Example: Acetic Acid - $K_a = 1.8 \times 10^{-5}$

(a) Calculate the pH and percent dissociation
of 0.01 M Acetic Acid - Result for Part (b) is on the slide

We will calculate both the exact and approximate solutions

$$ax^2 + bx + c = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$c_0 = 0.0M$

$$HAc \rightleftharpoons H^+ + Ac^-$$
$$c_0 - x \qquad 0.0 + x \quad 0 + x$$

Exact:
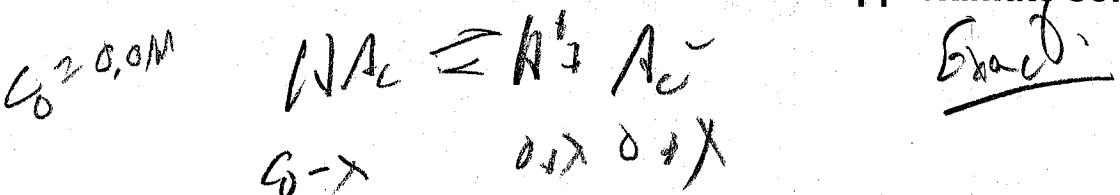
$$K_a = 1.8 \times 10^{-5} = \frac{x \cdot x}{c_0 - x} \rightarrow x^2 = c_0 K_a - K_a x$$

$$x^2 + K_a x - c_0 K_a = 0 \rightarrow [H^+] = 4.15 \times 10^{-4}$$

$$pH = 3.38$$

Approx

$$K_a = 1.8 \times 10^{-5} = \frac{x \cdot x}{c_0 - x} \approx \frac{x^2}{c_0}$$

$$x^2 = c_0 K_a = (0.01)(1.8 \times 10^{-5}) = 1.8 \times 10^{-7}$$

$$x = \sqrt{1.8 \times 10^{-7}} = 4.24 \times 10^{-4}$$

$$pH = -\log(4.24 \times 10^{-4})$$

$$= 3.37$$

% Diss on next pg

$$HAc \rightleftharpoons H^+_{0,Ac^-}$$
$$\qquad\quad x \quad x$$
$$C_0$$

$$x = 4.24 \times 10^{-4} M$$

$$\% \text{ Dissoc} = \frac{[A^-]}{[HA_0]_0} \times 100 = \frac{4.24 \times 10^{-4} M}{0.01 M} \times 100$$

$$= 4.24 \, g = 4.2 \, \%$$

Example: Pyridine - K_b = 1.8x10^{-9}

Pyr + H_2O ⇌ PyrH$^+$ + OH$^-$ $K_b = \dfrac{[PyrH^+][OH^-]}{[Pyr]}$
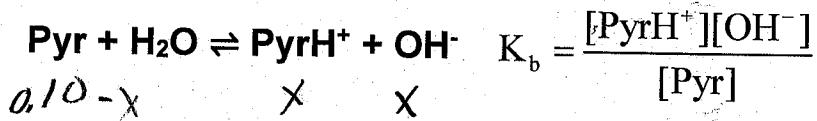
$0.10 - x$ x x

Calculate the pH and the percent of pyridine molecules which are protonated in a 0.10 M aqueous pyridine solution.

$$K_b = 1.8 \times 10^{-9} = \frac{[PyrH^+][OH^-]}{[Pyr]} = \frac{x \cdot x}{0.10 - x} \approx \frac{x^2}{0.1}$$

$$x^2 = 1.8 \times 10^{-9}(0.1) = 1.8 \times 10^{-10}$$

$$x = \sqrt{1.8 \times 10^{-10}} = 1.34 \times 10^{-5} = [OH^-] = [PyrH^+]$$

$$pOH = -\log(1.34 \times 10^{-5}) = 4.13$$

$$pH = 14 - 4.13 = 9.87$$

$$\% \ Prot. \ \frac{[PyrH^+]}{[Pyr]_0} \times 100 = \frac{1.34 \times 10^{-5}}{0.1} \times 100$$

$$= 0.0134\% \approx 0.013\%$$

Summary: Acid and Base Calculations
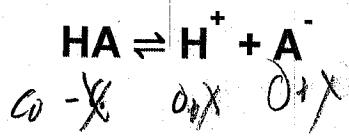
$$HA \rightleftharpoons H^+ + A^-$$
$$c_0 - x \quad 0 + x \quad 0 + x$$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{x \cdot x}{c_0 - x} \approx \frac{x^2}{c_0}$$

$$x = [H^+][A^-]$$

$$pH = -\log[H^+] = -\log(x)$$

$$\% \text{ Diss.} = \frac{[A^-]}{c_0} \times 100$$

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$
$$c_0 - x \qquad x \qquad x$$
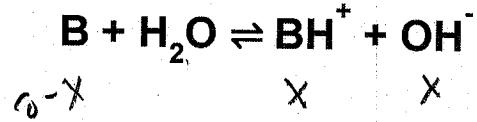
$$K_b = \frac{[BH^+][OH^-]}{[B]} = \frac{x \cdot x}{c_0 - x} = \frac{x^2}{c_0}$$

$$x = [BH^+] = [OH^-]$$

$$pOH = -\log(x) = -\log[OH^-]$$

$$pH = 14 - pOH$$

$$\% \text{ Prot.} = \frac{[BH^+]}{c_0} \times 100$$

Example: $K_a(HCN) = 4.9 \times 10^{-10}$
Calculate the pH of a 0.05 M solution of NaCN.

