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

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"

Neutralization

A solution can be neutralized
(equal amounts of H^+ and OH^-)
by adding an acid or base to the solution

As you are mixing two
solutions, it is generally easiest
to think in terms of moles
(rather than molarity)

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

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

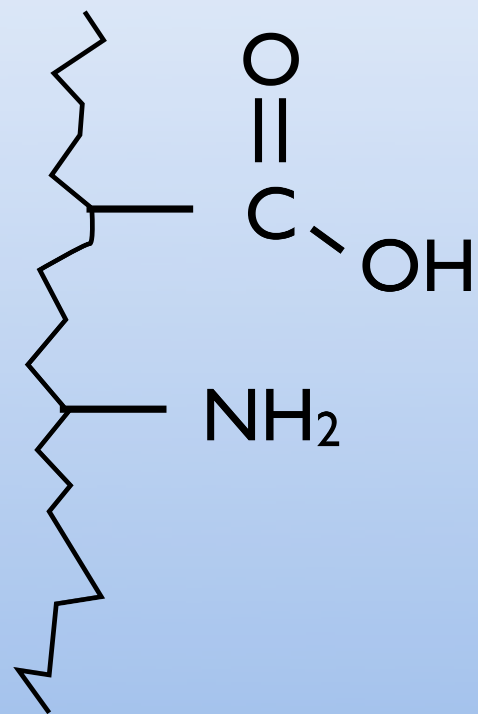
if $[H^+] = K_a$, then $[HA] = [A^-]$

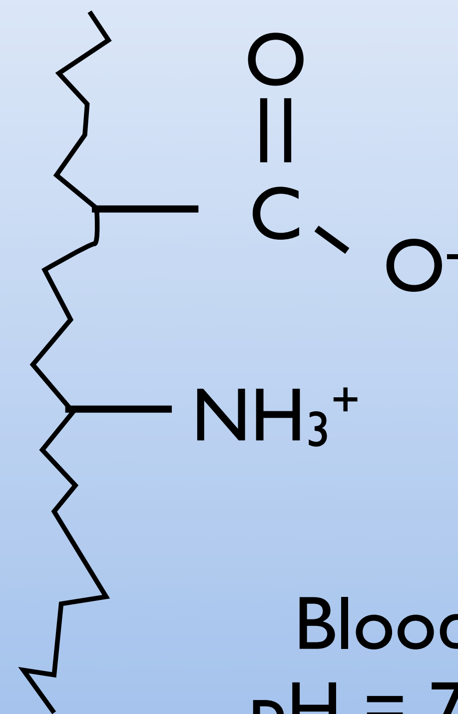
if $[H^+] > K_a$, then $[HA] > [A^-]$ "too many" protons

