

Section 2 Notes—Weak Acids and Bases

First, there is an ADDITION to Section 1 notes: Need for section 1 homework

The pH scale ranges from 0----14. The pOH scale ranges from 0----14.
These scales are only useful for $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ that are $\leq 1.0 \text{ M}$.

For concentrations greater than 1.0 M, the pH or pOH are less than 0 or negative. Even though we can work with these negative pH's and pOH's there is an easier way to find the $[\text{H}_3\text{O}^+]$ from the $[\text{OH}^-]$ and vs. We can go directly from one to the other.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \qquad K_w = 1.0 \times 10^{-14}$$

So if $[\text{H}_3\text{O}^+]$ was found to be **1.3 M**, I can find the $[\text{OH}^-]$ by the following method:

$$1.0 \times 10^{-14} = (1.3)[\text{OH}^-]$$

$$1.0 \times 10^{-14}/1.3 = [\text{OH}^-]$$

$$[\text{OH}^-] = 7.69 \times 10^{-15} \text{ M}$$

This is how to go from one concentration to the other without going through pH and pOH.

I could have done it the long way by the following:

$$[\text{H}_3\text{O}^+] = 1.3 \text{ M}$$

$$\text{pH} = -\log(1.3) = -.114$$

$$\text{pOH} = 14 - (-.114) = 14.114$$

$$[\text{OH}^-] = \text{inv log of } (-14.114) = 7.69 \times 10^{-15} \text{ M}$$

This process is much longer and there is more room for error so it is probably better to do the shorter form using the K_w for water. And, the pH and pOH are outside of the ranges for those scales.

Note: a HW problem has [12.0 M] for a strong acid. Therefore, the $[\text{H}_3\text{O}^+]$ is 12.0. You need to find the $[\text{OH}^-]$. Use the short method to do that.

For Section 2 of this Lesson, read and study the following Notes.

Please watch the video "Weak Acids and Weak Bases"

Please do HW problems #38, 64, 68, 74, 78, 80, 96, 98, 102.

Section 2 Notes—Weak Acids and Weak Bases

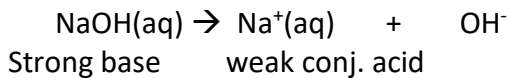
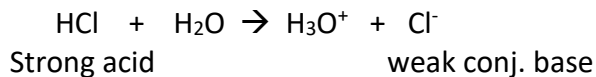
Pages covered in book are (p.666-670, and p.677-680) (also p.672 for percent ionization).

Weak acids and weak bases are ones that ionize “only slightly”. This is usually very, very slightly. So, the **[weak acid]** is **not equal** to **[H₃O⁺]**, and the **[weak base]** is **not equal** to **[OH⁻]** like they are with strong acids and strong bases.

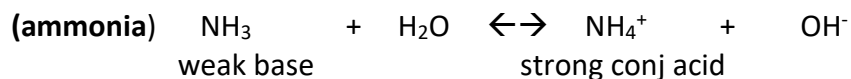
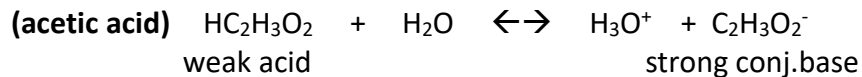
We will use the form of the Bronsted-Lowry acids and bases.

Strong acid form weak conjugate bases. Strong bases form weak conjugate acids.

Example: Strong A and B



Example: Weak A and B



The conjugates are strong because the conj. base takes a proton off the hydronium ion and reverses the products back to reactants. The strong conjugates acid donates its proton back to the hydroxide ion and reverses the products back to reactants. At equilibrium, very few hydronium or hydroxide ions are present.

The systems with weak acids and bases go to equilibrium. K_c for weak acid is now K_a and the K_c for weak base is now K_b .

The K_a for weak acids and K_b for weak bases are found in the appendix of your book (A22-A23).

The Equilibrium expressions for weak acids and bases are set up just like all equilibrium expressions.

$$K_c = \frac{[\text{product}]^{\text{coefficient}} [\text{product}]^{\text{coefficient}}}{[\text{reactant}]^{\text{coefficient}} [\text{reactant}]^{\text{coefficient}}}$$

In an acid equilibrium expression the products are the [hydronium ion] and [conj.base] and the product is always the [acid] so

$$K_a = \frac{[\text{H}_3\text{O}^+] [\text{conj base}]}{[\text{acid}]} \quad (\text{The } [\text{H}_2\text{O}] \text{ is not included because it is (l) or "1"})$$

$$K_b = \frac{[\text{OH}^-] [\text{conj acid}]}{[\text{base}]}$$

It is very important to remember that the **concentration of the hydronium and its conjugate base** are always the **same or equal**. Likewise, the **concentration of the hydroxide ion and its conjugate acid** are always the **same or equal**.

Example: I am doing this example using a RICE box. It will not be necessary to always use a RICE box in doing these equilibrium problems because the format is usually always the same. But I will do some examples using the RICE box and you can decide if you need to use one or not to keep your information organized.

Weak acid: What is the pH of a 0.15 M acetic acid solution? To find pH, I must first write the equation and find $[H_3O^+]$. The K_a for acetic acid, $HC_2H_3O_2$, is 1.8×10^{-5} .

