#17.32, p727

Find the pH and the degree of ionization for an 0.10 M solution of formic acid: $HCHO_2$

$K_a = 1.7 \times 10^{-4}$

$$HCHO_2 + H_2O \rightleftharpoons CHO_2^- + H_3O^+$$

$$K_a = 1.7 \times 10^{-4} = \frac{[CHO_2^-][H_3O^+]}{[HCHO_2]}$$

	Initial	Change	Equilibrium
$[CHO_2^-]$	0	$+x$	x
$[H_3O^+]$	0	$+x$	x
$[HCHO_2]$	0.10	$-x$	$0.10 - x$

After writing the equation for acid dissociation and the equilibrium expression, we write all concentrations in terms of one variable ... x

$$1.7 \times 10^{-4} = \frac{x^2}{0.10 - x}$$

$$1.7 \times 10^{-4} = \frac{x}{0.10 - x}$$

Assume x is much smaller than 0.10 ... so that 0.10 - x = 0.10

$$\downarrow$$

$$1.7 \times 10^{-4} = \frac{x^2}{0.10}$$

$$[H_3O^+] = x = 0.004123$$

x IS much smaller than 0.10

$$pH = -\log(0.004123) = 2.38 = pH$$

Degree of ionization? DEGREE OF IONIZATION is the fraction of a weak electrolyte (acid or base) that dissociates in water.

$$\frac{[CHO_2^-]}{[HCHO_2]} = \frac{[H_3O^+]}{[HCHO_2]} = \frac{0.004123}{.10} = 0.041 = DOI$$

Sometimes, we express degree of ionization as a percent ... PERCENT IONIZATION

$$\% = DOI \times 100\% = 4.1\% \text{ dissociated}$$

... so about 96% of this acid is present as UNDISSOCIATED molecules!

#17.46, p 728

An aqueous solution of 0.25 M trimethylamine has a pH of 11.63. What's the value of Kb?

$(CH_3)_3N$

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

$$K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$

	Initial	Change	Equilibrium
$[(CH_3)_3NH^+]$	0	$+x$	x
$[OH^-]$	0	$+x$	x
$[(CH_3)_3N]$	0.25	$-x$	$0.25-x$

$$\frac{x^2}{0.25-x} = K_b$$

For this problem, we need to first solve for x in order to find Kb. Use pH to find pOH, then use pOH to find [OH-] ... which equals x.

$$\frac{x^2}{0.25 - x} = K_b$$

$pH = 11,63$

$pOH = 14,00 - 11,63 = 2,37$

$x = [OH^-] = 10^{-2,37} = 4.265795 \times 10^{-3}$

$$K_b = \frac{(4.265795 \times 10^{-3})^2}{0,25 - (4.265795 \times 10^{-3})} = \boxed{7,4 \times 10^{-5}}$$