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

log → 10

ln → e

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

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

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

25°

$$\log(10^{-14}) = -pH - pOH$$

$$-14 = -pH - pOH$$

$$\Rightarrow 14 = pH + pOH$$

Weak Acid

$\_HA(aq) \longrightarrow \_H^+(aq) + \_A^-(aq)$ Dissociates →

	HA	H^+	A^-
I	C_A	0 (really 10^{-7})	0
C	-x	+x	+x
E	C_A-x	+x	+x

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(x)(x)}{C_A - x} \approx \frac{x^2}{C_A}$$

assuming x << C

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

This is a great simple result

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

small

$K_a > 10^{-7}$

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

conc. weak acid sol'n

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

A. 1

B. 3

C. 7

D. 8

E. 9

$$[H^+] = \sqrt{K_a C_A}$$

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

Weak Base

protonated

$$B(aq) + H_2O(l) \longrightarrow BH^+(aq) + OH^-(aq)$$

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

identical result as before (same assumptions)

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

BASE

BASE

Weak Acids

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

something else

with the proton
acid

without the proton
base

HA weak acid
A^- weak base

BASE

Weak Base

$$A^-(aq) + H_2O(l) \longrightarrow HA(aq) + OH^-(aq)$$

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

identical result as before (same assumptions)

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

Same with the base

$$B(aq) + H_2O(l) \longrightarrow BH^+(aq) + OH^-(aq)$$

without the proton
base

with the proton
acid

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

ACID

Weak acids

HA and BH^+ ← POSITIVE

Name is acid HA "acetic acid"

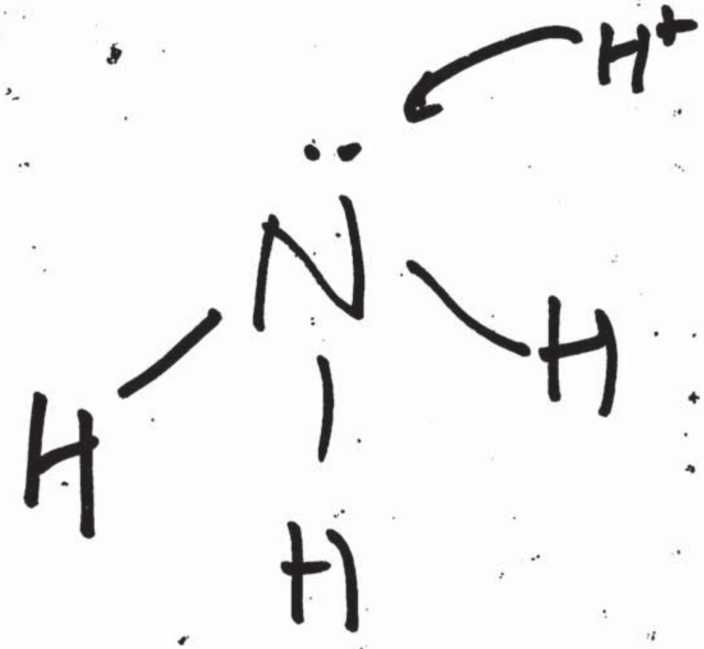
BH^+ has a positive charge
and an "extra" proton NH_4^+

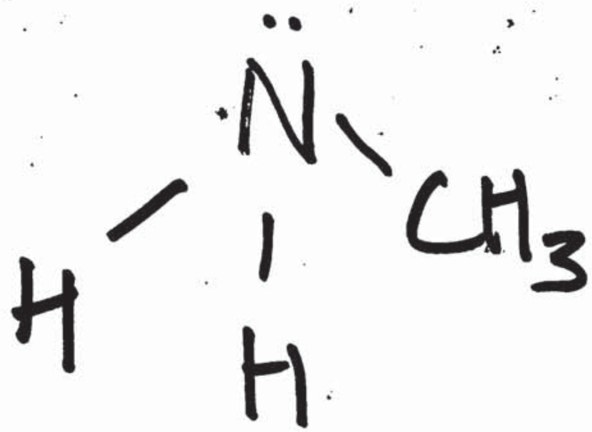
Weak bases

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 1M solution of sodium benzoate?

