Chemistry 132 NT

The face of a child can say it all. Especially the mouth part of the face.

Jack Handey

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Chem 132 NT

Acid-Base Equilibria

Module 1

Solutions of a Weak Acid

- **Acid-Ionization Equilibria**
- **Polyprotic Acids**

Reaction of zinc metal with hydrochloric acid.

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Review

- Identifying acid and base species
- Identifying Lewis acid and base species
- Deciding whether reactants or products are favored in an acid-base reaction
- Calculating the concentration of H^+ and OH^- in solutions of strong acid or base
- Calculating the pH from the hydrogen-ion concentration, or vice versa.

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Solutions of a Weak Acid or Base

- The simplest acid-base equilibria are those in which a single acid or base solute reacts with water.
  - In this chapter, we will first look at solutions of <u>weak</u> **acids** and **bases**.
  - We must also consider solutions of **salts**, which can have acidic or basic properties as a result of the reactions of their ions with water.

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Acid-Ionization Equilibria

- **Acid ionization** (or acid dissociation) is the reaction of an acid with water to produce hydronium ion (hydrogen ion) and the conjugate base anion.
  - When acetic acid is added to water it reacts as follows.

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

  - Because acetic acid is a weak electrolyte, it ionizes to a small extent in water.

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Acid-Ionization Equilibria

- For a weak acid, the equilibrium concentrations of ions in solution are determined by the **acid-ionization constant** (also called the acid-dissociation constant).
  - Consider the generic monoprotic acid, HA.

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

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Acid-Ionization Equilibria

- For a weak acid, the equilibrium concentrations of ions in solution are determined by the **acid-ionization constant** (also called the acid-dissociation constant).
  - The corresponding equilibrium expression is:

$$K_c = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

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Acid-Ionization Equilibria

- For a weak acid, the equilibrium concentrations of ions in solution are determined by the **acid-ionization constant** (also called the acid-dissociation constant).
  - Since the concentration of water remains relatively constant, we rearrange the equation to get:

$$K_a = [H_2O]K_c = \frac{[H_3O^+][A^-]}{[HA]}$$

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Acid-Ionization Equilibria

- For a weak acid, the equilibrium concentrations of ions in solution are determined by the **acid-ionization constant** (also called the acid-dissociation constant).
  - Thus, K_a **, the acid-ionization constant**, equals the constant $[H_2O]K_c$.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

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Acid-Ionization Equilibria

- For a weak acid, the equilibrium concentrations of ions in solution are determined by the **acid-ionization constant** (also called the acid-dissociation constant).
  - A table in your text lists acid-ionization constants for various weak acids. Here are a couple of examples.

Substance	Formula	K_a
Acetic acid	$HC_2H_3O_2$	1.7×10^{-5}
Hydrocyanic acid	HCN	4.9×10^{-10}

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Experimental Determination of K_a

- The **degree of ionization** of a weak electrolyte is the fraction of molecules that react with water to give ions.
  - Electrical conductivity or some other colligative property can be measured to determine the degree of ionization.
  - With weak acids, the pH can be used to determine the equilibrium composition of ions in the solution.

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - It is important to realize that the solution was made 0.012 M in nicotinic acid, however, some molecules ionize making the equilibrium concentration of nicotinic acid less than 0.012 M.
  - We will abbreviate the formula for nicotinic acid as HNic.

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - Let x be the moles per liter of product formed.

$$HNic(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Nic^-(aq)$$

	HNic		H_3O^+	Nic^-
Starting	0.012		0	0
Change	-x		+x	+x
Equilibrium	0.012-x		x	x

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - The equilibrium-constant expression is:

$$K_a = \frac{[H_3O^+][Nic^-]}{[HNic]}$$

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - Substituting the expressions for the equilibrium concentrations, we get

$$K_a = \frac{x^2}{(0.012 - x)}$$

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - We can obtain the value of x from the given pH.

$$x = [H_3O^+] = \text{antilog}(-pH)$$
$$x = \text{antilog}(-3.39)$$
$$x = 4.1 \times 10^{-4} = 0.00041$$

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - Substitute this value of x in our equilibrium expression.
  - Note first, however, that

$$(0.012 - x) = (0.012 - 0.00041) = 0.01159 \cong 0.012$$

the concentration of unionized acid remains virtually **unchanged**.

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - Substitute this value of x in our equilibrium expression.

$$K_a = \frac{x^2}{(0.012 - x)} \cong \frac{(0.00041)^2}{(0.012)} = 1.4 \times 10^{-5}$$

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A Problem To Consider

- Nicotinic acid is a weak monoprotic acid with the formula $HC_6H_4NO_2$. A 0.012 M solution of nicotinic acid has a pH of 3.39 at 25 °C. **Calculate the acid-ionization constant, K_a,** for this acid at 25 °C.
  - To obtain the degree of dissociation:

$$\text{Degree of dissociation} = \frac{0.00041}{0.012} = 0.034$$

  - The percent ionization is obtained by multiplying by 100, which gives 3.4%.

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Calculations With K_a

- Once you know the value of K_a, you can calculate the **equilibrium concentrations of species HA, A^-, and H_3O^+** for solutions of different molarities.
  - The general method for doing this was discussed in our chapter on equilibrium.

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Calculations With K_a

- Note that in our previous example, the degree of dissociation was so small that **"x" was negligible compared to the concentration of nicotinic acid**.
- It is the **small value of the degree of ionization that allowed us to ignore the subtracted x** in the denominator of our equilibrium expression.
- The degree of ionization of a weak acid depends on both the K_a and the concentration of the acid solution (see the figure on the next slide).

