Calculating the pH of a Weak Acid

Problem

Methanoic acid, HCOOH(ℓ), is present in the sting of certain ants. What is the pH of a 0.25 mol/L solution of methanoic acid?

What Is Required?

You need to calculate the pH of the solution.

What Is Given?

You know the concentration of methanoic acid:
[HCOOH(aq)] = 0.25 mol/L
The acid ionization constant for methanoic acid is listed in Appendix G: $K_a = 1.8 \times 10^{-4}$.

Plan Your Strategy

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

Step 2 Calculate the value of $\frac{[HCOOH]}{K_a}$ to make sure the amount that ionizes is not significant compared with the initial concentration of the acid.

Step 3 Write the equation for the acid ionization constant. Substitute equilibrium terms into the equation.

Step 4 Solve the equation for x.

Step 5 pH = –log $[H_3O^+(aq)]$

Calculating the pH of a Weak Acid (continued)

Act on Your Strategy

Step 1

$$HCOOH(aq) + H_2O(\ell) \leftrightarrow HCOO^-(aq) + H_3O^+(aq)$$

	[HCOOH(aq)] (mol/L)	[$H_2O(\ell)$] (mol/L)	[$HCOO^-(aq)$] (mol/L)	[$H_3O^+(aq)$] (mol/L)
Initial	0.025		0	~0
Change	$-x$		$+x$	$+x$
Equilibrium	$0.025 - x$		x	x

Step 2

$$\frac{[HCOOH]}{K_a} = \frac{0.25}{1.8\times10^{-4}}$$
$$= 1.4\times10^{3}$$

Since this value is greater than 1000, the amount that ionizes is not significant compared with the initial concentration of the acid. Thus, $(0.25 - x)$ is approximately 0.25.

Step 3

$$K_a = \frac{[HCOO^-][H_3O^+]}{[HCOOH]}$$
$$= \frac{(x)(x)}{0.25}$$
$$= 1.8\times10^{-4}$$

Step 4

$$1.8\times10^{-4} = \frac{x^2}{0.25}$$
$$x^2 = 4.5\times10^{-5}$$
$$x = 6.71\times10^{-3}$$
$$x = 6.71\times10^{-3}\ \frac{mol}{L} = [H_3O^+(aq)]$$

Step 5

$$pH = -\log 6.71 \times 10^{-3}$$
$$= 2.17$$

The pH of a solution of 0.25 mol/L methanoic acid is 2.17.

Check Your Solution

The pH indicates an acidic solution, as expected. Data that were given in the problem have two significant digits, and the pH has two digits following the decimal place.