Na^+ A^-

A. 4.9

B. 5.1

C. 6.2

D. 7

E. 9.09

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

$$HA \longleftrightarrow H^+ + A^- \qquad K_a$$

$$+ \quad A^- + H_2O \longrightarrow HA + OH^- \qquad K_b$$

$$H_2O \longrightarrow H^+ + OH^- \qquad K_w$$

$$K_a \times K_b = K_w$$

Strength of Acids

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

The larger the K_a the stronger the acid

Since $K_a \times K_b = K_w$

the larger the K_a the smaller the K_b

Which dissociates more?

$$HA \longrightarrow H^+ + A^-$$

more H^+ ⟶ stronger acid.

larger K_a ⟶ stronger acid.

$$K_w = K_a K_b$$

CONST

$Mg(OH)_2$ small K_{sp}.

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	$Sr(OH)_2$

2X

All Dissociate 100%

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

	HA	$\longleftrightarrow$	H^+	+	A^-
I	C_A		$\emptyset$		C_B
+ C	$-x$		$+x$		$+x$
E	$C_A - x$		$+x$		$C_B + x$

$$K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(x)(C_B + x)}{(C_A - x)} \approx \frac{(x)(C_B)}{(C_A)}$$

$[H^+]$

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 of sodium benzoate? $K_a = 6.5 \times 10^{-5}$

HA → benzoic acid; A⁻ → sodium benzoate; $> 10^{-7}$

A. 4.1

B. 8.2

C. 1.1

D. 7

E. 9.09

$10^{-4.1}$

$10^{-1.1}$

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

$K_a = [H^+]$

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

How are we going to control this equilibrium?

Add HA	shift to the "products"
Add A^-	shift to the "reactants"
Add H^+ (STRONG ACID)	shift to the "reactants"
Remove H^+ (STRONG BASE)	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 HCl solution of HCl to neutralize it?

A. 100 mL

B. 200 mL

C. 300 mL

D. 400 mL

E. 500 mL

$.2 L \times .2 M = .04$ moles H^+

$.04$ moles OH^-

$.1 M \times V = .04$

$V = .4 L$

How to deal with mixing acids and bases

First figure out what is in solution

Neutralize H^+ and any base ← A^-, OH^-, B

(neutralize OH^-) and any acid ← H^+, HA, BH^+

What is left?

H^+ → STRONG ACID

OH^- → STRONG BASE

HA → WEAK ACID

A^- → WEAK BASE

B → WEAK BASE

BH^+ → WEAK ACID

HA and A^-

BH^+ and B

} BUFFER

weak acid/strong base neutralization example

$\underline{A^-}$ ↙

100 mL of .1 M $Na(CH_3COO)$

+ 50 mL of ~~.22~~ 0.1 M ~~[illegible]~~ . HCl

pH = ???

$.1\,L \times .1\,M = .01$ moles A^-

$.05\,L \times .1\,M = .005$ moles H^+

$$A^- + H^+ \longrightarrow HA$$

	A^-	H^+	HA
i	.01	.005	0
c	−.005	−.005	+.005
	.005	0	.005

$$K_a \approx \frac{(x)C_B}{C_A}$$

$$[H^+] = x = K_A \frac{C_A}{C_B}$$

Buffer

$$C_A = C_B$$

$$[H^+] = K_a \frac{C_A}{C_B} \quad 1$$

$$\log[H^+] = \log K_a$$

$$-pH = -pK_a$$

$$pH = pK_a \leftarrow ?? \; -\log K_a$$

$$C_A = C_B$$

$$[H^+] = K_a \frac{C_A}{C_B}$$

$$\log[H^+] = \log K_a$$

$$-pH = -pK_a$$

$$pH = pK_a \leftarrow \therefore -\log K_a$$

$$HA \longrightarrow H^+ + A^-$$

protonated   deprotonated

$$[HA] = [A^-] \qquad pH = pK_a$$

$$[HA] > [A^-] \qquad pH < pK_a$$

$$[HA] < [A^-] \qquad pH > pK_a$$