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!

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

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

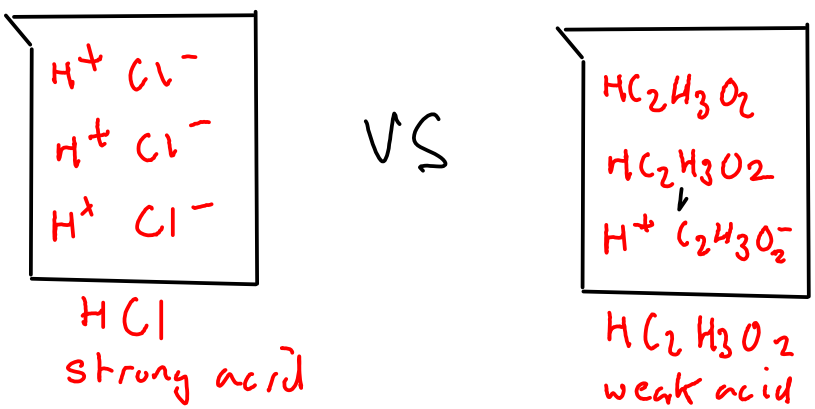
Values for Ka and Kb can often be found in data books / tables / or on the web.

In OpenStax, these constants are in Appendix H and Appendix I!

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)

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

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 4.6 \times 10^{-4}$$

See Appendix H in OpenStax for Ka values

What is the pH of the solution?

We will assume that all hydronium ion in solution comes from the acid (nitrous acid). While nitrous acid is a weak acid, it's FAR stonger than water itself! Set up an equilibrium chart.

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

Let "x" equal the change in hydronium ion concentration.

Now, plug in to the equilibrium expression (Ka):

$$\frac{(x)(x)}{(0.100 - x)} = 4.6 \times 10^{-4}$$

$$\frac{x^2}{0.100 - x} = 4.6 \times 10^{-4}$$

$$\frac{x^2}{0.100 - x} = 4.6 \times 10^{-4}$$

Assume "x" is very small,
x << 0.100

$0.100 - x \approx 0.100$

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

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

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

$$\frac{x^2}{0.100} = 4.6 \times 10^{-4}$$

$$x^2 = 4.6 \times 10^{-5}$$

$$x = 0.00678233$$

$$[H_3O^+] = 0.00678233\ M$$

$$pH = \boxed{2.17}$$

WHEN can we know that "x" is small enough to ignore in the subtraction:

1) Acid/base equilibria, a few solubility calculations
2) "x" is defined as representing products.
3) The Ka or Kb is 1000 times (or more ... more is better!) SMALLER than the initial concetration of acid or base!

Using this simplifying assumption, solving for "x" became really easy. But would we get the same thing using the quadratic eqation?

Calculating the pH using the quadratic equation to solve for "x" gave a pH of 2.18, a difference of 1 in the last significant figure and below the margin of error for a typical routine pH measurement.

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?

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

$$[H_3O^+] = [HNO_3]_{nominal} = 0.10\ M$$

$$pH = 1.00$$

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)

125 Consider an 0.100 M solution of the weak base ammonia:

NH_3 ; $K_b = 1.8 \times 10^{-5}$

See Appendix I in OpenStax for Kb values

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

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

What is the pH?

Set up an equilbrium chart to reduce the number of variables in the Kb expression!

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

Let "x" equal the change in hydroxide ion concentration.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(x)}{(0.100-x)} = 1.8 \times 10^{-5}$$

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

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

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

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

$$x^2 = 1.8 \times 10^{-6}$$

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

So, pOH $= -\log_{10}(0.0013416408) = 2.87$ $pOH = -\log_{10}[OH^-]$

$pH + pOH = 14.00$

$$pH = 14 - 2.87 = \boxed{11.13}$$

If we solve the problem using the quadratic equation to determine "x", we'll get a pH of 11.13, no different than the answer from the simplification.

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

$pH_{NH_3} = 11.13$

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

$[OH^-] = [NaOH]_{nominal} = 0.100 M$

$pOH = 1.00$

$pH = 13.00$

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)