(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)

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

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

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

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

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

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

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

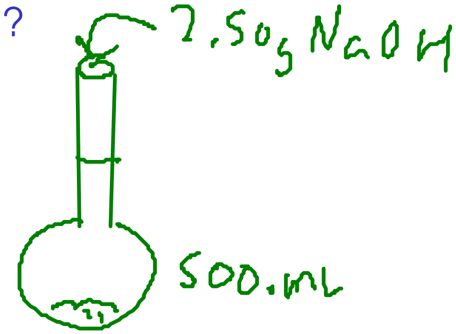
C) $NaOH \rightarrow Na^+ + \underline{\underline{OH^-}}$

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

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

$(0.00032\ M)$

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

$$M = \frac{\text{mol NaOH}}{\text{L solution}} \leftarrow 0.500\ \text{L}$$

Before we can figure out pH, we need to know MOLAR CONCENTRATION of the NaOH ...

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

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

Since NaOH is a strong base, it completely ionizes ...

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

$[OH^-] = 0.125\ M$

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

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

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

$\boxed{pH = 13.10}$

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!

HCl
strong acid

$HC_2H_3O_2$
weak acid

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 pH, we need to determine the concentration of hydronium ion at equilibrium. This time, we cannot assume all the acid ionizes. We will solve the equilibrium to find out how much acid ionizes.

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

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

Look familiar? Very similar to the equilibrium problems in chapter 14!

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

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

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

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

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

Ka is small. So, there will only be a small amount of acid that ionizes. That means 'x', which represents the amount of acid that ionizes, is also small. If 'x' is small relative to 0.100, then ...

$$0.100 - x \approx 0.100$$

When is it safe to assume 'x' is small? Look at the difference between the initial concentration and the K. If they differ by a factor of 1000 or more, the assumption is safe. If not, just solve the quadratic!

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

$$x^2 = 4.5 \times 10^{-5}$$

$$x = 0.006708204 = [H_3O^+]$$

$$\boxed{\text{So, pH} = 2.17}$$

(Solving the quadratic give pH = 2.19)