$$NaCN \rightarrow \cancel{Na^0} + CN^-$$

$$\underset{C_0}{CN^-} + H_2O \rightleftharpoons \underset{x}{HCN} + \underset{x}{OH^-}$$

similar to

$$[\; Ac^- + H_2O \rightleftharpoons HAc + OH^-]$$

$$[\; \text{similar to } NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-]$$

$$K_b = \frac{[HCN][OH^-]}{[CN^-]}$$
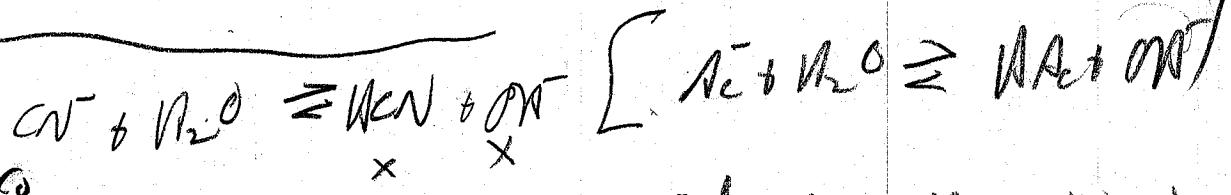
$$= \frac{x \cdot x}{C_0 - x} \approx \frac{x^2}{C_0}$$
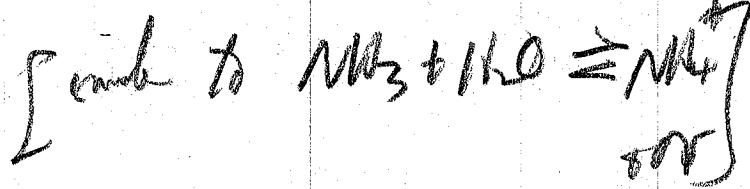
$$2 \times 10^{-5} = \frac{x^2}{0.05\,M}$$

$$K_b = \frac{10^{-14}}{K_a} = \frac{10^{-14}}{4.9 \times 10^{-10}}$$

$$= 2.0 \times 10^{-5}$$

$$x^2 = 2 \times 10^{-5}(0.05)$$

$$x^2 = 1 \times 10^{-6}$$

$$x = \sqrt{1 \times 10^{-6}} = 1.0 \times 10^{-3}$$

$$pOH = -\log(1.0 \times 10^{-3}) = 3.0$$

$$pH = 14 - 3.0 = 11.0$$

Example: $K_b(NH_3) = 1.8 \times 10^{-5}$

      Calculate the pH of a 0.20 M solution of NH₄Cl.

$$NH_4Cl \rightarrow NH_4^+ + Cl^-$$

$$NH_4^+ \rightleftharpoons NH_3 + H^+ \qquad \left[\text{like } HAc \rightleftharpoons H^+ + Ac^- \right]$$

$$c_0 - x \qquad x \qquad x$$
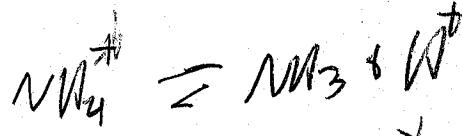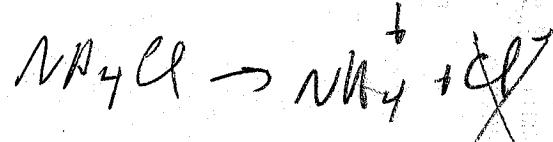
$$K_a = \frac{[NH_3][H^+]}{[NH_4^+]} = \frac{x \cdot x}{c_0 - x} = \frac{x^2}{c_0}$$

$$K_a = \frac{1 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$5.6 \times 10^{-10} = \frac{x^2}{0.2}$$

$$x^2 = 1.12 \times 10^{-10}$$

$$x = \sqrt{1.12 \times 10^{-10}} = 1.06 \times 10^{-5}$$

$$pH = -\log(1.06 \times 10^{-5}) = 4.98 = 5.0$$

pH Change of HAc/Ac⁻ Buffer after addition of H⁺ or OH⁻
Slide #35

Example: HAc - $K_a = 1.8 \times 10^{-5}$

$$pK_a = -\log(1.8 \times 10^{-5}) = 4.74$$

Consider 1 Liter of a solution containing
0.50 M HAc and 0.50 M NaAc. → Na^+ & Ac^- (0.5 M)

(a) Calculate the pH and $[H^+]$ of the above solution.

(b) Calculate the pH of the above solution after the addition of 0.01 mol of HCl.

(c) Calculate the pH of the above solution after the addition of 0.01 mol of NaOH.

ⓐ

$$[HAc] = 0.5 M$$
$$[Ac^-] = 0.5 M$$

$$pH = pK_a + \log \frac{[Ac^-]}{[HAc]} = 4.74 + \log\left(\frac{0.5}{0.5}\right)$$
$$= 4.74$$

$$[H^+] = 10^{-4.74} = 1.8 \times 10^{-5}$$

ⓑ int: $n_{HAc} = 0.5 \frac{mol}{L} \times 1L = 0.5 \, mol$

$$n_{Ac^-} = 0.5 \frac{mol}{L} \times 1L = 0.5 \, mol$$
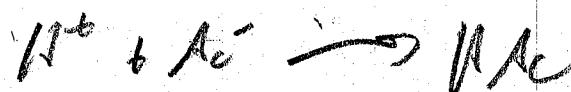
$$H^+ + Ac^- \longrightarrow HAc$$

|	H^+	Ac^-	HAc
int	0.01	0.50	0.50
chg	-0.01	-0.01	+0.01
fin	~0	0.49	0.51

$$[Ac^-] = 0.49 \quad [HAc] = 0.51$$

$$pH = pK_a + \log \frac{[Ac^-]}{[HAc]} = 4.74 + \log\left(\frac{0.49}{0.51}\right)$$

$$= 4.74 - 0.02 = 4.72$$

(c) imol

Sin 35 Sx. (Confirm)

