For a WEAK ACID, equilibrium does not lie far to the right. The ionization equilibrium of the acid itself is important!

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$\underbrace{K_a}_{\text{acid ionization constant}} = \frac{[H_3O^+][A^-]}{\underbrace{[HA]}_{\text{(HA) = concentration of undissociated acid}}}$$

Again, water's concentration will not change significantly, so it is folded into the ionization constant

For a WEAK BASE, equilibrium does not lie far to the right. The ionization equilibrium of the base itself is important!

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

$$\underbrace{K_b}_{\text{base ionization constant}} = \frac{[BH^+][OH^-]}{[B]}$$

Values for Ka and Kb can often be found in data books / tables / or on the web.

In Ebbing, this data is in the appendices, on pages A-13 and A-14

WEAK ELECTROLYTES

- In solutions of weak acids or bases, the UNDISSOCIATED form is present in significantly high concentration.

- The pH of a solution of weak acid will be HIGHER than the pH of a strong acid solution with the same nominal concentration!

- The pH of a solution of weak base will be LOWER than the pH of a strong base solution with the same nominal concentration!

Consider a 0.100M solution of the WEAK ACID HNO_2

$$HNO_2 + H_2O \rightleftharpoons H_3O^+ + NO_2^-$$

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 5.1 \times 10^{-4}$$

values for Ka are determined experimentally

What is the pH of the solution?

To find the pH, we need to determine the concnetration of hydronium, $[H_3O^+]$

... so we need to solve the equilibrium expression. But we don't know all of the concentrations AT EQUILIBRIUM to do so!

... but they ARE related!

We assume the amount of hydronium from the water is small enough to ignore

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
$[H_3O^+]$	0	$+x$	x
$[NO_2^-]$	0	$+x$	x
$[HNO_2]$	0.100	$-x$	$0.100 - x$

... this is similar to the problems from the equilibrium chapter!

$$5.1 \times 10^{-4} = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

$$5.1 \times 10^{-4} = \frac{(x)\,(x)}{(0.100 - x)}$$

$$5.1 \times 10^{-4} = \frac{x^2}{0.100 - x}$$

Quadratic equation!

$$ax^2 + bx + c = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Assume that x << 0.100

What's this?
For situations where the amount of dissociated acid or base is much, much smaller than the original amount, it's safe to assume that the amount of undissociated acid remains relatively constant.

In this case, 0.100 - x is essentially the same thing as 0.100 .

$$5.1 \times 10^{-4} = \frac{x^2}{0.100}$$

$$x^2 = 5.1 \times 10^{-5}$$

$$x = 7.14 \times 10^{-3} = [H_3O^+]$$

$$pH = \boxed{2.15}$$

... if we'd used the quadratic equation, our answer would have been pH = 2.16

Compare:

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

- Strong acid: pH of 0.10 M solution = 1.00

The stronger the acid, the lower the pH of a solution of given concentration will be!

Consider an 0.100 M solution of the weak base ammonia:

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

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

What is the pH?

$$K_b = 1,75 \times 10^{-5} = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

We must solve this ... for?

... for HYDROXIDE, since we can convert that to hydronium to find 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"!

$$1.75\times10^{-5} = \frac{x^2}{0.100-x}$$

This is a quadratic, BUT ... x should be small compared to 0.1

Assume that x << 0.100, so 0.100 -x = 0;100

$$1.75\times10^{-5} = \frac{x^2}{0.100}$$

$$0.00132288 = x = [OH^-]$$

This number, x, is the HYDROXIDE concentration.

$$-\log_{10}[OH^-] = 2.88 = pOH$$

Now, we can find pH ...

$$pH + pOH = 14.00$$

so,

$$\boxed{pH = 11.12}$$

If you had used the quadratic equation to solve this problem, you would have calculated the pH to be <u>11.12</u>

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

$pH_{NH_3} = 11.12$

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

$[OH^-] = 0.100$

$pOH = 1.00$

$pH + pOH = 14.00$

$\boxed{pH = 13.00}$

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