Acid-base Equilibrium

In Chem 1B we mostly use the Bronsted-Lowry **definitions of acids and bases**: a Bronsted acid is a source of H^+ while a Bronsted base is an acceptor of H^+. We are very interested in calculating aqueous concentrations of H^+ known as "H-plus" ions or "protons". We are also interested in calculating aqueous concentrations of OH^- known as "OH-minus" ions or "hydroxide" ions. Knowing concentrations of these ions has important ramifications in fields as wide-ranging as biotechnology, geology, oceanography, atmospheric science, and speculation about whether life could have existed on Mars.

A nice feature of acid-base equilibria is that reactions are usually simple homogeneous systems with stoichiometric coefficients of unity in dilute aqueous solution. So the resulting equilibrium expressions given by the Law of Mass Action are very simple:

generic acid dissociation: $HA <-> H^+ + A^-$ for which $Kc = Ka = [H^+][A^-]/[HA]$,

generic base hydrolysis: $A^- + H_2O <-> HA + OH^-$ for which $Kc = Kb = [HA][OH^-]/[A^-]$,

and

the auto-dissociation of water: $H_2O <-> H^+ + OH^-$ for which $Kc = Kw = [H^+][OH^-]$.

Note that whatever the substances are, HA and A^- are known as a **conjugate acid-base pair**. Also note that the third reaction (the auto-dissociation of water) is a sum of the first two. This implies that for any conjugate acid-base pair it must be that Ka x Kb = Kw.

Acid-base equilibria are principally about calculating equilibrium concentrations of acids, their conjugate bases, H^+ and OH^- ions by applying the techniques we already know about from Chapter 15 to the above reactions.