$n_{HAc} = 0.5$ mol. $n_{Ac^-} = 0.02$

$$HAc + OH^- \longrightarrow Ac^- + H_2O$$

	0.5	0.01	0.5
inol			$+0.01$ mol
chg.	-0.01	-0.01	0.51 mol.
	0.49 mol.		
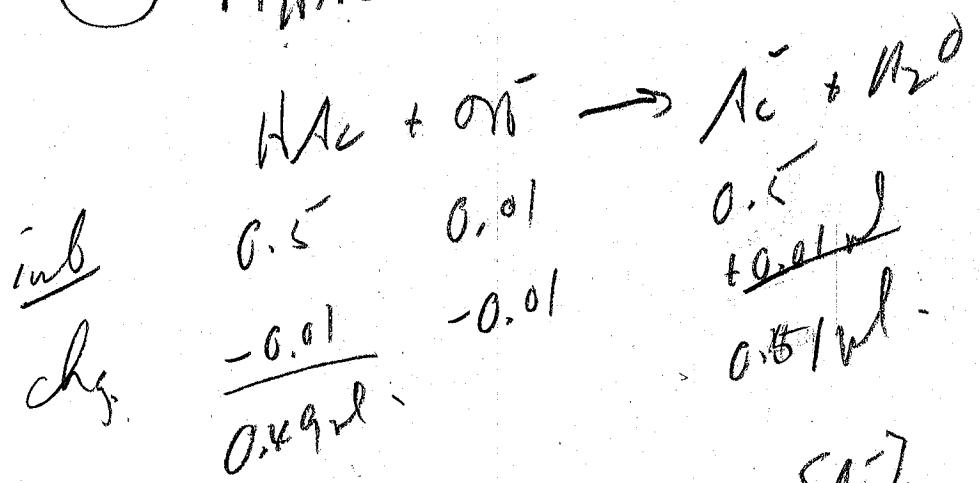
$$pH = pKa + \log\frac{[Ac^-]}{[HAc]} = 4.74 + \log\left(\frac{0.51}{0.49}\right)$$

$$= 4.74 + 0.02 = 4.76$$

NH₄⁺/NH₃ Buffer Example
Slide #40

Example: NH_3 - K_b = 1.8×10^{-5}

$pK_b = -\log(1.8 \times 10^{-5}) = 4.75$

$pKa = 14 - 4.75 = 9.25$

$\rightarrow NH_4^+ + Br^-$ ③

(A) Calculate the pH of a solution prepared by adding 100 mL of 2.0 M NaOH to 500 mL of 1.2 M NH_4Br

(B) Calculate the pH of a solution prepared by adding 100 mL of 2.0 HCl to 500 mL of 1.2 M NH_3.

(C) Calculate the ratio, $[NH_3] / [NH_4^+]$, of the above solution if the pH is adjusted to 9.85

Ind (A)

$n_{NH_4^+} = 1.2 \, m/L \times 0.50 \, L = 0.60 \, mol$

$n_{NH_3} = 0$

$n_{OH^-} = 2.0 \, m/L \times 0.10 \, L = 0.20 \, mol$

$V_{tot} = 100 + 500$
$= 600 \, mL$
$= 0.60 \, L$

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

int   0.60   0.2   0

chg. -0.2 -0.2 $+0.2$
        _____   _____
        0.40 mol NH_4^+ 0.2 mol NH_3
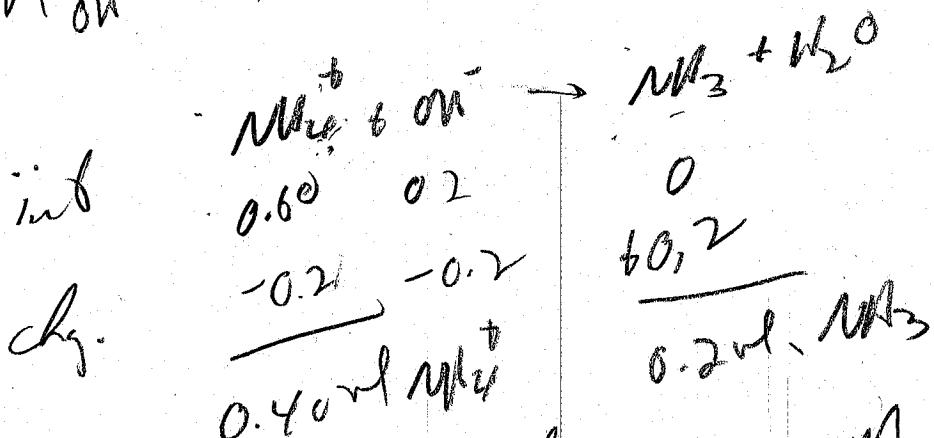
$[NH_4^+] = \dfrac{0.40 \, mol}{0.62} = 0.67 \, M$

$[NH_3] = \dfrac{0.2 \, mol}{0.62} = 0.33$

$pH = pKa + \log\dfrac{[NH_3]}{[NH_4^+]} = 9.25 + \log\left(\dfrac{0.33}{0.67}\right)$

$= 9.25 + \log(0.5) = 9.25 - 0.30 = \boxed{8.95}$

$\left(= 9.25 + \log\dfrac{n_{NH_3}}{n_{NH_4^+}} = 9.25 + \log\left(\dfrac{0.20}{0.40}\right) \right) = 8.95$

(B)

