A simple way to understand an ionization constant is to think of it in a clear-cut way: To what degree will a substance produce ions in water? In other words, to what extent will ions be formed?

**Introduction**

Water has a very low concentration of ions that are detectable. Water undergoes self-ionization, where two water molecules interact to form a hydronium ion and a hydroxide ion.

$$[H_2O + H_2O \leftrightarrow H_3O^+ + OH^-] \label{1}$$

Even though water does not form a lot of ions, the existence of these is evident in pure water by the electrical conductivity measurements. Water undergoes ionization because the powerfully electronegative oxygen atom takes the electron from a hydrogen atom, making the proton dissociate.

$$\text{H-O-H} \rightarrow H^+ + OH^- \label{2}$$

with two ions are formed:

- hydrogen ions \((H^+)\)
- hydroxyl (hydroxide) ions \((OH^-)\)

The hydrogen ions then react with water to form hydronium ions:

$$H^+ + H_2O \rightarrow H_3O^+ \label{3}$$

Typically, hydrogen atoms are bonded with another water molecule resulting in a hydration that forms a hydronium and hydroxide ion.

$$H_2O + H_2O \rightarrow H_3O^+ + OH^- \label{4}$$

In 1 L of pure water at 25 degrees Celsius, there is \(10^{-7}\) moles of hydronium ions or hydroxide ions at equilibrium. Let’s come back to the question of to what degree will a substance form ions in water? For water, we have learned that it will occur until \(10^{-7}\) moles of either ion will ionize at the previous given conditions. Since this is during equilibrium, a constant can be formed.

$$[H_2O \leftrightarrow H^+ + OH^-] \label{5}$$

$$K_{eq} = \dfrac{[H^+] [OH^-]}{[H_2O]} \label{6}$$

This equilibrium constant is commonly referred to as \(K_w\).

**What about acids and bases?**

Now that we have an idea of what an ionization constant is, let’s take a look at how acids and bases play in this scenario. Strong acids and bases are those that completely dissociate into ions once placed in solution. For example:
So, a 2 M solution of KOH would have 2 M concentration of OH⁻ ions (also 2 M concentration of K⁺).

Weak acids and bases, however do not behave the same way. Their amounts cannot be calculated as easily. This is because the ions do not fully dissociate in the solution. Weak acids have a higher pH than strong acids, because weak acids do not release all of its hydrogens. The acid dissociation constant tell us the extent to which an acid dissociates

\[\text{HCN}_{(aq)} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CN}^- \]

\[K_a = \dfrac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]} \]

This equation is used fairly often when looking at equilibrium reactions. During equilibrium, the rates of the forward and backward reaction are the same. However, the concentrations tend to be varied. Since concentration is what gives us an idea of how much substance has dissociated, we can relate concentration ratios to give us a constant. K is found by first finding out the molarity of each substance. Then, just as shown in the equation, we divide the products by the reactants, excluding solids and liquids. Also when there is more than one product or reactant, their concentrations must be multiplied together. Even though you will not see a multiplication sign, if there are two molecules associated, remember to multiply them. If there is a coefficient in front of a molecule, the concentration must be raised to that power in the calculations.

A weaker acid tends to have a smaller \(K_a\) because the concentration on the bottom of the equation is larger. There is more of the acid, and less of the ions. You can think of Ka as a way of relating concentration in order to find out other calculations, typically the pH of a substance. A pH tells you how basic or acidic something is, and as we have learned that depends on how much ions become dissociated.

Example \(\PageIndex{1}\): Calculate Ionization Constant

Calculate the ionization constant of a weak acid. Solve for \(K_a\) given 0.8 M of hydrogen cyanide and 0.0039 M for hyrodronium and cyanide ions:

\[\text{HCN}_{(aq)} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{CN}^-\]

Solution

• \([\text{HCN}] = 0.8 \text{ M}\)
• \([\text{H}_3\text{O}^+] = 0.0039 \text{ M}\)
• \([\text{CN}^-] = 0.0039 \text{ M}\).

We can assume here that \(\text{H}_3\text{O}^+ = \text{CN}^-\)

The equilibrium constant would be

\(K_a = \dfrac{0.0039^2}{0.8} = 1.9 \times 10^{-5}\)
We can use the similar method to find the $K_b$ constant for weak bases. Again, an ionization constant is a measure of the extent a base will dissociate. $K_b$ relates these molarity quantities. A smaller $K_b$ corresponds to a weaker base, as a higher one a stronger base. Some common weak bases: $\text{NH}_3$, $\text{NH}_4\text{OH}$, $\text{HS}^-$, $\text{C}_5\text{H}_5\text{N}$

$$\text{[C}_5\text{H}_5\text{N}_\text{(aq)} + \text{H}_2\text{O}_\text{(l)} \rightarrow \text{C}_5\text{H}_5\text{N}^+_{\text{(aq)}} + \text{OH}^\text{-}_{\text{(aq)}} \text{]} \label{10}$$

We would find $K_b$ the same way we did $K_a$. However, most problems are not as simple and obvious. Let’s do an example that is a little more challenging to help you understand this better.

Example $(\PageIndex{2})$: Calculate Concentration

Calculate the concentration of [OH$^-$] in a 3M Pyridine solution with $K_b= 1.5 \times 10^{-9}$.

**Solution**

Since we already have our equation, let’s write the expression for the constant as discussed earlier (concentration of products divided by concentration of reactants)

$$[K_b= \frac{[\text{OH}^-][\text{C}_5\text{H}_5\text{N}^+]}{[\text{C}_5\text{H}_5\text{N}]} = 1.5 \times 10^{-9}]$$

Because pyridine is a weak base, we can assume that not much of it will disassociate. And since we have 3M, let’s make 3M-X our calculation. Subtracting X is going to be the change in molarity due to dissociation. Since our ions are unknown and they are both one mole, we can solve for them as being X. THUS: $(X^2/3-X) = 1.5 \times 10^{-9}$

To make the calculation simple, we can estimate 3-X to be approximately 3 because of the weak base. Now our equation becomes: $(X^2/3) = 1.5 \times 10^{-9}$. When we solve for X our answer is $(4.7 \times 10^{-5})$ M.

This approximation was effective because x is small, which must be less than 5% of the initial concentration for that estimation to be justified.

**Summary**

We have learned that an ionization constant quantifies the degree a substance will ionize in a solution (typically water). $K_a$, $K_b$, and $K_w$ are constants for acids, bases, and finally water, respectively and are related by

$$[K_a \times K_b = K_w]$$

However, these constants are also used to find concentrations as well as pH.

**References**


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