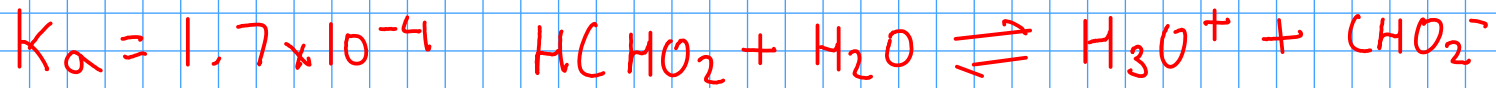


Find the pH and the degree of ionization for an 0.10 M solution of formic acid: HCHO_2



$$K_a = 1.7 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][\text{CHO}_2^-]}{[\text{HCHO}_2]}$$

	Initial	Δ	Equilibrium
$[\text{H}_3\text{O}^+]$	0	+X	X
$[\text{CHO}_2^-]$	0	+X	X
$[\text{HCHO}_2]$	0.10	-X	0.10 - X

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

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

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

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

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

$$x = 0.004123 = [\text{H}_3\text{O}^+] \quad \text{'x' is indeed much smaller than 0.10!}$$
$$\text{pH} = -\log_{10}(0.004123) = 2.38$$

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

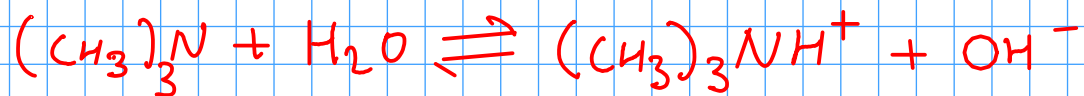
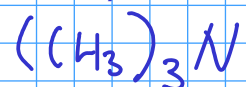
$$\frac{[\text{CHO}_2^-]}{[\text{HCHO}_2]} \approx \frac{[\text{H}_3\text{O}^+]}{[\text{HCHO}_2]} = \frac{0.004123}{0.10} = \boxed{0.041 = \text{DOI}}$$

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

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

... 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 K_b ?



$$K_b = \frac{[(\text{CH}_3)_3\text{NH}^+][\text{OH}^-]}{[(\text{CH}_3)_3\text{N}]}$$

	Initial	Δ	Equilibrium
$[(\text{CH}_3)_3\text{NH}^+]$	0	+ x	x
$[\text{OH}^-]$	0	+ x	x
$[(\text{CH}_3)_3\text{N}]$	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 K_b is!

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

$$pH = 11.63$$

$$[OH^-] = x$$

$$pH + pOH = 14.00$$

$$pOH = 2.37$$

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

$$x = 0.004266 \text{ M}$$

$$K_b = \frac{(0.004266)^2}{(0.25 - 0.004266)} = 7.4 \times 10^{-5} = K_b, (C_4H_9)_3N$$