Init: $n_{NH_3} = 1.2 \, mol/L \times 0.5 \, L = 0.60 \, mol$

$n_{NH_4^+} = 0$

$n_{H^+} = 2.0 \, mol/L \times 0.1 \, L = 0.20 \, mol$

	H^+	NH_3	$\rightarrow$	NH_4^+
init	$0.2 \, mol$	$0.6 \, mol$		0
chg	-0.2	-0.2		$+0.20$
fin	~ 0	$0.40 \, mol$		$0.20 \, mol$
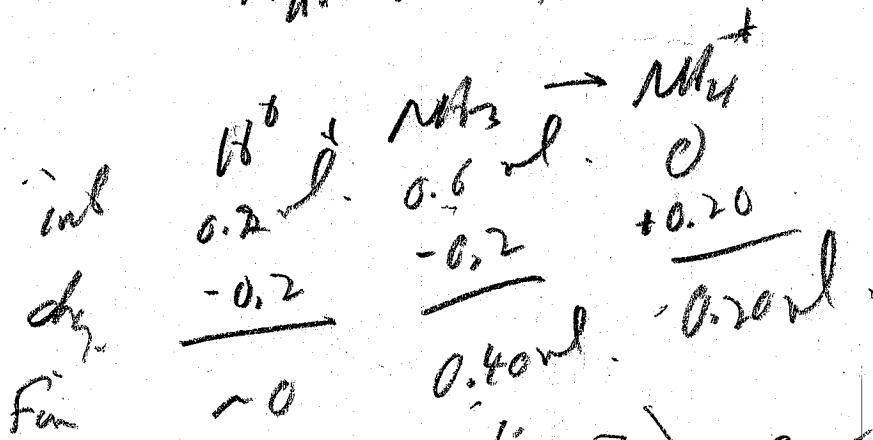
$pH = pK_a + \log\left(\frac{[NH_3]}{[NH_4^+]}\right) = 9.25 + \log\left(\frac{0.40}{0.20}\right)$

$= 9.25 + \log(2) = 9.25 + 0.30 = \underline{9.55}$

(c)

$$pH = pK_a + \log \frac{[NH_3]}{[NH_4]}$$

$$9.85 = 9.25 + \log \frac{[NH_3]}{[NH_4]}$$

$$\log \frac{[NH_3]}{[NH_4]} = 9.85 - 9.25 = +0.60$$

$$\frac{[NH_3]}{[NH_4]} = 10^{0.60} = 3.98 \approx \underline{4.0}$$

**REVIEW: Aniline Buffer** $AnilH^+$ $Anil$
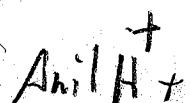
Review Example: Aniline [Anil] is a weak base with $K_b = 4.3 \times 10^{-10}$

(A) Calculate the pH of a solution prepared by adding
0.12 1.0 L of 3.0 M HCl to 3.0 of 0.4 M Anil

(B) Calculate the pH of a solution prepared by adding
1.0 L of 0.50 M NaOH to 3.0 L of 0.40 M
Anilinium Bromide [AnilHBr].

$$pK_b = -\log(4.3 \times 10^{-10})$$
$$= 9.37$$
$$pK_a = 14 - 9.37$$
$$= 4.63$$

(A)

$$n_{Anil} = 0.4 \text{ ml}/L \times 3L = 1.2 N$$

$$n_{H^+} = 0.1 L \times 3 N = 0.3 N$$

$$n_{AnilH^+} = 0$$

$$Anil + H^+ \rightarrow AnilH^+$$

in:
chg:

1.2 N	0.3 N	0
-0.3	-0.3	+0.3 N
0.90 N	~0	0.3 N
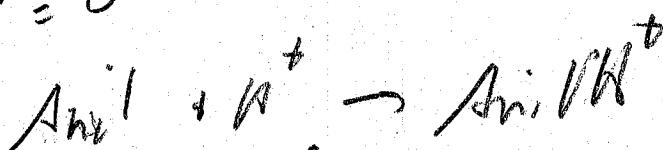
$$[Anil] = \frac{0.9 N}{3.1 L} \qquad [AnilH^+] = \frac{0.3 N}{3.1 L}$$

$$pH = pK_a + \log\left(\frac{[Anil]}{[AnilH^+]}\right) = 4.63 + \log\left(\frac{0.90}{0.3}\right)$$
$$= 4.63 + 0.48 = 5.11$$

(B) $AnilNHBr \rightarrow AnilNH^+ + Br^-$
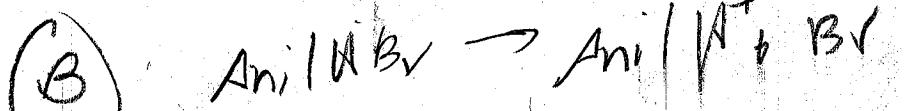
__int__ $0.40\ mol/L \times 3L$
  $= 1.2\ mol$ $1.2\ mol$

$OH^-\ n = 1\ L \times 0.5\ mol/L = 0.5\ mol$
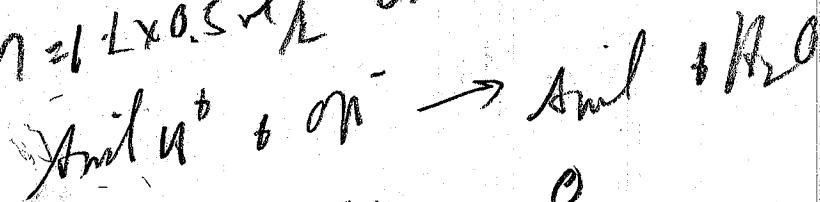
  $AnilNH^+ + OH^- \rightarrow Anil + H_2O$

           0

int   1.2   0.5   + 0.5

chg. $\dfrac{-0.5}{0.7\ mol}$ -0.5 $\dfrac{+0.5}{0.5\ mol}$
  $AnilNH^+$ $Anil$

$pH = p.k_a + \log \dfrac{[Anil]}{[AnilNH^+]} = 4.63 + \log\left(\dfrac{0.5}{0.7}\right)$

             $= 4.63 + (-0.15)$

            $= \boxed{4.48}$

$pK_a = -\log(1.8 \times 10^{-4}) = 3.74$

The acid dissociation constant of formic acid [HCOOH] is $K_a = 1.8 \times 10^{-4}$.

