(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, $[H_3O^+] = ?$

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

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

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

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

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

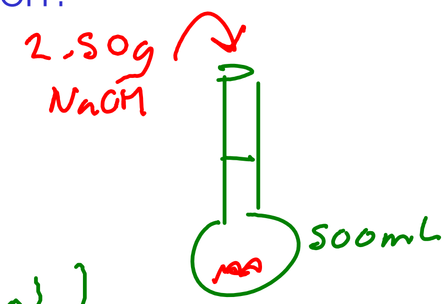
C) $NaOH \rightarrow Na^+ + OH^-$

Sodium hydroxide is a STRONG BASE and should ionize completely.

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

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 molar cocentration of the NaOH solution:

$$M = \frac{mol\ NaOH}{L\ solution} \leftarrow 0.5000\ L\ (500.0\ mL)$$

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

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

Sodium hydroxide is a STRONG BASE. We expect it to completely ionize in water - controlling the amount of hydroxide present.

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

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

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

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

$$\boxed{pH = 13.10}$$

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

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

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

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

acid 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

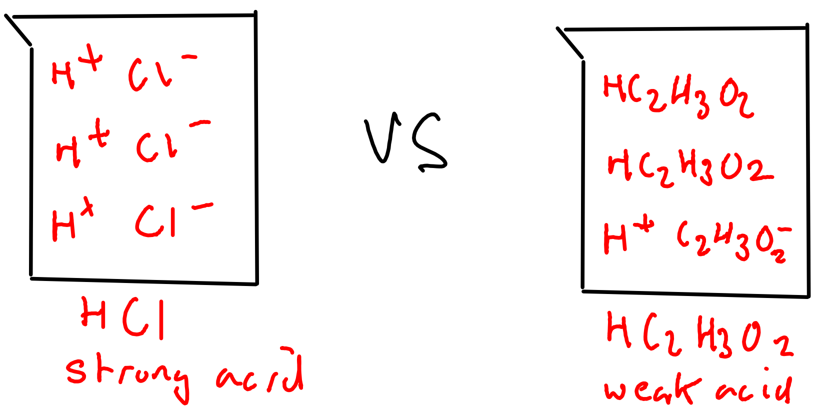
Values for Ka and Kb can often be found in data books / tables / or on the web.

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!

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:

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

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

We can solve this by the quadratic equation .... or ...

based on the value of the equilibrium constant, we assume that x is going to be very small compared to 0.100

If $x \ll 0.100$, then

$$0.100 - x \approx 0.100$$

How do we know when it's "safe" to assume that the difference is ignorable?

This assumption is usually OK provided the K value and the initial concentration differ by at least a factor of 1000. If that's not so - it's safer to just use the quadratic (which will always work...)

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

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

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

So, **pH = 2.15**

Solving this problem with the quadratic equation gives a pH of 2.16 ... which is not significantly different from the caclulated-with-assumption pH of 2.15.

Compare:

- Weak acid HNO_2 : pH of 0.10 M solution = 2.15

Let's compare the pH of the weak nitrous acid with the pH of a stopn acid like nitric acid:

0.10 m HNO_3, what is pH?

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

0.10 m HNO_3, $[H_3O^+] = 0.10$

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?

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.8 \times 10^{-5}$$

We need to solve this, but which one of these terms are we interested in?

We want to solve for HYDROXIDE ION, since it's closely related to pH.

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

Plus into the equilibrium expression:

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

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

This is a QUADRATIC EQUATION. We could use the quadratic formula, OR we could assume that 'x' is small relative to 0.100 ...

$$0.100 - x \approx 0.100$$

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

Be careful ... this is HYDROXIDE concentration, not hydronium concentration!

$$x = 0.0013416 \text{ M } OH^-$$

$$pOH = -\log_{10}(0.0013416) = 2.87$$

$$pH = 14.00 - 2.87 = \boxed{11.13}$$ *

* If we had solved the problem using the quadratic equation, we would have gotten an answer of 11.13. In other words, there's no difference to two significant figures.