

1 Learning Objectives

1. What are the physical and chemical properties of acids and bases?
2. How do we define what's an acid and what's a base?
3. How do we measure the strength of an acid?

2 Properties of Acids and Bases

There are numerous examples of acid-base chemistry in everyday life. Below you'll find a small table of applications and examples of where you might observe acid-base chemistry, along with links to some useful information.

Application	Examples
Household Cleaning Products	Bleach, Vinegar, Soap
Cooking and Food Preparation	Baking Soda, Lemon Juice
Medicine	Antacids, Antibiotics, Diagnostic Testing
Soil and Agriculture	Lime, Acid Rain, Fertilizer

The presence of acids and bases (also called alkalis) has been known even before chemistry was established as a field of science. During the time of the ancient Greeks, properties of acids and bases were vaguely understood, however, they realized that there were some fundamental differences between these substances.

Property	Acid	Base
Taste		
Touch		
pH		
Reaction with Metal		
Litmus Test		

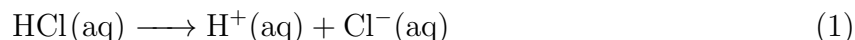
Acids and bases also share some unique properties such as: colorless, electrical conductivity, and reactive with indicators (changing their color). It wasn't until 1884 when Swedish chemist, Svante Arrhenius, provided a groundbreaking explanation for the categorization and behavior of acids and bases in aqueous solutions.

3 Acid-Base Theories

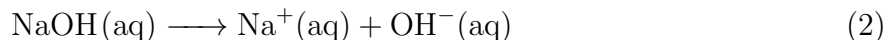
There are a handful of ways one can define if a compound acts like an acid or base. This section will cover the three main acid-base theories that are commonly discussed in chemistry (but acknowledge that there are others).

3.1 Arrhenius Theory

The Arrhenius definition is relatively straight forward but only applies to aqueous solutions. **Arrhenius acids** are substances that yield H^+ ions when dissolved in water, and **Arrhenius bases** yield OH^- ions. Lets look at a couple of examples.



HCl(aq) is considered an Arrhenius acid because when dissolved into water, HCl dissociates into $\text{H}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ ions.



NaOH(aq) is considered an Arrhenius base because when dissolved into water, NaOH dissociates into $\text{Na}^+(\text{aq})$ and $\text{OH}^-(\text{aq})$ ions.

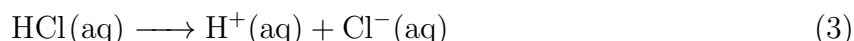
3.1.1 Exercise

Circle if the compound on the reactant side in the table is an Arrhenius Acid or Base.

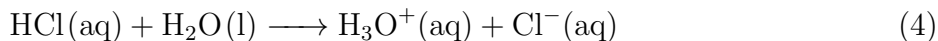
Compound	Arrhenius Acid/Base
$\text{KOH}(\text{aq}) \longrightarrow \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$	Acid / Base
$\text{HNO}_3(\text{aq}) \longrightarrow \text{H}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$	Acid / Base
$\text{Cu}(\text{OH})_2(\text{aq}) \longrightarrow \text{Cu}^{2+}(\text{aq}) + 2 \text{OH}^-(\text{aq})$	Acid / Base

3.2 Brønsted Theory

The Arrhenius definition helps identify many types of compounds which are acids and bases, however is limited because they only apply to aqueous solutions. A broader definition was proposed by Johannes Brønsted in 1932 to classify a larger group of compounds as acids or bases. A Brønsted acid is a proton donor and a Brønsted base is a proton acceptor. This definition does not require acids and bases to be in aqueous solution, just that acids donate a proton and bases accept one. Lets take our HCl example above:

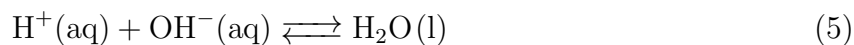


When we write the state of a compound, we typically write it as a solid (s), liquid (l), or gas (g). Here, our acids and bases are dissolved in water. To denote that, we say that it is in an aqueous (aq) state. In other words, in addition to having HCl as a reactant, we have water. So another way to write the reaction above is:

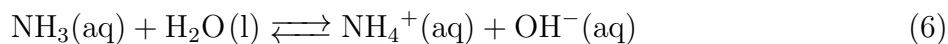


This reaction is more representative of what is happening in the reaction. H^+ is a very reactive atom, and will want to bond with another atom. Given the presence of water, the proton will stick to the negatively charge O atom of water forming a hydronium ion (H_3O^+). This equation shows the reaction between the Brønsted acid (HCl) and Brønsted base (H_2O) forming hydronium and chloride (Cl^-) ions.

