

10.5: Equilibria Involving Acids and Bases

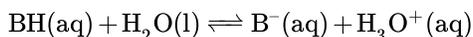
Consider a simple chemical system that is at equilibrium, such as dinitrogen tetroxide: nitrogen dioxide. The Law of Mass Action states that when this system reaches equilibrium, the ratio of the products and reactants (at a given temperature) will be defined by the equilibrium constant, K . Now imagine that, after equilibrium has been reached, *more* dinitrogen tetroxide is introduced into the container. In order for the ratio to remain constant (as defined by K) some of the N_2O_4 that you added must be converted to NO_2 . The addition of reactants or products to a system at equilibrium is commonly referred to as a “stress”. The response of the system to this stress is dictated by **Le Chatelier's Principle**.

Le Chatelier's Principle

Le Chatelier's Principle states that, if a “stress” is applied to a chemical reaction at equilibrium, the system will readjust in the direction that best *reduces* the stress imposed on the system. Again, *stress* refers to a change in **concentration**, a change in **pressure** or a change in **temperature**, depending on the system being examined. If pressure or temperature are changed, the numeric value K will change; **if only concentration changes are involved, K does not change**.

We will consider temperature and pressure effects in General Chemistry, but for now, remember; in a reaction at equilibrium, the introduction of more products will shift the mass balance towards more reactants, but the ratio of *Products/Reactants* (as defined by the equilibrium expression) does *not* change, hence, K is unchanged.

In [Chapter 8](#), we learned that a “weak acid” was only partially dissociated in solution, while a “strong acid” was fully dissociated. Now that we better understand the concept of equilibrium, these two classes of Brønsted acids can simply be differentiated based on their equilibrium constants. For an acid, BH , that dissociates in water to form B^- and hydronium ion:



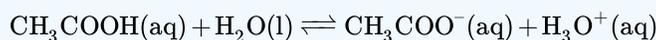
we can write a simple equilibrium expression, as follows:

$$K_C = \frac{[H_3O^+][B^-]}{[BH]} = K_a$$

You should note two things in this equation. Because the activity of water, as the solvent, is defined to have a value of 1, the activity for water does not affect the value of the equilibrium constant (remember, solids and liquids and solvents all have an activity of 1, and so do not affect the value of K) and the equilibrium constant for K_C is written as K_a to denote that this is an *acid dissociation equilibrium*. Now, as we learned in [Chapter 8](#), a strong acid is “fully dissociated”, which simply means that $[BH]$ is very, very small, thus K_a for a strong acid is very, very large. A weak acid is only “partially dissociated” which means that there are significant concentrations of both BH and B^- in solution, thus K_a for a weak acid is “small”. For most common weak acids, the values for K_a will be in the range of 10^{-3} to 10^{-6} .

Example 10.5.1:

Consider acetic acid (the acidic component of vinegar) where $K_a = 1.8 \times 10^{-5}$.



Solution

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]} = 1.8 \times 10^{-5}$$

Exercise 10.5.1

1. A series of acids have the following K_a values: rank these in descending order from the *strongest* acid to the *weakest* acid.

A. 6.6×10^{-4} B. 4.6×10^{-4} C. 9.1×10^{-8} D. 3.0×10^2

2. At 25.0 °C, the concentrations of H_3O^+ and OH^- in pure water are both 1.00×10^{-7} M, making $K_c = 1.00 \times 10^{-14}$ (recall that this equilibrium constant is generally referred to as K_w). At 60.0 °C, K_w increases to 1.00×10^{-13} . What is the pH of a sample of pure water at 60.0 °C?

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