# Acid-Base Review

Acid-Base Definitions

Initially acids and bases were **operationally defined** (based on observable qualitative properties). For example, **acids** (e.g. lemon juice, vinegar) react with metals to produce $H_{2(g)}$, turn litmus red, taste sour and etcetera while **bases** (e.g. antacids, soap) turn litmus blue, taste bitter, feels slippery and etcetera.

Later Swedish chemist, **Arrhenius**, defined acids and bases based on the ions they produce in water. This **conceptual definition** describes acids and bases on the particle level. An **acid** is any substance which produces $H^+_{(aq)}$ when dissolved in water (e.g. $HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$) while a **base** is any substance which provides $OH^-_{(aq)}$ when dissolved in water (e.g. $NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$). This allowed quantitative measurement of acidity/basicity of a solution. However, this definition misrepresents the fact that the small size and charge of the H^+ ion makes it is so strongly attracted to the partial negative oxygen in the polar water molecules that it is immediately bonds with water resulting in the formation of an **hydronium ion** $H_3O^+_{(aq)}$ (we often still write $H^+_{(aq)}$ to mean $H_3O^+_{(aq)}$).
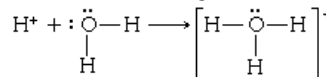
Acids can be identified by their **proticity**. It the acid releases 1 $H^+_{(aq)}$ or proton per molecule of acid, it is a **monoprotic** acid (e.g. HCl). **Polyprotic** acids release more than one $H^+_{(aq)}$ more molecule of acid. For example, **diprotic** acid releases 2 $H^+_{(aq)}$ (e.g. H_2SO_4) and **triprotic** acids release 3 $H^+_{(aq)}$. Like wise, bases can be defined as monobasic (e.g. NaOH), dibasic (e.g. $Ca(OH)_2$) and etc. depending upon the number of $OH^-_{(aq)}$ they release per compound of base.

Acid-Base Strength

Any solution which conducts electricity is called an **electrolyte** and the solution formed has free ions. A **strong electrolyte** releases a lot of ions when dissolved in water. If the substance is ionic we call this **dissociation** as ions already existed and just needed to be separated (e.g. $NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$).

NaCl(s)   **Na⁺(aq)**   **Cl⁻(aq)**
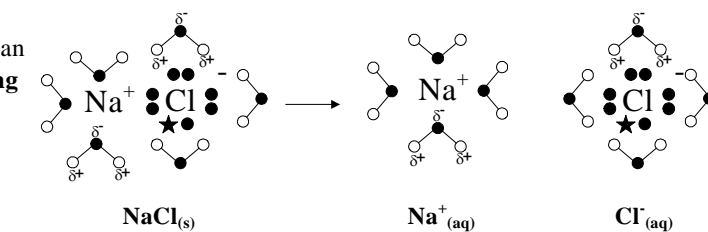
If the substance a covalent molecule, the process is called **ionization** as ions are created when the hydration of water overcomes the polar covalent bond holding the molecule together to pull the atoms apart (due to dipole-dipole interaction between water and the polar molecule (e.g. $HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$).

HF(l)   **H⁺(aq)**   **F⁻(aq)**
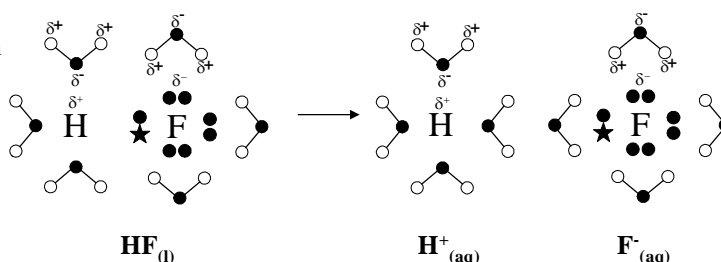
A **strong acid** is one that completely ionizes in water to produce hydronium ions and is thus a good proton donor (e.g. $HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$). A **strong base** is one that completely dissociates in water to produce hydroxide ions and readily accepts protons (e.g. $NaOH_{(s)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$).

A **weak electrolyte** dissolves but only forms a small amount of ions. It has little to do with solubility as something that is very soluble can be a strong or a weak electrolyte and something that is poorly soluble can be a weak or strong electrolyte. Thus a **weak acid** is simply one with incompletely ionizes in water (e.g. acetic acid). So, the actual strength of an acid or base can actually be calculated using the following formula (substitute the word, dissociation for ionization and base terms respectively as needed).

$$\% \text{ Ionization} = \frac{\text{concentration of } H^+}{\text{concentration of acid}} \, x \, 100$$

However, the total concentration of ions in the resulting solution is ultimately dependent upon both solubility and electrolyte strength.

= proton

= acid molecule

Weak Acid   **Strong Acid**

Prepared by K. Zuber