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-Lowery Definition**

**Edit section**

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:

<table>
<thead>
<tr>
<th>Acid</th>
<th>Base</th>
</tr>
</thead>
<tbody>
<tr>
<td>Donates hydrogen ions</td>
<td>Accepts hydrogen ions.</td>
</tr>
<tr>
<td>HCl(^+)</td>
<td>HOH →</td>
</tr>
<tr>
<td>HOH(^+)</td>
<td>NH(_3)→</td>
</tr>
<tr>
<td></td>
<td>H(_3)O(^+) + Cl(^-)</td>
</tr>
<tr>
<td></td>
<td>NH(_4)(^+) + OH(^-)</td>
</tr>
</tbody>
</table>

The determination of a substance as a Brønsted-Lowery 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 an Acid?
2. Why is a Base?
3. How can A⁻ be a base when HA was and Acid?
4. How can HZ⁺ be an acid when Z used to be a Base?
5. **Now that we understand the concept, let’s look at an 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 it's conjugate acid H₂O
• Cl⁻ is the conjugate base because it accepts an H from H₃O to return to it's 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 element 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

We discussed the concept of conjugate acid–base pairs in Chapter 4, using the reaction of ammonia, the base, with water, the acid, as an example. 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. For example, in the reaction of HCl with water (Equation 16.1), HCl, the parent acid, donates a proton to a water molecule, the parent base, thereby forming H⁺Cl⁻. Thus H⁺ 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⁺, forming HCl and H₂O. Thus H⁺ and Cl⁻ 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 H⁺Cl⁻ and HCl⁻(aq).

Note
All acid–base reactions contain two conjugate acid–base pairs.

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, \( \text{CH}_3\text{CO}_2\text{H}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{CH}_3\text{CO}_2\text{O}^-\text{(aq)} + \text{H}_3\text{O}^+(\text{aq}) \), acetic acid 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 (acetic acid and acetate ion) and the parent base and its conjugate acid (water and hydrogen ion).

In the reaction of ammonia with water to give ammonium ions and hydroxide ions (Equation 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 \( \text{NH}_3\text{(aq)} \) and \( \text{NH}_4^+(\text{aq}) \). Some common conjugate acid–base pairs are shown in Figure 16.2.

**Figure 16.2 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.

Contributors

Edit section

We’ll begin our discussion of acid-base chemistry with a couple of essential definitions. The first of these definitions was proposed in 1923 by the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry, and has come to be known as the Brønsted-Lowry definition of acids and bases. An acid, by the Brønsted-Lowry definition, is a species which is able to donate a proton (\(H^+\)), while a base is a proton acceptor. We have already discussed in the previous chapter one of the most familiar examples of a Brønsted-Lowry acid-base reaction, between hydrochloric acid and hydroxide ion:

In this reaction, a proton is transferred from HCl (the acid, or proton donor) to hydroxide (the base, or proton acceptor). As we learned in the previous chapter, curved arrows depict the movement of electrons in this bond-breaking and bond-forming process.

After a Brønsted-Lowry acid donates a proton, what remains – in this case, a chloride ion – is called the conjugate base. Chloride is thus the conjugate base of hydrochloric acid. Conversely, when a Brønsted-Lowry base accepts a proton it is converted into its conjugate acid form: water is thus the conjugate acid of hydroxide.

We can also talk about conjugate acid/base pairs: the two acid/base pairs involved in our first reaction are hydrochloric acid/chloride and hydroxide/water. In this next acid-base reaction, the two pairs involved are acetate/acetic acid and methyl ammonium/methylamine:

Throughout this text, we will often use the abbreviations HA and :B in order to refer in a general way to acidic and basic reactants:

*In order to act as a proton acceptor, a base must have a reactive pair of electrons.* In all of the examples we shall see in this chapter, this pair of electrons is a non-bonding lone pair, usually (but not always) on an oxygen, nitrogen, sulfur, or halogen atom. When acetate acts as a base in the reaction shown above, for example, one of its oxygen lone pairs is used to form a new bond to a proton. The same can be said for an amine acting as a base. Clearly, methyl ammonium ion cannot act as a base – it does not have a reactive pair of electrons with which to accept a new bond to a proton.
Later, in chapter 15, we will see several examples where the (relatively) reactive pair of electrons in a pi bond act in a basic fashion.

In this chapter, we will concentrate on those bases with non-bonding (lone pair) electrons.

