

If you know $[H^+]$ you know $[OH^-]$

$$K_w = [H^+][OH^-]$$

$$\log(K_w) = \log([H^+][OH^-])$$

$$\log(K_w) = \log[H^+] + \log[OH^-]$$

$$\log(10^{-14}) = \log[H^+] + \log[OH^-]$$

$$-14 = -\text{pH} - \text{pOH}$$

$$14 = \text{pH} + \text{pOH}$$

Weak Acid



	HA	H ⁺	A ⁻
I	C	0	0
C	-x	+x	+x
E	C-x	+x	+x

really 10⁻⁷



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

assuming $x \ll C$

$$x \sim \sqrt{K_a C}$$

This is a great simple result

$$[\text{H}^+] \approx \sqrt{K_a C_a}$$

C_a is the concentration of the acid
 K_a is the equilibrium constant for the acid

This assumes the concentration is large
and that K_a is small

What is the pH of a 1 M solution of weak acid with a $K_a = 10^{-6}$?

A. 1

B. 3

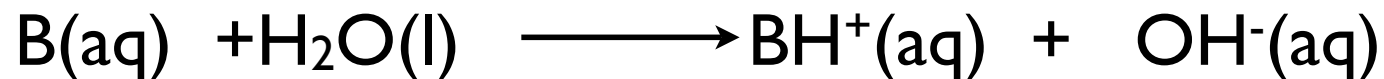
C. 7

D. 8

E. 9

$$[H^+] = \sqrt{1 \times 10^{-6}} = 10^{-3}$$

Weak Base

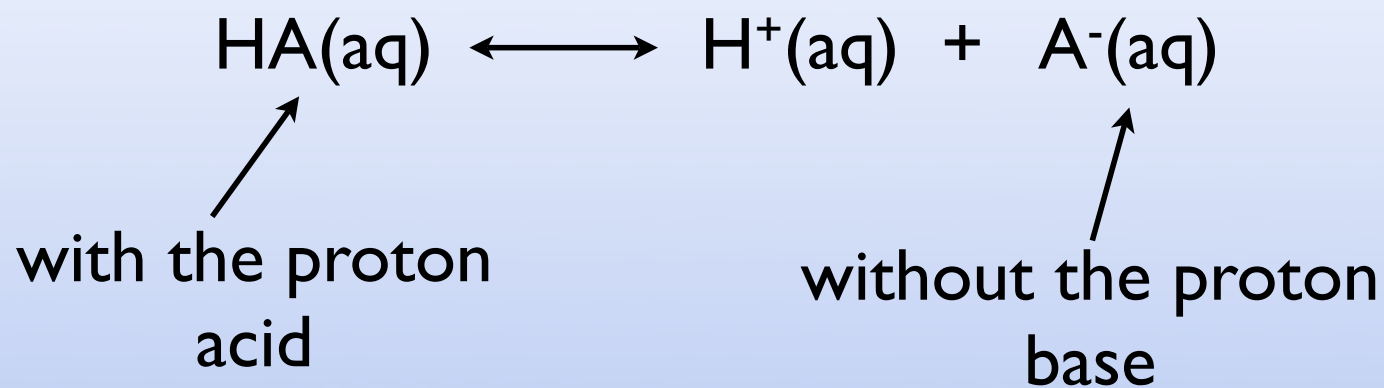


$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

identical result as before (same assumptions)

