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

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

Like before, we'll make an assumption. This time, we assume that all the HYDROXIDE in solution comes from the NaOH (the strong base). We expect the presence of hydroxide from the base to suppress the water equilibrium.

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

We want to know pH. First, find pOH.

$$pOH = -\log_{10}(0.0125) = 1.903$$

Now, use a pH identity to relate pH and pOH:

$$pH + pOH = 14.00$$

$$pH + 1.903 = 14.00$$

$$\boxed{pH = 12.10}$$

Let's check the concentration of HYDRONIUM ion, since the hydronium ions are produced by water self-ionizing ... and that lets us know how much hydroxide water creates, too!

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

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

... This number ALSO equals the concentration of HYDROXIDE produced by the water equilibrium - and it's a LOT smaller than 0.0125!

(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?

A) $pH = 10.50 \quad [H_3O^+] = ?$

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

$\boxed{[H_3O^+] = 3.2 \times 10^{-11} M \ H_3O^+}$

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

$(3.2 \times 10^{-11})[OH^-] = 1.0 \times 10^{-14}$

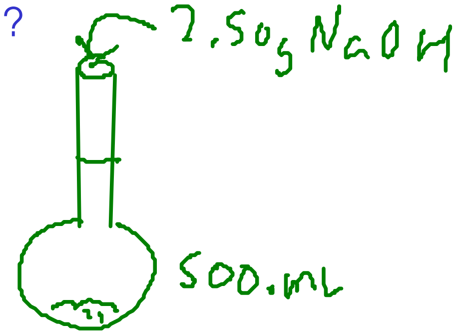
$\boxed{[OH^-] = 3.2 \times 10^{-4} M \ OH^-}$

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

1:1 ratio of NaOH to OH^-

$\boxed{[NaOH]_{nominal} = 3.2 \times 10^{-4} M}$ (0.00032 M)

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?



$NaOH: 40.00\ g/mol$

To solve this, we will need to find MOLARITY of the NaOH, since all the pH identities contain molairty as the concentration unit.

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

To find molarity of NaOH, we'll need to find out how many moles NaOH we dissolved in our 0.500 L ...

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

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

Since NaOH is a strong base, it will completely ionize and set the HYDROXIDE concentration of the solution.

$$NaOH \rightarrow Na^+ + OH^-, \quad so\ [OH^-] = [NaOH]_{nominal}$$

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

$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[H_3O^+](0.125) = 1.0 \times 10^{-14}$$

$$[H_3O^+] = 8.0 \times 10^{-14}$$

$$\boxed{pH = 13.10}$$

$$pH = -\log_{10}([H_3O^+])$$

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!

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

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

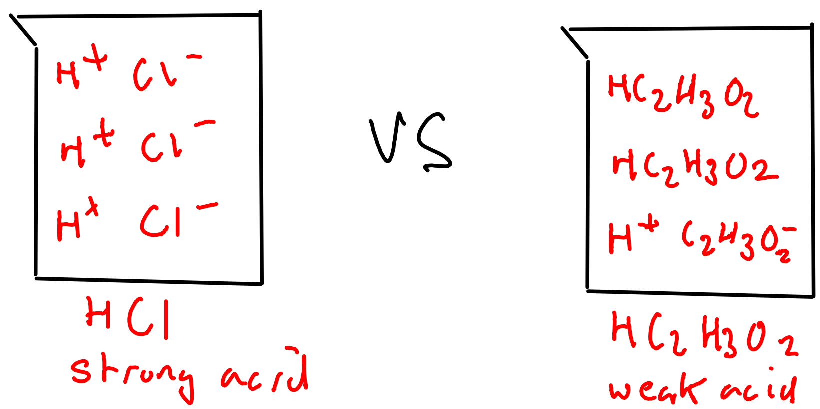
Values for Ka and Kb can often be found in data books / tables / or on the web.

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 the pH, we need to find out the HYDRONIUM ion concentration at equilibrium. This time, we can't just ASSUME all the acid ionizes ... because it doesn't. We need to solve the equilibrium of the acid's ionization.

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

Define "x" as the increase in hydronium ion concentration

$$\frac{[H_3O^+][NO_2^-]}{[HNO_2]} = \boxed{\frac{(x)(x)}{(0.100-x)} = 4.5 \times 10^{-4}}$$

This is very similar in appearance to our Chapter 14 problems!

$$\frac{(x)(x)}{(0.100-x)} = 4.5\times10^{-4}$$

$$\frac{x^2}{0.100-x} = 4.5\times10^{-4}$$

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

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

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

Ka is small, so there will be only a small amount of acid ionizing. That means our "x", which represents the amount of acid ionizing, is ALSO small!

If "x" is small relative to the original acid concentration (0.100 M). then ...

$$0.100 - x \approx 0.100$$

When is it safe to drop the "x" from the subtraction term? When the initial acid or base concentration is 1000 times larger than the value of "K", dropping the "x" is generally safe.

$$\frac{x^2}{0.100} = 4.5\times10^{-4}$$

$$x^2 = 4.5\times10^{-5}$$

$$x = 0.0067082039\ M = [H_3O^+]$$

So, $\boxed{pH = 2.17}$

(If you solve the quadratic rather than assuming "x" is small, you get pH = 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 strong acid like nitric acid:

0.10 M HNO_3, what is pH?

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

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 want to solve for HYDROXIDE concentration here, since we can relate hydroxide concentration to hydronium concentration (and pH) using the pH identities.

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

Let "x" equal the change in ammonium ion concentration

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

Solve for "x" ...

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

Solve using the quadratic equation if you like, OR simplify the equation to solve it faster ...

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

$x << 0.100$, and

$0.100 - x \approx 0.100$

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

be careful here! We have solved for "x" ... which is equal to the HYDROXIDE ion concentration, not the HYDRONIUM concentration, so we can't just take the negative logarithm of "x" and call it the answer!

$$x = 0.001341408 = [OH^-]$$

$$pOH = -\log(0.001341408) = 2.87$$

Since $pH + pOH = 14.00$,

$$pH + 2.87 = 14.00$$

$$\boxed{pH = 11.13}$$

If you'd solved this one with the quadratic equation, you would have found that pH = 11.13 ...

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

$pH_{, NH_3} = 11.13$

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

So, 0.100 m NaOH, $[OH^-] = 0.100$

$pOH = -\log_{10}(0.100) = 1.00$

$pH = 14.00 - 1.00 = \boxed{13.00}$

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)