

Write the generic equation that symbolized the dissociation of acetic acid in the ICE box.

Acetic acid, $K_a = 1.8 \times 10^{-5}$

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

1. Calculate the $[H^+]$ for a 0.10 M acetic acid solution.
 • Then calculate % ionization

R			
I			
C			
E			

$$1.8 \times 10^{-5} = \frac{[\quad] [\quad]}{[\quad]}$$

$$x = [H^+] =$$

$$\% \text{ ionization} =$$

$$\text{pH} =$$

2. Calculate the $[H^+]$ for a 0.010 M acetic acid solution.
 • Then calculate % ionization

R			
I			
C			
E			

$$1.8 \times 10^{-5} = \frac{[\quad] [\quad]}{[\quad]}$$

$$x = [H^+] =$$

$$\% \text{ ionization} =$$

$$\text{pH} =$$

3. Calculate the $[H^+]$ for a 0.0010 M acetic acid solution.
 • Then calculate % ionization

R			
I			
C			
E			

$$1.8 \times 10^{-5} = \frac{[\quad] [\quad]}{[\quad]}$$

$$x = [H^+] =$$

$$\% \text{ ionization} =$$

$$\text{pH} =$$

4. Calculate the $[H^+]$ for a 0.00010 M acetic acid solution.
 • Then calculate % ionization

R			
I			
C			
E			

$$1.8 \times 10^{-5} = \frac{[\quad] [\quad]}{[\quad]}$$

$$x = [H^+] =$$

$$\% \text{ ionization} =$$

$$\text{pH} =$$

5. What do these calculations tell us about percent ionization as it relates to the concentration of a weak acid?

Write the generic equation that symbolized the dissociation of acetic acid in the ICE box.
Acetic acid, $K_a = 1.8 \times 10^{-5}$

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

1. Calculate the $[H^+]$ for a 0.10 M acetic acid solution.

- Then calculate % ionization

R			
I	0.1	~0	~0
C	-X (too small)	+X	+X
E	0.1	X	X

$$1.8 \times 10^{-5} = \frac{[x][x]}{[.1]}$$

$$x = [H^+] = 1.3 \times 10^{-3}$$

$$\% \text{ ionization} = 1.3\%$$

$$\text{pH} = 2.88$$

2. Calculate the $[H^+]$ for a 0.010 M acetic acid solution.

- Then calculate % ionization

R			
I	0.01	~0	~0
C	-X (too small)	+X	+X
E	0.01	X	X

$$1.8 \times 10^{-5} = \frac{[x][x]}{[.01]}$$

$$x = [H^+] = 4.2 \times 10^{-4}$$

$$\% \text{ ionization} = 4\%$$

$$\text{pH} = 3.37$$

3. Calculate the $[H^+]$ for a 0.0010 M acetic acid solution.

- Then calculate % ionization

R			
I	0.001	~0	~0
C	-X (too small)	+X	+X
E	0.001	X	X

$$1.8 \times 10^{-5} = \frac{[x][x]}{[.001]}$$

$$x = [H^+] = 1.3 \times 10^{-4}$$

(1.3×10^{-4} with quadratic)

$$\% \text{ ionization} = 13\%$$

$$\text{pH} = 3.87$$

4. Calculate the $[H^+]$ for a 0.00010 M acetic acid solution.

- Then calculate % ionization

R			
I	0.0001	~0	~0
C	-X (not so small)	+X	+X
E	0.0001	X	X

$$1.8 \times 10^{-5} = \frac{[x][x]}{[.0001]}$$

$$x = [H^+] = 4.2 \times 10^{-5}$$

(3.4×10^{-5} using the quadratic)

$$\% \text{ ionization} = 42\% (34\%)$$

$$\text{pH} = 4.38 (4.47)$$

So the punch line is that the percent ionization of a particular weak acid increases as the acid concentration decreases. Note that this does not mean that the $[H^+]$ increases or the pH decreases for more dilute solutions of the same acid, because the amount of H^+ ions present will be less, but the amount of H^+ present will be a bigger percentage of the original, more dilute acid. For $HA \rightleftharpoons H^+ + A^-$ when diluting the solution, there will be less concentration of HA in the water *and* less concentration of H^+ and A^- in the water, thus the equilibrium will shift to products because one acid molecule can produce two ions, increasing overall concentration more.

This is analogous to the rational we used to explain why gas equilibria shift to the side with more gas molecules when we increase the volume of the container (decrease pressure). Decreasing concentration is analogous to increasing container size, and the shift its toward more gas molecules, just as the shift here is toward the side with more particles, the ions.