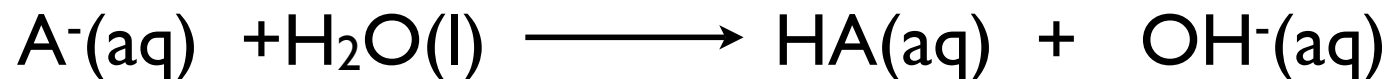
$$[\text{OH}^-] = \sqrt{K_b C_b}$$

Weak Acids



HA weak acid
A⁻ weak base

Weak Base

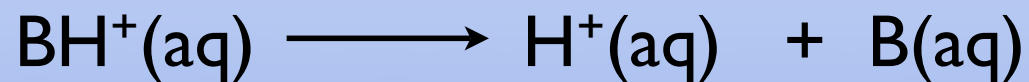
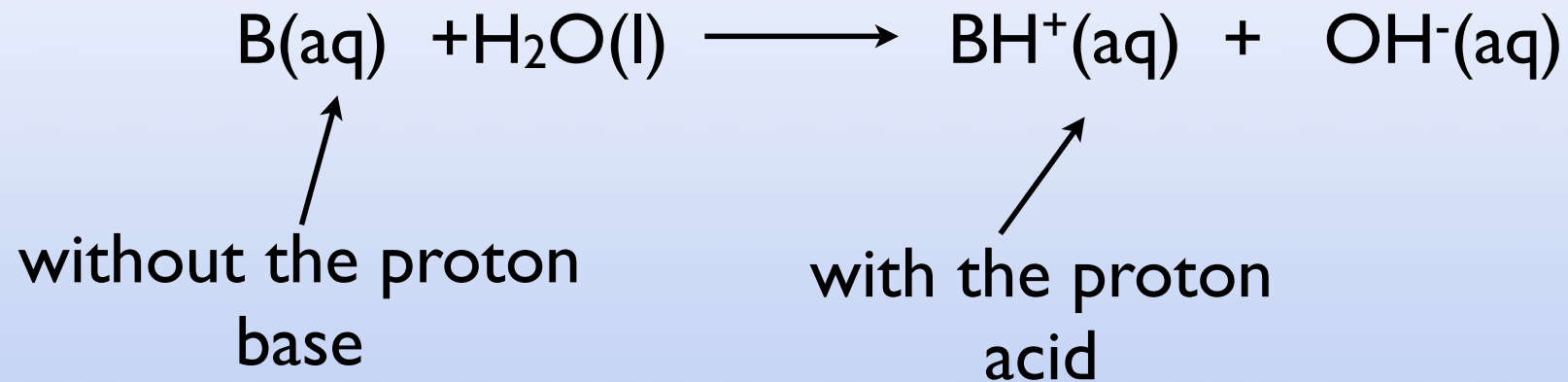


$$K_b = \frac{[\text{HA}][\text{OH}^{-}]}{[\text{A}^{-}]}$$

identical result as before (same assumptions)

$$[\text{OH}^{-}] = \sqrt{K_b C}$$

Same with the base



Weak acids

HA and BH^+

Name is acid HA “acetic acid”

BH^+ has a positive charge
and an “extra” proton NH_4^+

Weak bases

B and A^-

A^- is negative

usually name ends in “ate”

CH_3COO^- acetate

B is hardest to identify
it is not one of the other three
often it is an “amine”

What is the pH of a 1 M solution of sodium benzoate?

A. 4.9

B. 5.1

C. 6.2

D. 7

E. 9.09

benzoate is a weak base (A^-)
only pH of 9.09 is basic

How can I get K_b for benzoate
if I have K_a for benzoic acid?

Strength of Acids

The larger the K_a the more $[H^+]$

The larger the K_a the stronger the acid

$$\text{Since } K_a \times K_b = K_w$$

the larger the K_a the smaller the K_b

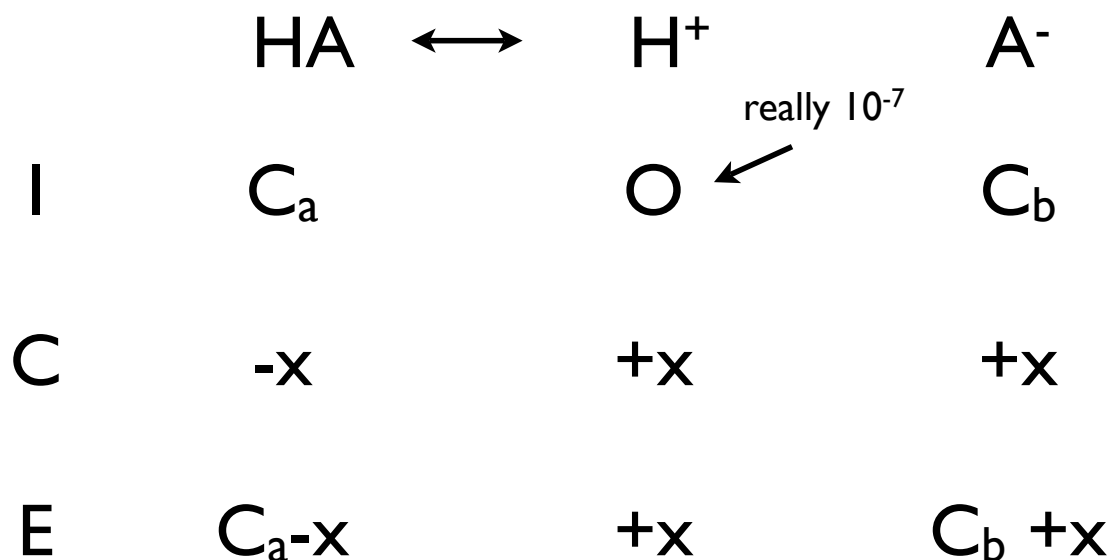
Strong Bases

Lithium Hydroxide	LiOH
Sodium Hydroxide	NaOH
Potassium Hydroxide	KOH
Rubidium Hydroxide	RbOH
Cesium Hydroxide	CsOH
Calcium Hydroxide	Ca(OH) ₂
Barium Hydroxide	Ba(OH) ₂
Strontium Hydroxide	Sr(OH) ₂