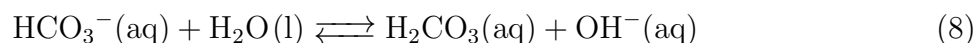
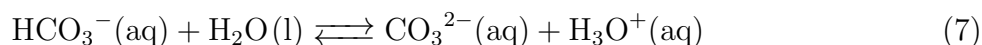
The hydroxide ion is classified as a Brønsted base because it can accept a base as follows:



Any Arrhenius base can be classified as a Brønsted base. However, not all Brønsted bases can be classified as an Arrhenius base. For example, ammonia (NH_3) is a Brønsted base because it can accept a proton from an acid or water.



Within Brønsted acid-base theory, some compounds can act as both an acid and a base. The compounds are called amphoteric. For example, bicarbonate ion can both donate and accept a proton.



The top reaction corresponds to bicarbonate acting as a Brønsted acid, where the bottom shows it as a Brønsted base.

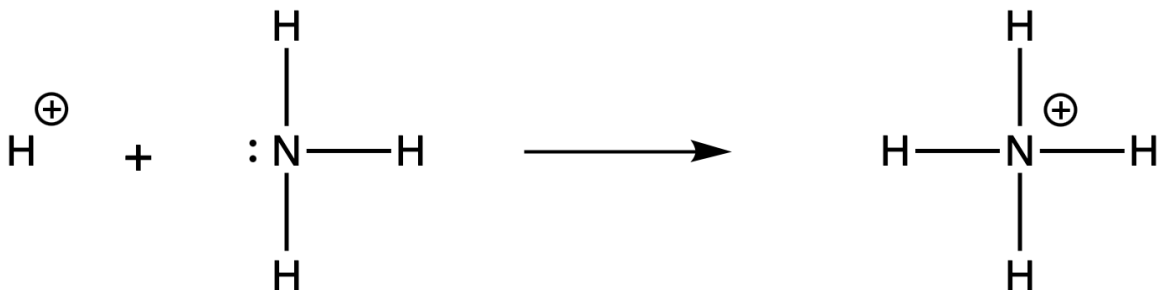
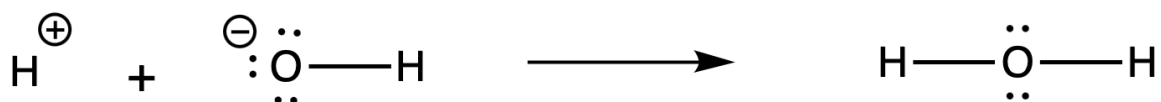
3.2.1 Exercise

Classify each of the following species as Brønsted acid or base if they react with water.

Compound	Brønsted Acid/Base/Both
HBr	Acid / Base / Both
NO_3^-	Acid / Base / Both
HCN	Acid / Base / Both
H_2O	Acid / Base / Both

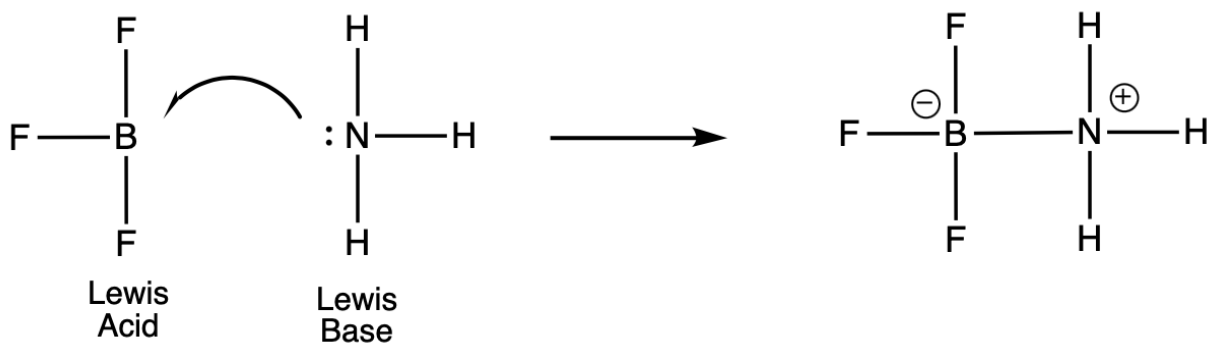
3.3 Lewis Theory

According to the Brønsted theory, a base must be able to accept protons. By this definition, both the hydroxide ion and ammonia are Brønsted bases:



In both cases, the atom which the proton becomes attached to possesses at least one unshared pair of electrons. The characteristic property of OH^- , NH_3 , and other Brønsted bases suggests a more general definition of acids and bases. In 1932, an American chemist named G. N. Lewis formulated a new definition of acids and bases. He defined a **Lewis base** as a substance that can *donate a pair of electrons* and a **Lewis acid** as a substance that can *accept a pair of electrons*. For example the protonation of NH_3 acts as a Lewis base because it donates a pair of electrons to the H^+ , which acts as a Lewis acid.

