

How are we going to control this equilibrium?

Add HA                    shift to the "products"

Add A<sup>-</sup>                    shift to the "reactants"

Add H<sup>+</sup>                    shift to the "reactants"

Remove H<sup>+</sup>                shift to the "products"

## Neutralization

A solution can be neutralized (equal amounts of H<sup>+</sup> and OH<sup>-</sup>) 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                    There are .04 moles of H<sup>+</sup> .2M x .2L to neutralize you'll need .04 moles of OH<sup>-</sup>
- B. 200 mL
- C. 300 mL                    For that you'll need .4L of a .1M solution
- D. 400 mL                    Or you can look at it as the acid is twice as concentrated as the base therefore you'll need twice as much
- E. 500 mL

## Back to Buffers

$$\text{pK}_a = \text{pH} - \log \frac{[\text{A}^{-}]}{[\text{HA}]}$$

$$\text{K}_a = \frac{[\text{H}^{\text{+}}][\text{A}^{-}]}{[\text{HA}]}$$

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

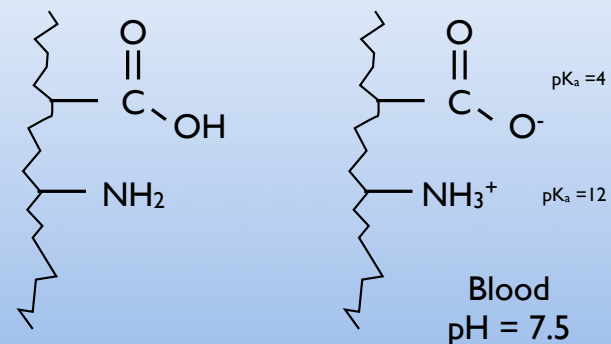
if  $[H^+] = K_a$ , then  $[HA] = [A^-]$

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



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<sup>-</sup>] = 0.1 M

add .001 moles of NaOH to 1L of solution

concentration of [HA] = .1 - .001 = 0.099  
concentration of [A<sup>-</sup>] = .1 + .001 = .101

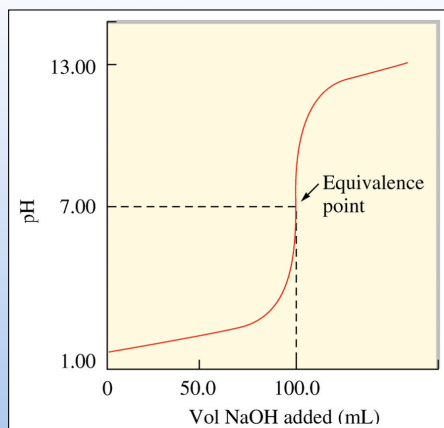
$$10^{-4.75} \rightarrow K_a = \frac{[H^+][A^-]}{[HA]} = \frac{[H^+](.101)}{0.099} \quad \text{pH} = 4.76$$

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

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

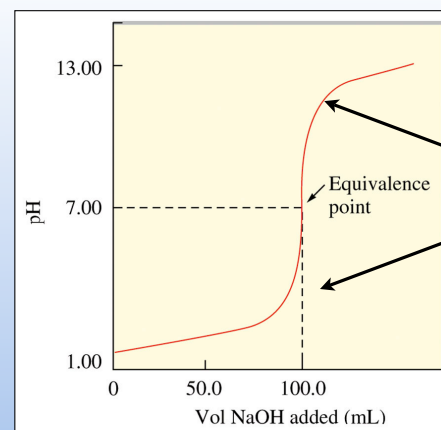
the only way to change the pH of the buffer system dramatically is to add enough acid or base to substantially change either the HA or A<sup>-</sup> concentrations

### 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<sup>-</sup>  
therefore the solution originally  
had 0.01 moles H<sup>+</sup>  
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<sup>+</sup> OH<sup>-</sup> is very small  
between pH 3 and pH 11