A solution is prepared with an initial concentration [HCOOH] = 0.5 M.
Sufficient KOH is added to the solution to raise the pH to 4.25

What are the concentrations, [HCOOH] & [HCOO⁻] of the solution at this pH?

$$\text{int.} \quad [HCOOH] = 0.50 \, M = [HCOOH] + [HCOO^-]$$

$$pH = pK_a + \log \frac{[HCOO^-]}{[HCOOH]}$$

$$4.25 = 3.74 + \log \frac{[HCOO^-]}{[HCOOH]}$$

$$\log \frac{[HCOO^-]}{[HCOOH]} = 4.25 - 3.74 = +0.51$$

$$\frac{[HCOO^-]}{[HCOOH]} = 10^{0.51} = 3.24$$

$$[HCOO^-] = 3.24 \, [HCOOH]$$

$$[HCOOH] + 3.24 \, [HCOOH] = 0.5$$

$$4.24 \, [HCOOH] = 0.5$$

$$[HCOOH] = \frac{0.5}{4.24} = 0.12 \, M$$

$$[HCOO^-] = 0.5 - 0.12 = 0.38 \, M$$

$H_2CO_3 \leftrightarrow H^+ + HCO_3^- \leftrightarrow 2H^+ + CO_3^{2-}$
    pKa'=6.38   pKa"=10.32
   Ka' = 4.2×10^{-7} Ka" = 4.8×10^{-11}
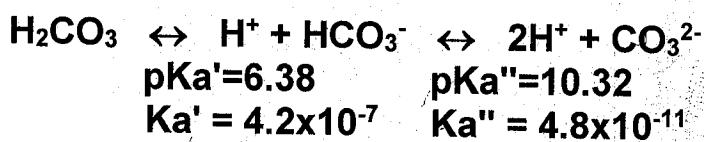
$$K_b'' = \frac{10^{-14}}{Ka''} = \frac{10^{-14}}{4.8 \times 10^{-11}} = 2.1 \times 10^{-4}$$

$$=$$

(a) What is the pH of 0.05 M H_2CO_3?

(b) What is the pH of 0.02 M Na_2CO_3?

(a)

$$H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$$
$$0.05 \qquad x \quad\;\; x$$

$$Ka' = 4.2 \times 10^{-7} = \frac{x \cdot x}{0.05}$$

$$x^2 = 2.1 \times 10^{-8} \qquad x = 1.45 \times 10^{-4}$$

$$pH \quad -\lg(1.45 \times 10^{-4}) = 3.84$$

$$Na_2CO_3 \rightarrow 2Na^+ + CO_3^{2-}$$

(b)

$$CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-$$
$$0.02 \qquad\qquad x \quad\;\; x$$
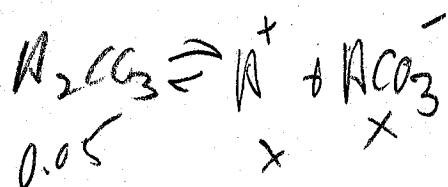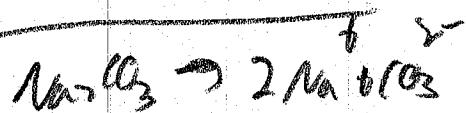
$$Ka'' = \frac{x^2}{0.02} = 2.1 \times 10^{-4}$$

$$x^2 = \cancel{x^2} = 4.2 \times 10^{-6}$$
$$x = 2.05 \times 10^{-3}$$

$$= -\log(2.05 \times 10^{-3})$$

$$pOH = 2.69$$

$$pH = 14 - 2.69 = 11.31$$

# Chap. 8   Carbonic Acid Buffer Questions
April 4   Slide #54

$H_2CO_3 \leftrightarrow H^+ + HCO_3^- \leftrightarrow 2H^+ + CO_3^{2-}$
   pKa'=6.38   pKa''=10.32
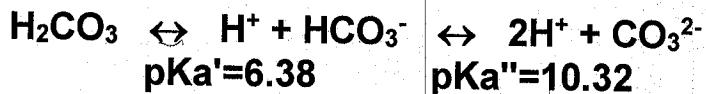
(1) Calculate the pH of a solution prepared by adding 200 mL of 0.50 M NaOH to 400 mL of 0.40 M H_2CO_2.

(2) Calculate the pH of a solution prepared by adding 500 mL of 0.70 M NaOH to 500 mL of 0.40 M H_2CO_2.

(3) What ratio of $[H_2CO_3]/[HCO_3^-]$ would be required for a buffer to have pH = 5.70

(1) init. $n_{H_2CO_3} = 0.4L \times 0.4 \, mol/L = 0.16 \, mol$.

$n_{OH^-} = 0.2L \times 0.5 \, mol/L = 0.10 \, mol$.

$n_{HCO_3^-} = 0$

$H_2CO_3 + OH^- \rightarrow HCO_3^- + H_2O$
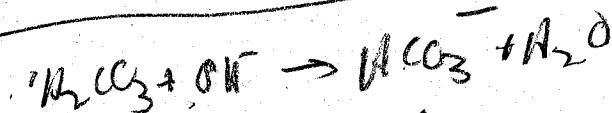
			0
init.	0.16	0.10	+0.10
chg.	-0.16	-0.10	
	0.06 mol		0.10 mol
			HCO_3^-

H_2CO_3

$pH = pKa' + log \dfrac{0.10}{0.06} = 6.38 + 0.22$

$= 6.60$

①

__init__ $n_{H_2CO_3} = 0.50\,L \times 0.4\,\frac{M}{L} = 0.20\,\text{mol}$

$n_{HCO_3^-} = 0$

$n_{NaOH} = 0.50\,L \times 0.70\,M/L = 0.35\,\text{mol}$

step 1
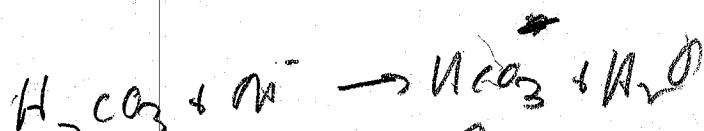
init $H_2CO_3 + OH^- \longrightarrow HCO_3^- + H_2O$

