Chemistry Lecture #93: Acid and Base Ionization Constants

Acetic acid, or $HC_2H_3O_2$, is a weak acid. An acetic acid solution has a lot of intact $HC_2H_3O_2$ molecules. Very little of it ionizes to form H^+ and $C_2H_3O_2^-$.

$$HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$$

lots   very little

The equilibrium constant expression and value for this reaction would be

$$Keq = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 1.75 \times 10^{-5}$$

Since this is the Keq for an (a)cid, we write

$$Ka = 1.75 \times 10^{-5}$$

Ka is the ionization constant for an acid. It tells us the strength of an acid. The larger the number, the stronger the acidic.

The Ka value of 1.75×10^{-5} is a very small number. Thus, acetic acid is a weak acid and there is very little H^+ in solution.

Sometimes the equilibrium reaction for acids is written using H_3O^+ instead of H^+.

$$HC_2H_3O_2 + H_2O \rightleftharpoons C_2H_3O_2^- + H_3O^+$$

$$Keq = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2][H_2O]}$$

$[H_2O]$, the concentration of water, doesn't really change very much since the reaction occurs in water solution. The number of water molecules that react compared to the total amount of water is so negligible that $[H_2O]$ can be considered as just an unchanging number or a constant. If we multiply both sides by $[H_2O]$, we get

$$Keq[H_2O] = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

We then substitute Ka in place of $Keq[H_2O]$.

$$Ka = = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]}$$

which is the same thing as

$$Keq = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 1.75 \times 10^{-5}$$

Kb is the ionization constant for a base. It tells us the strength of the base or the relative amount of OH^-.

For example, NH_3 reacts with water to produce NH_4^+ and OH^-.

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$Keq = \frac{[NH_4^+][OH^-]}{[NH_3][H_2O]}$$

$$Keq[H_2O] = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$Kb = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

The Kb for NH_3 is 1.8×10^{-5}. This is a very small number, so NH_3 is a weak base.

If you know the Ka or Kb, you can calculate the concentration of H^+ or OH^-.

What is the concentration of H^+ in a 0.0400 M solution of $HC_2H_3O_2$? Ka = 1.75×10^{-5}

Answer

If you had 1 L of 0.0400 M acetic acid, it means that you started with 0.0400 moles of $HC_2H_3O_2$, but some of it ionized and became H^+ and $C_2H_3O_2^-$. If 0.01 moles ionized you'd have 0.04-0.01 = 0.03 moles of $HC_2H_3O_2$ remaining and 0.01 moles of H^+ and $C_2H_3O_2^-$. But we don't know how much ionized, so let's call the amount x.

If x moles of $HC_2H_3O_2$ ionized you'd have 0.04 - x remaining.
For every x moles of $HC_2H_3O_3$ that ionized, you'll get x moles of H^+ and x moles of $C_2H_3O^-$.

$$\begin{array}{ccccc} 0.04 - x & & x & & x \\ HC_2H_3O_2 & \rightleftharpoons & H^+ & + & C_2H_3O_2^- \end{array}$$

$$Ka = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.75 \times 10^{-5} = \frac{[x][x]}{0.04 - x}$$

1.75×10^{-5} is such a small number, it means that x has to be very tiny. Thus, $0.04 - x \approx 0.04$

$$1.75 \times 10^{-5} = \frac{x^2}{0.04}$$

$x = 8.37 \times 10^{-4}$ M of H^+

The rule that I like to follow is that if Ka or Kb is about 10^{-4} or smaller, it is okay to approximate and say that x is a negligible value that can be disregarded in the denominator.

What is the $[OH^-]$ in a 0.0200 M solution of NH_3? What is the percent ionization? Kb = 1.80×10^{-5}

Answer

Percent ionization asks what percent of the original amount of NH_3 ionized to produce OH^- and NH_4^+. We first need to find out how much OH^- is produced.

$$\begin{array}{ccccccc} 0.02 - x & & & & x & & x \\ NH_3 & + & H_2O & \rightleftharpoons & NH_4^+ & + & OH^- \end{array}$$

$$Kb = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

$$1.8 \times 10^{-5} = \frac{[x][x]}{0.02 - x}$$ assume x is a negligible value

$$1.8 \times 10^{-5} = \frac{x^2}{0.02}$$

x = 6.00×10^{-4} M OH^-. This is also the amount of NH_3 that ionized.

$$\text{percent ionization} = \frac{6.00 \times 10^{-4}}{0.02} \times 100 = 3.00 \text{ percent}$$

Original amount of NH_3.