Consider an 0.100 M solution of the weak base ammonia:

NH_3 ; $K_b = 1.75 \times 10^{-5}$

What is the pH?

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$K_b = 1.75 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Which term in this expression are we really interested in? Solve to get the HYDROXIDE concentration, since it can be easily related to hydronium (and pH).

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

Plug into the equilibrium expression

$$1.75 \times 10^{-5} = \frac{(x)(x)}{(0.100 - x)} = \frac{x^2}{0.100 - x}$$

Solve for 'x':

$$1.75 \times 10^{-5} = \frac{x^2}{0.100 - x}$$

This is a QUADRATIC EQUATION. But, we expect that 'x' will be small compared to 0.100. So we can simplify this equation

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

$$1.75 \times 10^{-5} = \frac{x^2}{0.100}$$

$0.001322876 = x = [OH^-]$ HYDROXIDE ion concentration!

$$-\log_{10}(0.001322876) = 2.88 = pOH$$

$$pH + pOH = 14.00$$

$$\text{so, } pH = 14.00 - 2.88 = \boxed{11.12}$$

If you had used the quadratic equation to solve this problem, you would have gotten a pH of 11.12 - no difference from this method, at least to two significant figures!

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

$pH_{NH_3} = 11.12$

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

$[OH^-] = 0.100$

$pOH = 1.00$

$pH + pOH = 14.00$

$pH = 13.00$

The higher the Ka or Kb value, the stronger the acid or base!

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

$$HCHO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CHO_2^-(aq)$$

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

Constant's value at 25 C obtained from chart in textbook, page A-13

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

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

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

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

Assume that x is much smaller than 0.10

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

This number is indeed much smaller than 0.10

$$x = 0.004123106 = [H_3O^+]$$

$$pH = -\log_{10}(0.004123106) = 2.38 = pH$$

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

$$\frac{[CHO_2^-]}{[HCHO_2]} = \frac{[H_3O^+]}{[HCHO_2]} = \frac{0.004123106}{0.10} = 0.041 = D.O.I.$$

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

$$\% = DOI \times 100\% = 0.041 \times 100\% = 4.1\% \text{ ionized}$$

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

(WEAK acids exist in solution mostly as undissociated molecules!)

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

$(CH_3)_3N$

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

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

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

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

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

If we want to know what Kb is, we need to find the value of 'x', but NOT by solving this equation.

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

$x = [OH^-]$

... but concentration of hydroxide is related to pH

$$pH + pOH = 14.00$$

$$11.63 + pOH = 14.00$$

$$pOH = 2.37$$

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

$$[OH^-] = 0.004265795 2$$

So, $x = 0.0042657952$

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

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