

2.1: Brønsted–Lowry Acids and Bases

In 1923, chemists Johannes Brønsted and Martin Lowry independently developed definitions of acids and bases based on compounds abilities to either donate or accept protons (H^+ ions). Here, acids are defined as being able to donate protons in the form of hydrogen ions; whereas bases are defined as being able to accept protons. This took the Arrhenius definition one step further as water is no longer required to be present in the solution for acid and base reactions to occur.

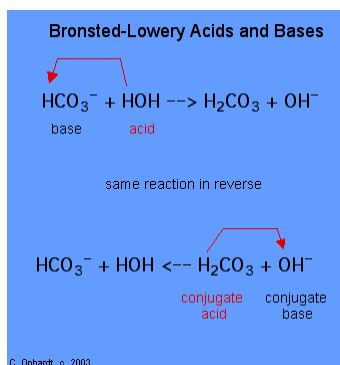
Brønsted-Lowry Definition

J.N. Brønsted and T.M. Lowry independently developed the theory of proton donors and proton acceptors in acid-base reactions, coincidentally in the same region and during the same year. The Arrhenius theory where acids and bases are defined by whether the molecule contains hydrogen and hydroxide ion is too limiting. The main effect of the Brønsted-Lowry definition is to identify the proton (H^+) transfer occurring in the acid-base reaction. This is best illustrated in the following equation:



Acid	Base	
Donates hydrogen ions	Accepts hydrogen ions.	
HCl^+	$HOH \rightarrow$	$H_3O^+ + Cl^-$
HOH^+	$NH_3 \rightarrow$	$NH_4^+ + OH^-$

The determination of a substance as a Brønsted-Lowry acid or base can only be done by observing the reaction. In the case of the HOH it is a base in the first case and an acid in the second case.



To determine whether a substance is an acid or a base, count the hydrogens on each substance before and after the reaction. If the number of hydrogens has decreased that substance is the acid (donates hydrogen ions). If the number of hydrogens has increased that substance is the base (accepts hydrogen ions). These definitions are normally applied to the reactants on the left. If the reaction is viewed in reverse a new acid and base can be identified. The substances on the right side of the equation are called conjugate acid and conjugate base compared to those on the left. Also note that the original acid turns in the conjugate base after the reaction is over.

Acids are Proton Donors and Bases are Proton Acceptors

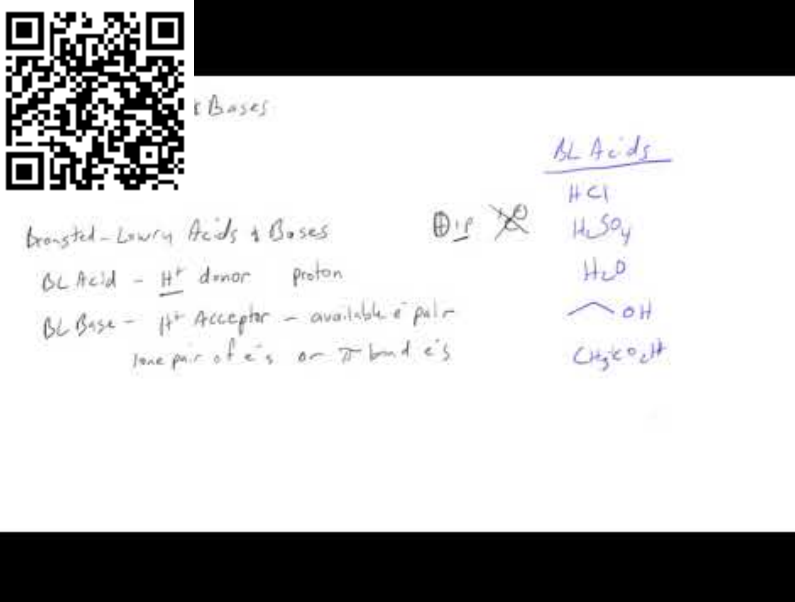
For a reaction to be in equilibrium a transfer of electrons needs to occur. The acid will give an electron away and the base will receive the electron. Acids and Bases that work together in this fashion are called a *conjugate pair* made up of *conjugate acids* and *conjugate bases*.



A stands for an Acidic compound and Z stands for a Basic compound

- A Donates H to form HZ^+ .
- Z Accepts H from A which forms HZ^+

- A^- becomes conjugate base of HA and in the reverse reaction it accepts a H from HZ to recreate HA in order to remain in equilibrium
- HZ^+ becomes a conjugate acid of Z and in the reverse reaction it donates a H to A^- recreating Z in order to remain in equilibrium



Questions

1. Why is HA an Acid?
2. Why is Z^- a Base?
3. How can A^- be a base when HA was an Acid?
4. How can HZ^+ be an acid when Z used to be a Base?