  $0.20\,\text{mol}\quad 0.35\,\text{mol}\quad\quad 0$

  $\dfrac{-0.2}{0}\quad\quad \dfrac{-0.2}{0.15}\quad\quad \dfrac{+0.2}{0.2\,\text{mol}}$

step 2 $HCO_3^- + OH^- \longrightarrow CO_3^{-2}$

init $0.2\quad\quad 0.15\quad\quad\quad 0$
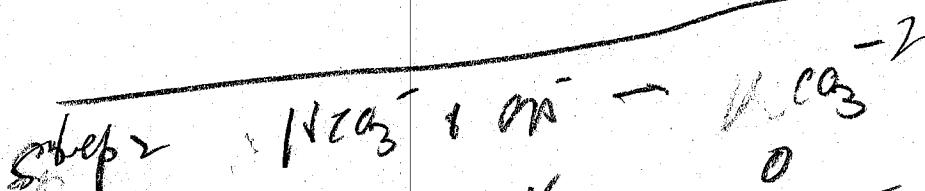
chg $\dfrac{-0.15}{0.05\,\text{mol}}\quad \dfrac{-0.15}{\sim 0}\quad \dfrac{+0.15}{0.15\,\text{mol}}$

$pH = pK_a'' + \log\dfrac{[CO_3^{2-}]}{[HCO_3^-]}$

$= 10.32 + \log\left(\dfrac{0.15}{0.05}\right)$

$= 10.32 + 0.48 = 10.80$

Carbonic Acid Buffer (cont.)

(C)

$$pH = pKa' + \log \frac{[HCO_3^-]}{[H_2CO_3]}$$

$$5.70 = 6.38 + \log \frac{[HCO_3^-]}{[H_2CO_3]}$$

$$\log \frac{[HCO_3^-]}{[H_2CO_3]} = 5.70 - 6.38 = -0.68$$

$$\frac{[HCO_3^-]}{[H_2CO_3]} = 10^{-0.68} = 0.21$$

$$\frac{[H_2CO_3]}{[HCO_3^-]} = \frac{1}{0.21} = 4.8$$

1 Equivalent Definition

Base: 1 equiv. = amount of OH⁻ with same number of moles as number of moles of Acid

Acid: 1 equiv. = amount of H⁺ with same number of moles as number of moles of Base

$1L$ $\Big/$ $0.5m$ $NAc = 0.5M$. $1Eq$ $OH^- = 0.5M$.

n $Na_2CO_3 = 0.5M$.

$1L$ $\Big/$ $0.5M$ $Ac^- = 0.5\,M$, $1Eq$ $H^+ = 0.5M$.

n $CO_3^{2-} = 0.5M$. $1Eq$ $H^+ = 0.5M$.
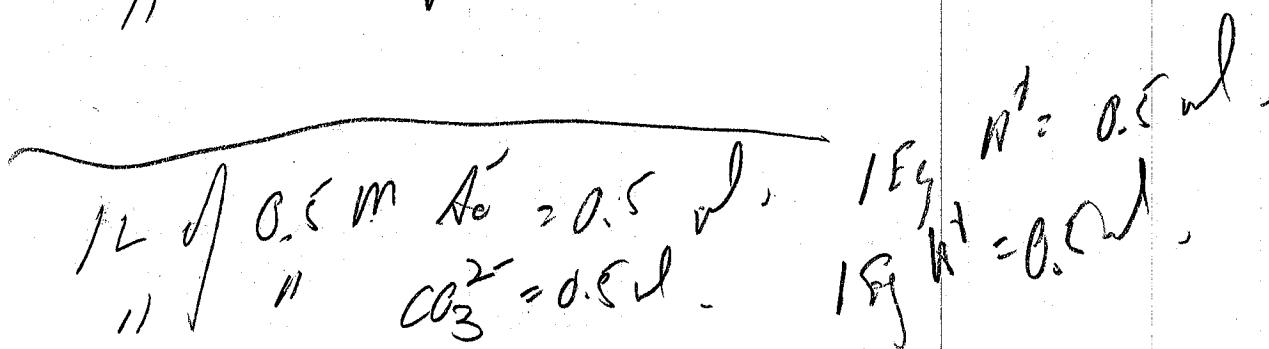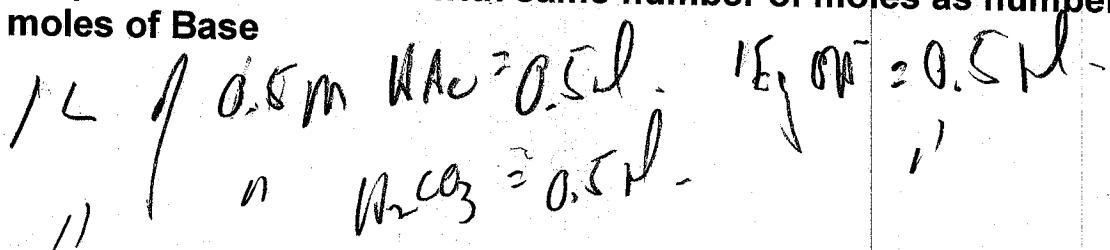
$pKa'_1 - \frac{1}{3}(6 \times 10^{-3}) = 2.22$

The following independent questions are on pH calculations in solutions of Arsenic Acid (H_3AsO_4) and its various anions.

