Compare:

- Weak acid HNO_2 : pH of 0.10 M solution = 2.17

Let's compare the pH of the weak nitrous acid with the pH of a strong acid like nitric acid:

0.10 M HNO_3, what is pH?

$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$

0.10 M HNO_3, $[H_3O^+] = 0.10$ M

pH = 1.00

The stronger the acid:

- the lower the pH of a solution of given concentration will be
- the higher the concentration of hydronium ion (when compared to the nominal acid concentration)

Consider an 0.100 M solution of the weak base ammonia:

NH_3 ; $K_b = 1.8 \times 10^{-5}$ (p A-14, Ebbing 9th)

What is the pH?

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^- \quad K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$$

Species	[Initial]	Δ	[Equilibrium]
NH_4^+	0	+X	X
OH^-	0	+X	X
NH_3	0.100	-X	0.100-X

Let "x" equal the change in ammonium concentration...

$$\frac{(x)(x)}{(0.100-x)} = 1.8 \times 10^{-5}$$

$$\frac{x^2}{0.100-x} = 1.8 \times 10^{-5}$$

$x << 0.100$, so

$0.100 - x \approx 0.100$

$$\frac{x^2}{0.100} = 1.8 \times 10^{-5}$$

$x = 0.001341640\,8 = [OH^-]$ (see chart)

$pOH = -\log(0.0013416408) = 2.87$

$pH + pOH = 14.00$, so

$pH + 2.87 = 14.00$

$$\boxed{pH = 11.13}$$

If you'd solved the quadratic for this one, you'd have gotten pH = 11.13 ...

Compare pH to the pH of an 0.100 M solution of the strong base NaOH:

$pH_{NH_3} = 11.13$

$NaOH \rightarrow Na^+ + OH^-$

So, 0.100 M NaOH, $[OH^-] = 0.100$

$pOH = -\log_{10}(0.100) = 1.00$

$pH = 14.00 - 1.00 = \boxed{13.00}$

The stronger the base:

- the higher the pH will be for a solution of given concentration
- the higher the HYDROXIDE concentration (compared to the nominal base concentration)

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

$HCHO_2 + H_2O \rightleftharpoons H_3O^+ + CHO_2^-$ $K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = 1.7 \times 10^{-4}$ ← Ka from page A-13 in E&G textbook...

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 concentration ...

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

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

$x << 0.10$, so

$0.10 - x \simeq 0.10$

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

$x = 0.004123105 6 = [H_3O^+]$

$$\boxed{pH = 2.38}$$

So the pH is 2.38 ... now what about degree of ionization?

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

What is degree of ionization? The fraction of a weak acod or base that ionizes in solution.

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

$$DOI \simeq \frac{0.0041231056}{0.10} \simeq \boxed{0.041 \text{ for } 0.10\,M\ HCHO_2}$$

Often, degree of ionization is expressed as a percentage ... called percent ionization:

$$\%\,\text{Ionization} = DOI \times 100$$

$$\%\,\text{Ionization} = 0.041 \times 100 = \boxed{4.1\%\ \text{ionized}}$$

When you do Experiment 16A. 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 experiment 16A data on acetic acid.

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

$(CH_3)_3N$

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^- \; ; \; K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$

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

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

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

Since x represents hydroxide ion concentration, we can find "x' by looking at the pH.

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

$pOH = 2.37$

$[OH^-] = 10^{-2.37} = 0.0042657952$

$x = 0.0042657952$

Plug value of "x" into the expression we wrote for Kb...

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

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

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