

## Chemistry Honors: Lesson 6

### Acids and Bases

#### Definitions

1. **Arrhenius** Definition – The first definition of acids and bases. Arrhenius defined an acid as a species that produces  $H^+$  (a proton) in an aqueous solution and a base as a species that produces  $OH^-$  (a hydroxide ion) in a aqueous solution.
2. **Bronsted-Lowry** Definition – A Bronsted-Lowry acid donates protons, while a Bronsted-Lowry base accepts protons. However, they can not be called Arrhenius bases since in aqueous solution they do not dissociate to form  $OH^-$ . The advantage of this definition is that it is not limited to aqueous solutions. Bronsted-Lowry acids and bases always occur in pairs called **conjugate acid-base pairs**. The pairs are related through the transfer of a proton.
3. **Lewis** Definition – Lewis defined an acid as an electron-pair acceptor and a base as a electron pair donor. This definition is the most inclusive and encompasses species not included in the Bronsted-Lowry definition.

#### Hydrogen Ion Equilibria

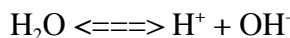
Hydrogen ion concentration,  $[H^+]$ , is a measure of both Arrhenius and Bronsted-Lowry defined acids. This is generally expressed as **pH**, where:

$$pH = -\log[H^+] = \log(1/[H^+])$$

Likewise, hydroxide ion concentration,  $[OH^-]$ , is measured as **pOH**.

$$pOH = -\log[OH^-] = \log(1/[OH^-])$$

In any aqueous solution, the  $H_2O$  solvent dissociates slightly:



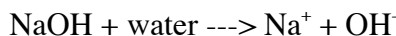
This dissociation is an equilibrium reaction and is therefore described by a constant,  $K_w$ , the water dissociation constant.  $K_w = [H^+][OH^-] = 10^{-14}$

Rewriting this equation using logarithms show:  $pH + pOH = 14$

In pure  $H_2O$ ,  $[H^+]$  is equal to  $[OH^-]$ , since for every one  $H_2O$  that dissociates, one  $H^+$  and  $OH^-$  is made. A solution with equal  $H^+$  and  $OH^-$  is **neutral** and has a pH of 7. A pH below 7 indicates an excess of  $H^+$  and thus an acidic solution. A pH above 7 indicates a relative excess of  $OH^-$  ions and therefore a basic solution.

### Strong Acids and Bases

Strong acids and bases are those that completely dissociate into their component ions in aqueous solution. For example, NaOH is added to water, it dissociates completely:

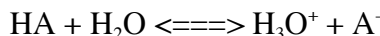


In a 1 M solution of NaOH, complete dissociation gives 1 M of OH<sup>-</sup>:

$$\text{pH} = 14 - (-\log[\text{OH}^-]) = 14 + \log[1] = 14$$

### Weak Acids and Bases

Weak acids and bases are those that only partially dissociate in aqueous solution. A weak **monoprotic acid**, HA in aqueous solution will achieve the following equilibrium after dissociation:



The **acid dissociation constant, Ka**, is a measure of the how much the acid dissociates.

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

Similarly, a weak monovalent base, BOH, dissociates to give B<sup>+</sup> and OH<sup>-</sup>. The **base dissociation constant, Kb**, is a measure of the degree that the base dissociates.

$$K_b = \frac{[\text{B}^+][\text{OH}^-]}{[\text{BOH}]}$$

A **conjugate acid** is defined as the acid formed when a base gains a proton. Similarly, a **conjugate base** is formed when an acid loses a proton. For example, with HCO<sub>3</sub><sup>-</sup>/CO<sub>3</sub><sup>2-</sup> conjugate acid/base pair, the CO<sub>3</sub><sup>2-</sup> is the conjugate base and HCO<sub>3</sub><sup>-</sup> is the conjugate acid. By adding the reactions of the conjugate acid and conjugate base with water, the net reaction is simply the dissociation of water. Thus if you know the dissociation constant for one, the other can be determined:

$$K_a \times K_b = K_w = 1 \times 10^{-14}$$

Thus Ka and Kb are inversely related. In other words, if Ka is large (the acid is strong), then Kb (base is weak) will be small and vice versa. From this relationship, one can see that when a conjugate acid/base pair form from a weak acid, the conjugate base is stronger than the acid.

## **Titration**

**Titration** is a procedure used to determine the concentration of an acid or base. This is accomplished by reacting a known volume of a solution of an unknown concentration with a known volume of a solution of known concentration. When the number of acid molecules equal the number of base molecules added (or vice versa), the **equivalence point** is reached.

The equivalence point in a titration is estimated in two common ways.: either by using a graphical method, plotting the pH of the solution as a function of added titrant by using a **pH meter** or by watching for a color change of an added **indicator**. Indicators are weak organic acids or bases that have different colors in their undissociated and dissociated states. These states are dependent on pH.

The **Henderson-Hasselbach equation** is used to estimate the pH of a solution:

$$\text{pH} = \text{pK}_a + \log\left[\frac{\text{conjugate base}}{\text{weak acid}}\right]$$

Note that when  $[\text{conjugate base}] = [\text{weak acid}]$  the  $\text{pH} = \text{pK}_a$  because the  $\log 1 = 0$ . In a titration, half-way to the equivalent point is when  $[\text{conjugate base}] = [\text{weak acid}]$ .