H_3AsO_4 is a triprotic acid with Acid Dissociation Constants:

$K_a' = 6.0 \times 10^{-3}$, $K_a'' = 1.0 \times 10^{-7}$, $K_a''' = 3.2 \times 10^{-12}$

$Na_3AsO_4 \rightarrow 3 Na^+ + AsO_4^{3-}$
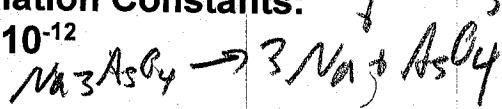
(C) Calculate the pH of a solution prepared by mixing 3.0 L of 0.40 M Na_3AsO_4 with 1.0 L of 0.80 M HNO_3.

(D) Calculate the pH of a solution prepared by mixing 3.0 L of 0.40 M Na_3AsO_4 with 2.0 L of 0.80 M HNO_3.

(E) What is the pH of a solution in which the following ratio is $[HAsO_4^{2-}] / [H_2AsO_4^-] = 1.50$

$H_3AsO_4 \underset{-H^+}{\rightleftharpoons} H_2AsO_4^- \underset{-H^+}{\rightleftharpoons} HAsO_4^{2-} \rightleftharpoons AsO_4^{3-}$

$pKa' = 2.22$ $pKa'' = 7.00$ $pKa''' = 11.49$

(C) $n_{AsO_4} = 0.4 \text{ M} \times 3L = 1.2 \text{ mol}$

$n_{H^+} = 0.8 \text{ M} \times 1L = 0.8 \text{ mol}$

$AsO_4^{3-} + H^+ \rightarrow HAsO_4^{2-}$

1.2 mol	0.8 mol	0
-0.8	-0.8	+0.8
0.4 mol		0.8 mol

$pH = pKa''' + \log \frac{[AsO_4^{3-}]}{[HAsO_4]} = 11.49 + \log\left(\frac{0.40}{0.8}\right)$

$= 11.49 + (-0.30) = 11.19$

(D)

$n_{AsO_4^{3-}} = 0.4 \text{M} \times 3L = 1.2 \text{mol}$

$n_{H^+} = 0.8 \times 2 = 1.6 \text{mol}$

Step 1

$$AsO_4^{2-} + H^+ \rightarrow HAsO_4^{2-}$$

1.2	1.6	0
-1.2	-1.2	+1.2
0	0.4 mol	1.2 mol HAsO_4^{2-}

Step 2

$$HAsO_4^{2-} + H^+ \rightarrow H_2AsO_4^{-}$$

1.2	0.4	0.4
-0.4	-0.4	+0.4
0.8 mol		0.4 mol H_2AsO_4^{-}

$$pH = pKa'' + \lg \frac{[HAsO_4^{2-}]}{[H_2AsO_4^{-}]}$$
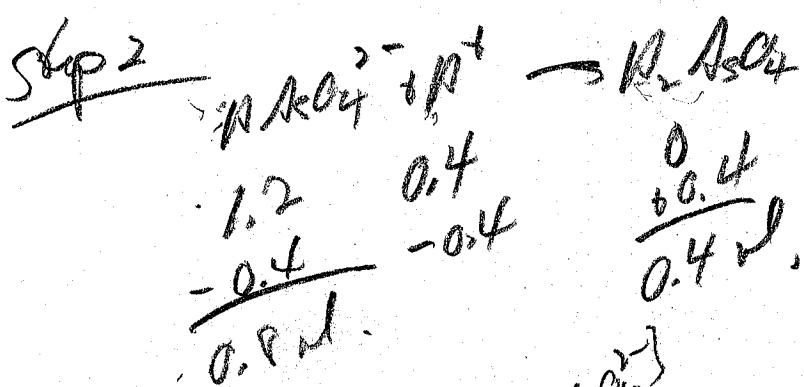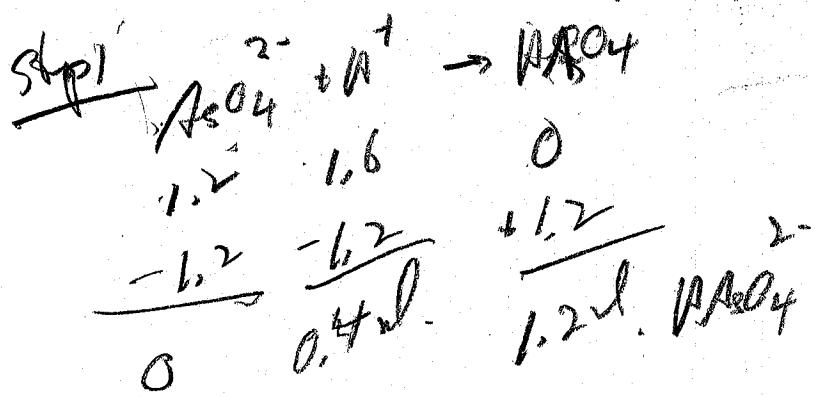
$$= 7.00 + \lg \left(\frac{0.8}{0.4} \right)$$

$$= 7.00 + 0.30$$

$$= 7.30$$

(E) $\dfrac{[H_2AsO_4^-]}{[HAsO_4^{2-}]} = 1.5 \longrightarrow \dfrac{[HAsO_4^{2-}]}{[H_2AsO_4^-]} = \dfrac{1}{1.5}$

$pH = pK_a'' + \log \dfrac{[HAsO_4^{2-}]}{[H_2AsO_4^-]}$

$= 7.00 + \log(1.5) = 7.00 + 0.18$

$= 7.18 \approx 7.2$

Slide #55 Σx

$\frac{1}{2}(2.12+7.21)=4.67$
$\frac{1}{2}(7.21+12.32)=9.77$

$\frac{\quad}{4.67} \qquad 9.77 = \frac{1}{2}(7.21+12.32)=9.77$

$H_3PO_4 \rightleftharpoons H_2PO_4 \rightleftharpoons HPO_4 \rightleftharpoons PO_4^{3-}$

