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

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

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = 1.8\times10^{-4} \text{ (Appendix H, OpenStax)}$$

Set up an equilibrium chart to reduce the number of variables in the Ka expression.

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

Let "x" equal the change in hydronium ion concentration.

$$\frac{(x)(x)}{(0.10-x)} = 1.8\times10^{-4}$$

Assume $x \ll 0.10$

$0.10 - x \approx 0.10$

$$\frac{x^2}{0.10} = 1.8\times10^{-4}$$

$x = 0.004242640687 = [H_3O^+]$

So, $\boxed{pH = 2.37}$

... but what about degree of ionization?

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

Degree of ionization is the fraction of a weak acid or base that ionizes in solution.
Here, that's:

$$\frac{[CHO_2^-]}{[HCHO_2]_{nominal}} = \frac{[H_3O^+]}{[HCHO_2]_{nominal}}$$

$$\frac{0.004242640\mathrm{7}\ M\ H_3O^+}{0.10\ M\ HCHO_2} = \boxed{0.042 = D.O.I\ \text{for}\ 0.10\ M\ HCHO_2}$$

Often, we express DOI as a percentage ... call PERCENT IONIZATION;

% ionization = DOI x100 = 4.2% ionized

From the pH lab: By Le Chateleir's Principle, adding water to the equilibrium should force it to the right - meaning that more acid will ionize - even as the pH goes up!. Therefore, the degree of (or percent) ionization should INCREASES as the concentration of the acid DECREASES. Check this with your experimental data on acetic acid solutions.

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

$(CH_3)_3N$

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

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

We don't know Kb ... so how do we proceed? Let's make an equilibrium chart to reduce the number of variables!

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

Let "x" equal the change in hydroxide ion concentration

Plug back into Kb expression:

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

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

Now what? We have fewer variables now, but we still have two!

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

If we had some way to determine "x", we could use it to find Kb! We can use pH identities and the solution's measured pH of 11.63 to find the hydroxide ion concentration ... which is equal to "x".

$$pH + pOH = 14.00$$
$$11.63 + pOH = 14.00$$
$$pOH = 2.37$$

$$[OH^-] = 10^{-pH}$$
$$[OH^-] = 10^{-2.37}$$
$$[OH^-] = 0.004265795 2 = x$$

We've found "x" ... just plug back into the Kb expression to find our answer, Kb.

$$K_b = \frac{(x)(x)}{(0.25 - x)} = \frac{x^2}{0.25 - x} = \frac{(0.0042657952)^2}{0.25 - 0.0042657952}$$

$$= \boxed{7.4 \times 10^{-5} = K_b}$$