**Arrhenius acids and bases**

According to the Arrhenius definition, an acid is any substance, which when dissolved in water, tends to increase the amount of H+. An example is HCl:

\begin{displaymath}
{\rm HCl}(g)\longrightarrow {\rm H}^+(aq) + {\rm Cl}^-(aq)
\end{displaymath}

An Arrhenius base is any substance, which when dissolved in water, tends to increase the amount of OH-. An example is NaOH:

\begin{displaymath}
{\rm NaOH}(s) \longrightarrow {\rm Na}^+(aq) + {\rm OH}^-
\end{displaymath}

These definitions are correct but not general enough to include the wide range of acid and base substances which are known to exist. In addition, they rely on the use of water as a solvent, which is also too narrow.

**Bronsted-Lowry acids and bases**

The Bronsted-Lowry definition is named for Johannes Bronsted and Thomas Lowry, who independently proposed it in 1923. A Bronsted-Lowry (BL) acid is defined as any substance that can donate a hydrogen ion (proton) and a Bronsted-Lowry base is any substance that can accept a hydrogen ion (proton). Thus, according to the BL definition, acids and bases must come in what is called *conjugate pairs*. For example, consider acetic acid dissolved in water:

\begin{displaymath}
{\rm CH}_3{\rm COOH}(aq) + {\rm H}_2{\rm O}(l) \rightleftharpoons
{\rm H}_3{\rm O}^+(aq) + {\rm CH}_3{\rm COO}^-(aq)
\end{displaymath}

Notice that we have written H2O(l) explicitly in these reactions. The reason is that acid/base dissociation occurs by a *proton transfer reaction* from an acid species to a specific water molecule. The transfer occurs through a hydrogen bond between the acid molecule and a solvating water molecule.

Here, CH3COOH is a BL acid because it can donate a proton, and CH3COO- its *conjugate* base because it can accept a proton. Note that H2O and H3O+ also form such a conjugate pair.

Note that the H3O+ rather than H+ has been used to denote the nature of H+ ions in water in the above reaction. This is really only a *very crude* representation of the true nature of solvated H+ ions.   
  
Similarly when ammonia is dissolved in water, one has

\begin{displaymath}
{\rm H}_2{\rm O}(l) + {\rm N}{\rm H}_3(aq)\rightleftharpoons {\rm N}{\rm H}_4^+(aq) + {\rm OH}^-(aq)
\end{displaymath}

Here, NH3 is the BL base and its conjugate acid is NH4+. Similarly, H2O acts as a BL acid and ${\rm OH}^-$acts as a BL base.

Another important advantage of the BL definition is that we are not limited to water as the solvent. Consider the reaction that occurs when HCl is dissolved in ammonia:

\begin{displaymath}
{\rm HCl}({\rm in}\; {\rm N}{\rm H}_3) + {\rm N}{\rm H}_3(l)...
...; {\rm N}{\rm H}_3) +
{\rm Cl}^-({\rm in}\; {\rm N}{\rm H}_3)
\end{displaymath}

Here, HCl acts as a BL acid with Cl- as its conjugate base. Also, NH3 acts as a BL base with NH4+ as its conjugate acid.

Some species can act either as a BL acid *or* a BL base. Such beasts are called ***amphoteric***.

Eg. hydrogen carbonate ion, HCO3-. When dissolved in water, two possible reactions can occur:

\begin{displaymath}
{\rm HCO}_3^-(aq) + {\rm H}_2{\rm O}(l)\rightleftharpoons {\rm H}_3{\rm O}^+(aq) + {\rm CO}_3^{2-}(aq)
\end{displaymath}

or

\begin{displaymath}
{\rm HCO}_3^-(aq) + {\rm H}_2{\rm O}(l)\rightleftharpoons {\rm H}_2{\rm CO}_3(aq) + {\rm OH}^-(aq)
\end{displaymath}

In the first of these, HCO3- acts as a BL acid with CO32- as its conjugate base, while in the second it acts as a BL base with H2CO3 as its conjugate acid. The treatment of amphoteric reactions is mathematically a little more **hideous** than you might think… We don’t need to go into it just yet… Few!!