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

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

Like before, let's assume that all of the sodium hydroxide ionizes, and that all the hydroxide ion in solution comes from the sodium hydroxide.

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

We'd like to know the pH. First, find pOH:

$$pOH = -\log_{10}(0.0125) = 1.90$$

pOH is related to pH, very simply.

$$pH + pOH = 14.00$$

$$pH + 1.90 = 14.00$$

$$\boxed{pH = 12.10}$$

Let's find out hydronium ion concentration. We expect it to be very small, and we'd like to know how much water has self-ionized, snce we assumed that amount was insignificant.

$$[H_3O^+] = 10^{-pH} = 10^{-12.10}$$

$$[H_3O^+] = 7.9 \times 10^{-13}$$

... which is indeed MUCH smaller than 0.0125 (1.25×10^{-2})

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

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

$[H_3O^+] = 10^{-10.50} = 3.2 \times 10^{-11}$ M H_3O^+

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

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

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

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

Sodium hydroxide is a STRONG BASE and should completely ionize in water.

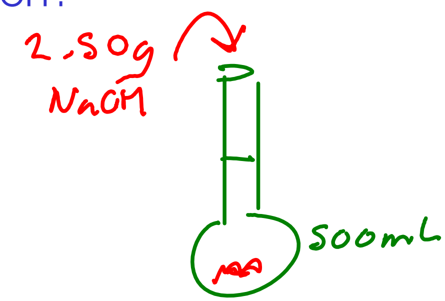
1:1 ratio of $NaOH : OH^-$

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

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



Find the molar concentration (M) of the NaOH solution:

$$M = \frac{\text{mol NaOH}}{\text{L solution}} \leftarrow 0.5000\,L$$

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

$$M = \frac{\text{mol NaOH}}{\text{L solution}} = \frac{0.0625\ \text{mol NaOH}}{0.5000\,L} = 0.125\ \underline{M}\ NaOH$$

Sodium hydroxide is a STRONG BASE, so we expect it to ionize completely in water. It will control the amount of hydroxide ion present.

$$NaOH \rightarrow Na^+ + OH^- \ ;\ [OH^-] = [NaOH]_{nominal} = 0.125\ \underline{M}\ OH^-$$

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

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

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

$$\boxed{pH = 13.10}$$

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

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

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

acid ionization constant

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!

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

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

base ionization constant

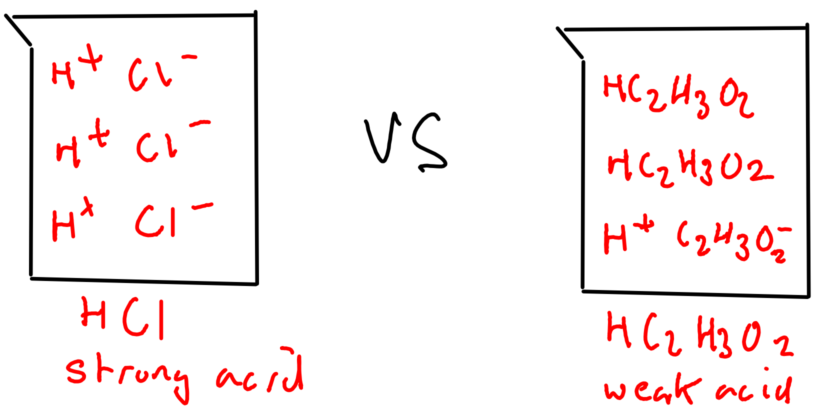
Values for Ka and Kb can often be found in data books / tables / or on the web.

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!