The ionization of an acid, HA, in water is depicted in the chemical equation shown below:

HA(aq) + H2O(l) ⇌H3O+(aq) + A-(aq).

The equilibrium constant for the ionization of an acid in water is called “Ka”:

$K\_{a}=\frac{\left[H\_{3}O^{+}\right][A^{-}]}{[HA]}$ .

The higher the Ka of an acid, the stronger the acid and the more it tends to ionize in water to form H3O+ and its conjugate base.

1. Why is H2O(l) not included in the Ka expression?
2. At equilibrium, the following concentrations of species are found in a sample of the acid HF. [HF] = 0.10 M, [H3O+] = 8.0\*10-3 M, [F-] = 8.0\*10-3 M

Based on these values, calculate the Ka of HF at this temperature.

1. The hydronium ion, H3O+, is the strongest acid that exists in water. Acids that are stronger than H3O+ are called strong acids. **Strong acids are assumed to completely ionize in solution to generate hydronium ions and their conjugate base.** For example, HCl is a strong acid. If 1M HCl is dissolved in water, it will completely ionize to produce 1M H3O+(aq) and 1M Cl-(aq), and no intact HCl will remain in solution.
2. Write out the chemical equation for the ionization of HCl in water.
3. What is the Ka value for a strong acid such as HCl?
4. Suppose 5 molecules of HCl are dissolved in water. In the beaker below, show the number of H3O+, Cl-, and HCl particles in the solution.



1. Suppose 0.10 M HCl is dissolved in water. What will the concentrations of H3O+, Cl-, and HCl be in solution?
2. pH = -log[H3O+]. If 0.10 M HCl is dissolved in water, what is the pH of the solution?
3. If 0.50 M HCl is dissolved in water, what is the pH of the solution?
4. HNO3 is another strong acid. If 0.50 M HNO3 is dissolved in water, what species will be present, and what is the pH of the solution?
5. **Weak acids only partly ionize in water** **to generate hydronium ions and their conjugate base. The higher the Ka of the acid, the more it will ionize in solution.** Because weak acids only partly ionize in solution, RICE tables are useful tools to keep track of the changes in concentrations when a weak acid solution comes to equilibrium.
6. The make-believe acid HX has Ka=0.25. Suppose 5 molecules of HX are dissolved in water. In the beaker below, show the number of H3O+, X-, and HX in the solution.

*(Hint: From the reaction HX(aq) + H2O(l) ⇌ H3O+(aq) + X-(aq), H3O+ and X- will be formed in a 1:1 ratio, so assume these are equal to each other for now.* *What ratio of products to reactants will give Ka=0.25?)*



1. Suppose 0.10 M HX (Ka=0.25) is dissolved in water. Fill in the RICE table below, then use the Ka value to determine the concentrations of H3O+, X-, and HX in solution at equilibrium. Assume H3O+ and X- are not present initially.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Reaction |  |  |  |  |
| I (M) |  |  |  |  |
| C (M) |  |  |  |  |
| E (M) |  |  |  |  |

1. What is the pH of the 0.10 M HX solution?

How does it compare to the pH of the 0.10 M HCl solution? [3e, on p. 2]

Why do the pH values of the solutions differ?

1. Acetic acid (CH3COOH), the acid in vinegar, is another weak acid, with Ka = 1.8\*10-5. The ionization reaction is shown below:

CH3COOH(*aq*) + H2O(*l*) ⇌ CH3COO-(*aq*) + H3O+(*aq*)

1. Before doing any calculations, how would you expect the [H3O+] *and* pH of a 0.10 M aqueous acetic acid solution to compare to the 0.10 M HCl and 0.10 M HX solutions?
2. Suppose 0.10 M CH3COOH is dissolved in water. What will the concentrations of H3O+, CH3COO-, and CH3COOH be in the solution at equilibrium*?* Assume no H3O+ or CH3COO- are initially present. *(Since the equilibrium constant is very small, what approximation might you try to simplify the calculation?)*
3. What is the pH of the 0.10 M CH3COOH solution? Was your prediction from b correct?
4. The percent ionization of a weak acid is the ratio of the concentration of the ionized acid to the initial acid concentration, times 100:

% ionization = $\frac{[conjugate base]eq}{\left[acid\right]initial}\*100\%$

What is the percent ionization of the 0.10 M CH3COOH(aq) solution in Question 5?

1. A solution is prepared by dissolving 0.250 mole acetic acid in enough water to make a 0.500 L solution. Use a RICE table to calculate the equilibrium concentrations, then determine the pH of the solution and the percent ionization of the acid at equilibrium.
2. In the acetic acid solutions above, how do the (1) pH of the solution and

 (2) percent ionization of the acid

change as the concentration of the weak acid in solution increases?