**Strong and Weak Acids and Bases**

Acids are classified as either strong or weak, based on the extent to which they ionize in water. A **strong acid** *is an acid which is completely ionized in aqueous solution*. A **weak acid** *is an acid that ionizes only slightly in aqueous solution*. Acetic acid, the acid found in vinegar, is a very common weak acid. Its ionization is shown below.

CH3COOH(aq) ⇌ H+(aq) + CH3COO−(aq)

The ionization of acetic acid is incomplete, and so the equation is shown with a double arrow, indicating equilibrium between the reactant and products.

**The Acid Ionization Constant, Ka**

The ionization for a generic weak acid, HA, can be written in one of two ways.

HA(aq) + H2O(l) ⇌ H3O+(aq) + A−(aq)

HA(aq) ⇌ H+(aq) + A−(aq)

The water molecule is omitted for simplicity in the second case. This will be the way that acid ionization equations are written in many texts (because it is simpler), although the first equation shows the bronstead-lowry version of acid dissociation. Because the acid is weak and an equilibrium is established, an equilibrium constant expression can be written. As this Equilibrium constant is for an acid the K value is referred to as an **acid ionization constant (Ka)**.

| **Acid Ionization Constants at 25°C** |
| --- |
| **Name of Acid** | **Ionization Equation** | **Ka** |
| Sulfuric acid | H2SO4 ⇌ H+ + HSO4−HSO4− ⇌ H+ + SO42- | very large1.3 × 10−2 |
| Phosphoric acid | H3PO4 ⇌ H+ + H2PO4−H2PO4− ⇌ H+ + HPO42-HPO42- ⇌ H+ + PO43- | 7.5 × 10−36.2 × 10−84.8 × 10−13 |
| Acetic acid | CH3COOH ⇌ H+ + CH3COO− | 1.8 × 10−5 |
| Carbonic acid | H2CO3 ⇌ H+ + HCO3−HCO3− ⇌ H+ + CO32- | 4.2 × 10−74.8 × 10−11 |

$$K\_{a}= \frac{\left[H^{+}\right] [A^{-}]}{[HA]}$$

The higher the Ka value the stronger the acid. Weak acids have low Ka values. Because strong acids are essentially 100% ionized, the concentration of the acid in the denominator is nearly zero and the Ka value approaches infinity. For this reason, Ka values are generally reported only for weak acids.

**The Base Ionization Constant, Kb**

As with acids, bases can be either strong or weak, depending on their extent to which they accept a proton. This can be shown as an ionization reaction in water. A **strong base** *is a base that ionizes completely in aqueous solution*. A **weak base** *is a base that ionizes only slightly in aqueous solution*. An example is the dissociation of ammonia:

NH3(aq) + H2O(l) ⇌ NH4+(aq) + OH−(aq)

This equilibrium greatly favours the reactants, and the extent of ionization of the ammonia molecule is very small.

A **base ionization constant (Kb)** *is the equilibrium constant for the ionization of a base*. For ammonia, the expression is:

$$K\_{b}= \frac{\left[NH\_{4}^{+}\right] [OH^{-}]}{[NH\_{3}]}$$

The numerical value of Kb is a reflection of the strength of the base.

Notice that the conjugate base of a weak acid is also a weak base. For example, the acetate ion has a small tendency to accept a hydrogen ion from water to form acetic acid and the hydroxide ion.

**Calculations with Ka and Kb**

The numerical value of Ka or Kb can be determined by experiment. A solution of known concentration is prepared, and its pH is measured with an instrument called a pH meter.

**Sample Problem: Calculation of an Acid Ionization Constant for Formic Acid**

A 0.500 M solution of formic acid is prepared, and its pH is measured to be 2.04. Determine the Ka for formic acid.

* initial [HCOOH] = 0.500 M
* pH = 2.04
* Ka = ?

Calculating H+ concentration from pH:

 [H+] = 10-pH = 10-2.04 = 9.12 × 10−3

HCOOH ⇌ H+ + HCOO−

| **Concentrations** | **[HCOOH]** | **[H+]** | **[HCOO−]** |
| --- | --- | --- | --- |
| Initial | 0.500 | 0 # | 0 |
| Change | −9.12 × 10−3 | +9.12 × 10−3 | +9.12 × 10−3 |
| Equilibrium | 0.491 | 9.12 × 10−3 | 9.12 × 10−3 |

#Although the concentration of H+ would technically be 1.0 × 10−7 before any dissociation occurs, we can ignore this very small amount because, in this case, it does not make a noticeable contribution to the final value of [H+].

Substituting the equilibrium values into the Ka expression gives the following:

$$K\_{a}= \frac{\left[H^{+}\right] [HCOO^{-}]}{[HCOOH]} K\_{a}= \frac{\left[9.12×10^{-3}\right] [9.12×10^{--3}]}{[0.491]} K\_{a}= 1.7×10^{-4}$$

***Practice Problems***

1. Hypochlorous acid (HClO) is a weak acid. What is its Ka if a 0.250 M solution of hypochlorous acid has a pH of 4.07?
2. Calculate the pH of a 0.20 M solution of hydrocyanic acid.

**Problems**

1. A 1.25 M solution of a certain unknown acid has a pH of 4.73. Calculate the Ka of the acid.
2. A 0.350 M solution of a certain unknown base has a pH of 11.22. Calculate the Kb of the base.
3. Determine the pH and pOH of the following solutions.
	1. 2.40 M benzoic acid
	2. 0.745 M pyridine
4. What concentration of a solution of hydrofluoric acid should be prepared in order to have a pH of 2.00?