

## 1.18: BRØNSTED-LOWRY ACIDS AND BASES (REVIEW)

### learning objective

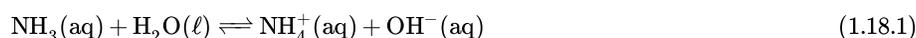
- recognize acids and bases

### THE BRØNSTED-LOWRY THEORY OF ACIDS AND BASES

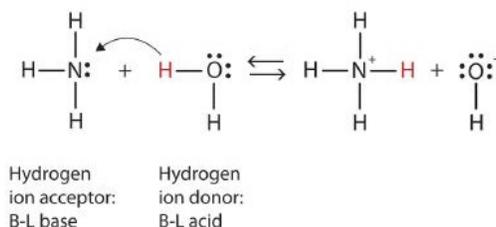
In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton ( $\text{H}^+$ ) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, a **Brønsted-Lowry acid** is a **proton donor (PD)**, while a **Brønsted-Lowry base** is a **proton acceptor (PA)**.

*A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.*

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:



What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows:

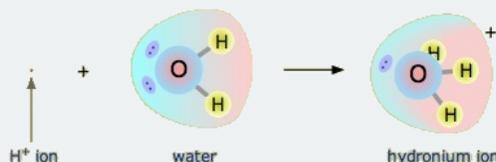


Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

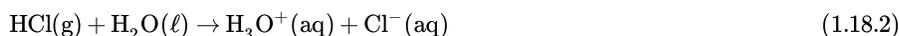
Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the  $\text{H}^+$  ion attaches itself to  $\text{H}_2\text{O}$  to make  $\text{H}_3\text{O}^+$ , which is called the *hydronium ion*. For most purposes,  $\text{H}^+$  and  $\text{H}_3\text{O}^+$  represent the same species, but writing  $\text{H}_3\text{O}^+$  instead of  $\text{H}^+$  shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

### The Hydronium Ion

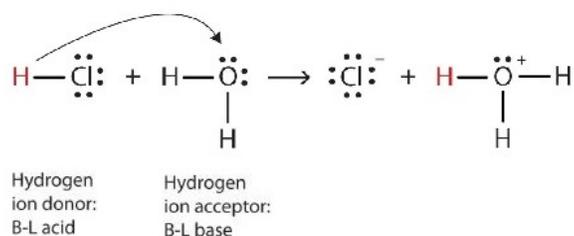
A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like  $\text{H}_5\text{O}_2^+$  or  $\text{H}_9\text{O}_4^+$  rather than  $\text{H}_3\text{O}^+$ . It is simpler, however, to use  $\text{H}_3\text{O}^+$  to represent the hydronium ion.



With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in  $\text{H}_2\text{O}$ :



We can depict this process using Lewis electron dot diagrams:



Now we see that a hydrogen ion is transferred from the HCl molecule to the H<sub>2</sub>O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H<sub>2</sub>O is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H<sub>2</sub>O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H<sub>2</sub>O a base in this circumstance.

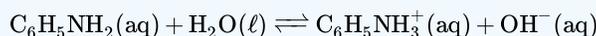
- A Brønsted-Lowry acid is a proton (hydrogen ion) donor.
- A Brønsted-Lowry base is a proton (hydrogen ion) acceptor.
- All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

### Example 1.18.1

Aniline (C<sub>6</sub>H<sub>5</sub>NH<sub>2</sub>) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

#### Solution

C<sub>6</sub>H<sub>5</sub>NH<sub>2</sub> and H<sub>2</sub>O are the reactants. When C<sub>6</sub>H<sub>5</sub>NH<sub>2</sub> accepts a proton from H<sub>2</sub>O, it gains an extra H and a positive charge and leaves an OH<sup>-</sup> ion behind. The reaction is as follows:



Because C<sub>6</sub>H<sub>5</sub>NH<sub>2</sub> accepts a proton, it is the Brønsted-Lowry base. The H<sub>2</sub>O molecule, because it donates a proton, is the Brønsted-Lowry acid.

