction, 0.0100 moles of ClO<sup>-</sup> are  
formed (the rxn is 1:1 HClO:ClO<sup>-</sup>)

$$[\text{ClO}^-] = \frac{0.0100 \text{ mol}}{.09545 \text{ L}} = 0.1048 \text{ M ClO}^-$$



$$x = \sqrt{(2.86 \times 10^{-7})(.1048)} = 1.7 \times 10^{-4} \text{ M OH}^-$$

$$\text{pOH} = 3.76$$

$$\text{pH} = 10.24$$

#3 continued

p. 5

c. at the halfway point: this is when  $[HA] = [A^-]$

so  $K_a = \frac{[H_3O^+][A^-]}{[HA]}$

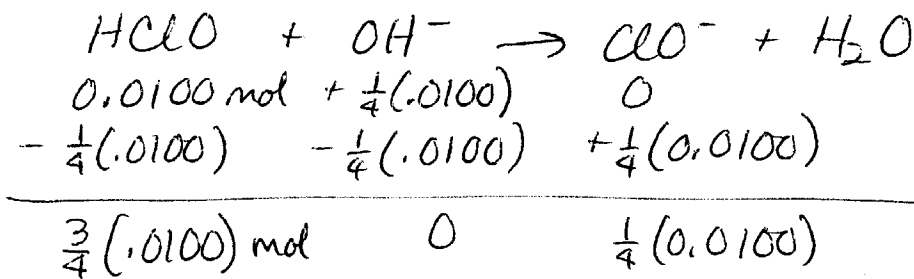
$K_a = [H_3O^+]$

and  $pH = pK_a$

$[H_3O^+] = 3.5 \times 10^{-8}$   
 $pH = 7.46$

d. at the 1/4 way point:

1/4 of the original HClO has been converted to ClO<sup>-</sup>. 3/4 of the HClO is left. Ratio is all that is important!



buffer.  $K_a = [H^+] \left( \frac{\text{moles } A^-}{\text{moles } HA} \right)$        $H^+ = K_a \left( \frac{\text{moles } HA}{\text{moles } A^-} \right)$

$[H^+] = 3.5 \times 10^{-8} \left( \frac{3/4}{1/4} \right) = 3.5 \times 10^{-8} \left( \frac{3}{1} \right) = 1.05 \times 10^{-7} M$   
 $H^+$

$pH = 6.98$

e. at 2/3 way point

2/3 of the original HClO has been converted to ClO<sup>-</sup>, and 1/3 of it is still left. It's a buffer.

$[H^+] = K_a \left( \frac{\text{moles } HA}{\text{moles } A^-} \right) = 3.5 \times 10^{-8} \left( \frac{1/3}{2/3} \right) = 3.5 \times 10^{-8} \left( \frac{1}{2} \right)$

$[H^+] = 1.75 \times 10^{-8}$

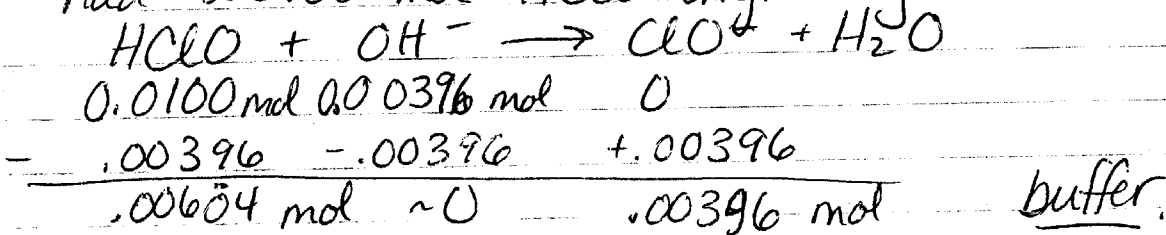
$pH = 7.76$

#3 continued

p. 6

f. after 18.0 mL .220 M NaOH has been added:  
 $(.0180 \text{ L}) \left( \frac{.220 \text{ mol}}{\text{L}} \right) = .00396 \text{ mol NaOH added.}$

had 0.0100 mol HClO originally.



