Neutralization
Buffers
Titration

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

How are we going to control this equilibrium?

Add H^+	shift to the "reactants"
Remove H^+ (Add OH^-)	shift to the "products"

Strong Acid/Strong Base (only H^+ and OH^-)

What volume of a 0.5 M NaOH will you need to add to 200 mL of a 0.5 M solution of HCl to neutralize it?

A. 100 mL

B. 200 mL

C. 300 mL

D. 400 mL

E. 500 mL

What is the pH of a solution that is made of equal moles of a HF and NaOH?

A. neutral (pH 7)

B. acidic (pH < 7)

C. basic (pH > 7)

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

They will react
Neutralization reactions

$$H^+(aq) \quad + \quad OH^-(aq) \longleftrightarrow$$

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

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

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

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

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-

$OH^-(aq)$	+	$HA(aq)$	$\longleftrightarrow$	$A^-(aq)$
0.5 moles	+	I mole		0 moles

neutralize

First Neutralize
Second Solve the Equilibirium

Now I have a solution which initially contains both

HA and A^-

pH in a buffer solution

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

we have approximated a small change

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

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

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

initial conjugate base → $[A^-]_0$

initial weak acid → $[HA]_0$

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 $[H^+] = K_a$

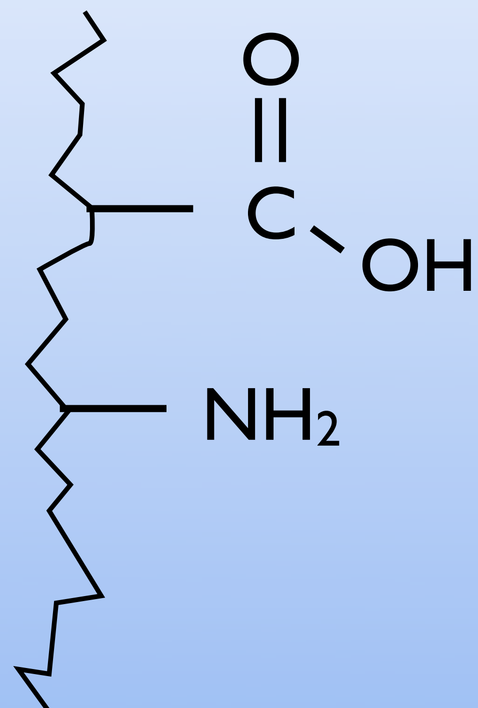
or we could look at it as

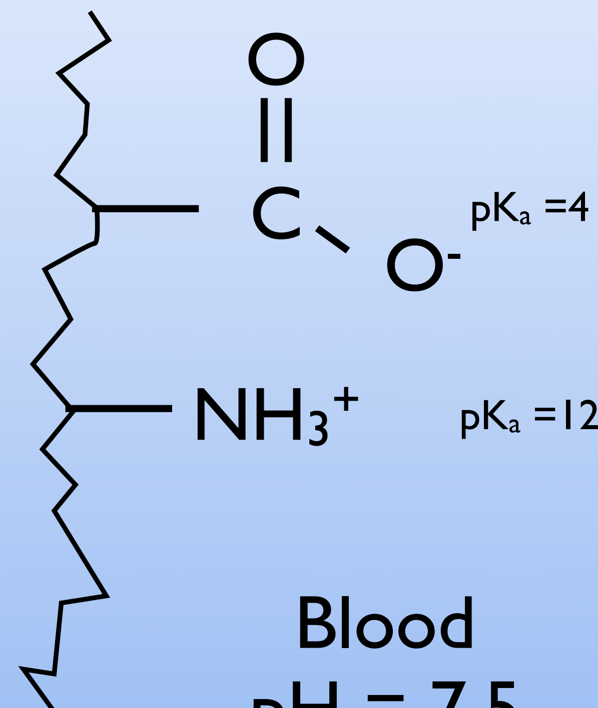
if $[H^+] = K_a$, then $[HA] = [A^-]$

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





Blood
pH = 7.5

Why should I care

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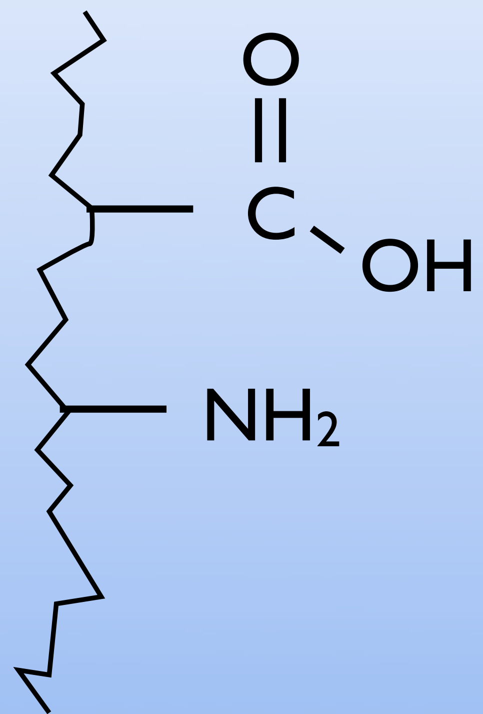

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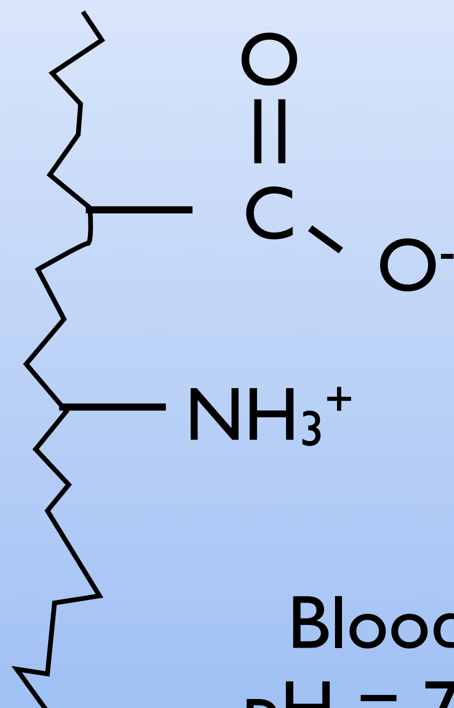


$pK_a = 4$

$pK_a = 12$

Blood
pH = 7.5

The pK_a of acetic acid is 4.75. What will the pH be for a solution that has equal moles sodium acetate and acetic acid?

A. much less than 4.75

B. about 4.75

C. much higher than 4.75

Let's give it a try

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

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

Buffer before adding NaOH   pH = 4.75
after adding NaOH 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

Buffer capacity = amount of acid(base) the buffer can "absorb".
More HA/A^- in solution = larger buffer capacity

Titration

Slow addition of strong base(acid)
to a solution of an acid(base)

while measuring the pH

Imagine titration of 100 mL of 0.1 M acetic acid with 0.1 M NaOH

Initial point (no base added)

