Skills to Develop

- Explain conjugate acids of bases.
- Evaluate $K_a$ of the conjugate acid of a base.
- Treat the conjugate acid of a base as an acid in numerical calculations.
- Reverse the role of acid and base for the previous skills.

Conjugate Acids of Bases

The conjugation of acids and bases has been discussed earlier. After losing a proton, the acid species becomes the conjugate base. A base and its protonated partner also form a conjugated acid-base pair. These relationships have been represented by

$$\text{H}^+ + \text{[Base]} = \text{[Conjugate acid of Base]}^+$$

$$\text{[Acid]} = \text{H}^+ + \text{[Conjugate base of Acid]}^-$$

For example:

$$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$$

$$\text{HAc} \rightleftharpoons \text{H}^+ + \text{Ac}^-$$

Thus, $\text{NH}_4^+$ and $\text{NH}_3$ are a pair of conjugate acids and bases, as are $\text{HAc}$ and $\text{Ac}^-$.

$K_a$ Values of Conjugate Acids of Bases

We have used $K_a$ and $K_b$ as the acidic and basic constants of acids and bases. Can an acidic constant, $K_a$, be assigned to the conjugate acid of a base? If so, what is the relationship between $K_a$ of the conjugate acid and $K_b$ of the base? We are going to derive the relationship here. Note that water always plays a role in the conjugation acid-base pair.

Let $\text{BH}^+$ be the conjugate acid of a base. The expression for the acidic constant $K_a$ for the conjugate acid

$$\text{BH}^+ = \text{[B]} + \text{H}^+$$

can be written as

$$K_a = \frac{[\text{B}][\text{H}^+]}{[\text{BH}^+]}$$

$$= \frac{[\text{B}][\text{H}^+]}{[\text{BH}^+][\text{OH}^-]}$$

$$= \frac{1}{K_b}K_w$$

where $K_w$ is the ion product of water.
Thus,

\[
\text{\color{Red} K_a} \text{ \color{Blue} K_b} = \text{K_w}
\]

Furthermore,

\[
-\log (\text{\color{Red} K_a}) - \log (\text{\color{Blue} K_b}) = -\log (\text{K_w})
\]

and at 298 K, we have

\[
p \text{\color{Red} K_a} + p \text{\color{Blue} K_b} = 14
\]

Example 1

The \(\text{K}_{a}\) for \(\ce{HCO3-}\) is 4.7E-11; what is the conjugate base and its \(\text{K}_{b}\)?

**Solution**

The conjugate base is \(\ce{CO3^2-}\).

\[
\text{K}_b = \dfrac{1E{-}14}{4.7E{-}11} = 2.1E{-}4
\]

**Discussion**

The \(\text{K}_b\) so calculated is for the reaction,

\[
\ce{CO_3^{2-} + H_2O \rightleftharpoons HCO_3^- + OH^-}
\]

\[
\text{K}_b = \dfrac{[HCO_3^-] [OH^-]}{[CO_3^{2-}]}
\]

The anion \(\ce{CO3^2-}\) is a rather strong base, and the large value calculated for \(\text{K}_b\) agrees with the fact.

Example 2

The \(\text{K}_b\) for the anion of oxalic acid, \(\ce{COO^- | COOH}\) is 1.8E-10. What is the \(\text{K}_a\) for the oxalic acid \(\text{K}_{(COOH)}\)?

**Solution**

The \(\text{K}_a\) for oxalic acid is

\[
\text{K}_a = \dfrac{1E{-}14}{1.8E{-}10} = 5.6E{-}5
\]

**Discussion**

The calculation regarding \(\text{K}_a\) and \(\text{K}_b\) conversion is simple, but understanding what problems require this type of conversion is difficult. The concept is rather useful, and it further broadens the concept of acid and base.

**K\textsubscript{b} Values of Conjugate Bases of Acids**

We can also calculate the \(\text{K}_b\) value of the conjugate base from the \(\text{K}_a\) value of its conjugate acid. The principle is the same.
as that used to calculate the $K_a$ values of the conjugate acid of a base as we have just discussed. Let $\text{A}^-$ be the conjugate base of an acid $\text{HA}$. Then the expression for the equilibrium constant for the reaction

$$\text{A}^- + \text{H}_2\text{O} \rightleftharpoons \text{HA} + \text{OH}^-$$

can be written as

$$K_b = \frac{[\text{HA}] [\text{OH}^-]}{[\text{A}^-]}$$

Multiplying the numerator and denominator with $\frac{[\text{H}^+]}{[\text{H}^+]}$ leads to

$$K_b = \frac{[\text{HA}] [\text{OH}^-]}{[\text{A}^-] \frac{[\text{H}^+]}{[\text{H}^+]}}$$

Rearrangement gives

$$\begin{align}
K_b &= \frac{[\text{HA}]}{[\text{A}^-] [\text{H}^+]} \frac{[\text{OH}^-] [\text{H}^+]}{[\text{H}^+]} \\
&= \frac{[\text{HA}]}{[\text{A}^-] [\text{H}^+]} K_w \\
&= \frac{K_w}{K_a}
\end{align}$$

Thus,

$$\text{K}_a \text{K}_b = K_w$$

and this formula is the same as the one derived for the conjugate acid of a base. Again, at 298 K, we have

$$\text{K}_a \text{K}_b = 1E^{-14}$$

and the value for $K_w$ is larger than 1E-14 at higher temperatures. $K_w$ is smaller at temperature less than 298 K.

**Applications**

The concept of conjugate acid and base pairs is very useful for the consideration of acidity and basicity of salts. The applications of the relationship

$$\text{K}_a \text{K}_b = K_w$$

are further illustrated on the topic of Hydrolysis. Hydrolysis reactions are reactions of cations or anions of salts with water. As a result of these reactions, a salt solution is either acidic or basic.

**Confidence Building Questions**

1. **Calculate $K_b$ for the acetate ion from the $K_a$ for acetic acid of 1.8E-5.**

   Answer 5.6E-10

   Consider...

   $$K_b = \frac{1e^{-14}}{1.8e^{-5}} = 5.6E{-}10$$
If $K_b$ for the acetate ion is $5.6 \times 10^{-10}$, what is $K_a$ for acetic acid?

2. The $K_a$ for trimethylammonium ion $\text{(CH}_3\text{)}_3\text{NH}^+$ is $1.6 \times 10^{-10}$. Calculate $K_b$ for its conjugate base.

   **Answer** $6.25 \times 10^{-5}$

   Consider... \[ K_a = \frac{1 \times 10^{-14}}{1.6 \times 10^{-10}} = 6.25 \times 10^{-5} \]

   You know all about conjugate acid-base pairs now. Learning is a pleasure.

3. At some temperature, $K_w = 1 \times 10^{-13}$. Calculate the $K_b$ value for the acetate ion. ($K_a$ for acetic acid is $9.5 \times 10^{-5}$ at the same temperature).

   **Answer** $1.05 \times 10^{-9}$

   Consider... \[ K_b = \frac{1 \times 10^{-13}}{9.5 \times 10^{-5}} \]

   The acidic constants are dependent on temperature.

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**Contributors**

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