R	$HC_2H_3O_2$	\leftrightarrow	H_3O^+	+	$C_2H_3O_2^-$	
I	0.15 M		0		0	
C	-X		+X		+X	(The -X under $HC_2H_3O_2$ can be ignored because K_a has an exponent ≤ -5). The exponent needs to be ≤ -5 for acids and bases, not -4.)
E	0.15		X		X	
E	0.15M		$1.6 \times 10^{-3} M$		$1.6 \times 10^{-3} M$	

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} \quad 1.8 \times 10^{-5} = \frac{(X)(X)}{(0.15)} \quad (\text{cross multiply})$$

$$X^2 = 2.7 \times 10^{-6} \quad (\text{square root both sides})$$

$$X = 1.6 \times 10^{-3} M$$

So $[H_3O^+] = 1.6 \times 10^{-3} M$
Then $pH = -\log(1.6 \times 10^{-3}) = 2.80$

The format for the equilibrium expression will always be $K_a = X^2 / [\text{acid}]$ and $K_b = X^2 / [\text{base}]$ with X being the $[H_3O^+]$ for an weak acid problem and X being $[OH^-]$ for a weak base.

I will now do an example where the $K_a > 10^{-5}$ and the X under the acid concentration cannot be dropped.

What is the pH of a solution of 0.15M $HC_2H_2ClO_2$ $K_a = 1.35 \times 10^{-3}$

R	$HC_2H_2ClO_2$	\leftrightarrow	H_3O^+	+	$C_2H_2ClO_2^-$
I	0.15		0		0
C	-X		+X		+X
E	0.15-X		X		X
E	0.136 M		0.014 M		0.0143M

$$1.35 \times 10^{-3} = \frac{(X)^2}{(0.15-X)} \quad (\text{cross multiply, distributing})$$

$$X^2 = 2.03 \times 10^{-4} - 1.35 \times 10^{-3} X \quad (\text{Rearrange the equation to } = 0)$$

$$X^2 + 1.35 \times 10^{-3} X - 2.03 \times 10^{-4} = 0 \quad (\text{Apply the quadratic to the equation})$$

a b c

$$X = \frac{-1.35 \times 10^{-3} \pm \sqrt{(1.35 \times 10^{-3})^2 - (4)(1)(-2.03 \times 10^{-4})}}{2(1)} \quad (\text{Eliminate the negative answer})$$

$$X = 0.014 M \quad pH = -\log(0.014) = 1.85$$

Example: This is an example of a weak base. I will not use a RICE box with this example to show that it is not necessary to always use a RICE box as long as you remember the form of the equilibrium expression.

What is the **pH** of a weak base solution of **0.18 M NH₃**? The K_b of NH₃ is 1.8×10^{-5} . (Just happens to be the same as acetic acid)

Since $K_b \leq 10^{-5}$, I can drop the X in the denominator of the expression.

$$K_b = \frac{X^2}{[\text{base}]} \quad 1.8 \times 10^{-5} = \frac{X^2}{(0.18)} \quad (\text{Basic Form})$$

$$X^2 = 3.24 \times 10^{-5}$$

$$X = 5.70 \times 10^{-3} \text{ M} \quad (\text{Remember X is } [\text{OH}^-] \text{ so this will give me pOH not pH.})$$

$$[\text{OH}^-] = 5.7 \times 10^{-3} \text{ M} \quad [\text{NH}_4^+] = 5.7 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log(5.7 \times 10^{-3}) = 2.24$$

$$\text{pH} = 14 - 2.24 = 11.76$$

Percent Ionization (dissociation) (p.672)

Percent ionization is the percent of hydronium ion or hydroxide ion formed when a weak acid or a weak base goes into solution.

$$\text{Percent ionization of a weak acid} = \frac{[\text{H}_3\text{O}^+]}{[\text{acid}]} \times 100$$

Example: **0.15 M acetic acid**, the $[\text{H}_3\text{O}^+]$ is $1.6 \times 10^{-3} \text{ M}$.

$$\text{The percent ionization} = \frac{1.6 \times 10^{-3}}{0.15} \times 100 = 1.1\% \quad \text{Very little acid ionized into hydronium ion.}$$

$$\text{Percent ionization of a weak base} = \frac{[\text{OH}^-]}{[\text{base}]} \times 100$$

Example: **.18 M ammonium**, the $[\text{OH}^-]$ is $5.7 \times 10^{-3} \text{ M}$.

$$\text{The percent ionization} = \frac{5.7 \times 10^{-3}}{0.18} \times 100 = 3.2\% \quad \text{Also, very little base ionized to hydroxide.}$$

You can calculate K_a or K_b from the percent ionization (dissociation). P. 674

$$[\text{acid}] (\% \text{ ionization}/100) = K_a \quad [\text{base}] (\% \text{ ionization}/100) = K_b$$

There are times when you need to use the pK_a of an acid instead of the K_a and the pK_b of a base instead of the K_b . As stated earlier in the notes...

$$\text{pK}_a = -\log K_a \quad \text{pK}_b = -\log K_b$$