Pearson

16.3: Definitions of Acids and Bases

Key Concept Video Definitions of Acids and Bases

In this section, we examine two different definitions of acids and bases: the Arrhenius definition and the Brønsted–Lowry definition (later in the chapter we examine a third one called the Lewis definition). Why three definitions, and which one is correct? As Lewis himself noted in the opening quote of this chapter, each definition has limits, and no single definition is "correct." Rather, each definition is useful in a given instance.

The Arrhenius Definition

In the 1880s, Swedish chemist Svante Arrhenius (1859–1927) proposed the following molecular definitions of acids and bases:

Acid—A substance that produces H^+ ions in aqueous solution

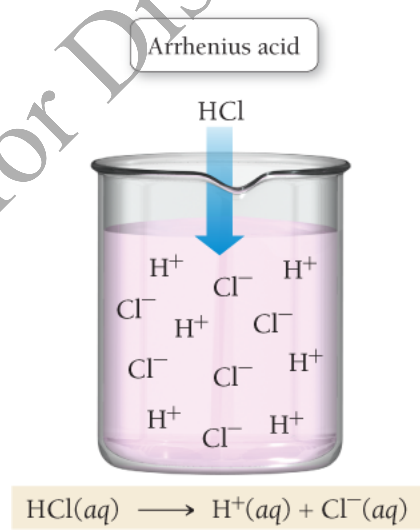
Base—A substance that produces OH^- ions in aqueous solution

According to the **Arrhenius definition**, HCl is an acid because it produces H^+ ions in solution (Figure 16.1).

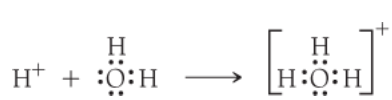
$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Figure 16.1 Arrhenius Acid

An Arrhenius acid produces H^+ ions in solution.



Hydrogen monochloride (HCl) is a covalent compound and does not contain ions. However, in water it *ionizes* completely to form $H^+(aq)$ ions and $Cl^-(aq)$ ions. The H^+ ions are highly reactive. In aqueous solution, the ions bond to water to form H_3O^+:



The H_3O^+ ion is called the **hydronium ion**. In water, H^+ ions always associate with H_2O molecules to form hydronium ions and other associated species with the general formula $H(H_2O)_n^+$. For example, an H^+ ion can associate with two water molecules to form $H(H_2O)_2^+$, with three to form $H(H_2O)_3^+$, and so on. Chemists use $H^+(aq)$ and $H_3O^+(aq)$ interchangeably to mean the same thing—an H^+ ion that has been solvated (or dissolved) in water.

Recall from Section 9.4 that the strength of an electrolyte (a substance, such as an acid, that forms ions in solution) depends on the extent of the electrolyte's dissociation into its component ions in solution. A *strong electrolyte* completely dissociates into ions in solution, whereas a *weak electrolyte* only partially dissociates. We define strong and weak acids accordingly. A **strong acid** completely ionizes in solution, whereas a **weak acid** only partially ionizes. We represent the ionization of a strong acid with a single arrow and that of a weak acid with an equilibrium arrow. In this notation, HA represents a generic acid with the anion A^-.

Single arrow indicates complete ionization.

$$HA(aq) \longrightarrow H^+(aq) + A^-(aq)$$

Strong acid

Double arrow indicates partial ionization.

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

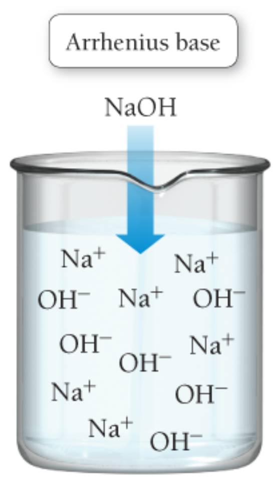
Weak acid

According to the Arrhenius definition, NaOH is a base because it produces OH^- ions in solution (Figure 16.2).