Imagine titration of 100 mL of 0.1 M acetic acid with 0.1 M NaOH

after the addition of 10 mL of base

Imagine titration of 100 mL of 0.1 M acetic acid with 0.1 M NaOH

after the addition of 50 mL of base

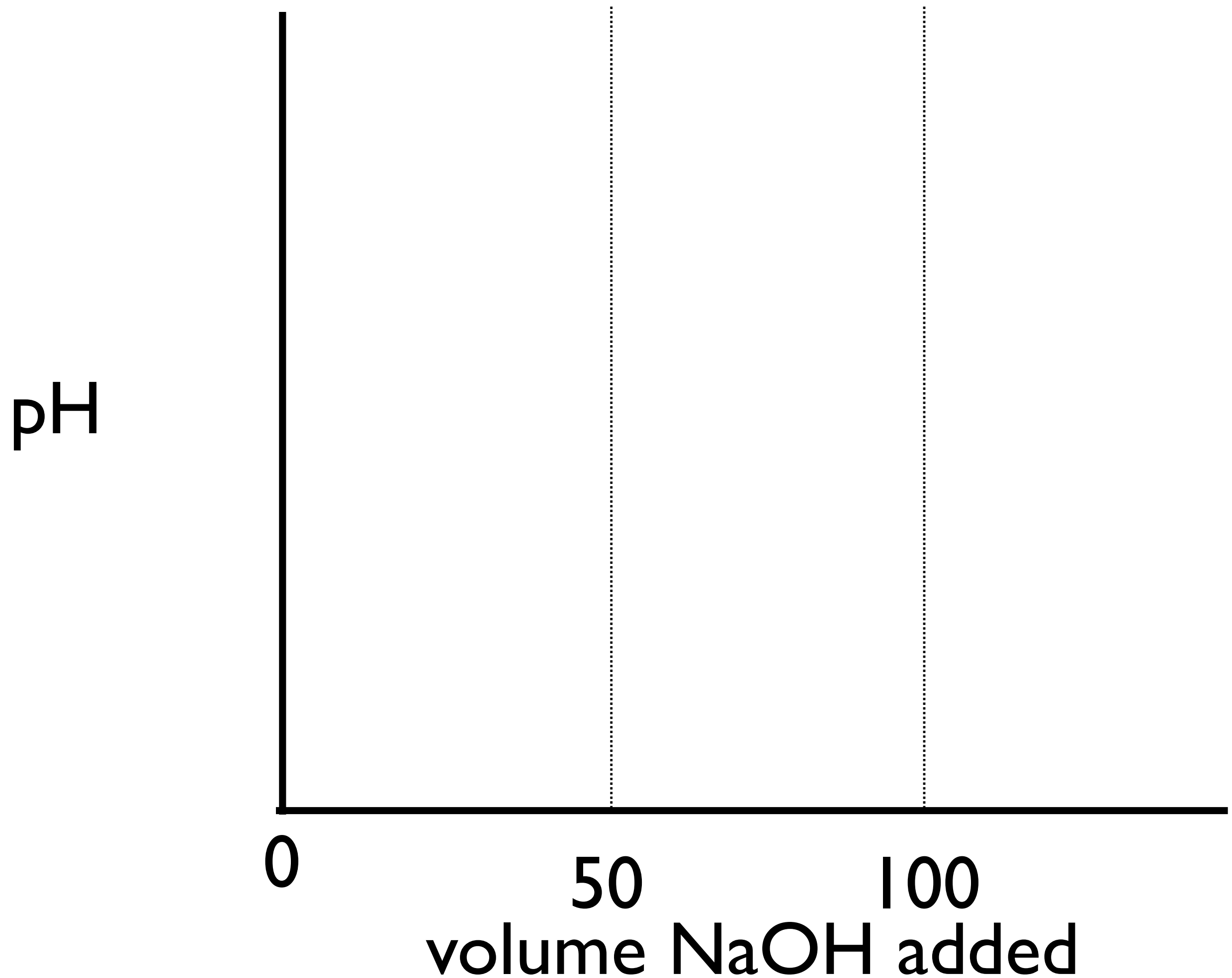