**Example**

**Exercise 7.1**: Draw structures for the missing conjugate acids or conjugate bases in the reactions below.

**Solution**

**Contributors**
You are no doubt aware that some acids are stronger than others. Sulfuric acid is strong enough to be used as a drain cleaner, as it will rapidly dissolve clogs of hair and other organic material.

\[
\begin{array}{c}
\text{O} \\
\text{HO–S–OH} \\
\text{O}
\end{array}
\]

sulfuric acid

Not surprisingly, concentrated sulfuric acid will also cause painful burns if it touches your skin, and permanent damage if it gets in your eyes (there’s a good reason for those safety goggles you wear in chemistry lab!). Acetic acid (vinegar), will also burn your skin and eyes, but is not nearly strong enough to make an effective drain cleaner. Water, which we know can act as a proton donor, is obviously not a very strong acid. Even hydroxide ion could \textit{theoretically} act as an acid – it has, after all, a proton to donate – but this is not a reaction that we would normally consider to be relevant in anything but the most extreme conditions.

The relative acidity of different compounds or functional groups – in other words, their relative capacity to donate a proton to a common base under identical conditions – is quantified by a number called the \textit{dissociation constant}, abbreviated \(K_a\). The common base chosen for comparison is water.

We will consider acetic acid as our first example. When a small amount of acetic acid is added to water, a proton-transfer event (acid-base reaction) occurs to some extent.

Notice the phrase ‘to some extent’ – this reaction does \textit{not} run to completion, with all of the acetic acid converted to acetate, its conjugate base. Rather, a \textit{dynamic equilibrium} is reached, with proton transfer going in both directions (thus the two-way arrows) and finite concentrations of all four species in play. The nature of this equilibrium situation, as you recall from General Chemistry, is expressed by an equilibrium constant, \(K_{eq}\). The equilibrium constant is actually a ratio of activities (represented by the symbol \(a(a)\)), but activities are rarely used in courses other than analytical or physical chemistry. To simplify the discussion for general chemistry and organic chemistry courses, the activities of all of the solutes are replaced with molarities, and the activity of the solvent (usually water) is defined as having the value of 1.

In our example, we added a small amount of acetic acid to a large amount of water: water is the \textit{solvent} for this reaction. Therefore, in the course of the reaction, the concentration of water changes very little, and the water can be treated as a pure solvent, which is always assigned an activity of 1. The acetic acid, acetate ion and hydronium ion are all \textit{solute}s, and so their activities are approximated with molarities. The acid dissociation constant, or \(K_a\), for acetic acid is therefore defined as:

\[
K_{eq} = \frac{a_{\text{CH}_3\text{COO}^-} \cdot a_{\text{H}_3\text{O}^+}}{a_{\text{CH}_3\text{COOH}} \cdot a_{\text{H}_2\text{O}}} \approx \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}][1]}
\]
Because dividing by 1 does not change the value of the constant, the "1" is usually not written, and $K_a$ is written as:

$$K_{eq} = K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} = 1.75 \times 10^{-5}$$

In more general terms, the dissociation constant for a given acid is expressed as:

$$K_a = \frac{[A^-][\text{H}_3\text{O}^+]}{[\text{HA}]} \quad \text{(First)}$$

or

$$K_a = \frac{[A][\text{H}_3\text{O}^+]}{[\text{HA}^+]} \quad \text{(Second)}$$

The first expression applies to a neutral acid such as like HCl or acetic acid, while the second applies to a cationic acid like ammonium ($\text{NH}_4^+$).

The value of $K_a = 1.75 \times 10^{-5}$ for acetic acid is very small - this means that very little dissociation actually takes place, and there is much more acetic acid in solution at equilibrium than there is acetate ion. Acetic acid is a relatively weak acid, at least when compared to sulfuric acid ($K_a = 10^9$) or hydrochloric acid ($K_a = 10^7$), both of which undergo essentially complete dissociation in water.

A number like $1.75 \times 10^{-5}$ is not very easy either to say or to remember. Chemists have therefore come up with a more convenient term to express relative acidity: the $pK_a$ value.

$$pK_a = -\log K_a$$

Doing the math, we find that the $pK_a$ of acetic acid is 4.8. The use of $pK_a$ values allows us to express the acidity of common compounds and functional groups on a numerical scale of about –10 (very strong acid) to 50 (not acidic at all). Table 7 at the end of the text lists exact or approximate $pK_a$ values for different types of protons that you are likely to encounter in your study of organic and biological chemistry. Looking at Table 7, you see that the $pK_a$ of carboxylic acids are in the 4-5 range, the $pK_a$ of sulfuric acid is –10, and the $pK_a$ of water is 14. Alkenes and alkanes, which are not acidic at all, have $pK_a$ values above 30. The lower the $pK_a$ value, the stronger the acid.

It is important to realize that $pK_a$ is not at all the same thing as pH: the former is an inherent property of a compound or functional group, while the latter is the measure of the hydronium ion concentration in a particular aqueous solution:

$$pH = -\log [\text{H}_3\text{O}^+]$$

Any particular acid will always have the same $pK_a$ (assuming that we are talking about an aqueous solution at room temperature) but different aqueous solutions of the acid could have different pH values, depending on how much acid is added to how much water.
Our table of pK\textsubscript{a} values will also allow us to compare the strengths of different bases by comparing the pK\textsubscript{a} values of their conjugate acids. The key idea to remember is this: \textit{the stronger the conjugate acid, the weaker the conjugate base.} Sulfuric acid is the strongest acid on our list with a pK\textsubscript{a} value of −10, so HSO\textsubscript{4}\textsuperscript{−} is the weakest conjugate base. You can see that hydroxide ion is a stronger base than ammonia (NH\textsubscript{3}), because ammonium (NH\textsubscript{4}\textsuperscript{+}, pK\textsubscript{a} = 9.2) is a stronger acid than water (pK\textsubscript{a} = 14.0).