$$K_a = [\text{H}^+] \left( \frac{\text{mol A}^-}{\text{mol HA}} \right) \quad [\text{H}^+] = K_a \left( \frac{\text{mol HA}}{\text{mol A}^-} \right)$$

$$[\text{H}^+] = 3.5 \times 10^{-8} \left( \frac{.00604 \text{ mol HClO}}{.00396 \text{ mol ClO}^-} \right) = 5.3 \times 10^{-8}$$

pH = 7.27

g. 10.0 mL past the endpoint:

have an extra  $(0.0100 \text{ L}) \left( \frac{.220 \text{ mol}}{\text{L}} \right) = .0022 \text{ mol OH}^-$

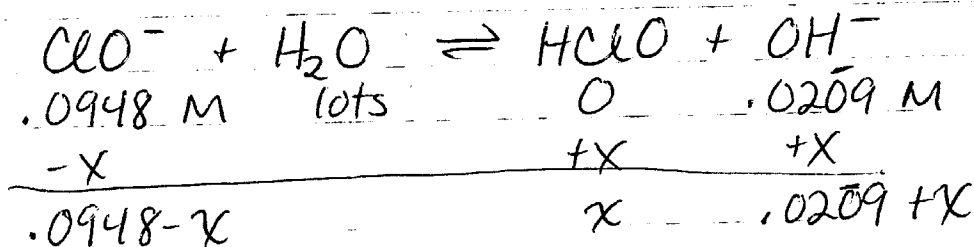
present in the solution. Total volume is

$$95.45 \text{ mL} + 10.0 \text{ mL} = 105.45 \text{ mL}$$

Also have 0.0100 mol ClO<sup>-</sup> present (see part b)

Type: weak base in the presence of a common ion.

$$[\text{ClO}^-] = \frac{.0100 \text{ mol}}{.10545 \text{ L}} = .0948 \text{ M} \quad [\text{OH}^-] = \frac{.0022 \text{ mol}}{.10545 \text{ L}} = .0209 \text{ M}$$



$$K_b = \frac{(x)(.0209 + x)}{(.0948 - x)} = 2.86 \times 10^{-7} \quad \text{neglect } x\text{'s.}$$

#3 continued

p. 7

g.  $x = 2.86 \times 10^{-7} \left( \frac{.0948}{.0209} \right) = 1.3 \times 10^{-6} \text{ M}$   
represents  $[\text{HClO}]$  at equilibrium.

$x$  really is negligible compared to  $[\text{ClO}^-]$  and  $[\text{OH}^-]$   
pH depends on  $[\text{OH}^-]$ .  $[\text{OH}^-] = .0209 \text{ M}$

$\text{pOH} = 1.68$

$\text{pH} = 12.32$

4. Good choice of indicator: must change color close to the pH of the equivalence point.

From part b, pH @ eq. pt = 10.24.

Hmm. Tough choice. It looks like phenolphthalein changes color too soon (pH 8-10) and alizarin yellow changes color too late, (pH 10-12)

Probably alizarin yellow would be better than phenolphthalein. Maybe universal indicator would work.

5. Have 80.0 mL 0.200 M HCl  $(.0800 \text{ L})(.200 \frac{\text{mol}}{\text{L}}) = .0160 \text{ mol HCl}$

want a buffer, pH = 7.75

mixing  $\text{Na}_3\text{PO}_4 + \text{HCl}$

looking @  $\text{H}_3\text{PO}_4$  ionizations,  $K_{a2}$  is best:

$\text{p}K_{a2} = 7.21$ , which is within 1 pH unit of the desired pH.  $K_{a2}$  corresponds to

the equilibrium  $\text{H}_2\text{PO}_4^- + \text{H}_2\text{O} \rightleftharpoons \text{HPO}_4^{2-} + \text{H}_3\text{O}^+$

so the conjugate acid will be  $\text{H}_2\text{PO}_4^-$  and

the conjugate base will be  $\text{HPO}_4^{2-}$ .

In the buffer,  $[\text{H}_3\text{O}^+] = 1.78 \times 10^{-8} \text{ M}$  (from desired pH). We can use this and the  $K_a$

to get the ratio of acid:base in the buffer.