Imagine titration of 100 mL of 0.1 M acetic acid with 0.1 M NaOH

after the addition of 100 mL of base

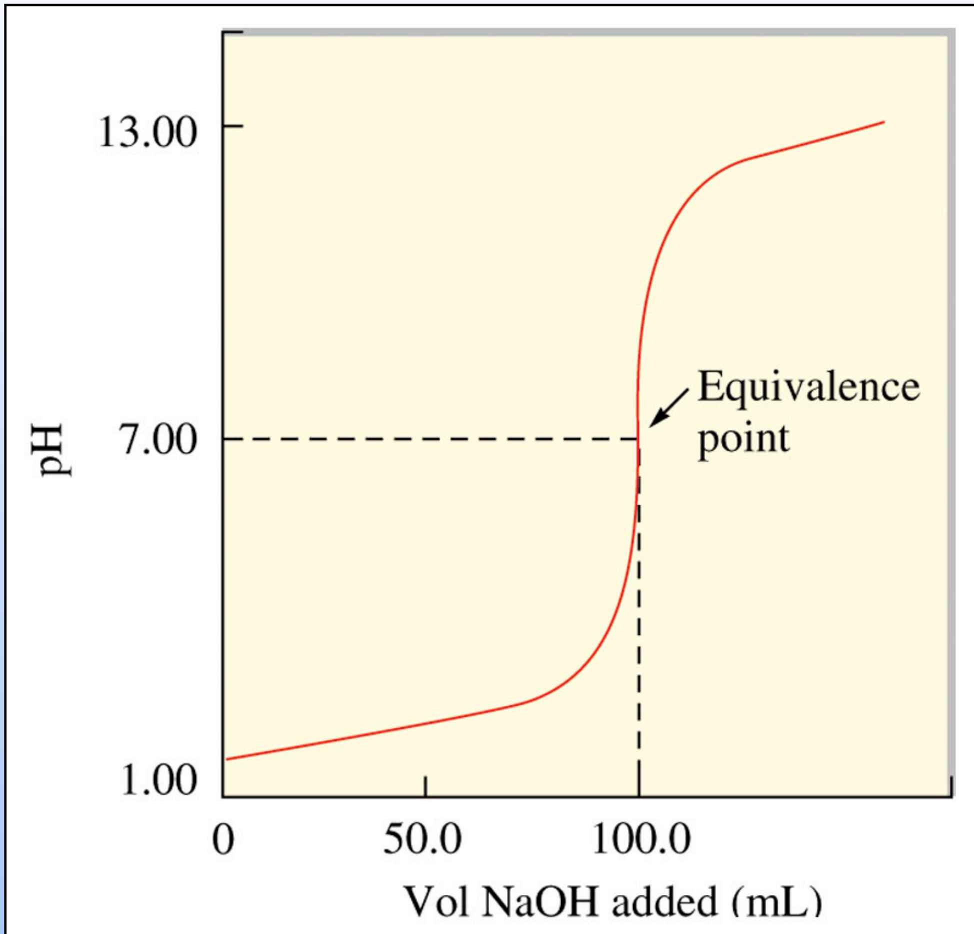
Imagine titration of 100 mL of 0.1 M acetic acid with 0.1 M NaOH

