Arrhenius and Brønsted-Lowry Acid/Bases (SC7b)

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:

$$HCl(aq) + H_2O_+(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

Incomplete Ionization/Disassociation (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.

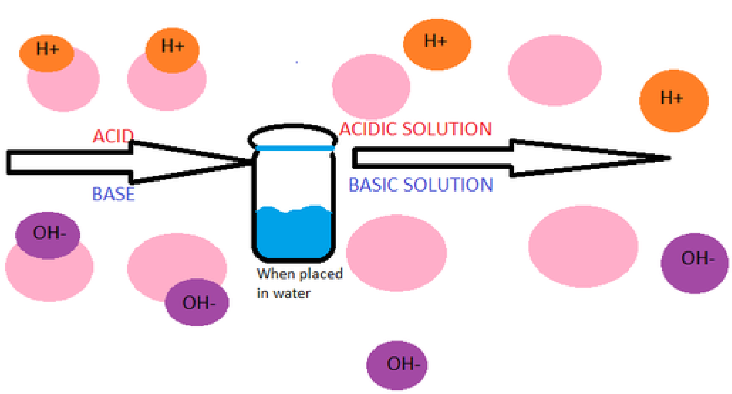

Figure 1. Arrhenius acids dissociate to form aqueous H^+ ions and Arrhenius bases dissociate to form aqueous OH^- ions. The stronger the acid and base, the more dissociation will occur.

Incomplete Ionization/Disassociation (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.

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:

$$HCl(aq)+NH3(aq)\rightarrow NH+4(aq)+Cl-(aq)$$

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:

$$\mathbf{pH = -log[H^+]}$$

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

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.

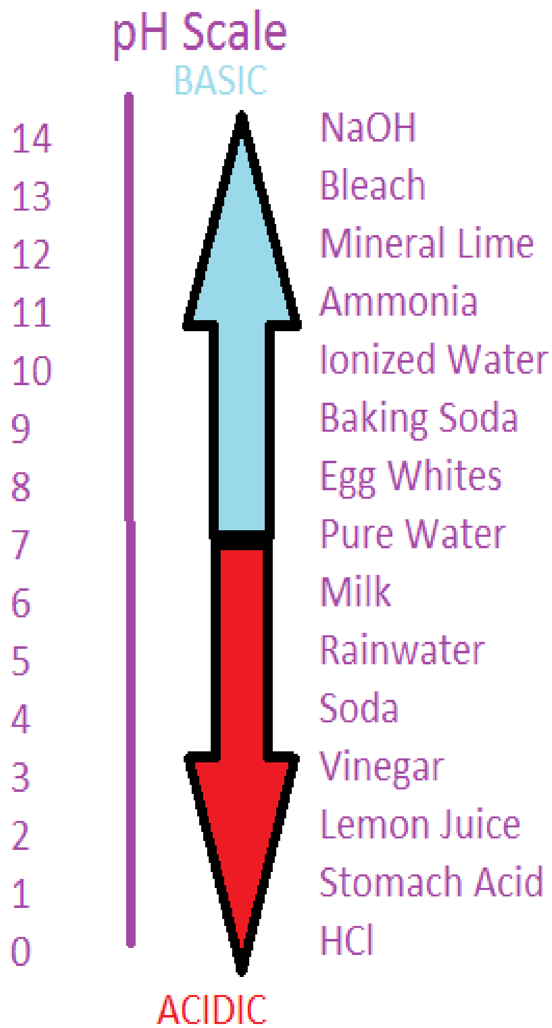

Figure 4. The pH scale shows that substances with a pH greater than 7 are basic and a pH less than 7 are acidic.

Definition of Brønsted-Lowry acid and base:

- An acid is a proton donor.
- A base is a proton acceptor.

Contrast this to the Arrhenius definitions:

- An acid dissolves in water to form hydronium ions.
- A base dissolves in water to form hydroxide ions.