### Exercise 1.18.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.



#### Answer:

Brønsted-Lowry acid: H<sub>2</sub>PO<sub>4</sub><sup>-</sup>; Brønsted-Lowry base: H<sub>2</sub>O

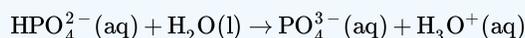
### Exercise 1.18.2

Which of the following compounds is a Brønsted-Lowry base?

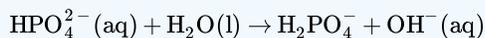
- HCl
- HPO<sub>4</sub><sup>2-</sup>
- H<sub>3</sub>PO<sub>4</sub>
- NH<sub>4</sub><sup>+</sup>
- CH<sub>3</sub>NH<sub>3</sub><sup>+</sup>

#### Answer:

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an H<sup>+</sup>. This eliminates HCl, H<sub>3</sub>PO<sub>4</sub>, NH<sub>4</sub><sup>+</sup> and CH<sub>3</sub>NH<sub>3</sub><sup>+</sup> because they are Brønsted-Lowry acids. They all give away protons. In the case of HPO<sub>4</sub><sup>2-</sup>, consider the following equation:



Here, it is clear that  $\text{HPO}_4^{2-}$  is the acid since it donates a proton to water to make  $\text{H}_3\text{O}^+$  and  $\text{PO}_4^{3-}$ . Now consider the following equation:

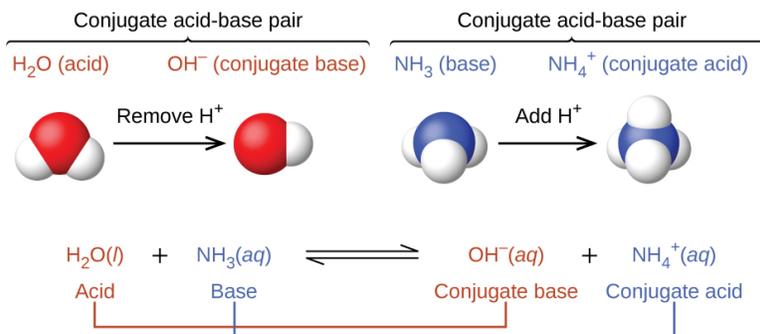


In this case,  $\text{HPO}_4^{2-}$  is the base since it accepts a proton from water to form  $\text{H}_2\text{PO}_4^-$  and  $\text{OH}^-$ . Thus,  $\text{HPO}_4^{2-}$  is an acid and base together, making it amphoteric.

Since  $\text{HPO}_4^{2-}$  is the only compound from the options that can act as a base, the answer is **(b)  $\text{HPO}_4^{2-}$** .

## CONJUGATE ACID-BASE PAIR

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,  $\text{NH}_4^+$ :



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  $\text{H}_2\text{O}/\text{OH}^-$ .

