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

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

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

B) $pH + pOH = 14.00$

$pOH = 14.00 - 10.50 = 3.50$

$[OH^-] = 10^{-3.50} = \boxed{3.2 \times 10^{-4}\ M\ OH^-}$

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

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

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

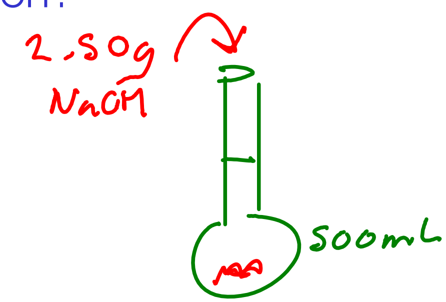
C) Sodium hydroxide is a STRONG BASE

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

1:1 ratio so $3.2 \times 10^{-4}\ \underline{M}\ NaOH$

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?

NaOH: 40.00 g/mol



Find molarity of the NaOH:

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

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

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

Sodium hydroxide is a strong base, so we expect it to completely ionize in water. The hydroxide concentration will equal the nominal sodium hydroxide concentration.

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

$[OH^-] = 0.125$

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

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

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

$\boxed{pH = 13.10}$

$pOH = -\log_{10}(0.125) = 0.90$

$pH + pOH = 14.00$

$pH + 0.90 = 14.00$

$\boxed{pH = 13.10}$

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

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

$$\underbrace{K_a}_{\text{acid ionization constant}} = \frac{[H_3O^+][A^-]}{\underbrace{[HA]}_{\text{(HA) = concentration of undissociated acid}}}$$

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.

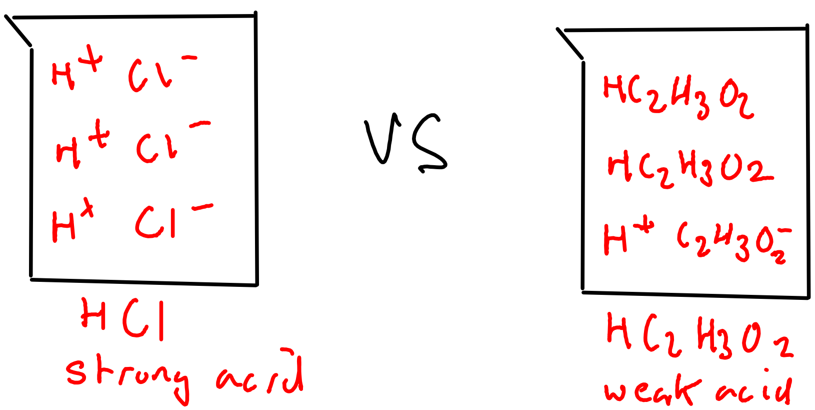
In Ebbing, this data is in the appendices, on pages A-13 and A-14

END OF MATERIAL FOR TEST 3 (7/16/2010)

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)

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

$$K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]} = 5.1 \times 10^{-4}$$

values for Ka are determined experimentally

(We look this number up in a table of acid ionization constants)

What is the pH of the solution?

To find the pH, we need to determine the concentration of hydronium, $[H_3O^+]$

... so we need to solve the equilibrium expression. But we don't know all of the concentrations AT EQUILIBRIUM to do so!

... but they ARE related!

We assume the amount of hydronium from the water is small enough to ignore

SPECIES	INITIAL CONC	CHANGE	EQUILIBRIUM CONC
$[H_3O^+]$	0	$+x$	x
$[NO_2^-]$	0	$+x$	x
$[HNO_2]$	0.100	$-x$	$0.100 - x$

... this is similar to the problems from the equilibrium chapter!

$$5.1\times10^{-4} = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

$$5.1\times10^{-4} = \frac{(x)(x)}{(0.100-x}$$

$$5.1\times10^{-4} = \frac{x^2}{0.100-x}$$

Quadratic equation:

$ax^2+bx+c=0$

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

This assumption is usually safe if the K is at least 1000 times smaller than the initial acid or base concentration

Assume that x << 0.100

$(0.100-x \approx 0.100)$

What's this?

For situations where the amount of dissociated acid or base is much smaller than the original amount, it's safe to assume that the amount of undissocated acid or base remains essentially constant.

In this example, 0.100-x is essentially the same number as 0.100

$$5.1\times10^{-4} = \frac{x^2}{0.100}$$

$$5.1\times10^{-5} = x^2$$

$$x = 0.00714$$

$$= 7.14\times10^{-3} = [H_3O^+]$$

$$\boxed{pH = 2.15}\ ^*$$

* Solving the quadratic gives pH = 2.16 (not much error by using the assumption)