Acid/Base Basics

How does one define acids and bases? In chemistry, acids and bases have been defined differently by three sets of theories. One is the Arrhenius definition, which revolves around the idea that acids are substances that ionize (break off) in an aqueous solution to produce hydrogen (H^+) ions while bases produce hydroxide (OH^-) ions in solution. On the other hand, the Bronsted-Lowry definition defines acids as substances that donate protons (H^+) whereas bases are substances that accept protons. Also, the Lewis theory of acids and bases states that acids are electron pair acceptors while bases are electron pair donors. Acids and bases can be defined by their physical and chemical observations.

Introduction

Acids and bases are common solutions that exist everywhere. Almost every liquid that we encounter in our daily lives consists of acidic and basic properties, with the exception of water. They have completely different properties and are able to neutralize to form H_2O, which will be discussed later in a subsection. The table below compares the different properties between them:

Table 1.

ACIDS	BASES
produce a piercing pain in a wound.	give a slippery feel.
taste sour.	taste bitter.
are colorless when placed in phenolphthalein (an indicator).	are pink when placed in phenolphthalein (an indicator).
are red on blue litmus paper (a pH indicator).	are blue on red litmus paper (a pH indicator).
have a pH<7.	have a pH>7.
produce hydrogen gas when reacted with metals.	
produce carbon dioxide when reacted with carbonates.	
Common examples: Lemons, oranges, vinegar, urine, sulfuric acid, hyrdrochoric acid	Common Examples: Soap, toothpaste, bleach, cleaning agents, limewater, ammonia water, sodium hydroxide.

The Arrhenius Definition

In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution.

Arrhenius Acids

An Arrhenius acid is a compound that increases the concentration of **H^+ ions** that are present when added to water. These H^+ ions form the hydronium ion (H_3O^+) when they combine with water molecules. This process is represented in a chemical equation by adding H_2O to the reactants side.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

In this reaction, hydrochloric acid (HCl) dissociates into hydrogen (H^+) and chlorine (Cl^-) ions when dissolved in water, thereby releasing H^+ ions into solution. Formation of the hydronium ion equation:

$$\boldsymbol{HCl(aq) + H_3O^+(l) \rightarrow H_3O^+(aq) + Cl^-(aq)}$$

Incomplete Ionization (Weak Acids)

Strong acids are molecular compounds that essentially ionize to completion in aqueous solution, disassociating into H^+ ions and the additional anion; there are very few common strong acids. All other acids are "weak acids" that incompletely ionized in aqueous solution.

Strong Acids	HCl, HNO_3, H_2SO_4, HBr, HI, $HClO_4$
Weak Acids	All other acids, such as HCN, HF, H_2S, HCOOH

Arrhenius Bases

An Arrhenius base is a compound that increases the concentration of OH^- ions that are present when added to water. The dissociation is represented by the following equation:

$$NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$$

In this reaction, sodium hydroxide (NaOH) disassociates into sodium (Na^+) and hydroxide (OH^-) ions when dissolved in water, thereby releasing OH^- ions into solution.

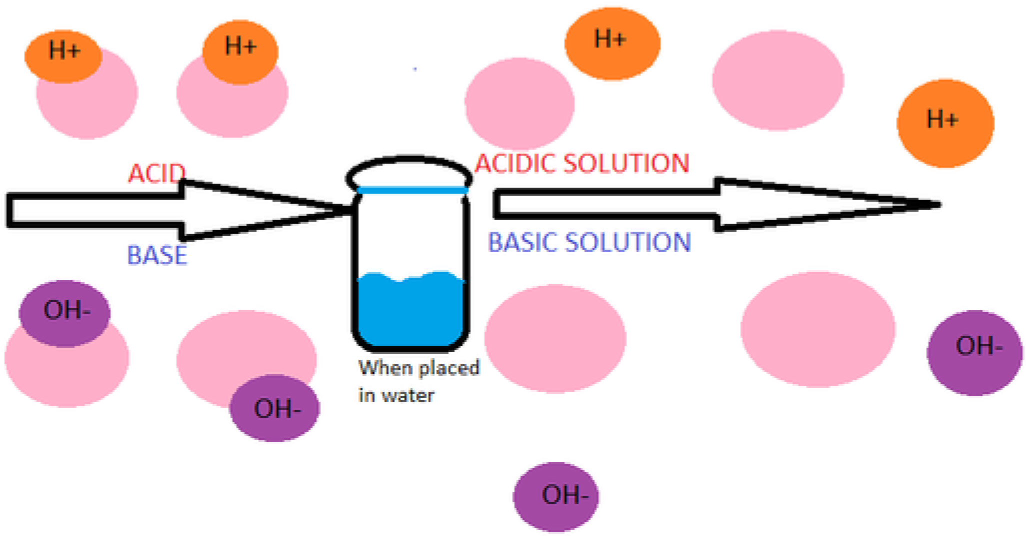


Figure 1. Arrhenius acids dissociate to form aqueous H^+ ions and Arrhenius bases dissociate to form aqueous OH^- ions.