#5

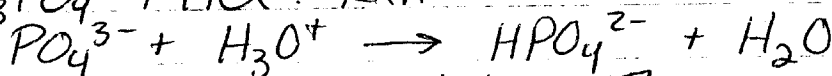
p.8

$$\frac{K_a}{[H^+]} = \frac{[A^-]}{[HA]} = \frac{6.2 \times 10^{-8}}{1.78 \times 10^{-8}} = \frac{3.49}{1} \frac{A^-}{HA}$$

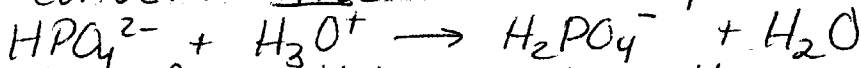
In the buffer, the mole ratio will be 3.49 mol  $HPO_4^{2-}$  to 1 mol HA.

Now - to get that buffer: we have

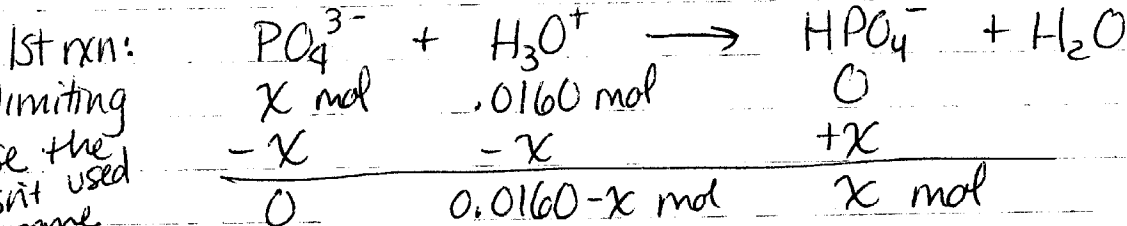
$Na_3PO_4 + HCl$  Rxn:



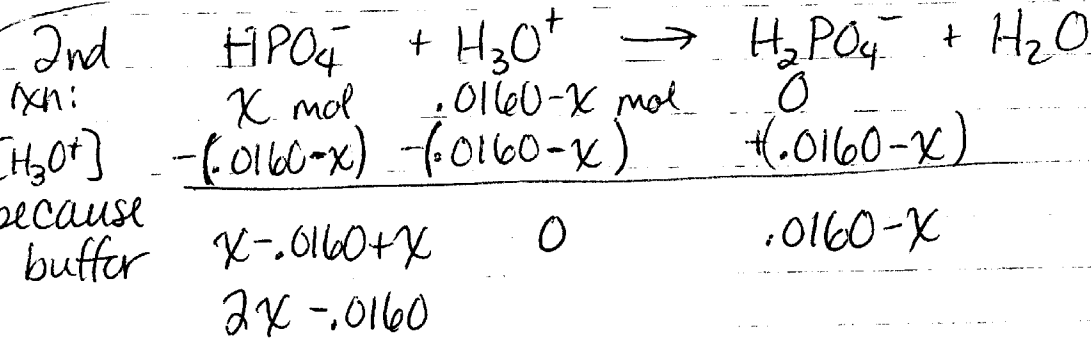
this will go to completion. Then we want to convert some of the  $HPO_4^{2-}$  to  $H_2PO_4^-$ :



and after this reaction, the ratio of  $HPO_4^{2-} : H_2PO_4^-$  is 3.49 : 1.



$(PO_4^{3-})$  is limiting because the  $H_3O^+$  isn't used up yet - some is needed for rxn 2.



this time,  $[H_3O^+]$  is limiting because we want a buffer @ the end.

$$\text{now } \frac{\text{moles } HPO_4^{2-}}{\text{moles } H_2PO_4^-} = \frac{3.49}{1} = \frac{2x - .0160}{0.0160 - x}$$

$$\begin{array}{r} 3.49(.0160 - x) = 2x - .0160 \\ 0.0558 - 3.49x = 2x - .0160 \\ +.0160 \quad +3.49x \quad +3.49x \quad +.0160 \\ \hline .0718 \qquad \qquad \qquad = 5.49x \end{array}$$

$$x = \frac{.0718}{5.49} = .0131$$



#5 continued

$x$  represents the initial # moles  $\text{PO}_4^{3-}$  present. p. 9

$x = .0131$  moles  $\text{PO}_4^{3-}$  what mass  $\text{Na}_3\text{PO}_4$ ?

