Consider a solution of 0.0125 M sodium hydroxide (a strong base):

We'll make a similar assumption to the one we use for the acid; in this case, we expect essentially all of the HYDROXIDE ion in solution to come from the NaOH. That means the hydroxide ion concentration just equals the hydroxide ion concentration make by the NaOH ... which equals the nominal NaOH concentration.

$$[OH^-] = [NaOH]_{nominal} = 0.0125\,M\ OH^-$$

We'd like to find pH. Several ways to do this! Here's one. Find pOH.

$$pOH = -\log_{10}[OH^-] = 1.90$$

... then use another identity to find pH.

$$pH + pOH = 14.00\ ;\ pH + 1.90 = 14.00\ ;\ \boxed{pH = 12.10}$$

Like before, we can check our assumption that the amount of hydroxide ion produced by the water was small enough to ignore. Check by finding the HYDRONIUM concentration in solution, since only water produces hydronium in this setup - and every time the water produces a hydronium ion it must also produce a hydroxide ion!

$$[H_3O^+] = 10^{-pH} = 10^{-12.10} = 8.0\times10^{-13}\,M\ H_3O^+$$

$\hookrightarrow = 8.0\times10^{-13}\,M\ OH^-$ made by water itself!

8x10^-13 M is far smaller than the amount of hydroxide made by the base, so we're justified in ignoring water's contribution to hydroxide concentration.

(A) What is the concentration of hydronium ion in an aqueous solution whose pH is 10.50? (B) What is the hydroxide ion concentration? (C) What molar concentration of sodium hydroxide solution would provide this pH?

Use pH identities to answer the questions!

A) $[H_3O^+] = 10^{-10.50}$ $[H_3O^+] = 10^{-pH}$

$= 3.16 \times 10^{-11}$ M H_3O^+

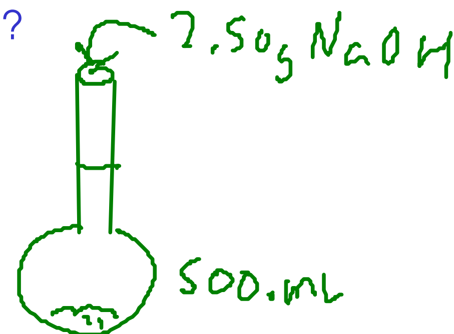
B) $(3.16 \times 10^{-11})[OH^-] = 1.0 \times 10^{-14}$ $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$

$[OH^-] = 3.16 \times 10^{-4}$ M OH^-

C) Since NaOH is a strong base, $NaOH \rightarrow Na^+ + OH^-$ ... and we'll need the same nominal (NaOH) as the hydroxide concentration we want!

$[NaOH] = 3.16 \times 10^{-4}$ M NaOH

What is the pH of a sodium hydroxide solution made from dissolving 2.50 g of sodium hydroxide in enough water to make 500.0 mL of solution?



First, we should try to find the MOLARITY of this solution, since pH calculations are always done based on molarity.

$$M = \frac{\text{mol NaOH}}{\text{L solution}} \leftarrow 0.500\,L \quad (500\,mL)$$

We already know the volume, so we just need to find moles NaOH.

$$2.50\,g\ NaOH \times \frac{\text{mol NaOH}}{40.00\,g\ NaOH} = 0.0625\ \text{mol NaOH}$$

Divide to get MOLARITY.

$$\frac{0.0625\ \text{mol NaOH}}{0.500\,L} = 0.125\,M\ NaOH$$

Since it's a strong base, and $NaOH \rightarrow Na^+ + OH^-$,

$$[OH^-] = 0.125\,M\ OH^-$$

Find pOH...

$$pOH = 0.90$$

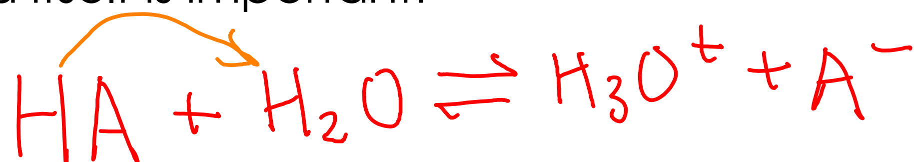
$$pOH = -\log_{10}[OH^-]$$

... and pH

$$pH = 14.00 - 0.90 = \boxed{13.10}$$

$$pH + pOH = 14.00$$

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



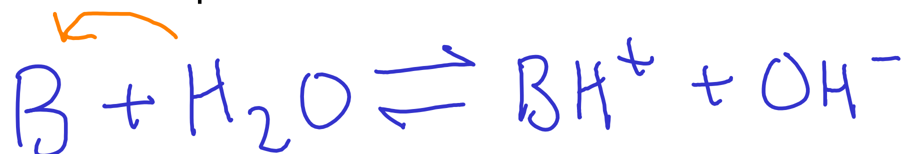
$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

acid ionization constant (K_a)

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

(HA) = concentration of undissociated acid

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



$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

base ionization constant (K_b)

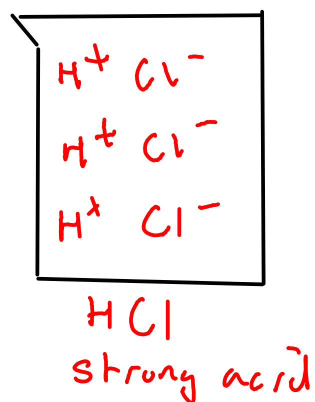
Values for Ka and Kb can often be found in data books / tables / or on the web.

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

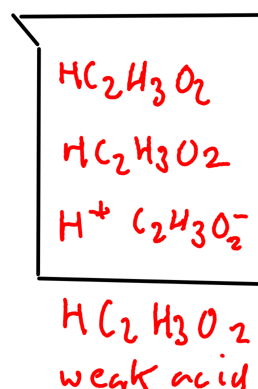
WEAK ELECTROLYTES

- 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!



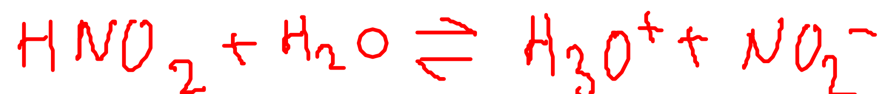
VS



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)



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

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?

As with the strong acid, we'll assume that the acid itself will set the hydronium ion concentration. Even though this acid is WEAK, it's still FAR stronger an acid than water itself. We'll need to solve the equilibrium to find out the hydronium concentration, since we cannot assume all the acid ionizes. 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

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

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

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

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

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

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

Since x should be much smaller than 0.100, then ...

$$0.100 - x \simeq 0.100$$

When can we assume that x is small relative to the initial concentration?

It's usually safe IF the value of the initial concentration is at least 1000x larger than the value of K.

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

$$x = 0.0067082039 \text{ M } H_3O^+$$

$$pH = 2.17$$

(Solving the quadratic would give you a pH of about 2.19)