$$HA(aq) \longleftrightarrow H^+(aq) + A^-(aq)$$

How are we going to
control this equilibrium?

Add HA	shift to the "products"
Add A^-	shift to the "reactants"
Add H^+	shift to the "reactants"
Remove H^+	shift to the "products"

I can't have acid and OH^-
I can't have base and H^+

They will react
Neutralization reactions

$$H^+(aq) + OH^-(aq) \longleftrightarrow H_2O(l)$$

$$H^+(aq) + A^-(aq) \longleftrightarrow HA(aq)$$

$$H^+(aq) + B(aq) \longleftrightarrow BH^+(aq)$$

$$OH^-(aq) + HA(aq) \longleftrightarrow A^-(aq)$$

$$OH^-(aq) + BH^+(aq) \longleftrightarrow B(aq)$$

What volume of a 0.1 M NaOH will you need to add to 200 mL of a 0.2 M solution of HCl to neutralize it?

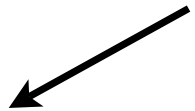
A. 100 mL

B. 200 mL

C. 300 mL

D. 400 mL

E. 500 mL



There are .04 moles of H^+ .2M x .2L
to neutralize you'll need .04 moles of OH^-

For that you'll need .4L of a .1M solution

Or you can look at it as the acid is twice as concentrated as the base
therefore you'll need twice as much

What can I have at the same time
in a solution?

A weak acid and conjugate base HA and A^-

A weak base and its conjugate acid B and BH^+

What will have in solution if initially I have
I mole of acetic acid and I add
0.5 mole of NaOH?

A. I mole acetic acid and 0.5 mole OH^-

B. I mole H^+ and 0.5 moles OH^-

C. 0.5 moles H^+

D. 0.5 mole HA and 0.5 moles of A^-

E. I mole of Ha and 0.5 moles of A-

OH^- will react with HA until one of them is gone

	$OH^-(aq)$	+	HA(aq)	⟷	$A^-(aq)$
	0.5 moles	+	I mole		0 moles
neutralize	-0.5 moles	+	-0.5 moles		+0.5 moles
	0 moles	+	0.5 moles		0.5 moless

First Neutralize
Second Solve the Equilibrium

First find out what is mostly in solution
Second solve for the little concentrations

Now I have a solution which initially contains both
HA and A^-

Buffer I have BOTH HA and A-

$$HA(aq) \longrightarrow H^+(aq) + A^-(aq)$$

	HA	H^+	A^-
I	$[HA]_0$	0 (really 10^{-7})	$[A^-]_0$
C	-x	+x	+x
E	$[HA]_0$ -x	+x	$[A^-]_0$ + x

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(x)([A^-]_0+x)}{[HA]_0 - x} = \frac{(x)([A^-]_0)}{[HA]_0}$$ assuming x << C

pH in a buffer solution

$$K_a \approx \frac{[H^+][A^-]_0}{[HA]_0}$$

we have approximated a small change

$$\log(K_a) \approx \log \frac{[H^+][A^-]_0}{[HA]_0} = \log[H^+] + \log \frac{[A^-]_0}{[HA]_0}$$

$$pKa = pH - \log \frac{[A^-]_0}{[HA]_0}$$

initial conjugate base

$$pKa = pH - \log \frac{[A^-]_0}{[HA]_0}$$

initial weak acid

$[A^-]_0 = [HA]_0$	$[A^-]_0 < [HA]_0$	$[A^-]_0 > [HA]_0$
equal acid/base	more acid	more base
$-\log \frac{[A^-]_0}{[HA]_0} = 0$	$-\log \frac{[A^-]_0}{[HA]_0} > 0$	$-\log \frac{[A^-]_0}{[HA]_0} < 0$
pH = pKa	pH < pKa	pH > pKa

The pK_a of HF is 3.18. What is the pH of solution of 100 mL of 0.1 M HF and 100 mL of a 0.2 M NaF?

A. slightly less than 3.18

B. 3.18

C. slightly more than 3.18

initial conjugate base

$$pKa = pH - \log \frac{[A^-]_0}{[HA]_0}$$

initial weak acid

if the initial acid and base are similar in concentration than the pH is close to the pKa

For the pH to be 1 unit different than the pKa the difference in concentrations must be at least 10 X!

Back to Buffers

$$pKa = pH - \log \frac{[A^-]}{[HA]}$$

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

This is the same equation!

