Learning Objectives

- Identify acids, bases, and conjugate acid-base pairs according to the Brønsted-Lowry definition
- Write equations for acid and base ionization reactions
- Use the ion-product constant for water to calculate hydronium and hydroxide ion concentrations
- Describe the acid-base behavior of amphiprotic substances

The Arrhenius definition of an acid as a compound that dissolves in water to yield hydronium ions (H$_3$O$^+$) and a base as a compound that dissolves in water to yield hydroxide ions (\(\ce{OH-}\)). This definition is not wrong; it is simply limited. We extended the definition of an acid or a base using the more general definition proposed in 1923 by the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry. Their definition centers on the proton, \(\ce{H^+}\). A proton is what remains when a normal hydrogen atom, \(\ce{^1_1H}\), loses an electron. A compound that donates a proton to another compound is called a Brønsted-Lowry acid, and a compound that accepts a proton is called a Brønsted-Lowry base. An acid-base reaction is the transfer of a proton from a proton donor (acid) to a proton acceptor (base). In a subsequent chapter of this text we will introduce the most general model of acid-base behavior introduced by the American chemist G. N. Lewis.

Note: Brønsted-Lowry Definitions

- A compound that **donates** a proton to another compound is called a Brønsted-Lowry acid.
- A compound that **accepts** a proton is called a Brønsted-Lowry base.

Acids may be compounds such as HCl or H$_2$SO$_4$, organic acids like acetic acid (\(\ce{CH_3COOH}\)) or ascorbic acid (vitamin C), or H$_2$O. Anions (such as \(\ce{HSO_4^-}\), \(\ce{H_2PO_4^-}\), \(\ce{HS^-}\), and \(\ce{HCO_3^-}\)) and cations (such as \(\ce{H_3O^+}\), \(\ce{NH_4^+}\), and \(\ce{[Al(H_2O)_6]^{3+}}\)) may also act as acids. Bases fall into the same three categories. Bases may be neutral molecules (such as \(\ce{H_2O}\), \(\ce{NH_3}\), and \(\ce{CH_3NH_2}\)), anions (such as \(\ce{OH^-}\), \(\ce{HS^-}\), \(\ce{HCO_3^-}\), \(\ce{CO_3^{2−}}\), \(\ce{F^-}\), and \(\ce{PO_4^{3−}}\)), or cations (such as \(\ce{[Al(H_2O)_5OH]^{2+}}\)). The most familiar bases are ionic compounds such as \(\ce{NaOH}\) and \(\ce{Ca(OH)_2}\), which contain the hydroxide ion, \(\ce{OH^-}\). The hydroxide ion in these compounds accepts a proton from acids to form water:

\[
\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \quad \text{(14.2.1)}
\]

We call the product that remains after an acid donates a proton the **conjugate base** of the acid. This species is a base because it can accept a proton (to re-form the acid):

\[
\text{acid} \rightleftharpoons \text{proton} + \text{conjugate base} \quad \text{(14.2.2a)}
\]

\[
\text{HF} \rightleftharpoons \text{H}^+ + \text{F}^- \quad \text{(14.2.2b)}
\]

\[
\text{H}_2\text{SO}_4 \rightleftharpoons \text{H}^+ + \text{HSO}_4^- \quad \text{(14.2.2c)}
\]

\[
\text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{OH}^- \quad \text{(14.2.2d)}
\]

\[
\text{HSO}_4^- \rightleftharpoons \text{H}^+ + \text{SO}_4^{2−} \quad \text{(14.2.2e)}
\]
We call the product that results when a base accepts a proton the base's **conjugate acid**. This species is an acid because it can give up a proton (and thus re-form the base):

\[
\text{base} + \text{proton} \rightleftharpoons \text{conjugate acid} \label{14.2.3a}\]

\[
\text{OH}^- + \text{H}^+ \rightleftharpoons \text{H}_2\text{O} \label{14.2.3b}\]

\[
\text{H}_2\text{O} + \text{H}^+ \rightleftharpoons \text{H}_3\text{O}^+ \label{14.2.3c}\]

\[
\text{NH}_3 + \text{H}^+ \rightleftharpoons \text{NH}_4^+ \label{14.2.3d}\]

\[
\text{S}^{2-} + \text{H}^+ \rightleftharpoons \text{HS}^- \label{14.2.3e}\]

\[
\text{CO}_3^{2-} + \text{H}^+ \rightleftharpoons \text{HCO}_3^- \label{14.2.3f}\]

\[
\text{F}^- + \text{H}^+ \rightleftharpoons \text{HF} \label{14.2.3g}\]

In these two sets of equations, the behaviors of acids as proton donors and bases as proton acceptors are represented in isolation. In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, \(\text{OH}^-\), and the conjugate acid of ammonia, \(\text{NH}_4^+\):

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

Similarly, in the reaction of acetic acid with water, acetic acid donates a proton to water, which acts as the base. In the reverse reaction, \(\text{H}_3\text{O}^+\) is the acid that donates a proton to the acetate ion, which acts as the base. Once again, we have two conjugate acid–base pairs: the parent acid and its conjugate base (\((\text{CH}_3\text{CO}_2\text{H}/\text{CH}_3\text{CO}_2^−)\)) and the parent base and its conjugate acid (\((\text{H}_3\text{O}^+/\text{H}_2\text{O})\)).