Now that we understand the concept, let's look at an example with actual compounds!



- HCl is the acid because it is donating a proton to H₂O
- H₂O is the base because H₂O is accepting a proton from HCl
- H₃O⁺ is the conjugate acid because it is donating an acid to Cl⁻ turn into its conjugate acid H₂O
- Cl⁻ is the conjugate base because it accepts an H from H₃O to return to its conjugate acid HCl

How can H₂O be a base? I thought it was neutral?

Answers

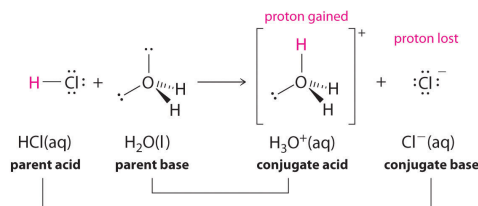
1. It has a proton that can be transferred
2. It receives a proton from HA
3. A^- is a conjugate base because it is in need of a H in order to remain in equilibrium and return to HA
4. HZ^+ is a conjugate acid because it needs to donate or give away its proton in order to return to its previous state of Z
5. In the Brønsted-Lowry Theory what makes a compound an acid or a base is whether or not it donates or accepts protons. If the H₂O was in a different problem and was instead donating an H rather than accepting an H it would be an acid!

Conjugate Acid-Base Pairs

In aqueous solutions, acids and bases can be defined in terms of the transfer of a proton from an acid to a base. Thus for every acidic species in an aqueous solution, there exists a species derived from the acid by the loss of a proton. These two species that differ by only a proton constitute a conjugate acid-base pair.

All acid-base reactions contain two conjugate acid-base pairs.

For example, in the reaction of HCl with water, HCl, the parent acid, donates a proton to a water molecule, the parent base, thereby forming Cl^- . Thus HCl and Cl^- constitute a conjugate acid–base pair. By convention, we always write a conjugate acid–base pair as the acid followed by its conjugate base. In the reverse reaction, the Cl^- ion in solution acts as a base to accept a proton from H_2O and HCl. Thus H_3O^+ and H_2O constitute a second conjugate acid–base pair. In general, any acid–base reaction must contain two conjugate acid–base pairs, which in this case are HCl/Cl^- and $\text{H}_3\text{O}^+/\text{H}_2\text{O}$.



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, H_3O^+ 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 ($\text{H}_3\text{O}^+/\text{H}_2\text{O}$).

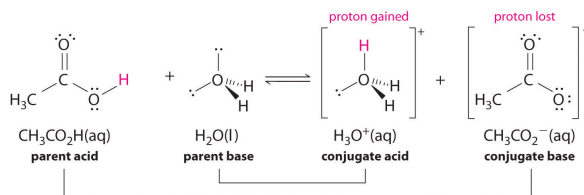


Figure 2.1.2

In the reaction of ammonia with water to give ammonium ions and hydroxide ions (Figure 16.3), 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 $\text{H}_2\text{O}/\text{OH}^-$.

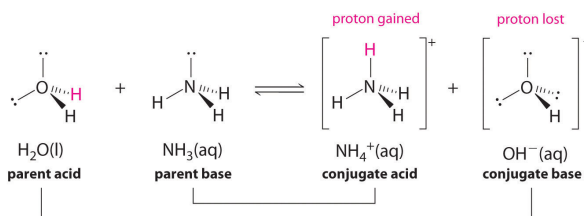


Figure 2.1.3: The Relative Strengths of Some Common Conjugate Acid–Base Pairs

Some common conjugate acid–base pairs are shown in Figure 2.1.4.

ACID		BASE	
negligible	OH^-	O^{2-}	strong
	HS^-	S^{2-}	
weak	H_2O	OH^-	weak
	HPO_4^{2-}	PO_4^{3-}	
	HCO_3^-	CO_3^{2-}	
	NH_4^+	NH_3	
	HCN	CN^-	
	H_2PO_4^-	HPO_4^{2-}	
	HSO_3^-	SO_3^{2-}	
	H_2S	HS^-	
	H_2CO_3	HCO_3^-	
	$\text{C}_5\text{H}_5\text{NH}^+$	$\text{C}_5\text{H}_5\text{N}$	
	$\text{CH}_3\text{CO}_2\text{H}$	CH_3CO_2^-	
	HF	F^-	
	H_3PO_4	H_2PO_4^-	
	H_2SO_3	HSO_3^-	
	HSO_4^-	SO_4^{2-}	
strong	H_3O^+	H_2O	negligible
	HNO_3	NO_3^-	
	H_2SO_4	HSO_4^-	
	HCl	Cl^-	
	HBr	Br^-	

← Relative acid strength increasing

Relative base strength increasing →

Figure 2.1.4: 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.

Contributors

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