after the addition of 110 mL of base

Draw pH Curve



Strong Acid/Strong Base Titration



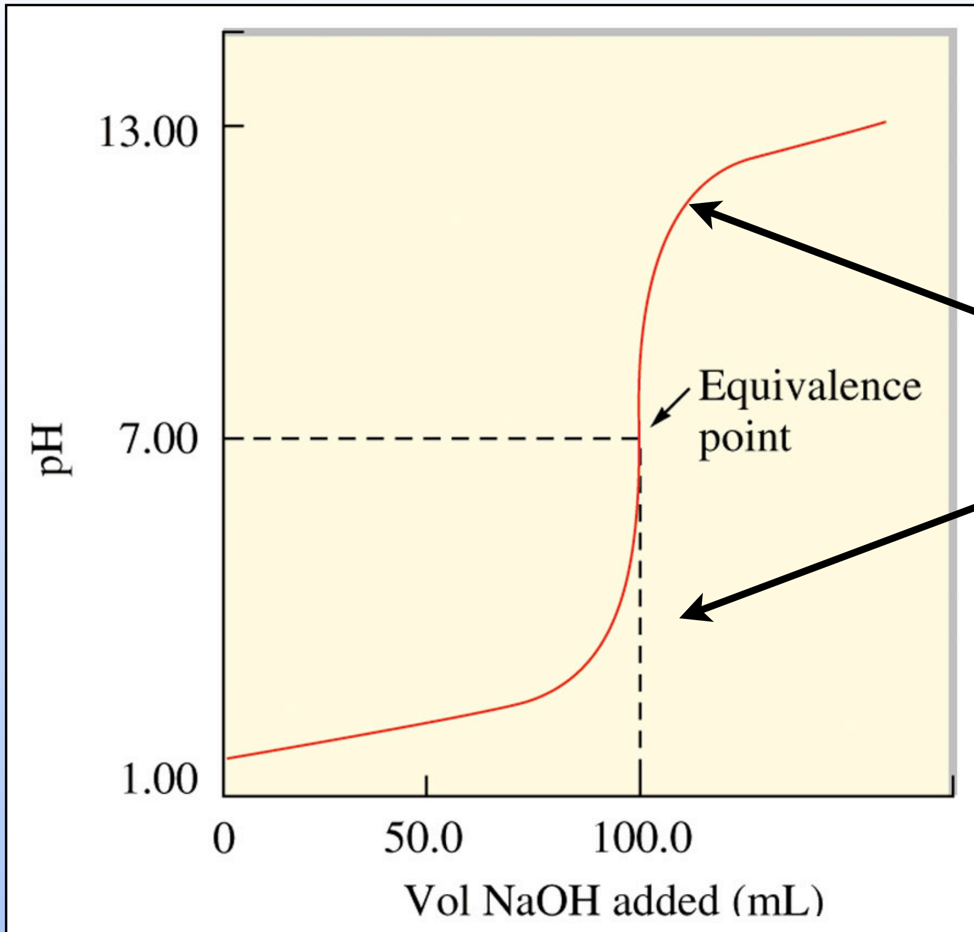
original solution 50 mL HCl
adding .1 M NaOH
at equivalence point

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