All Dissociate 100%

Identify Stuff from Cart

What if we have a solution with both
HA and A⁻ (or B and BH⁺)



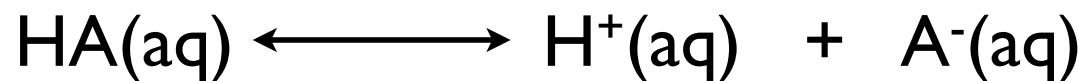
$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(x)(C_b+x)}{C_a-x} = \frac{(x)(C_b)}{C_a} \quad x \sim K_a \frac{C_a}{C_b}$$

assuming $x \ll C$

A solution with both the protonated and deprotonated form of a weak acid or weak base is a buffer (HA & A⁻ or BH⁺ & B)

What is the pH of a buffer that contains 1 mole of benzoic acid and 1 mole of sodium benzoate? $K_a = 6.5 \times 10^{-5}$

- A. 4.1
- B. 8.2
- C. 1.1
- D. 7
- E. 9.09



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^+ $.2\text{M} \times .2\text{L}$
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



How to deal with mixing acids and bases

First figure out what is in solution

Neutralize H^+ and any base
(neutralize OH^-) and any acid

What is left?

H^+

OH^-

HA

A^-

B

BH^+

HA and A^-

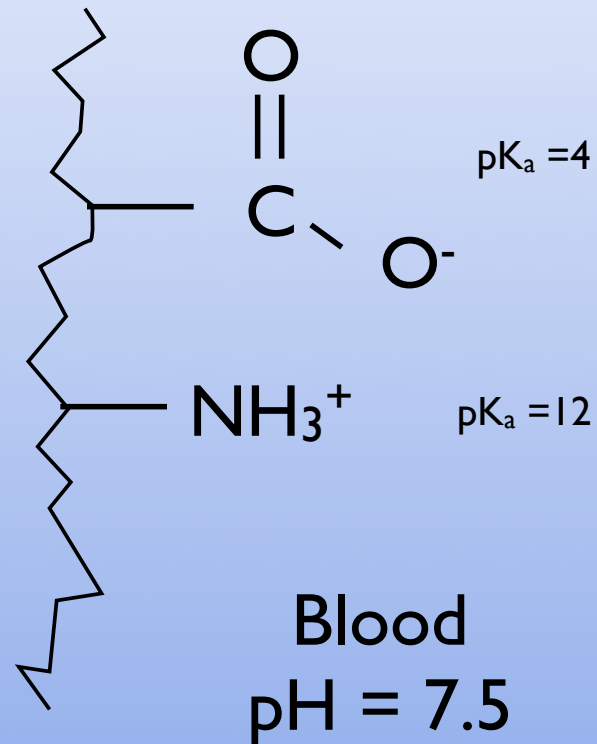
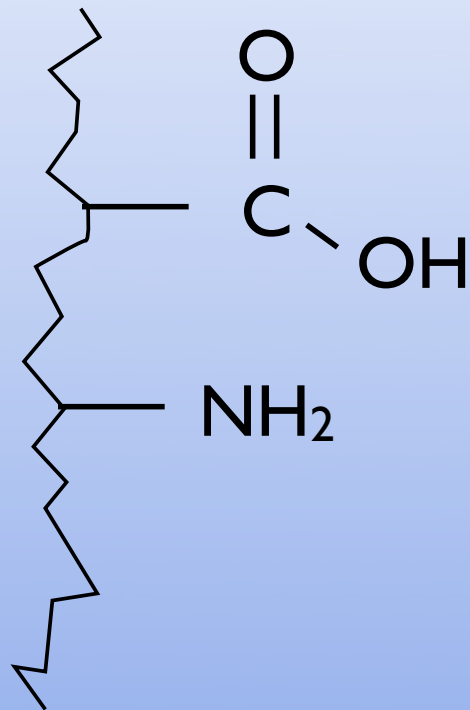
BH^+ and B

strong acid/strong base
neutralization example

weak acid/strong base
neutralization example

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

NaOH added to buffer

initial concentration of $[HA] = 0.1 \text{ M}$

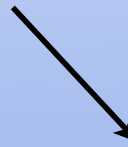
initial concentration of $[A^-] = 0.1 \text{ M}$

add .001 moles of NaOH to 1L of solution

concentration of $[HA] = .1 - .001 = 0.099$

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

$10^{-4.75}$


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

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⁻ concentrations