Compare:

- Weak acid HNO_2 : pH of 0.10 M solution = 2.15

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

0.10 M HNO_3, what is pH?

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

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

$$pH = -\log_{10}(0.10) = 1.00$$

The stronger the acid:

- the lower the pH of a solution of given concentration will be
- the higher the concentration of hydronium ion .

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?

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

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

We beed to solve this, but which term are we most interested in?

We want to solve for the hydroxide concentration, because it can be converted to hydronium concentration (and pH)!

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
NH_4^+	0	$+x$	x
OH^-	0	$+x$	x
NH_3	0.100	$-x$	$0.100 - x$

Plug in to the equilibrium expression:

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

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

This is a QUADRATIC EQUATION. We can solve it with the quadratic formula, OR we can notice that that the value of 'x' should be much smaller than 0.100!

Assume $x << 0.100$, so $0.100 - x \approx 0.100$

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

$x = 0.001341640 8 = [OH^-]$

This is HYDROXIDE ion concentration, *NOT* hydronium!

$pOH = -\log_{10}(0.0013416408) = 2.87$

$pH = 14.00 - 2.87$

$$\boxed{pH = 11.13}$$

If you had solved this problem with the quardatic equation, you would have obtained a pH of 11.13 - no difference at all to two significant figures!

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

$pH_{NH_3} = 11.13$

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

0.100 M NaOH ; $[OH^-] = 0.100$

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

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

The stronger the base:

- the higher the pH will be for a solution of given concentration
- the higher the HYDROXIDE concentration
- the lower the HYDRONIUM concentration

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

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

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

Constant's value at 25C obtained from the chart on p A-13, Ebbing 9th ed

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
H_3O^+	0	$+x$	x
CHO_2^-	0	$+x$	x
$HCHO_2$	0.10	$-x$	$0.10 - x$

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

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

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

$$\Big\downarrow \; x << 0.10$$

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

$$x = 0.004123 1056 = [H_3O^+]$$

$$pH = -\log_{10}(0.004123 1056) = \boxed{2.38 = pH}$$

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

$$\frac{[CHO_2^-]}{[HCHO_2]_{orig}} = \frac{[H_3O^+]}{[HCHO_2]_{orig}} = \frac{0.004123 1056}{0.10} = \boxed{0.041 = D.O.I.}$$

Sometimes, we express degerr of ionization as a percent and call it PERCENT IONIZATION:

$$\% = D.O.I. \times 100\% = \boxed{4.1\% \text{ ionized}}$$

Check this in lab - expt 16A: A more dilute acid solution should have a HIGHER degree of ionozationthanks to Le Chateleir's Principle - even if the dilution causes the pH of the acid solution to decrease overall!

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 CONC	CHANGE	EQUILIBRIUM CONC
$(CH_3)_3NH^+$	0	+ x	x
OH^-	0	+ x	x
$(CH_3)_3N$	0.25	− x	0.25 − x

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

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

If we want to find the value of Kb, then we need to come up with some other way (than solving the quadratic equation) of finding 'x'.

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

We know from our setup that 'x' equals the hydroxide ion concentration. Since concentration of hydroxide is reated to pH, we can find 'x' through the pH

$$x = [OH^-]$$

$$pH + pOH = 14.00$$

$$11.63 + pOH = 14.00$$

$$pOH = 2.37$$

$$x = 10^{-2.37} = 0.0042657952$$

Now, plug 'x' back into the Kb expression:

$$Kb = \frac{x^2}{(0.25 - x)} = \frac{(0.0042657952)^2}{0.25 - 0.0042657952} = \boxed{7.4 \times 10^{-5} = Kb}$$