$$(.0131 \text{ mol Na}_3\text{PO}_4) \left( \frac{163.94 \text{ g}}{1 \text{ mol}} \right) = 2.1 \text{ g Na}_3\text{PO}_4$$

add this to HCl  
to get buffer, pH = 7.75

6. a.  $\text{HNO}_2$ ,  $\text{NO}_2^-$  pH = 3.00  $[\text{H}^+] = 1.0 \times 10^{-3} \text{ M}$   
 $\text{HNO}_2$   $K_a = 4.5 \times 10^{-4}$

$$K_a = \frac{[\text{H}^+][\text{NO}_2^-]}{[\text{HNO}_2]} \quad \frac{K_a}{[\text{H}^+]} = \frac{[\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{4.5 \times 10^{-4}}{1.0 \times 10^{-3}} = 0.45$$

ratio is 0.45 mol  $\text{NO}_2^-$  : 1 mol  $\text{HNO}_2$   
so  $\text{HNO}_2$  ~~is~~ has the greater concentration  
in the buffer.

Another way to tell: compare pH and pKa  
pH = 3.00, pKa = 3.35 If the # moles HA = # moles A<sup>-</sup>,  
the pH would be 3.35. It's lower than that,  
so that must mean there's a higher proportion  
of weak acid in this buffer.

6 b.  $\text{HNO}_2$  0.400 M, 200. mL  
 $(.200 \text{ L}) \left( \frac{.400 \text{ mol}}{\text{L}} \right) = .0800 \text{ mol HNO}_2$  in buffer.

and ratio is  $\frac{0.45 \text{ NO}_2^-}{1 \text{ HNO}_2}$  so find moles  $\text{NO}_2^-$   
in buffer.

$$\frac{.45 \text{ NO}_2^-}{1 \text{ HNO}_2} = \frac{x}{.0800 \text{ mol HNO}_2}$$

$x = .036 \text{ mol}$   
 $\text{NO}_2^-$  in 200. mL  
buffer.

#6 continued

p. 10

So: at pH = 3.00, there are .0800 mol  $\text{HNO}_2$  and 0.036 mol  $\text{NO}_2^-$ . What would we add to get pH = 3.50? It's getting more basic, so we would need to add  $\text{NO}_2^-$ . New ratio needed:

$$\text{pH} = 3.50 \quad (\text{H}^+) = 3.16 \times 10^{-4} \text{ M}$$
$$\frac{K_a}{[\text{H}^+]} = \frac{[\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{4.5 \times 10^{-4}}{3.16 \times 10^{-4}} = 1.42 = \frac{1.42 \text{ mol NO}_2^-}{1 \text{ mol HNO}_2}$$

So... we're adding  $\text{NO}_2^-$  to the original buffer to get this new ratio.  $x$  = amt of  $\text{NO}_2^-$  to add (in moles).

$$\frac{1.42 \text{ mol NO}_2^-}{1 \text{ mol HNO}_2} = \frac{.036 + x}{.0800}$$

$$(.0800)(1.42) = .036 + x$$

$$.114 = .036 + x$$

$$-.036 \quad -.036$$

$$.0779 = x$$

one sig fig?!

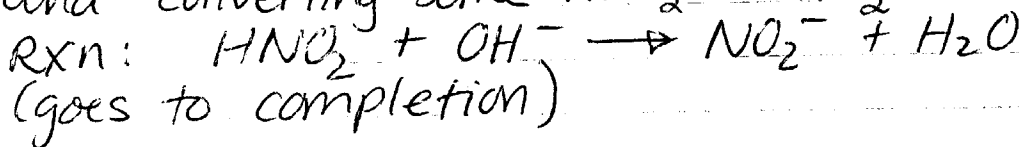
$$x = .078 \text{ mol NO}_2^- \text{ to add.}$$

If it's in the form of  $\text{NaNO}_2$  (MM = 69.00 g/mol)

$$(.078 \text{ mol}) \left( \frac{69.00 \text{ g}}{1 \text{ mol}} \right) = 5.4 \text{ g}$$

add 5.4 g  $\text{NaNO}_2$  to buffer.