NOTE: The stronger the acid and base, the more dissociation will occur.

Incomplete Ionization (Weak Bases)

Like acids, strong and weak bases are classified by the extent of their ionization. Strong bases disassociate almost or entirely to completion in aqueous solution. Similar to strong acids, there are very few common strong bases. Weak bases are molecular compounds where the ionization is not complete.

Table 2. The strong and weak acids and bases.

STRONG BASES	The hydroxides of the Group I andGroup II metals such as LiOH, NaOH, KOH, RbOH, CsOH
WEAK BASES	All other bases, such as NH_3, CH_3NH_2, C_5H_5N

Limitations to the Arrhenius Theory

The Arrhenius theory has many more limitations than the other two theories. The theory suggests that in order for a substance to release either H^+ or OH^- ions, it must contain that particular ion. However, this does not explain the weak base ammonia (NH_3), which in the presence of water, releases hydroxide ions into solution, but does not contain OH- itself.

Hydrochloric acid is neutralised by both sodium hydroxide solution and ammonia solution. In both cases, you get a colourless solution which you can crystallise to get a white salt - either sodium chloride or ammonium chloride. These are clearly very similar reactions. The full equations are:

$$NaOH(aq)+HCl(aq)\rightarrow NaCl(aq)+H_2O(l)$$

$$NH_3(aq)+HCl(aq)\rightarrow NH_4Cl(aq)$$

In the sodium hydroxide case, hydrogen ions from the acid are reacting with hydroxide ions from the sodium hydroxide - in line with the Arrhenius theory. However, in the ammonia case, there are no hydroxide ions!

You can get around this by saying that the ammonia reacts with the water it is dissolved in to produce ammonium ions and hydroxide ions:

$$\mathbf{NH_3(aq)+H_2O(l)\rightleftharpoons NH_4^+(aq)+OH^-(aq)}$$

This is a reversible reaction, and in a typical dilute ammonia solution, about 99% of the ammonia remains as ammonia molecules. Nevertheless, there are hydroxide ions there, and we can squeeze this into the Arrhenius theory. However, this same reaction also happens between ammonia gas and hydrogen chloride gas.

$$NH_3(g)+HCl(g)\rightarrow NH_4Cl(s)$$

In this case, there are not any hydrogen ions or hydroxide ions in solution - because there isn't any solution. The Arrhenius theory wouldn't count this as an acid-base reaction, despite the fact that it is producing the same product as when the two substances were in solution. Because of this short-coming, later theories sought to better explain the behavior of acids and bases in a new manner.

The Brønsted-Lowry Definition

In 1923, British chemists Johannes Nicolaus Brønsted and Thomas Martin Lowry independently developed definitions of acids and bases based on the compounds' abilties to either donate or accept protons (H^+ ions). In this theory, acids are defined as **proton donors**; whereas bases are defined as **proton acceptors**. A compound that acts as both a Bronsted-Lowry acid and base together is called**amphoteric**.This took the Arrhenius definition one step further, as a substance no longer needed to be composed of hydrogen (H^+) or hydroxide (OH^-) ions in order to be classified as an acid or base.

Consider the following chemical equation:

$$HCl(aq)+NH_3(aq)\rightarrow NH_4^+(aq)+Cl^-(aq)$$

Here, hydrochloric acid (HCl) "donates" a proton (H^+) to ammonia (NH_3) which "accepts" it , forming a positively charged ammonium ion (NH_4^+) and a negatively charged chloride ion (Cl^-). Therefore, HCl is a Brønsted-Lowry acid (donates a proton) while the ammonia is a Bronsted-Lowry base (accepts a proton). Also, Cl^- is called the **conjugate base** of the acid HCl and NH_4^+ is called the **conjugate acid** of the base NH_3.

pH Scale

Since acids increase the amount of H^+ ions present and bases increase the amount of OH^- ions, under the pH scale, the strength of acidity and basicity can be measured by its concentration of H^+ ions. This scale is shown by the following formula:

$$pH = -\log[H^+]$$

with $[H^+]$ being the concentration of H^+ ions.

To see how these calculations are done, refer to Calculating the pH of the solution of a Polyprotic Base/Acid

The pH scale is often measured on a 1 to 14 range, but this is incorrect (see pH for more details). Something with a pH less than 7 indicates acidic properties and greater than 7 indicates basic properties. A pH at exactly 7 is neutral. The higher the $[H^+]$, the lower the pH.

