This is BCE#21.

I recommend you print out this page and bring it to class. Click here to show a set of five BCE21 student responses randomly selected from all of the student responses thus far in a new window.

John , here are your responses to the BCE and the Expert's response.

As we expand our application of aqueous acid-base chemistry we will look at the next type of acid-base aqueous equilibrium system which is called common ion.

1. A solution in a container is 0.500 M $HC_2H_3O_2$ and 0.400 M $NaC_2H_3O_2$. Answer the following questions about this solution.

a) What is the initial concentration of $HC_2H_3O_2$ in the solution?

[$HC_2H_3O_2$] = 0.500 M 87%

The initial concentration of $HC_2H_3O_2$ is 0.500 M.

b) What is the initial concentration of $C_2H_3O_2^-$ (from the $NaC_2H_3O_2$) in the solution?

[$C_2H_3O_2^-$] = 0.400 M 78%

The initial concentration of $C_2H_3O_2^-$ is also 0.400 M. In this case the assumption is that the $NaC_2H_3O_2$ completely dissociates into Na^+ and $C_2H_3O_2^-$ ions.

c) Complete the ICE table below

Reaction	$HC_2H_3O_2$	$\rightleftarrows$	H^+	+ $C_2H_3O_2^-$
Initial	0.500 M 87% 0.500 M		0 M 94% ~ 0 M	0.400 M 67% 0.400 M
Change	-x M 83%		+x M 87%	+x M 87%

	- x		+ x	+ x
Equilibrium	0.500-x M 0.500 - x 83%		0+x M 0 + x 83%	0.400+x M 0.400 + x 67%

17%
0±x

To compete the ICE table we know the initial concentration of all the species in the reaction. This system is interesting because there is initial amounts of both the weak acid, $HC_2H_3O_2$, and its conjugate base, $C_2H_3O_2^-$. The reaction must still proceed from left to right to establish equilibrium because the concentration of H^+ is assumed to be nearly zero. Notice that in the equilibrium row we are subtracting the amount of the weak acid that dissociates from the initial amount of the weak acid and adding it to the initial amount of the conjugate base.

2. The K_a for $HC_2H_3O_2$ is 1.8 x 10^{-5}. Calculate the pH of the common ion solution above using the information from your ICE table.

pH = 4.65 44%

22% $HC_2H_3O_2$

$K_a = 1.8 \times 10^{-5} = [C_2H_3O_2^-][H^+]/[HC_2H_3O_2]$

$1.8 \times 10^{-5} = (0.400 + x)(x)/(0.500 - x)$

Assume x <<<< 0.400 M because K_a is very small compared to the initial concentration of the weak acid. So 0.500 - x reduces to 0.500 because x is so small compared to 0.500. ALSO 0.400 + x reduces to 0.400 because x will also be small compared to 0.400.

$1.8 \times 10^{-5} = (0.400)(x)/(0.500)$

2.1×10^{-5} M = x = $[H^+]$

pH = -log $[H^+]$

pH = -log (2.1×10^{-5}) = 4.67

	After completing the calculation play the movie and observe the measured pH of a solution that is 0.500 M $HC_2H_3O_2$ and 0.400 M $NaC_2H_3O_2$.

3. The pH of a solution that only contains acetic acid (0.500 M $HC_2H_3O_2$) is 2.54. How do you explain the difference in pH of these solution?

The solution in Q1 contains some of the conjugate base initially, so one would expect the pH to be more basic compared to a solution with only the acid.

The pH of a solution that only contains 0.500 M $HC_2H_3O_2$ is 2.54, lower than the pH of a solution that contains 0.500 M $HC_2H_3O_2$ and 0.400 M $C_2H_3O_2^-$. We can understand why the this is the case because in the solution that is only 0.500 M $HC_2H_3O_2$ there is very little $C_2H_3O_2^-$ (a base). In the solution with 0.500 M $HC_2H_3O_2$ and 0.400 M $C_2H_3O_2^-$ there is as much base as there is of the acid so it makes the solution more basic.

An intereting question might be why is a solution that contains the weak acid $HC_2H_3O_2$ and the weak base $C_2H_3O_2^-$ acidic and not basic? The reason is that the K_a for $HC_2H_3O_2$ is 1.8×10^{-5} is much larger than the K_b for $C_2H_3O_2^-$ which is $1.0 \times 10^{-14}/1.8 \times 10^{-5} = 5.9 \times 10^{-10}$

4. Is there anything about the questions that you feel you do not understand? List your concerns/questions.

nothing

5. If there is one question you would like to have answered in lecture, what would that question be?

nothing

h3>