Let's look at the second one

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

If $[HA] = [A^-]$, then $[H^+] = K_a$

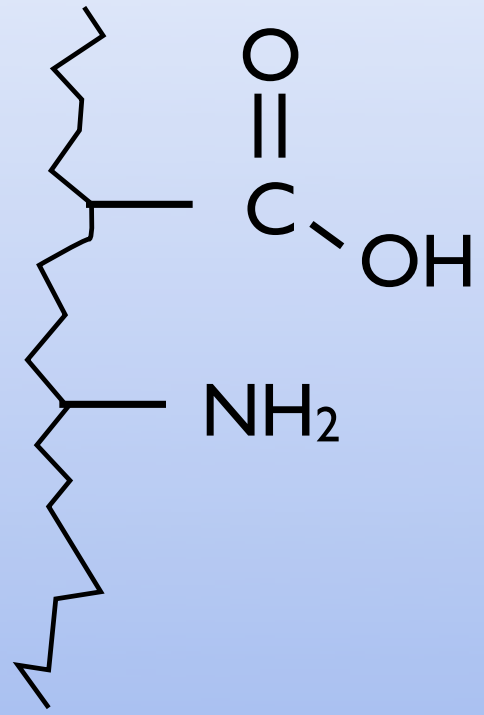
or we could look at it as

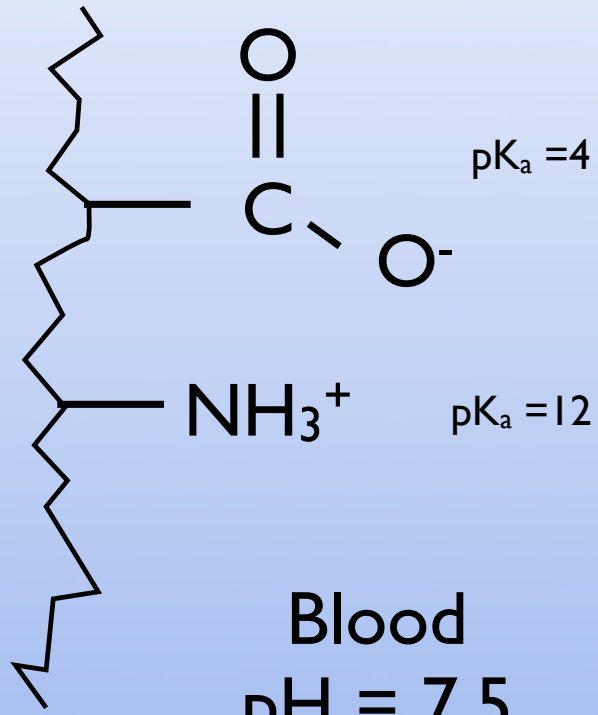
if $[H^+] = K_a$, then $[HA] = [A^-]$

if $[H^+] > K_a$, then $[HA] > [A^-]$ "too many" protons
if $[H^+] < K_a$, then $[HA] < [A^-]$ "too few" protons

Why should I care

Proteins have lots of acid and base groups





$pK_a = 4$

$pK_a = 12$

Blood
pH = 7.5

We want to "Buffer" against pH change

demo

Add NaOH to water and the pH shoots up to 12

Add NaOH to mixture of acetic acid and sodium acetate and the pH doesn't change at all

NaOH added to water

Water. Add 10^{-3} moles of OH^- to the solution

The $[OH^-] = 10^{-3}$ pOH = 3 pH =11

NaOH added to buffer

initial concentration of [HA] = 0.1 M
initial concentration of $[A^-]$ = 0.1 M

add .001 moles of NaOH to 1L of solution

concentration of [HA] = .1 -.001 = 0.099
concentration of $[A^-]$ = .1 + .001 = .101

$10^{-4.75}$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[H^+](.101)}{0.099}$$

