Symbols   HA = weak acid H^+ = strong acid
            A^- = conj base **BUFFERS** OH^- = strong base
            of weak acid (so, weak base)

In some situations, a change in pH can cause extreme reactions to occur. For example, the pH of blood should be from 7.38-7.42. If it changes to 7 or 8, we would die. We must therefore have a mechanism to protect against even a slight change in pH. The chemical that does this is called a __buffer__.
Buffers maintain pH, they do not make solutions neutral.

Buffer Components: Buffers are made of 2 solutes that contain an acid/base __conjugate pair__ (HA/A^-)

Ex. $(HC_2H_3O_2$ & $NaC_2H_3O_2)$ $(H_2CO_3$ & __HCO_3^-__) $(NH_4^+$ & __NH_3__)
   * in blood

How Buffers Work:

A buffer must neutralize any strong acid or base added so that __the original pH remains constant__.
One solute in the buffer will neutralize the acid, the other solute neutralizes the base.

Consider a buffer of (HA & A^-). What happens if a strong acid is added to this solution?

$$H^+ + A^- \longrightarrow \boxed{HA} \quad \text{(not equil)}$$
strong acid
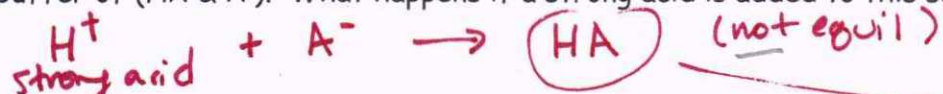
What happens if a strong base is added to this solution?

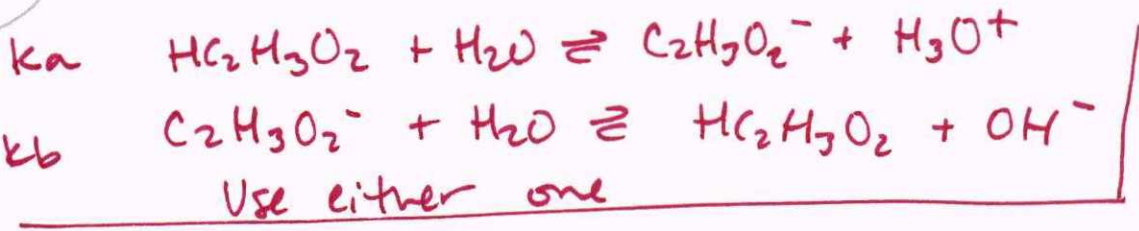$$OH^- + HA \longrightarrow \boxed{H_2O} + \boxed{A^-}$$
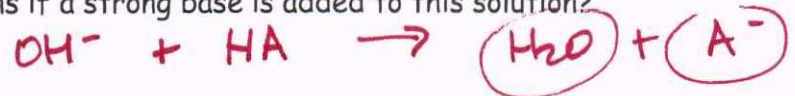
none of these species change the PH !

Calculating the pH of a Buffered Solution:
To calculate pH, write either the k_a or the k_b reactions that are occurring. Use the equation to determine __the $[H^+]$__ which can be used to calculate pH.

Ex. 1: A buffer solution contains 0.090 M acetic acid and 0.11 M sodium acetate. What is the pH?

Do #82 $HC_2H_3O_2 / C_2H_3O_2^-$

k_a $HC_2H_3O_2 + H_2O \rightleftharpoons C_2H_3O_2^- + H_3O^+$

k_b $C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$
          Use either one

	HA $\rightleftharpoons$	$C_2H_3O_2^-$ +	H^+
i	.090	.11	0
D	$-x$	$+x$	$+x$
e	$.09-x$	$.11 + x$	x

$$k_a = 1.8 \times 10^{-5} = \frac{(.11 + x)(x)}{.09 - x}$$

$.11 + x \to 0$ $.09 - x \to 0$

$$x = 1.5 \times 10^{-5} = [H^+]$$

$$\boxed{pH = 4.82}$$

The pH of a buffer depends on 1) k_a or k_b and 2) the ratio of $[HA]/[A^-]$.
Diluting a buffer *(does, ~~does not~~)* change its pH because it does not change ↑ or k_a.

