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!

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

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

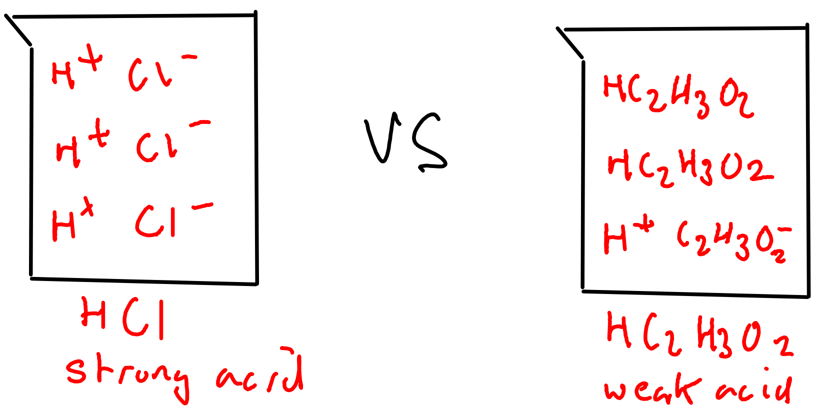
Values for Ka and Kb can often be found in data books / tables / or on the web.

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!



Fewer molecules of the weak acid ionize, so the concnetration of hydrogen/hydronium ion is lower, meaning a higher pH!

- 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 nitrous acid, a WEAK ACID (HNO_2)

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

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 4.5 \times 10^{-4}$$

Found on page A-14 in Ebbing 10th edition. These K values are determined experimentally like other equlibrium constants.

What is the pH of the solution?

To find pH, we need to determine the concentration of hydronium ion at equilibrium, but this time we can't assume all the acid ionizes! So we need to solve the equilibrium expression for the acid.

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	$+x$	x
NO_2^-	0	$+x$	x
HNO_2	0.100	$-x$	$0.100 - x$

$$\frac{(x)(x)}{(0.100 - x)} = 4.5 \times 10^{-4}$$

Similar to other equilibrium problems we've seen in Chapter 14!

$$\frac{(x)(x)}{(0.100-x)} = 4.5 \times 10^{-4}$$

This is a quadratic, We can solve it with the quadratic equation:

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

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

$$\frac{x^2}{0.100-x} = 4.5 \times 10^{-4}$$

Ka is small, so only a small fraction of the acid will ionize. That means 'x', which represents how much acid ionizes, will be small relative to 0.1. If that's true, then ...

$$0.100 - x \approx 0.100$$

When is it safe to assume 'x' is small?
When the initial concentration and the K differ by about 1000x.
ex: Ka/(HA) >= 1000

$$\frac{x^2}{0.100} = 4.5 \times 10^{-4}$$

$$x^2 = 4.5 \times 10^{-5}$$

$$x = 0.006708204 = [H_3O^+]$$

$$\boxed{pH = 2.17}$$

(Solving the quadratic gives a pH of 2.19)

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 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 = 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?

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

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

We need to solve this expression, but which term will help us find pH?

We want to solve for HYDROXIDE concentration, as it can be easily converted to pH.

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

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

$$1.8\times10^{-5} = \frac{x^2}{0.100 - x}$$

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

This is a quadratic, but it can be simplified since 'x' is small compared to 0.100 ...

$$1.8\times10^{-5} = \frac{x^2}{0.100}$$

$$1.8\times10^{-6} = x^2$$

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

Be careful here! We have calculated HYDROXIDE concentration, not hydronium, so we can't just take the negative log and say we're done! We'll find pOH and convert to pH...

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

$$pOH = 2.87$$

Since $pH + pOH = 14.00$

So, $pH = 14.00 - 2.87 = \boxed{11.13}$

Solving this with the full quadratic equation gets you a pH of ... 11.13