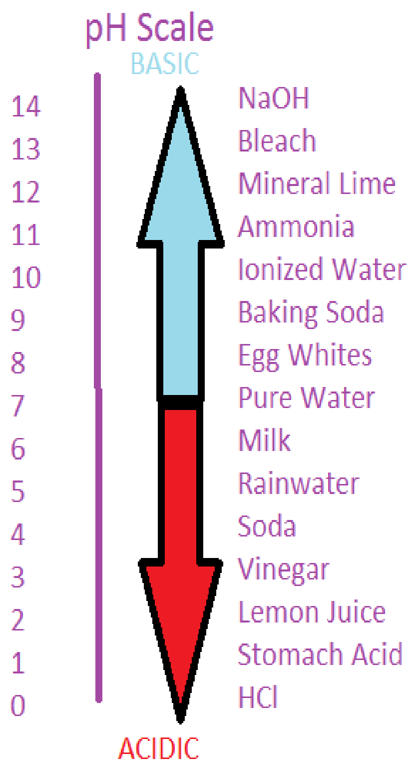


Figure 4. The pH scale shows that substances with a pH greater than 7 are basic and a pH less than 7 are acidic.

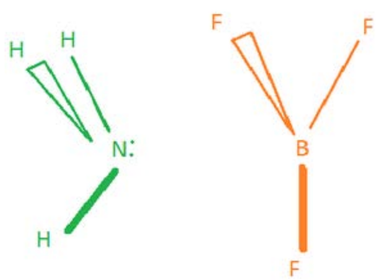
Lewis Theory

The Lewis theory of acids and bases states that acids act as **electron pair acceptors** and bases act as **electron pair doners**. This definition doesn't mention anything about the hydrogen atom at all, unlike the other definitions. It only talks about the transfer of electron pairs. To demonstrate this theory, consider the following example.

$$NH_{3(g)} + BF_{3(g)} \rightleftharpoons H_3NBF_{3(g)}$$

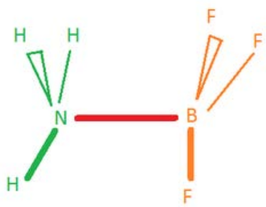
This is a reaction between ammonia (NH_3) and boron trifluoride (BF_3). Since there is no transfer of hydrogen atoms here, it is clear that this is a Lewis acid-base reaction. In this reaction, NH_3 has a lone pair of electrons and BF_3 has an incomplete octet, since boron doesn't have enough electrons around it to form an octet.

Figure 2. The Lewis structures of ammonia and boron trifluoride.



Because boron only has 6 electrons around it, it can hold 2 more. BF_3 can act as an acid and accept the pair of electrons from the nitrogen in NH_3, which will then form a bond between the nitrogen and the boron.

Figure 3. The Lewis structure of H_3NBF_3, which resulted from the bond between nitrogen and boron.



This is considered an acid-base reaction where NH_3 (base) is donating the pair of electrons to BF_3. BF_3 (acid) is accepting those electrons to form a new compound, H_3NBF_3.

Neutralization

A special property of acids and bases is their ability to neutralize the other's properties. In an acid-base (or neutralization) reaction, the H^+ ions from the acid and the OH^- ions from the base react to create water (H_2O). Another product of a neutralization reaction is an ionic compound called a salt. Therefore, the general form of an acid-base reaction is:

$$Acid_{(aq)} + Base_{(aq)} \rightarrow \ H_2O_{(l)} + Salt_{(s)}$$

The following are examples of neutralization reactions:

1. $HCl_{(aq)} + NaOH_{(aq)} \rightarrow \ H_2O_{(l)} + NaCl_{(s)}$

(NOTE: To see this reaction done experimentally, refer to the YouTube video link under the section "References".)

2. $H_2SO_{4(aq)} + 2NH_4OH_{(aq)} \rightarrow \ H_2O_{(l)} + (NH_4)_2SO_{4(s)}$

Titrations

Titrations are performed with acids and bases to determine their concentrations. At the equivalence point, the number of moles of the acid will equal the number of moles of the base. This indicates that the reaction has been neutralized.

Neutralization: moles of acid = moles of base

Here's how the calculations are done:

For instance, hydrochloric acid is titrated with sodium hydroxide:

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow \ H_2O_{(l)} + NaCl_{(s)}$$

For instance, 30 mL of 1.00 M NaOH is needed to titrate 60 mL of an HCl solution. The concentration of HCl needs to be determined. At the eqivalence point:

moles of HCl = moles of NaOH

$$(Molarity_{acid})(Volume_{acid}) = (Molarity_{base})(Volume_{base})$$

$$(Molarity_{HCl})(Volume_{HCl}) = (Molarity_{NaOH})(Volume_{NaOH})$$

To solve for the molarity of HCl, plug in the given data into the equation above.

M_{HCl}(60 mL HCl) = (1.00 M NaOH)(30 mL NaOH)

M_{HCl}=0.5 M

The concentration of HCl is **0.5 M**.