● The basic k_a expression can be rearranged to another useful equation known as **Henderson-Hasselbalch Eqn.**

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

take log of the eqn.

$$\longrightarrow \boxed{pH = pka + \log\left(\frac{[A^-]}{[HA]}\right)}$$

or $pOH = pkb + \log\left(\frac{A}{B}\right)$

According to the HH Eqn, when [acid] = [base], $\underline{pH = pka}$.
Simplifications (neglecting x) will always work for buffers, therefore the amount of acid or base ionized when a strong acid or base is added to a buffer can be **neglected**. Therefore, the initial concentrations can be plugged straight into **the HH Eqn.** in order to determine $[H^+]$ & pH.

Ex. 2: A buffer solution contains 0.090 M acetic acid and 0.11 M sodium acetate. What is the pH?
(This is the same as Ex. 1. Solve this now using the HH Eqn.)
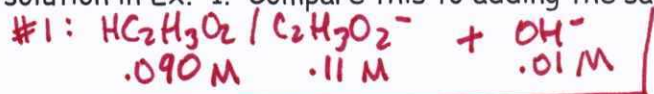
$$pH = pka + \log\left(\frac{[A^-]}{[HA]}\right)$$

$$pH = -\log(1.8\times10^{-5}) + \log\left(\frac{.11}{.09}\right)$$

$$4.74 \qquad + \qquad .08715 \qquad = \boxed{4.83}$$

● **Calculating ΔpH when a Strong Acid or Base is Added to a Buffer System:**
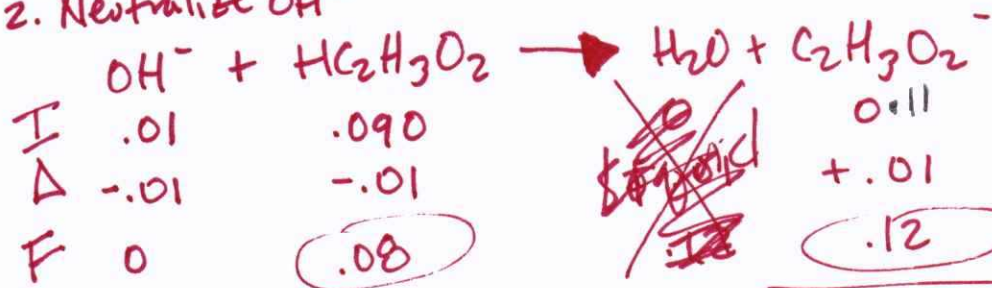Do #98
To determine changes in pH, you must first determine the original $\underline{pH}$ of the buffered solution. Then you must neutralize the strong base or acid that has been added. Last, to figure out the pH, you must determine the amount of $\underline{[H^+]}$ that remains.

Ex. 3:   Calculate the ΔpH that occurs when 0.010 moles of NaOH is added to 1.0 L of the buffered solution in Ex. 1. Compare this to adding the same amount of NaOH to pure water.

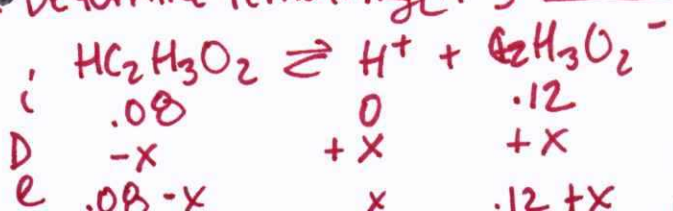#1: $HC_2H_3O_2 / C_2H_3O_2^-$ + OH^-
    .090 M   .11 M   .01 M

1. Original pH = 4.83

2. Neutralize OH^-

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

				0.11
I	.01	.090		
Δ	-.01	-.01		+.01
F	0	.08		.12

OR.

