CHAPTER 17		BLM 17.3.1
OVERHEAD	# Determining K_a and Percent Ionization	

Problem

Propanoic acid, $CH_3CH_2COOH(aq)$, is a weak monoprotic acid that is used to inhibit mould formation in bread. A student prepared a 0.10 mol/L solution of propanoic acid and found the pH was 2.96. What is the acid ionization constant for propanoic acid? What percentage of its molecules were ionized in the solution?

What Is Required?

You need to find K_a and the percent ionization for propanoic acid.

What Is Given?

You have the following data:
Initial $[C_2H_5COOH(aq)] = 0.10$ mol/L
pH = 2.96

Plan Your Strategy

Step 1 Write the equation for the ionization equilibrium of propanoic acid in water. Then set up an ICE table.

Step 2 Write the expression for the acid ionization constant.

Step 3 Calculate $[H_3O^+(aq)]$ using $[H_3O^+(aq)] = 10^{-pH}$.

Note that $[C_2H_5COO^-] = [H_3O^+]$.

Step 4 Use the stoichiometry of the equation and $[H_3O^+(aq)]$ to complete the ICE table. Calculate K_a.

Step 5 Calculate the percent ionization by expressing the fraction of molecules that ionize out of 100.

Determining K_a and Percent Ionization (continued)

Act on Your Strategy

Step 1 Use the equation for the ionization equilibrium of propanoic acid in water to set up an ICE table.

$C_2H_5COOH(aq) + H_2O(\ell) \leftrightarrow C_2H_5COO^-(aq) + H_3O^+(aq)$

	$[C_2H_5COOH(aq)]$ (mol/L)	$[H_2O(\ell)]$ (mol/L)	$[C_2H_5COO^-(aq)]$ (mol/L)	$[H_3O^+(aq)]$ (mol/L)
Initial	0.10		0	~0
Change	-1.1×10^{-3}		$+1.1 \times 10^{-3}$	$+1.1 \times 10^{-3}$
Equilibrium	$0.10 - 1.1 \times 10^{-3}$		1.1×10^{-3}	1.1×10^{-3}

Step 2 Remember, pure water is not included in the expression for K_a.

$$K_a = \frac{[C_2H_5COO^-][H_3O^+]}{[C_2H_5COOH]}$$

Step 3 Calculate the value of $[H_3O^+(aq)]$.

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

$[H_3O^+(aq)] = 10^{-2.96}$

$[H_3O^+] = 1.1 \times 10^{-3}$ mol/L

Step 4 Recall that $[C_2H_5COO^-] = [H_3O^+]$

$$K_a = \frac{(1.1 \times 10^{-3})^2}{0.10 - (1.1 \times 10^{-3})}$$

$$= 1.2 \times 10^{-5}$$

Step 5

$$\text{Percent ionization} = \frac{1.1 \times 10^{-3} \cancel{\frac{mol}{L}}}{0.10 \cancel{\frac{mol}{L}}} \times 100\%$$

$$= 1.1\%$$

Check Your Solution

The value for K_a and the percent ionization are reasonable for a weak acid.

MathTip

As you know, powers of ten do not count in the number of significant digits in a measurement. For this reason, the part of a logarithm that depends on the power of ten does not count in the number of significant digits in the logarithm. For example, the number 3.57×10^4 has three significant digits. The logarithm of 3.57×10^4 is 4.5527. The numeral 4 to the left of the decimal represents the exponent of 10, so it does not count in the number of significant digits. The numerals to the right of the decimal determine the 3.57, so they count as significant digits. (**Hint:** Find the logarithm of 3.57×10^8.) Therefore, the logarithm of 3.57×10^4 with the correct number of significant digits is 4.553. If the concentration of hydronium ion is 1.1×10^{-5} mol/L, there are two significant digits. The pH is 4.96, correct to two significant digits.