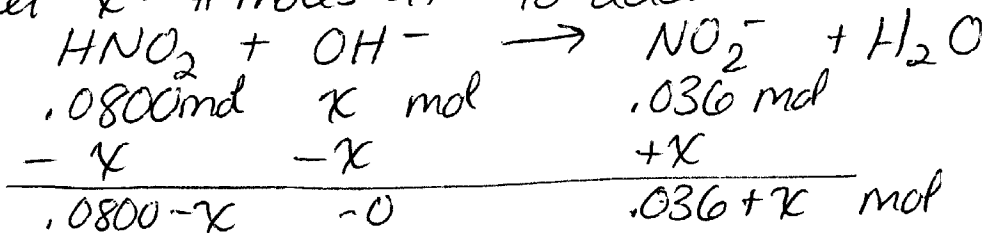
C. OK - In this case, we still have 0.0800 mol  $\text{HNO}_2$  and 0.036 mol  $\text{NO}_2^-$  in the buffer. We still want a pH of 3.50, so our desired mole ratio is still 1.42  $\text{NO}_2^-$  : 1  $\text{HNO}_2$ . But this time, we need to add  $\text{NaOH}$  (because it's getting more basic) and it will be converting some  $\text{HNO}_2$  to  $\text{NO}_2^-$ .



#6 continued

p. 11

c. let  $x = \# \text{ moles OH}^-$  to add.



now substitute into ratio, solve for  $x$ .

$$\frac{1.42 \text{ mol NO}_2^-}{1 \text{ mol HNO}_2} = \frac{.036 + x}{.0800 - x}$$

$$1.42(.0800 - x) = .036 + x$$

$$.1136 - 1.42x = .036 + x$$

$$\begin{array}{r}
 -.036 + 1.42x \\
 \hline
 .0776
 \end{array}
 = \begin{array}{r}
 -.036 + 1.42x \\
 \hline
 2.42x
 \end{array}$$

$$x = \frac{.0776}{2.42} = .0321 \text{ mol OH}^-$$

to add.

$$(.0321 \text{ mol OH}^-) \left( \frac{1 \text{ L NaOH}}{1.50 \text{ mol NaOH}} \right) = .0214 \text{ L}$$

add 21.4 mL NaOH to the buffer (-20 mL)

d. Again, were starting with the  $\text{pH} = 3.00$  buffer, so there are  $0.0800 \text{ mol HNO}_2$  and  $0.036 \text{ mol NO}_2^-$  present initially. This time, we want  $\text{pH} = 2.75$ , so it's getting more acidic. We need to add  $\text{HNO}_2$ .

First, find ratio needed in this new buffer.

$$[\text{H}^+] = 1.78 \times 10^{-3} \text{ M if } \text{pH} = 2.75$$

$$\frac{K_a}{[\text{H}^+]} = \frac{[\text{NO}_2^-]}{[\text{HNO}_2]} = \frac{4.5 \times 10^{-4}}{1.78 \times 10^{-3}} = .253 = \frac{.253 \text{ mol NO}_2^-}{1 \text{ mol HNO}_2}$$

let  $x = \text{moles of HNO}_2$  to add

#6 continued

p.12

$$d. \frac{.253 \text{ NO}_2^-}{1 \text{ HNO}_2} = \frac{.036}{.0800 + x}$$

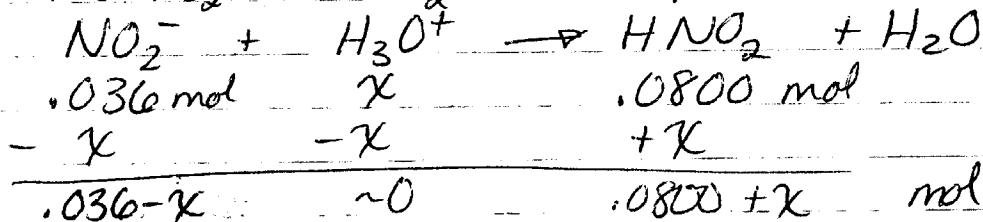
$$\begin{aligned} .253(.0800 + x) &= .036 \\ .0202 + .253x &= .036 \\ - .0202 & \quad - .0202 \end{aligned}$$

$$.253x = .0158$$

$$x = \frac{.0158}{.253} = .0623 \text{ moles HNO}_2 \text{ to add}$$

$$(.0623 \text{ mol HNO}_2) \left( \frac{47.02 \text{ g}}{1 \text{ mol}} \right) = \text{add } 2.9 \text{ g HNO}_2 \text{ to buffer.}$$

e. Start: 0.0800 mol HNO<sub>2</sub>, 0.036 mol NO<sub>2</sub><sup>-</sup>.  
Want it to get more acidic than pH=3.00,  
so add HCl. When HCl is added, it will convert  
some NO<sub>2</sub><sup>-</sup> to HNO<sub>2</sub>. Let x = # moles HCl to add.