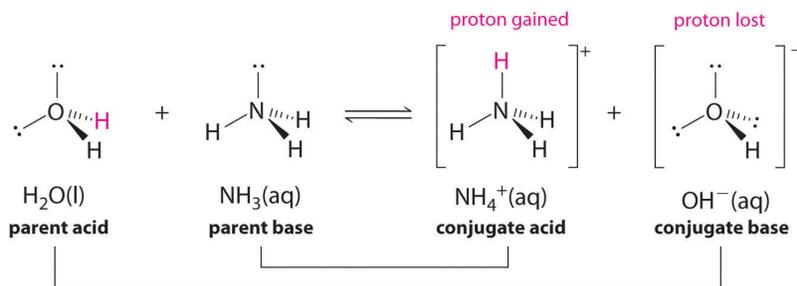


Figure 1.18.1. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

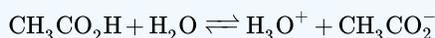
ACID		BASE	
negligible	OH <sup>-</sup> HS <sup>-</sup>	O <sup>2-</sup> S <sup>2-</sup>	strong
← Relative acid strength increasing	H <sub>2</sub> O	OH <sup>-</sup>	Relative base strength increasing →
	HPO <sub>4</sub> <sup>2-</sup>	PO <sub>4</sub> <sup>3-</sup>	
	HCO <sub>3</sub> <sup>-</sup>	CO <sub>3</sub> <sup>2-</sup>	
	NH <sub>4</sub> <sup>+</sup>	NH <sub>3</sub>	
	HCN	CN <sup>-</sup>	
	H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	HPO <sub>4</sub> <sup>2-</sup>	
	HSO <sub>3</sub> <sup>-</sup>	SO <sub>3</sub> <sup>2-</sup>	
	H <sub>2</sub> S	HS <sup>-</sup>	
	H <sub>2</sub> CO <sub>3</sub>	HCO <sub>3</sub> <sup>-</sup>	
	C <sub>5</sub> H <sub>5</sub> NH <sup>+</sup>	C <sub>5</sub> H <sub>5</sub> N	
weak	CH <sub>3</sub> CO <sub>2</sub> H	CH <sub>3</sub> CO <sub>2</sub> <sup>-</sup>	
	HF	F <sup>-</sup>	
	H <sub>3</sub> PO <sub>4</sub>	H <sub>2</sub> PO <sub>4</sub> <sup>-</sup>	
	H <sub>2</sub> SO <sub>3</sub>	HSO <sub>3</sub> <sup>-</sup>	
	HSO <sub>4</sub> <sup>-</sup>	SO <sub>4</sub> <sup>2-</sup>	
	strong	H <sub>3</sub> O <sup>+</sup>	H <sub>2</sub> O
		HNO <sub>3</sub>	NO <sub>3</sub> <sup>-</sup>
		H <sub>2</sub> SO <sub>4</sub>	HSO <sub>4</sub> <sup>-</sup>
		HCl	Cl <sup>-</sup>
		HBr	Br <sup>-</sup>
			negligible

Figure 1.18.1: 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.

### Example 1.18.2

Identify the conjugate acid–base pairs in this equilibrium.

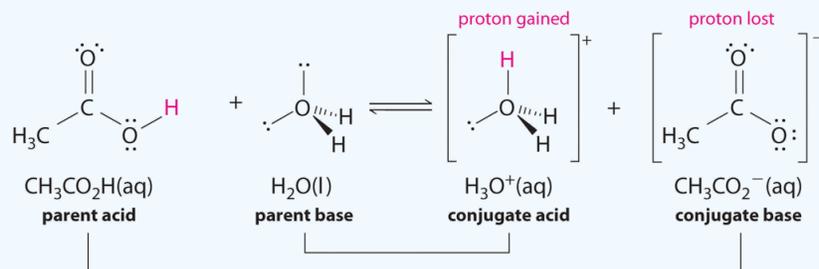


#### Solution

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.

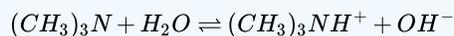
Once again, we have two conjugate acid–base pairs:

- 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 ( $\text{H}_3\text{O}^+/\text{H}_2\text{O}$ ).



**Example 1.18.3**

Identify the conjugate acid-base pairs in this equilibrium.

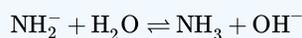
**Solution**

One pair is  $H_2O$  and  $OH^-$ , where  $H_2O$  has one more  $H^+$  and is the conjugate acid, while  $OH^-$  has one less  $H^+$  and is the conjugate base.

The other pair consists of  $(CH_3)_3N$  and  $(CH_3)_3NH^+$ , where  $(CH_3)_3NH^+$  is the conjugate acid (it has an additional proton) and  $(CH_3)_3N$  is the conjugate base.

**Exercise 1.18.3**

Identify the conjugate acid-base pairs in this equilibrium.

**Answer:**

$H_2O$  (acid) and  $OH^-$  (base);  $NH_2^-$  (base) and  $NH_3$  (acid)

**CONTRIBUTORS AND ATTRIBUTIONS**

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