

AP Chemistry Unit 7 Review

Acids and Bases

Honors Lecture Notes: Unit 7

Definitions

An Arrhenius acid is a H^+ donor, and an Arrhenius base is an OH^- donor. These definitions are not always accurate, though, though, and so instead we use the Bronsted-Lowry definitions. A Bronsted-Lowry acid is the same as an Arrhenius acid- a proton donor. However, a Bronsted-Lowry base is a proton acceptor. By accepting protons from water, hydroxide ions are created, causing the solution to be basic.

Substances that can act as either acids or bases (such as water) are called amphoteric. Whether an amphoteric substance acts as an acid or as a base depends on what it is combined with.

A conjugate pair is two species that differ from each other by a single proton. Examples: $\text{NH}_3/\text{NH}_4^+$, $\text{HClO}_4/\text{ClO}_4^-$. The stronger the acid or base is, the weaker the conjugate is.

Ions In Solution

Typically, most anions create basic solutions, as due to their negative charge they will accept protons from water readily. The exception to this rule is the conjugates of the strong acids. As strong acids dissociate completely, their conjugate ions tend to be ineffective bases. The strong acids are: HCl , HNO_3 , H_2SO_4 , HI , HBr , and HClO_4 .

Ex: Na_2CO_3 is basic in solution because:



NaCl , however, is neutral in solution, as Cl^- is the conjugate of HCl and thus will not accept protons from water.

Unlike anions, cations are generally not considered to be acidic as they don't often have protons to donate (notable exception: NH_4^+). Strong bases are defined as those which donate 100% of their OH^- ions in solution. Strong bases are any alkali hydroxide as well as $\text{Ba}(\text{OH})_2$ and $\text{Sr}(\text{OH})_2$.

pH/pOH

To determine the relative acidity/basicity of a solution, the pH/pOH is used. $\text{pH} = -\log[\text{H}^+]$ and $\text{pOH} = -\log[\text{OH}^-]$. At 25°C , $\text{pH} + \text{pOH} = 14$ and $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ for any solution. Thus, as $[\text{H}^+]$ goes up, pH goes down. Likewise, as pH goes down, pOH goes up.

Any solution with a pH of less than 7 is considered acidic, and a pH of above 7 indicates a basic solution. Note that pH/pOH are NOT limited to a 0-14 scale. Sufficiently concentrated strong acids can have negative pHs, and concentrated strong bases can have pHs in excess of 14.

Auto-ionization of Water



For water at 25°C , $K_w = 1.0 \times 10^{-14}$, and $[\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}$ M. The K_w of water (like any equilibrium constant) changes with temperature. The auto-ionization of water is an endothermic process, and as such, if temperature increases K_w increases. This means that at temperatures above 25°C , the pH of pure water is NOT 7.0. However, neither is the pOH. For instance, at 50°C , $K_w = 5.5 \times 10^{-14}$, and thus $[\text{H}^+] = [\text{OH}^-] = 2.3 \times 10^{-7}$ M. At 50°C , $\text{pH} = \text{pOH} = 6.63$, and $\text{pH} + \text{pOH} = 13.2$.

K_a/K_b

The strength of a weak acid or base is defined by the equilibrium constant. These equilibrium constants are calculated the same way as any other equilibrium constant (products/reactants), but as it is always the loss/gain of a single proton, there are never any exponents. K_a/K_b values are often used to calculate the pH of a solution containing a weak acid or base.

Ex: Calculate the pH of 1.5 M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$. ($K_a = 1.8 \times 10^{-5}$).



$$K_a = \frac{[\text{H}^+][\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

As H^+ and $\text{C}_2\text{H}_3\text{O}_2^-$ are in a 1:1 ratio, both values can be replaced by x. So: $1.8 \times 10^{-5} = \frac{x^2}{1.5}$

$$x = [\text{H}^+] = 0.0051 \text{ M} \quad -\log(0.0051 \text{ M}) = 2.28$$

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