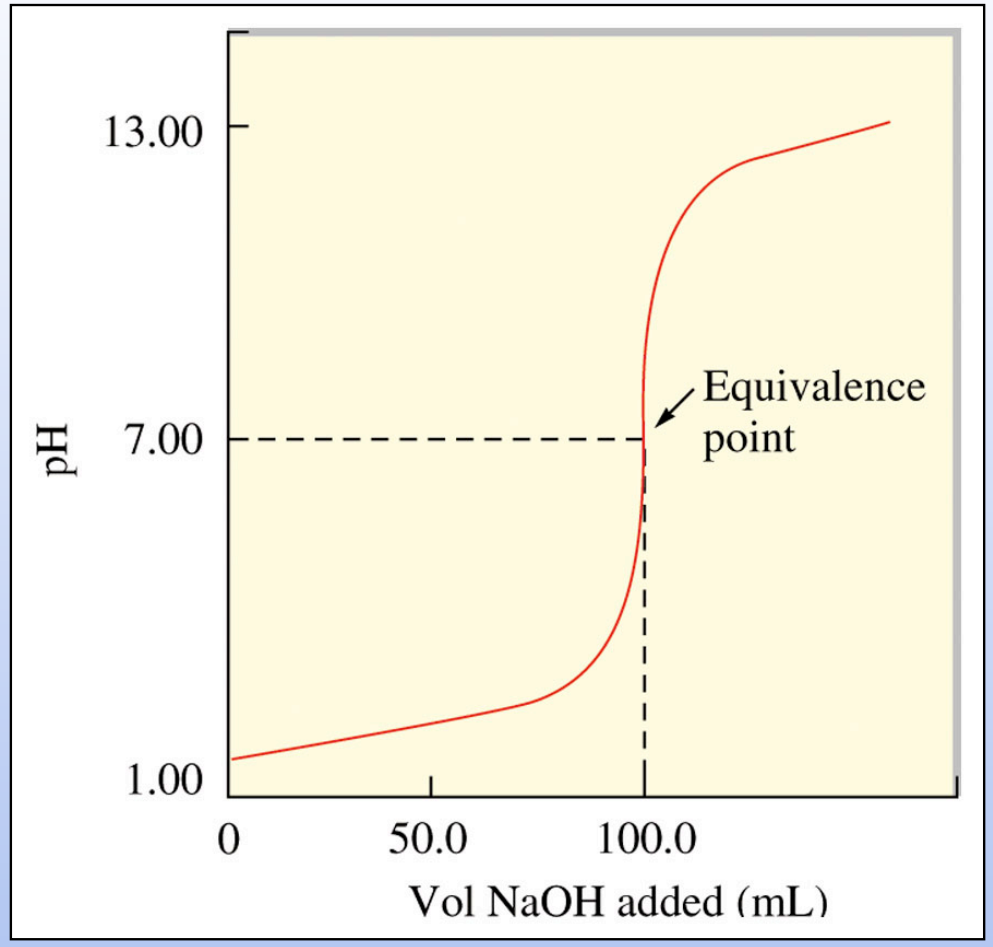
pH = 4.76

Water before adding NaOH pH = 7
after adding NaOH pH = 3

Buffer before adding NaOH pH = 4.75
after adding NaOH pH = 4.76

the only way to change the pH of the buffer system dramatically is to add enough acid or base to substantially change either the HA or A^- concentrations

Strong Acid/Strong Base Titration



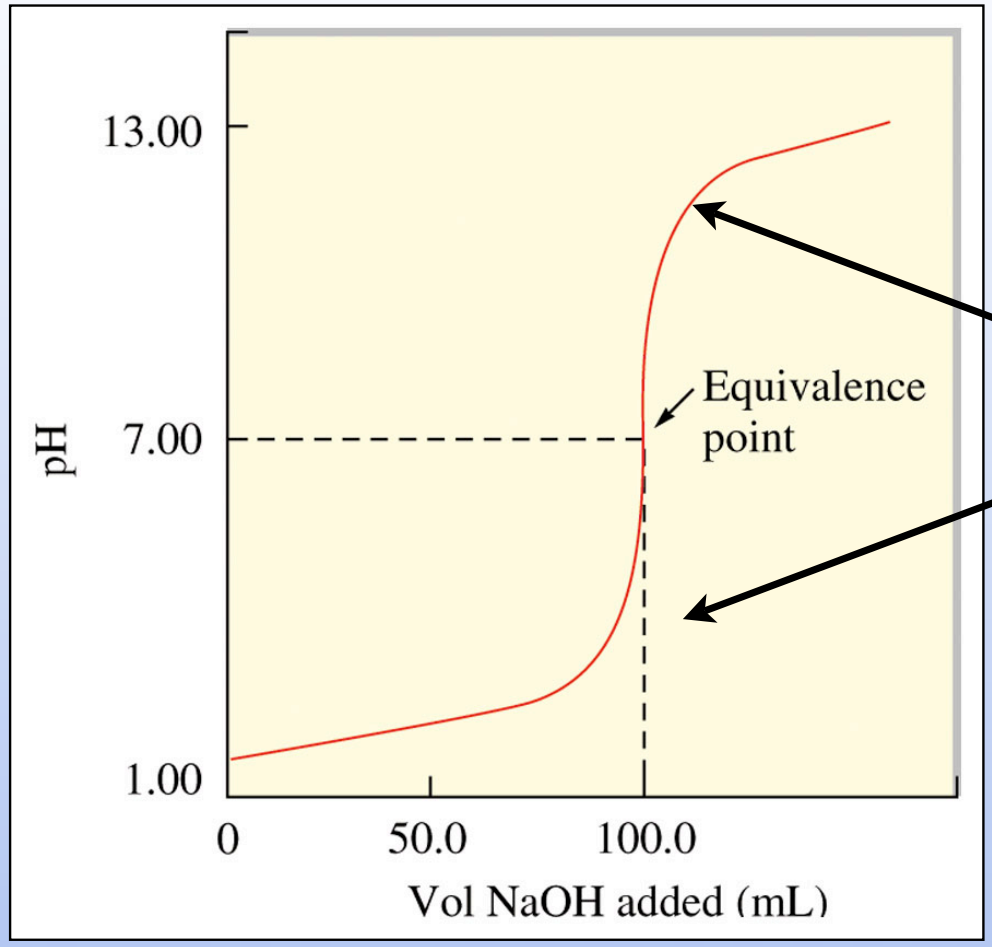
original solution 50 mL HCl
adding .1 M NaOH
at equivalence point

same number of moles of base
.1L x .1M = 0.01 moles OH^-

therefore the solution originally
had 0.01 moles H^+

concentration was .2 M

at the equivalence point we have
equal number of moles of acid and base



pH changes rapidly because the total amount of H^+ OH^- is very small between pH 3 and pH 11