In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting
a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The conjugate acid–base pairs for this reaction are \(\text{NH}_4^+/\text{NH}_3\) and \(\text{H}_2\text{O}/\text{OH}^-\).

Figure 16.2.1. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

**Figure 16.2.1: The Relative Strengths of Some Common Conjugate Acid–Base Pairs**

The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong...
Acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

**Amphiprotic Species**

Like water, many molecules and ions may either gain or lose a proton under the appropriate conditions. Such species are said to be amphiprotic. Another term used to describe such species is amphoteric, which is a more general term for a species that may act either as an acid or a base by any definition (not just the Brønsted-Lowry one). Consider, for example, the bicarbonate ion, which may either donate or accept a proton as shown here:

\[
\ce{HCO^-}_{3(aq)} + \ce{H_2O}_{(l)} \rightleftharpoons \ce{CO^{2-}}_{3(aq)} + \ce{H_3O^+}_{(aq)} \tag{14.2.5a}
\]

\[
\ce{HCO^-}_{3(aq)} + \ce{H_2O}_{(l)} \rightleftharpoons \ce{H_2CO}_{3(aq)} + \ce{OH^-}_{(aq)} \tag{14.2.5b}
\]

**Example 14.2.3: The Acid-Base Behavior of an Amphoteric Substance**

Write separate equations representing the reaction of \(\ce{HSO3^-}\)

- as an acid with \(\ce{OH^-}\)
- as a base with HI

**Solution**

- \(\ce{HSO3^-}(aq)+\ce{OH^-}(aq)\rightleftharpoons\ce{SO3^2-}(aq)+\ce{H_2O}_{(l)}\)
- \(\ce{HSO3^-}(aq)+\ce{HI}(aq)\rightleftharpoons\ce{H2SO3}(aq)+\ce{I^-}(aq)\)

**Exercise 14.2.3**

Write separate equations representing the reaction of \(\ce{H2PO4^-}\)

- as a base with HBr
- as an acid with \(\ce{OH^-}\)

**Answer**

- \(\ce{H2PO4^-}(aq)+\ce{HBr}(aq)\rightleftharpoons\ce{H3PO4}(aq)+\ce{Br^-}(aq)\)
- \(\ce{H2PO4^-}(aq)+\ce{OH^-}(aq)\rightleftharpoons\ce{HPO4^{2-}}(aq)+\ce{H_2O}_{(l)}\)

**Key Concepts and Summary**

A compound that can donate a proton (a hydrogen ion) to another compound is called a Brønsted-Lowry acid. The compound that accepts the proton is called a Brønsted-Lowry base. The species remaining after a Brønsted-Lowry acid has lost a proton is the conjugate base of the acid. The species formed when a Brønsted-Lowry base gains a proton is the conjugate acid of the base. Thus, an acid-base reaction occurs when a proton is transferred from an acid to a base, with formation of the conjugate base of the reactant acid and formation of the conjugate acid of the reactant base. Amphiprotic species can act as both proton donors and proton acceptors. Water is the most important amphiprotic species. It can form
both the hydronium ion, \( \text{H}_3\text{O}^+ \), and the hydroxide ion, \( \text{OH}^- \) when it undergoes autoionization:

\[
\text{\small{\text{\(2 \text{H}_2\text{O}}\text{_{(l)} = \text{H}_3\text{O}^+\text{(aq)} + \text{OH}^-\text{(aq)}\)}}}
\]

The ion product of water, \( K_w \) is the equilibrium constant for the autoionization reaction:

\[
\text{\small{\( K_w = \text{[H}_2\text{O}^+]\text{[OH}^-\text{]} = 1.0 \times 10^{−14} \text{ at 25°C}\)}}
\]

**Key Equations**

- \( K_w = \text{[H}_3\text{O}^+]\text{[OH}^-\text{]} = 1.0 \times 10^{−14} \text{ at 25°C}\)

**Glossary**

- **acid ionization**
  - reaction involving the transfer of a proton from an acid to water, yielding hydronium ions and the conjugate base of the acid

- **amphiprotic**
  - species that may either gain or lose a proton in a reaction

- **amphoteric**
  - species that can act as either an acid or a base

- **autoionization**
  - reaction between identical species yielding ionic products; for water, this reaction involves transfer of protons to yield hydronium and hydroxide ions

- **base ionization**
  - reaction involving the transfer of a proton from water to a base, yielding hydroxide ions and the conjugate acid of the base

- **Brønsted-Lowry acid**
  - proton donor

- **Brønsted-Lowry base**
  - proton acceptor

- **conjugate acid**
  - substance formed when a base gains a proton

- **conjugate base**
  - substance formed when an acid loses a proton

- **ion-product constant for water (K_w)**
  - equilibrium constant for the autoionization of water
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