3. Use HH Eqn.
$$pH = pka + \log\left(\frac{B}{A}\right)$$
$$= -\log 1.8\times10^{-5} + \log\left(\frac{.12}{.08}\right)$$
$$\boxed{pH = 4.9}$$

NaOH in H_2O
$[OH^-] = .01$ $pOH = 2$
$\boxed{pH = 12}$

●3. Determine remaining $[H^+]$

$$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$$

i	.08	0	.12
D	-x	+x	+x
e	.08-x	x	.12+x

$$k_a = 1.8\times10^{-5} = \frac{(x)(.12+x)}{.08-x}$$

$$x = 1.2\times10^{-5} M = [H^+]$$
$$\boxed{pH = 4.9}$$

Preparing a Buffer of a Given pH:

Buffers are made so that the ratio of ___HA___/___A⁻___ is near 1. This makes the pH = ___pka___. *from HH Eqn.*

   1) Choose a weak acid with a pk_a near the needed ___pH___. (pH = $pk_a \pm$ ___1___)

   2) Adjust the ratio to get the exact desired ___pH___.
          of [HA]/[HA]

Ex. 4: A solution buffered at pH = 5 is needed. Can we use $HC_2H_3O_2$/$NaC_2H_3O_2$ & in what ratio?

1) Check to see if pH = pka ± 1 :

    $Ka = 1.8 \times 10^{-5}$ pka = 4.7   so **Yes**

2) Adjust the ratio :

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

rearrange to solve for ratio $\frac{[A]}{[HA]}$

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

pH = 5
$[H^+] = 10^{-5}$

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

or   **$\dfrac{.18\ M}{.1\ M}$**

Buffer Capacity:

There is a limit to the amount of strong base or acid that a buffer can ___neutralize___. This limit is known as ___buffer capacity___. Buffer capacity is determined by ___concentrations___, NOT by ___the ratio___.   (or amounts) of HA & A⁻
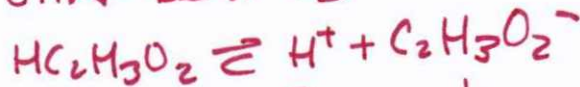
Ex. 5: A student prepared 1 L of a 1 M $HC_2H_3O_2$ and 1 M $NaC_2H_3O_2$ buffer for an experiment where 0.15 mol of H^+ was to be generated. If there enough buffer capacity to neutralize the H^+ without changing the pH by more than .1? — So calculate pH of buffer & pH w/ H⁺ added

② Neutralize:

	H^+	+	$C_2H_3O_2^-$	→	$HC_2H_3O_2$
i	.15M		1 M		1M
A	-.15		-.15		+.15
e	0		.85		1.15
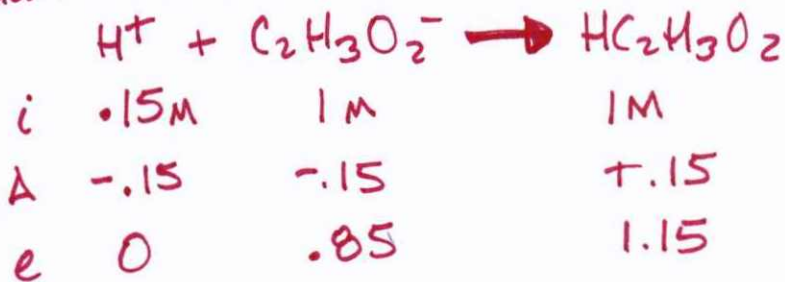
① pH of Buffer:
Either ka or kb

	$HC_2H_3O_2$	⇌	H^+	+	$C_2H_3O_2^-$
i	1		0		1
D	-x		+x		+x
e	1-x		x		1+x

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

Calculate new H^+ w/ ka (use HH)

$$pH = pk_a + \log\left(\frac{B}{A}\right)$$
$$-\log 1.8 \times 10^{-5} + \log\left(\frac{.85}{1.15}\right)$$

pH = 4.61

③ $\Delta pH = .13$ **NO**

②