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

$K_a = 1.7 \times 10^{-4}$ $\quad HCHO_2 + H_2O \rightleftharpoons H_3O^+ + CHO_2^-$

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

	Initial	Δ	Equilibrium
$[H_3O^+]$	0	$+X$	X
$[CHO_2^-]$	0	$+X$	X
$[HCHO_2]$	0.10	$-X$	$0.10 - X$

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

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

$$1.7\times10^{-4} = \frac{x^2}{0.10 - x}$$

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

$$1.7\times10^{-4} = \frac{x^2}{0.10}$$

$$x = 0.004123 = [H_3O^+]$$ 'x' is indeed much smaller than 0.10!

$$pH = -\log_{10}(0.004123) = 2.38$$

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

$$\frac{[CHO_2^-]}{[HCHO_2]} = \frac{[H_3O^+]}{[HCHO_2]} = \frac{0.004123}{0.10} = \boxed{0.041 = DOI}$$

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

$$\% = DOI \times 100\% = 0.04123 \times 100\% = \boxed{4.1\% \text{ dissociated}}$$

... so about 96% of this acid exists in solution as undissociated formic acid molecules.

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

$$K_b = \frac{(x)(x)}{0.25-x}$$

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

So we need to find 'x' if we want to know what Kb is!

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

$pH = 11.63$ $[OH^-] = x$

$pH + pOH = 14.00$

$pOH = 2.37$

so, $[OH^-] = 10^{-2.37} = 0.004266\ M$

$x = 0.004266\ M$

$$K_b = \frac{(0.004266)^2}{(0.25 - 0.004266)} = \boxed{7.4 \times 10^{-5} = K_{b,(CH_3)_3N}}$$