16.11 LEWIS ACIDS AND BASES

For a substance to be a proton acceptor (a Brønsted–Lowry base), it must have an unshared pair of electrons for binding the proton. \( \text{NH}_3 \), for example, acts as a proton acceptor. Using Lewis structures, we can write the reaction between \( \text{H}^+ \) and \( \text{NH}_3 \) as follows:

\[
\begin{align*}
\text{H}^+ + \text{N} & \quad \text{H} \\
\text{H} & \quad \text{N} \\
\text{H} & \quad \text{H}
\end{align*}
\]

G. N. Lewis was the first to notice this aspect of acid–base reactions. He proposed a definition of acid and base that emphasizes the shared electron pair: A **Lewis acid** is an electron-pair acceptor, and a **Lewis base** is an electron-pair donor.

Every base that we have discussed thus far—whether it be \( \text{OH}^- \), \( \text{H}_2\text{O} \), an amine, or an anion—is an electron-pair donor. Everything that is a base in the Brønsted–Lowry sense (a proton acceptor) is also a base in the Lewis sense (an electron-pair donor). In the Lewis theory, however, a base can donate its electron pair to something other than \( \text{H}^+ \). The Lewis definition therefore greatly increases the number of species that can be considered acids; \( \text{H}^+ \) is a Lewis acid, but not the only one. For example, consider the reaction between \( \text{NH}_3 \) and \( \text{BF}_3 \). This reaction occurs because \( \text{BF}_3 \) has a vacant orbital in its valence shell. (Section 8.7) It therefore acts as an electron-pair acceptor (a Lewis acid) toward \( \text{NH}_3 \), which donates the electron pair. The curved arrow shows the donation of a pair of electrons from \( \text{N} \) to \( \text{B} \) to form a covalent bond:

\[
\begin{align*}
\text{H} & \quad \text{F} \\
\text{H} & \quad \text{N} \\
\text{H} & \quad \text{F}
\end{align*}
\]

\[
\begin{align*}
\text{H} & \quad \text{F} \\
\text{H} & \quad \text{N} \\
\text{H} & \quad \text{F}
\end{align*}
\]

**GIVE IT SOME THOUGHT**

What feature must any molecule or ion have to act as a Lewis base?

Our emphasis throughout this chapter has been on water as the solvent and on the proton as the source of acidic properties. In such cases we find the Brønsted–Lowry definition of acids and bases to be the most useful. In fact, when we speak of a substance as being acidic or basic, we are usually thinking of aqueous solutions and using these terms in the Arrhenius or Brønsted–Lowry sense. The advantage of the Lewis theory is that it allows us to treat a wider variety of reactions, including those that do not involve proton transfer, as acid–base reactions. To avoid confusion, a substance such as \( \text{BF}_3 \) is rarely called an acid unless it is clear from the context that we are using the term in the sense of the Lewis definition. Instead, substances that function as electron-pair acceptors are referred to explicitly as “Lewis acids.”

Lewis acids include molecules that, like \( \text{BF}_3 \), have an incomplete octet of electrons. In addition, many simple cations can function as Lewis acids. For example, \( \text{Fe}^{3+} \) interacts strongly with cyanide ions to form the ferricyanide ion, \( \text{Fe(CN)}_6^{3-} \):

\[
\text{Fe}^{3+} + 6\text{[C≡N]}^- \rightarrow \text{[Fe(C≡N)}_6]^{3-}
\]

The \( \text{Fe}^{3+} \) ion has vacant orbitals that accept the electron pairs donated by the cyanide ions; we will learn more in Chapter 24 about just which orbitals are used by the \( \text{Fe}^{3+} \) ion. The metal ion is highly charged, too, which contributes to the interaction with \( \text{CN}^- \) ions.