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

Figure 16.2 Arrhenius Base

An Arrhenius base produces OH^- ions in solution.



$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

NaOH is an ionic compound and therefore contains Na^+ and OH^- ions. When NaOH is added to water, it *dissociates* or breaks apart into its component ions. NaOH is an example of a **strong base**, one that completely dissociates in solution (analogous to a strong acid). A **weak base** is analogous to a weak acid. Unlike strong bases that contain OH^- and *dissociate* in water, the most common weak bases produce OH^- by accepting a proton from water and ionizing water to form OH^- according to the general equation:

proton from water and ionizing water to form OH^- according to the general equation:

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

In this equation, B is a generic symbol for a weak base. Ammonia, for example, ionizes water as follows:

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

We examine weak bases more thoroughly in Section 16.8.

Under the Arrhenius definition, acids and bases combine to form water, neutralizing each other in the process:

$$H^+\,(aq) + OH^-\,(aq) \rightarrow H_2O\,(l)$$

The Brønsted–Lowry Definition

A second, more widely applicable definition of acids and bases, called the **Brønsted–Lowry definition**, was introduced in 1923. This definition focuses on the *transfer of* H^+ *ions* in an acid–base reaction. Since an H^+ ion is a proton—a hydrogen atom without its electron—this definition focuses on the idea of a proton donor and a proton acceptor:

Acid—Proton $\left(H^+\text{ion}\right)$ *donor*

Base—Proton $\left(H^+\text{ion}\right)$ *acceptor*

According to this definition, HCl is an acid because, in solution, it donates a proton to water:

$$HCl\,(aq) + H_2O\,(l) \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)$$

This definition clearly describes what happens to the H^+ ion from an acid—it associates with a water molecule to form H_3O^+ (a hydronium ion). The Brønsted–Lowry definition also applies nicely to bases (such as NH_3) that do not inherently contain OH^- ions but still produce OH^- ions in solution. According to the Brønsted–Lowry definition, NH_3 is a base because it accepts a proton from water:

$$NH_3\,(aq) + H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)$$

> All Arrhenius acids and bases are acids and bases under the Brønsted–Lowry definition. However, some Brønsted–Lowry acids and bases cannot be classified as Arrhenius acids and bases.

According to the Brønsted–Lowry definition, acids (proton donors) and bases (proton acceptors) always occur together. In the reaction between HCl and H_2O, HCl is the proton donor (acid) and H_2O is the proton acceptor (base):

$$\begin{array}{ccl} HCl\,(aq) & + & H_2O\,(l) \rightarrow H_3O^+\,(aq) + Cl^-\,(aq) \\ \text{acid} & & \text{base} \\ \text{(proton donor)} & & \text{(proton acceptor)} \end{array}$$

In the reaction between NH_3 and H_2O, H_2O is the proton donor (acid) and NH_3 is the proton acceptor (base):

$$\begin{array}{ccl} NH_3\,(aq) & + & H_2O\,(l) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq) \\ \text{base} & & \text{acid} \\ \text{(proton acceptor)} & & \text{(proton donor)} \end{array}$$

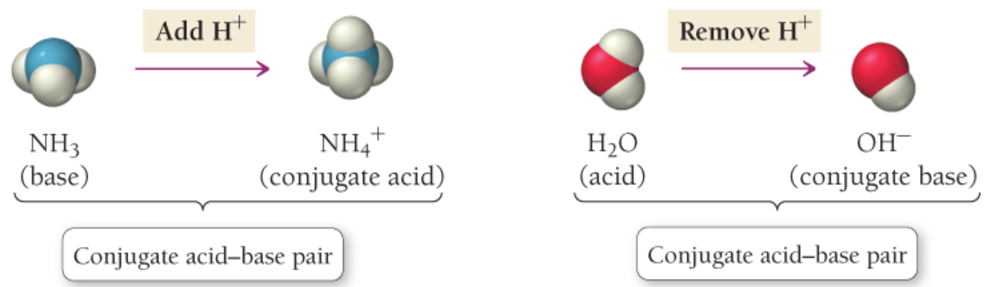
According to the Brønsted–Lowry definition, some substances—such as water in the previous two equations—can act as acids or bases. Substances that can act as acids or bases are **amphoteric**. Notice what happens when we reverse an equation representing Brønsted–Lowry acid–base behavior:

$$\begin{array}{ccl} NH_4^+\,(aq) & + & OH^-\,(aq) \rightleftharpoons NH_3\,(aq) + H_2O\,(l) \\ \text{acid} & & \text{base} \\ \text{(proton donor)} & & \text{(proton acceptor)} \end{array}$$

In this reaction, NH_4^+ is the proton donor (acid) and OH^- is the proton acceptor (base). The substance that was the base (NH_3) has become the acid (NH_4^+) and vice versa. NH_4^+ and NH_3 are often referred to as a **conjugate acid–base pair**, two substances related to each other by the transfer of a proton (Figure 16.3). A **conjugate acid** is any base to which a proton has been added, and a **conjugate base** is any acid from which a proton has been removed. Going back to the original forward reaction, we can identify the conjugate acid–base pairs:

Figure 16.3 Conjugate Acid–Base Pairs

A conjugate acid–base pair consists of two substances related to each other by the transfer of a proton.



$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Base — Acid — Conjugate acid — Conjugate base

Summarizing the Brønsted–Lowry Definition of an Acid–Base Reaction:

- A base accepts a proton and becomes a conjugate acid.
- An acid donates a proton and becomes a conjugate base.

Example 16.1 Identifying Brønsted–Lowry Acids and Bases and Their Conjugates

In each reaction, identify the Brønsted–Lowry acid, the Brønsted–Lowry base, the conjugate acid, and the conjugate base.

a. $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$

b. $HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$

SOLUTION

a. Because H_2SO_4 donates a proton to H_2O in this reaction, it is the acid (proton donor). After H_2SO_4 donates the proton, it becomes HSO_4^-, the conjugate base. Because H_2O accepts a proton, it is the base (proton acceptor). After H_2O accepts the proton, it becomes H_3O^+, the conjugate acid.

$$H_2SO_4(aq) + H_2O(l) \longrightarrow HSO_4^-(aq) + H_3O^+(aq)$$

$$H_2SO_4(aq) + H_2O(l) \longrightarrow HSO_4^-(aq) + H_3O^+(aq)$$

Acid — Base — Conjugate base — Conjugate acid

b. Because H_2O donates a proton to HCO_3^- in this reaction, it is the acid (proton donor). After H_2O donates the proton, it becomes OH^-, the conjugate base. Because HCO_3^- accepts a proton, it is the base (proton acceptor). After HCO_3^- accepts the proton, it becomes H_2CO_3, the conjugate acid.

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$$

$$\underset{\text{Base}}{HCO_3^-(aq)} + \underset{\text{Acid}}{H_2O(l)} \rightleftharpoons \underset{\substack{\text{Conjugate}\\ \text{acid}}}{H_2CO_3(aq)} + \underset{\substack{\text{Conjugate}\\ \text{base}}}{OH^-(aq)}$$

FOR PRACTICE 16.1 In each reaction, identify the Brønsted–Lowry acid, the Brønsted–Lowry base, the conjugate acid, and the conjugate base.

a. $C_5H_5N\,(aq) + H_2O\,(l) \rightleftharpoons C_5H_5NH^+\,(aq) + OH^-\,(aq)$

b. $HNO_3\,(aq) + H_2O\,(l) \rightleftharpoons H_3O^+\,(aq) + NO_3^-\,(aq)$

Interactive Worked Example 16.1 **Identifying Brønsted-Lowry Acids and Bases and Their Conjugates**

Conceptual Connection 16.1 **Conjugate Acid–Base Pairs**

Interactive