now, plug into ratio needed (from part d)

$$\frac{.253 \text{ mol NO}_2^-}{1 \text{ mol HNO}_2} = \frac{.036 - x}{.0800 + x}$$

$$\begin{aligned} .253(.0800 + x) &= .036 - x \\ .0202 + .253x &= .036 - x \\ - .0202 & \quad + x \quad - .0202 + x \end{aligned}$$

$$1.253x = .0158$$

$$x = \frac{.0158}{1.253} = 0.0126 \text{ mol H}^+ \text{ to add.}$$

#6 continued

$$(.0126 \text{ mol HCl}) \left( \frac{1 \text{ L}}{2.00 \text{ mol HCl}} \right) = .00629 \text{ L to add}$$

p. 13

add 6.3 mL HCl to the buffer

7. 1st: choose the buffer system.  $pK_a$  must be close to the desired pH.

acids:  $\text{HC}_2\text{H}_3\text{O}_2$   $pK_a = 4.74$

$\text{HClO}$   $pK_a = 7.46$

$\text{H}_3\text{BO}_3$   $pK_a = 9.23$

← this is closest to the desired pH of 9.00.

The system is:  $\text{H}_3\text{BO}_3$  and  $\text{H}_2\text{BO}_3^-$ .

Now, find ratio needed. If  $\text{pH} = 9.00$ ,  $[\text{H}^+] = 1.0 \times 10^{-9} \text{ M}$

$$\frac{K_a}{[\text{H}^+]} = \frac{[\text{A}^-]}{[\text{HA}]} = \frac{5.9 \times 10^{-10}}{1.0 \times 10^{-9}} = 0.59 = \frac{0.59 \text{ mol } \text{H}_2\text{BO}_3^-}{1 \text{ mol } \text{H}_3\text{BO}_3}$$

also: want the most concentrated component to be 0.500 M. According to the above ratio, the most concentrated component is the  $\text{H}_3\text{BO}_3$ .

So  $[\text{H}_3\text{BO}_3] = .500 \text{ M}$  in the buffer.

$(.500 \text{ L}) \left( \frac{.500 \text{ mol}}{\text{L}} \right) = .25 \text{ mol } \text{H}_3\text{BO}_3$  needed in the buffer.

find # moles  $\text{H}_2\text{BO}_3^-$  needed:

$$\frac{0.59 \text{ mol } \text{H}_2\text{BO}_3^-}{1 \text{ mol } \text{H}_3\text{BO}_3} = \frac{x \text{ mol } \text{H}_2\text{BO}_3^-}{.25 \text{ mol } \text{H}_3\text{BO}_3}$$

$$x = (.25)(.59) = .1475 \text{ mol } \text{H}_2\text{BO}_3^- \text{ needed.}$$

1st way of making buffer: mix 0.25 mol  $\text{H}_3\text{BO}_3$  and .1475 mol  $\text{H}_2\text{BO}_3^-$ . This corresponds to:

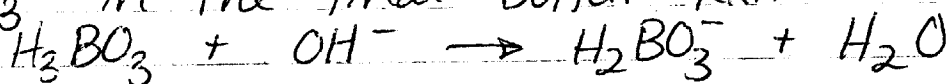
#7 continued

$$(.25 \text{ mol } H_3BO_3) \left( \frac{61.83 \text{ g } H_3BO_3}{1 \text{ mol}} \right) = 15.5 \text{ g } H_3BO_3 \quad (15.46 \text{ g}) \quad p.14$$

