

8.S: Acids, Bases and pH (Summary)

- A bond that is formed from a hydrogen atom, which is part of a polar covalent bond (such as the O—H bond) to another, more electronegative atom (that has at least one unshared pair of electrons in its valence shell) is called a **hydrogen bond**. Hydrogen bonds are *weak*, partially covalent bonds. The **bond dissociation energy** of the O—H covalent bond 464 kJ/mole; the bond dissociation energy an O—H•••O hydrogen bond is about 21 kJ/mole.
- Even though hydrogen bonds are relatively weak, the *vast network* of hydrogen bonds in water makes the energy significant, and hydrogen bonding is generally used to explain the high boiling point of water (100 °C), relative to molecules of similar mass that *cannot* hydrogen bond. The extra energy represents the energy required to break down the hydrogen bonding network.
- Polar molecules, such as **acids**, strongly hydrogen bond to water. This hydrogen bonding not only *stabilizes* the molecular dipoles, but also *weakens* the H—A covalent bond (A represents the acid molecule). As a result of this weakening, the H—A bond in these acids stretches (the bond length increases) and then fully breaks. The hydrogen that was hydrogen-bonded to the water molecule now becomes *fully* bonded to the oxygen, forming the species H_3O^+ (the **hydronium ion**) and the acid now exists as an anion (A^-); this is the process of **acid dissociation**.
- The hydronium ion and the acid anion that are formed in an acid dissociation can react to re-form the original acid. This represents a set of forward- and back-reactions that occur together on a *very fast* time-scale; this type of a set of reactions is called an **equilibrium** and a double arrow is used in the chemical reaction to show this. This type of reaction is referred to as an **acid dissociation equilibrium**. $\text{HA}(\text{aq}) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{A}^-(\text{aq})$
- For any equilibrium, an **equilibrium constant** can be written that describes whether the *products* or the *reactants* will be the predominant species in solution. For the dissociation of the simple acid, HA, the equilibrium constant, K_a , is simply given by the ratio of the concentrations of the products and the reactants, **remembering that the molarity of the solutes have been used to approximate their activity, and that solvents, such as water, have an activity of 1**. Thus for the ionization of HA;

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

- According to the **Brønsted Acid-Base Theory**, any substance that ionizes in water to form hydronium ions (a **proton donor**) is called an **acid**; any substance that *accepts* a proton from a hydronium ion is a **base**. In an acid-base equilibrium, the **conjugate acid** is defined as the species that *donates* a hydrogen (a proton) in the *forward* reaction, and the **conjugate base** is the species that *accepts* a hydrogen (a proton) the *reverse* reaction. Thus for the ionization of HCl, HCl is the conjugate acid and Cl^- is the conjugate base. $\text{HCl}(\text{aq}) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
- Metal hydroxides, such as NaOH, dissolve in water to form metal cations and hydroxide anion. Hydroxide anion is a strong **Brønsted base** and, therefore, hydroxide anion accepts a proton from the hydronium ion to form *two moles of water*. The reaction of a Brønsted acid with a Brønsted base to form water is the process of **neutralization**.
- Just like water can promote the ionization of acids, water can also promote the ionization of *itself*. This equilibrium process occurs very rapidly in pure water and any sample of pure water will *always* contain a small concentration of hydronium and hydroxide ions. In pure water, at 25 °C, the concentration of hydronium ions ($[\text{H}_3\text{O}^+]$) and hydroxide ions ($[\text{HO}^-]$) will both be equal to exactly $1 \times 10^{-7} \text{ M}$. This is referred to as the **autoprotolysis** of water.
- The equilibrium for the autoprotolysis of water is defined as K_w , according to the equation shown below:

$$K_w = [\text{H}_3\text{O}^+][\text{HO}^-]$$

and at neutrality, $[\text{H}_3\text{O}^+]$ and $[\text{HO}^-]$ are *both* $1 \times 10^{-7} \text{ M}$, making the value of K_w

$$K_w = [1 \times 10^{-7}][1 \times 10^{-7}] = 1 \times 10^{-14}$$

- Based on the autoprotolysis equilibrium, acidic, basic and neutral solutions can be defined as:
 - A solution is **acidic** if $[\text{H}_3\text{O}^+] > 1 \times 10^{-7} \text{ M}$.
 - A solution is **basic** if $[\text{H}_3\text{O}^+] < 1 \times 10^{-7} \text{ M}$.
 - A solution is **neutral** if $[\text{H}_3\text{O}^+] = 1 \times 10^{-7} \text{ M}$.
- A **pH** value is simply the *negative* of the logarithm of the hydronium ion concentration ($-\log[\text{H}_3\text{O}^+]$).
- Remember that a logarithm consists of two sets of numbers; the digits to the left of the decimal point (the **characteristic**) reflect the integral power of 10, and are *not included* when you count significant figures. The numbers *after* the decimal (the **mantissa**) have the same significance as your experimental number, thus a logarithm of 4.15482 represents *five* significant figures.

- In an **acid-base titration** a solution of acid or base with a known concentration is slowly add to an acid or base solution with an *unknown* concentration, using a volumetric **burette**, to until neutrality has been achieved. Typically an indicator or a pH meter is used to signify neutrality.
- At neutrality, the *volume* and *concentration* of the reactant you have added is known, which means that you can calculate the number of *moles* that you added (remember, **concentration** × **volume** = **moles**). Based on the stoichiometry of your neutralization reaction, you then know how many moles of acid or base were in the unknown sample.

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