The Lewis definition of acids and bases is more general than the Brønsted and Arrhenius definitions. Consider the reaction between BF_3 and NH_3 :



The vacant p_z orbital accepts an electron pair from the NH_3 . BF_3 functions as an acid according to the Lewis theory, even though it does not contain an ionizable proton.

3.3.1 Exercise

Write the reaction of two water molecules forming a hydronium and hydroxide ion. Highlight which water is acting as the Lewis acid and which water is the Lewis base. Draw the Lewis dot diagrams and show the movement of the lone pair electrons from the Lewis acid to the Lewis base.

4 Acid-Base Properties of Aqueous Solutions are Governed by the Autoionization of Water

In the previous section, we showed that water can act as an acid or a base depending on what it's reacting with. Water is a weak electrolyte but does undergo ionization to a small extent (as demonstrated in the previous example). This reaction is sometimes known as **autoionization of water**.

The reaction as written in the exercise above indicates that there is an equilibrium between liquid water (reactants) and the hydronium and hydroxide ions (products). Therefore, we can express the equilibrium constant for the autoionization of water:

$$K_w = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]^2} \quad (9)$$

The concentration of a pure substance can be taken to be 1. So the equation above can be simplified to:

$$K_w = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{1M^2} \quad (10)$$

or

$$K_w = \frac{[\text{H}^+][\text{OH}^-]}{1M^2} \quad (11)$$

K_w is known as the **ion-product constant** which is the product of the molar concentrations of H^+ and OH^- ions at a particular temperature. In pure water at 25°C , the concentration of H^+ and OH^- are equal and found to be $[\text{H}^+] = [\text{OH}^-] = 1 \times 10^{-7}M$. This means that K_w is:

$$K_w = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]} = \frac{[1 \times 10^{-7}M][1 \times 10^{-7}M]}{1M^2} = 1 \times 10^{-14} \quad (12)$$

The value of K_w is fixed at 25°C regardless if we have a pure solution of water or an aqueous solution of dissolved species. Now given an arbitrary solution, we can now determine how acidic or basic it is.

4.1 Measuring the Acidity or Alkalinity of a Solution

Whenever the $[\text{H}^+] = [\text{OH}^-]$, the aqueous solution is said to be neutral. In acidic solutions, there is an excess of H^+ (or H_3O^+) ions and $[\text{H}^+] > [\text{OH}^-]$. In basic solutions, there is an excess of OH^- ions and $[\text{H}^+] < [\text{OH}^-]$. In practice, we can change the $[\text{H}^+]$ and $[\text{OH}^-]$, but we cannot vary them independently. If we adjust the solutions $[\text{H}^+]$, the $[\text{OH}^-]$ *must* change to.

4.1.1 Exercise

If we add an acid to a beaker of pure water such that the concentration of $[\text{H}^+]$ increases to $1 \times 10^{-6}M$, what is the concentration of $[\text{OH}^-]$?

4.2 pH - a Measure of Acidity

We typically deal with very small concentrations of $[\text{H}^+]$ and $[\text{OH}^-]$ ions. A more practical measure of acidity (and alkalinity) is called **pH**. The **pH** of a solution is defined as the *negative logarithm (base 10) of the hydrogen ion activity*:

$$\text{pH} = -\log([\text{H}^+]) \quad (13)$$

or

$$\text{pH} = -\log([\text{H}_3\text{O}^+]) \quad (14)$$

$[\text{H}^+]$ and $[\text{H}_3\text{O}^+]$ is understood to be the numerical value of the hydrogen or hydronium concentration in mol L^{-1} . When calculating the pH of a solution, we traditionally drop the units. The equation above is simply designed to give a convenient number to work with. The negative logarithm gives us a positive number for pH. Like the equilibrium constant, the pH of a solution is a dimensionless quantity.

Because pH is simply a way to express hydrogen ion concentration, acidic and basic solution can be identified based on their pH.

	$[\text{H}^+]$	pH
Acidic		
Basic		
Neutral		

4.2.1 Exercise

The pH of a swimming pool is measured to be 7.6 at a temperature of 25°C . Calculate the concentration of hydroxide ions $[\text{OH}^-]$ in the pool water.