

8.7: Lewis Acids and Bases

Learning Objectives

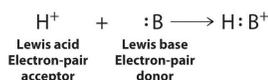
- To identify Lewis acids and bases.

A **Lewis acid** is a compound with a strong tendency to *accept* an additional pair of electrons from a **Lewis base**, which can *donate* a pair of electrons. Such an acid–base reaction forms an **adduct**, which is a compound with a *coordinate covalent bond* in which both electrons are provided by only one of the atoms. **Electron-deficient molecules**, which have less than an octet of electrons around one atom, are relatively common. They tend to acquire an octet electron configuration by reacting with an atom having a lone pair of electrons. Learning Objective is to identify Lewis acids and bases.

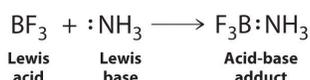
Introduction

The Brønsted–Lowry concept of acids and bases defines a base as any species that can accept a proton, and an acid as any substance that can donate a proton. Lewis proposed an alternative definition that focuses on *pairs of electrons* instead. A Lewis base is defined as any species that can *donate* a pair of electrons, and a Lewis acid is any species that can *accept* a pair of electrons. All Brønsted–Lowry bases (proton acceptors), such as OH^- , H_2O , and NH_3 , are also electron-pair donors. Thus the Lewis definition of acids and bases does not contradict the Brønsted–Lowry definition. Rather, it expands the definition of acids to include substances other than the H^+ ion.

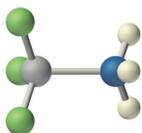
Electron-deficient molecules, such as BCl_3 , contain less than an octet of electrons around one atom and have a strong tendency to gain an additional pair of electrons by reacting with substances that possess a lone pair of electrons. Lewis's definition, which is less restrictive than either the Brønsted–Lowry or the Arrhenius definition, grew out of his observation of this tendency. A general Brønsted–Lowry acid–base reaction can be depicted in Lewis electron symbols as follows:



The proton (H^+), which has no valence electrons, is a Lewis acid because it accepts a lone pair of electrons on the base to form a bond. The proton, however, is just one of many electron-deficient species that are known to react with bases. For example, neutral compounds of boron, aluminum, and the other **Group 13 elements**, which possess only six valence electrons, have a very strong tendency to gain an additional electron pair. Such compounds are therefore potent Lewis acids that react with an electron-pair donor such as ammonia to form an acid–base adduct, a new covalent bond, as shown here for boron trifluoride (BF_3):



The bond formed between a Lewis acid and a Lewis base is a *coordinate covalent bond* because both electrons are provided by only one of the atoms (N, in the case of $\text{F}_3\text{B:NH}_3$). After it is formed, however, a coordinate covalent bond behaves like any other covalent single bond.



Boron trifluoride-ammonia adduct

Species that are very weak Brønsted–Lowry bases can be relatively strong Lewis bases. For example, many of the group 13 trihalides are highly soluble in ethers ($\text{R-O-R}'$) because the oxygen atom in the ether contains two lone pairs of electrons, just as in H_2O . Hence the predominant species in solutions of electron-deficient trihalides in ether solvents is a Lewis acid–base adduct. A reaction of this type is shown in Figure 8.7.1 for boron trichloride and diethyl ether:

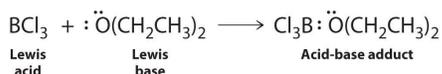
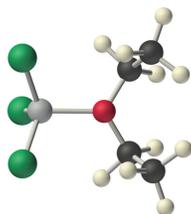


Figure 8.7.1: Lewis Acid/Base reaction of boron trichloride and diethyl ether reaction

Many molecules with multiple bonds can act as Lewis acids. In these cases, the Lewis base typically donates a pair of electrons to form a bond to the central atom of the molecule, while a pair of electrons displaced from the multiple bond becomes a lone pair on a terminal atom.



Boron trichloride-diethyl ether adduct

A typical example is the reaction of the hydroxide ion with carbon dioxide to give the bicarbonate ion, as shown in Figure 8.7.2. The highly electronegative oxygen atoms pull electron density away from carbon, so the carbon atom acts as a Lewis acid. Arrows indicate the direction of electron flow.

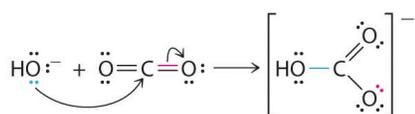


Figure 8.7.2: Lewis Acid/Base reaction of the hydroxide ion with carbon dioxide

✓ Example 8.7.1

Identify the acid and the base in each Lewis acid–base reaction.

- $\text{BH}_3 + (\text{CH}_3)_2\text{S} \rightarrow \text{H}_3\text{B}:\text{S}(\text{CH}_3)_2$
- $\text{CaO} + \text{CO}_2 \rightarrow \text{CaCO}_3$
- $\text{BeCl}_2 + 2 \text{Cl}^- \rightarrow \text{BeCl}_4^{2-}$

Given: reactants and products

Asked for: identity of Lewis acid and Lewis base

Strategy:

In each equation, identify the reactant that is electron deficient and the reactant that is an electron-pair donor. The electron-deficient compound is the Lewis acid, whereas the other is the Lewis base.

Solution:

- In BH_3 , boron has only six valence electrons. It is therefore electron deficient and can accept a lone pair. Like oxygen, the sulfur atom in $(\text{CH}_3)_2\text{S}$ has two lone pairs. Thus $(\text{CH}_3)_2\text{S}$ donates an electron pair on sulfur to the boron atom of BH_3 . The Lewis base is $(\text{CH}_3)_2\text{S}$, and the Lewis acid is BH_3 .
- As in the reaction shown in Equation 8.21, CO_2 accepts a pair of electrons from the O^{2-} ion in CaO to form the carbonate ion. The oxygen in CaO is an electron-pair donor, so CaO is the Lewis base. Carbon accepts a pair of electrons, so CO_2 is the Lewis acid.
- The chloride ion contains four lone pairs. In this reaction, each chloride ion donates one lone pair to BeCl_2 , which has only four electrons around Be. Thus the chloride ions are Lewis bases, and BeCl_2 is the Lewis acid.

? Exercise 8.7.1

Identify the acid and the base in each Lewis acid–base reaction.

- $(\text{CH}_3)_2\text{O} + \text{BF}_3 \rightarrow (\text{CH}_3)_2\text{O}:\text{BF}_3$
- $\text{H}_2\text{O} + \text{SO}_3 \rightarrow \text{H}_2\text{SO}_4$

Answer

- Lewis base: $(\text{CH}_3)_2\text{O}$; Lewis acid: BF_3

b. Lewis base: H_2O ; Lewis acid: SO_3

 Note

- Electron-deficient molecules (those with less than an octet of electrons) are Lewis acids.
- The acid-base behavior of many compounds can be explained by their Lewis electron structures.

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