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

$10^{-4.75}$

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

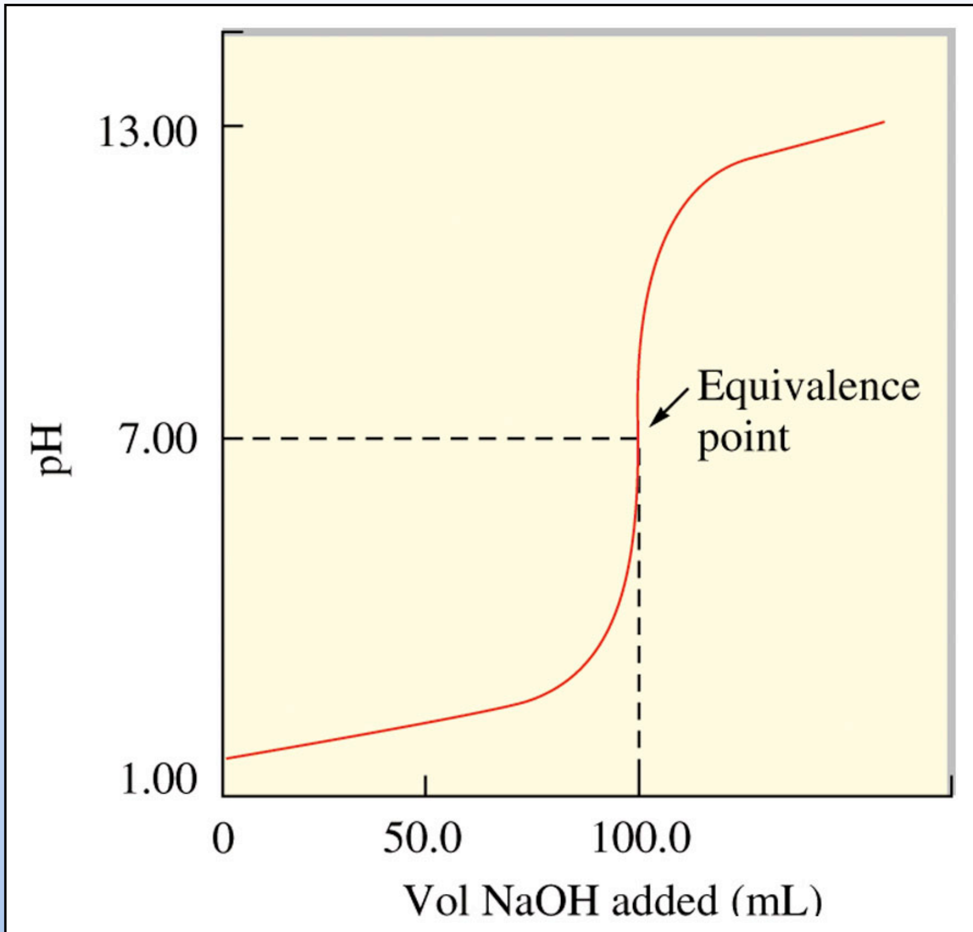
pH = 4.76

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



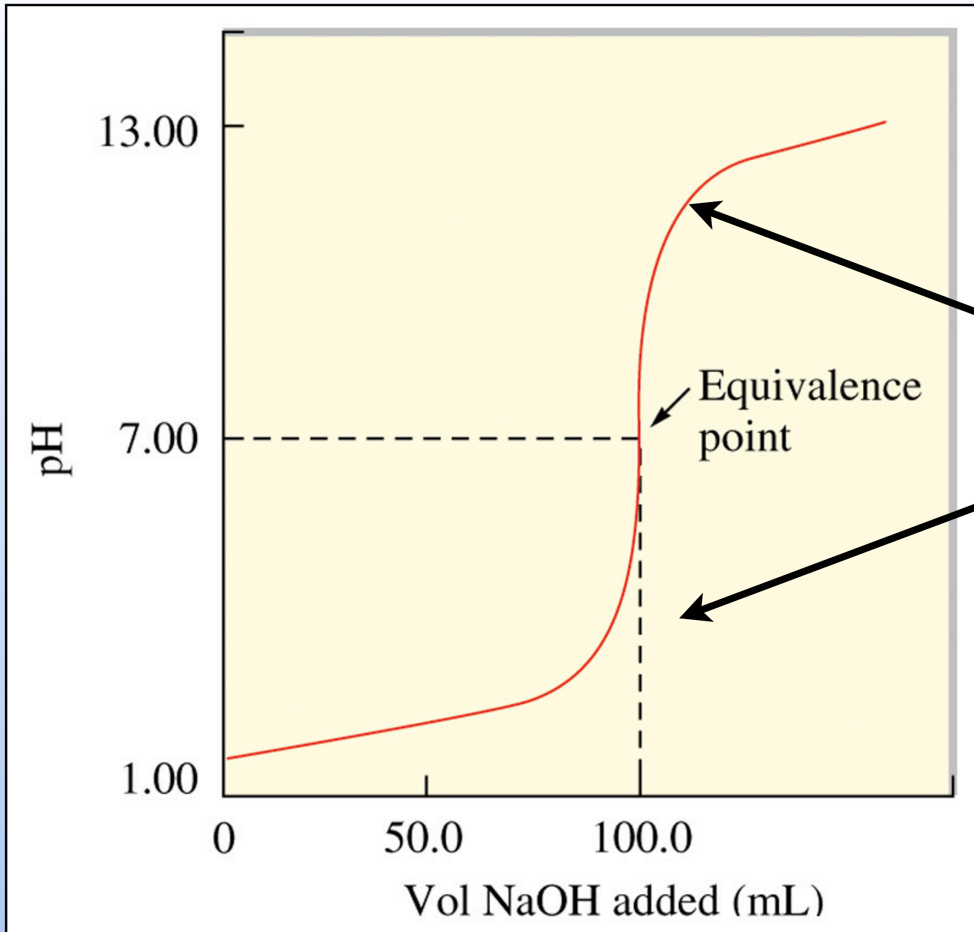
original solution 50 mL HCl
adding .1 M NaOH
at equivalence point

same number of moles of base
$.1L \times .1M = 0.01$ 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