$\quad 2.12 \qquad\qquad 7.21 \qquad\qquad 12.32$
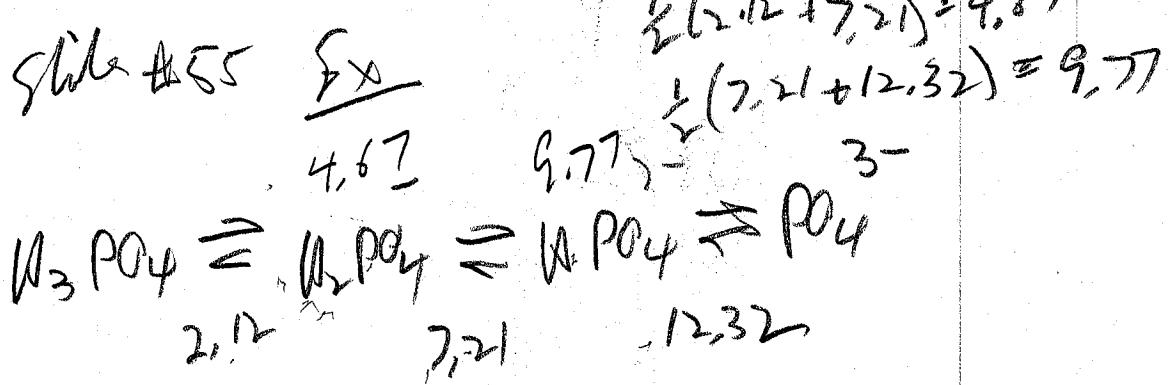
(a) $pH = 4.67$

(b) $pH = 9.77$

(c) $pH = pK_a' + \log\frac{[H_2PO_4^{2-}]}{[H_3PO_4]} = 2.12 + \log(1) = 2.12$

(d) $pH = pK_a''' + \log\frac{[PO_4^{3-}]}{[HPO_4^{2-}]}$

$\qquad = 12.32 + \log 1 = 12.32$

HF Example: Slide #62

$pKa = -\log(7.2 \times 10^{-4})$

$= 3.14$

Hydrofluoric Acid, HF, is a weak acid with Ka = 7.2x10⁻⁴.
**One initially has 600 mL of 0.30 M HF,**

(a) How many mL of 0.50 M NaOH are required to reach half-way to the equivalence point?

(b) What is the pH of the solution half-way to the equivalence point.

(c) If one adds an additional half equivalence of NaOH, will the pH higher than 7.0, lower than 7.0, or equal to 7.0?
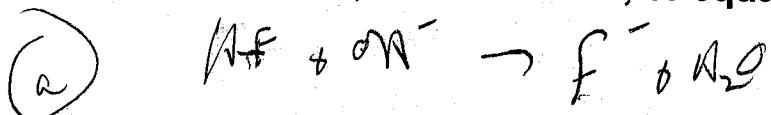
(a) $HF + OH^- \rightarrow F^- + H_2O$

$n_{HF} = 0.30 \text{ mol}/L \times 0.6 L = 0.18 \text{ mol}$

$n_{OH^-} = \frac{1}{2} n_{HF} = 0.09 \text{ mol}$

$C_{OH^-} = \frac{n_{OH^-}}{V} \rightarrow V = \frac{n_{OH^-}}{C_{OH^-}} = \frac{0.09 \text{ mol}}{0.5 \text{ mol}/L}$

$= 0.18 L = 180 \text{ mL}$

(b) $pH = pKa + \log\frac{[F^-]}{[HF]} =$

$= 3.14 + \log(1) = 3.14$

(c) $pH > 7$

pH Dependence of Species Concentrations
Glycine: Slide #70

$$\tfrac{1}{2}(pK_a' + pK_a'')$$
$$AA^+ \leftrightarrow AA \leftrightarrow AA^-$$
$$pK_a' \qquad pK_a''$$
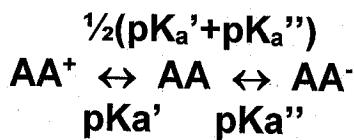
Example: Glycine - $pK_a' = 2.34$ $\qquad$ $pK_a'' = 9.60$

(A) What is the pH of 0.10 M neutral Glycine?

(B) If the pH of the above solution is adjusted to 9.0,
what species are present and what are their concentrations?

(C) What are the species present if the pH is adjusted to 4.5?

$$pH = \tfrac{1}{2}(2.34 + 9.60)$$
$$= 5.97$$

5.97

$$Gly^+ \rightleftharpoons Gly \rightleftharpoons Gly^-$$
$$2.34 \qquad 9.60$$

(A) $\quad pH = 5.97$

$$[Gly] \,\&\, [Gly^-] = 0.10\,M$$

$$9.00 = pH = pKa + \log \frac{[Gly^-]}{[Gly]} = 9.60 + \log \frac{[Gly^-]}{[Gly]}$$

$$\log \frac{[Gly^-]}{[Gly]} = 9.00 - 9.60 = -0.60$$

$$\frac{[Gly^-]}{[Gly]} = 10^{-0.60} = 0.25 \rightarrow [Gly^-] = 0.25\,[Gly]$$

$$[Gly] + 0.25\,[Gly] = 0.10 = 1.25\,[Gly]$$

$$[Gly] = \frac{0.1}{1.25} = 0.08\,M \qquad [Gly^-] = 0.1 - 0.08 = 0.02$$

$$Lys^{+2} \geq Lys^{+1} \geq Lys^{0} < Lys^{-1}$$

5.57   9.74

2.18   8.96   10.53

Add #E gion of OH⁻

1 eq - pH = 5.57   Avg. Chg = +1

1.5   pK = 8.96   Avg. Chg = 0½

2.5   pK = 10,53   Avg. Chg = -½

pI = 9.74

3 ? eq OH⁻ need to get Avg. Chg = -1