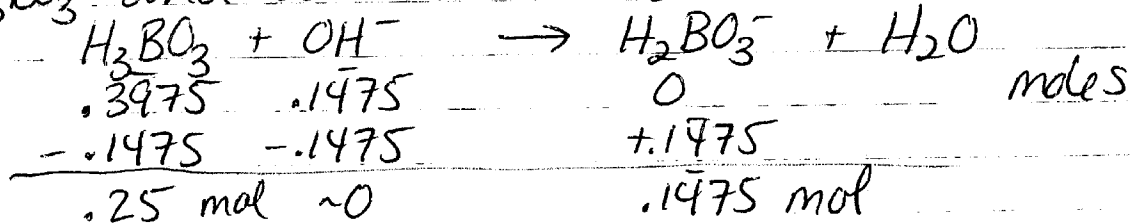
$$(.1475 \text{ mol } NaH_2BO_3) \left( \frac{83.81 \text{ g}}{1 \text{ mol}} \right) = 12.4 \text{ g } NaH_2BO_3$$

Buffer ① Mix 15.5 g  $H_3BO_3$  + 12.4 g  $NaH_2BO_3$ .  
add water until  $V_{total} = 500. \text{ mL}$ , and  
mix well!

2nd option: add  $NaOH$  to  $H_3BO_3$  until  
the ratio of  $H_2BO_3^- : H_3BO_3$  is  $0.59:1$ . We  
still have .25 moles  $H_3BO_3$  and .1475 moles  
 $H_2BO_3^-$  in the final buffer. Rxn:



any  $NaOH$  added converts  $H_3BO_3$  to  $H_2BO_3^-$ . All the  
 $H_2BO_3^-$  that is formed comes from  $H_3BO_3$ . So we  
have to start with  $0.25 + .1475 = .3975$  moles  
 $H_3BO_3$  and add .1475 moles  $NaOH$  to it.



$$\text{so: } (.3975 \text{ mol } H_3BO_3) \left( \frac{61.83 \text{ g } H_3BO_3}{1 \text{ mol}} \right) = 24.6 \text{ g } H_3BO_3 \text{ needed}$$

$$(.1475 \text{ mol } NaOH) \left( \frac{1 \text{ L}}{2.5 \text{ mol}} \right) = .059 \text{ L } NaOH \text{ needed}$$

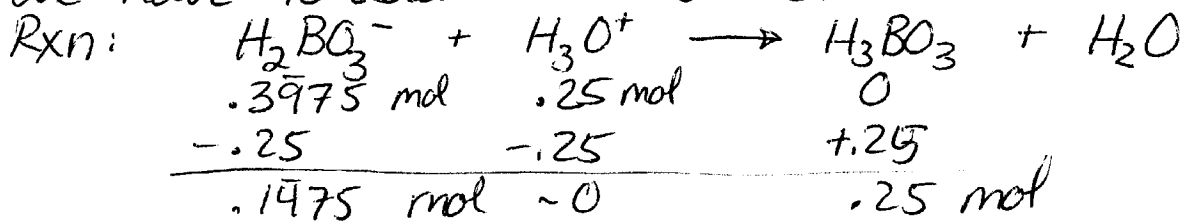
Buffer ② To make the buffer: mix 24.6 g  $H_3BO_3$  and  
59 mL  $NaOH$ , dilute to a total volume of  
500. mL.

#7 continued

p.15

3rd option: add HCl to  $\text{NaH}_2\text{BO}_3$  until the ratio of  $\text{A}^-/\text{HA}$  is 0.59/1.

Again, we need 0.25 mol  $\text{H}_3\text{BO}_3$  and 0.1475 mol  $\text{NaH}_2\text{BO}_3$  in the final buffer. Any HCl added converts  $\text{H}_2\text{BO}_3^-$  to  $\text{H}_3\text{BO}_3$ , so we have to start with  $0.25 + 0.1475 \text{ mol} = 0.3975 \text{ mol NaH}_2\text{BO}_3$ . We have to add 0.25 mol HCl.



$$.3975 \text{ mol NaH}_2\text{BO}_3 \left( \frac{83.816 \text{ g}}{1 \text{ mol}} \right) = 33.3 \text{ g NaH}_2\text{BO}_3$$

$$(.25 \text{ mol HCl}) \left( \frac{1 \text{ L}}{2.5 \text{ mol}} \right) = 0.10 \text{ L HCl}$$

Buffer  
3

To make the buffer: mix 33.3 g  $\text{NaH}_2\text{BO}_3$  and 100 mL HCl. Dilute to a total volume of 500. mL.