While Table 7 provides the pK\textsubscript{a} values of only a limited number of compounds, it can be very useful as a starting point for estimating the acidity or basicity of just about any organic molecule. Here is where your familiarity with organic functional groups will come in very handy. What, for example, is the pK\textsubscript{a} of cyclohexanol? It is not on the table, but as it is an alcohol it is probably somewhere near that of ethanol (pK\textsubscript{a} = 16). Likewise, we can use Table 7 to predict that para-hydroxyphenyl acetaldehyde, an intermediate compound in the biosynthesis of morphine, has a pK\textsubscript{a} in the neighborhood of 10, close to that of our reference compound, phenol.

\[
\begin{align*}
\text{HO} & \quad \text{O} \\
& \quad \text{C} \\
& \quad \text{H}
\end{align*}
\]

\textit{p=hydroxyphenyl acetaldehyde}

Notice in this example that we need to evaluate the potential acidity at \textit{four} different locations on the molecule.

Aldehyde and aromatic protons are not at all acidic (pK\textsubscript{a} values are above 40 – not on our table). The two protons on the carbon next to the carbonyl are slightly acidic, with pK\textsubscript{a} values around 19-20 according to the table. The most acidic proton is on the phenol group, so if the compound were to be subjected to a single molar equivalent of strong base, this is the proton that would be donated.
As you continue your study of organic chemistry, it will be a very good idea to commit to memory the approximate pKₐ ranges of some important functional groups, including water, alcohols, phenols, ammonium, thiols, phosphates, carboxylic acids and carbons next to carbonyl groups (so-called α-carnons). These are the groups that you are most likely to see acting as acids or bases in biological organic reactions.

A word of caution: when using the pKₐ table, be absolutely sure that you are considering the correct conjugate acid/base pair. If you are asked to say something about the basicity of ammonia (NH₃) compared to that of ethoxide ion (CH₃CH₂O⁻), for example, the relevant pKₐ values to consider are 9.2 (the pKₐ of ammonium ion) and 16 (the pKₐ of ethanol). From these numbers, you know that ethoxide is the stronger base. Do not make the mistake of using the pKₐ value of 38: this is the pKₐ of ammonia acting as an acid, and tells you how basic the NH₂⁻ ion is (very basic!)

<table>
<thead>
<tr>
<th>Compound</th>
<th>pKₐ Values</th>
</tr>
</thead>
<tbody>
<tr>
<td>Phosphate monoester</td>
<td>pKₐ ~ 1.65</td>
</tr>
<tr>
<td>Phosphate diester</td>
<td>pKₐ ~ 1.5</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>pKₐ = 2.2, 7.2, 12.3</td>
</tr>
<tr>
<td>Aniline</td>
<td>pKₐ = 4.6</td>
</tr>
<tr>
<td>Carboxylic acid</td>
<td>pKₐ ~ 4.5</td>
</tr>
<tr>
<td>Pyridinium</td>
<td>pKₐ = 5.3</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>pKₐ = 6.4, 10.3</td>
</tr>
<tr>
<td>Hydrogen cyanide</td>
<td>pKₐ = 9.2</td>
</tr>
<tr>
<td>Ammonium</td>
<td>pKₐ = 9.2</td>
</tr>
<tr>
<td>Phenol</td>
<td>pKₐ = 9.9</td>
</tr>
</tbody>
</table>

- RO-P-OH
  - phosphate monoester (pKₐ ~ 1.65)
- RO-P-OH
  - phosphate diester (pKₐ ~ 1.5)
- HO-P-OH
  - phosphoric acid (pKₐ = 2.2, 7.2, 12.3)
- R-NH₃⁺
  - aniline (pKₐ = 4.6)
- O-C=O
  - carboxylic acid (pKₐ ~ 4.5)
- N⁺H
  - pyridinium (pKₐ = 5.3)
- O-C=O
  - carbonic acid (pKₐ = 6.4, 10.3)
- N≡C-H
  - hydrogen cyanide (pKₐ = 9.2)
- NH₄⁺
  - ammonium (pKₐ = 9.2)
- O
  - phenol (pKₐ = 9.9)
Example

Exercise 7.2: Using the pK<sub>a</sub> table, estimate pK<sub>a</sub> values for the most acidic group on the compounds below, and draw the structure of the conjugate base that results when this group donates a proton.

Solution