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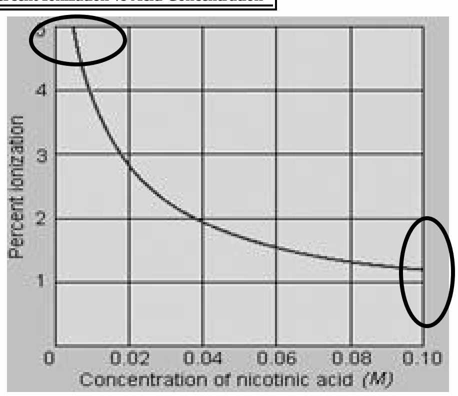
Percent Ionization vs Acid Concentration



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Calculations With K_a

- **How do you know when you can use this simplifying assumption?**
  - It can be shown that if the acid concentration, C_a, divided by the K_a exceeds 100, that is,

$$\text{if} \quad C_a / K_a > 100$$

then this simplifying assumption of **ignoring the subtracted x gives an <u>acceptable</u> error of less than 5%.**

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Calculations With K_a

- **How do you know when you can use this simplifying assumption?**
  - If the simplifying assumption **is valid**, you can approximate the equilibrium calculations by ignoring added or subtracted "x"s in the equilibrium expression.
  - The next example illustrates this with a solution of nicotinic acid, again.

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4×10^{-5}.
  - As in our previous example, the solution was made 0.10 M in nicotinic acid, however, some molecules ionize making the equilibrium concentration of nicotinic acid less than 0.10 M.
  - We will again abbreviate the formula for nicotinic acid as HNic.

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4×10^{-5}.
  - Let x be the moles per liter of product formed.

$$HNic(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Nic^-(aq)$$

	HNic	H_3O^+	Nic^-
Starting	0.10	0	0
Change	-x	+x	+x
Equilibrium	0.10-x	x	x

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4 x 10^{-5}.
  - The equilibrium-constant expression is:

$$K_a = \frac{[H_3O^+][Nic^-]}{[HNic]}$$

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4 x 10^{-5}.
  - Substituting the expressions for the equilibrium concentrations, we get

$$K_a = \frac{x^2}{(0.10 - x)} = 1.4 \times 10^{-5}$$

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4 x 10^{-5}.
  - Here is where we must decide whether we can simplify the equation or if we must solve the quadratic.

$$K_a = \frac{x^2}{(0.10 - x)} = 1.4 \times 10^{-5}$$

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4×10^{-5}.
  - Divide the acid concentration by the K_a.

$$C_a / K_a = 0.10\ M\ Hnic / 1.4 \times 10^{-5} = 7143$$

which is much greater than 100, so **we can use our simplifying assumption that x will be negligible**.

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4×10^{-5}.
  - This means that "x" is so much smaller than the 0.10 M HNic concentration, that subtracting it from 0.10 won't change the value.
  - That is,

$$0.10 - x \approx 0.10$$

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4×10^{-5}.
  - Going back to our equilibrium expression, we see

$$K_a = \frac{x^2}{(0.10 - x)} = 1.4 \times 10^{-5} \approx \frac{x^2}{(0.10)}$$

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4 x 10^{-5}.
  - Solving for x

$$x^2 \cong (1.4 \times 10^{-5}) \times (0.10) \cong 1.4 \times 10^{-6}$$

$$x \cong 1.2 \times 10^{-3} = 0.0012$$

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4 x 10^{-5}.
  - Checking to see if our simplifying assumption was correct, we see that

0.10 - x = (0.10 - 0.0012) = 0.10

(to two significant figures)

the assumption was indeed valid.

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4 x 10^{-5}.
  - We can now find the equilibrium concentrations of all the species by substituting x into the last line of our original table.
  - The concentrations of **HNic**, $\mathbf{H^+}$, and $\mathbf{Nic^-}$ are **0.10 M, 0.0012 M,** and **0.0012 M**, respectively.

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A Problem To Consider

- What are the concentrations of **nicotinic acid**, **hydrogen ion**, and **nicotinate ion** in a solution of 0.10 M nicotinic acid, $HC_6H_4NO_2$, at 25°C? What is the **pH** of the solution? The acid ionization constant, K_a, is 1.4 x 10^{-5}.
  - The pH of the solution is

$$pH = -\log[H^+] = -\log(0.0012) = 2.92$$

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Calculations With K_a

- **How do you know when you can use this simplifying assumption?**
  - If the simplifying assumption **is not valid**, you can solve the equilibrium equation exactly by using the **quadratic equation**.
  - The next example illustrates this with a solution of aspirin (acetylsalicylic acid), $HC_9H_7O_4$, a common headache remedy.

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - The molar mass of $HC_9H_7O_4$ is 180.2 g.

    From this we find that the sample contained 0.00180 mol of the acid.

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - The molar mass of $HC_9H_7O_4$ is 180.2 g.

    Hence, the concentration of the acetylsalicylic acid is 0.00180 mol/0.500 L = **0.0036 M** (Retain two significant figures, the same number of significant figures in K_a).

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - Note that

$$C_a / K_a = 0.0036 / 3.3 \times 10^{-4} = 11$$

which is less than 100, so **we must solve the equilibrium equation exactly**.

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - We will abbreviate the formula for acetylsalicylic acid as **HAcs** and let **x** be the amount of H_3O^+ formed per liter.
  - The amount of acetylsalicylate ion is also **x** mol; the amount of nonionized acetylsalicylic acid is **(0.0036-x) mol**.

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - These data are summarized below.

$$HAcs(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Acs^-(aq)$$

	HAcs	H_3O^+	Acs^-
Starting	0.0036	0	0
Change	-x	+x	+x
Equilibrium	0.0036-x	x	x

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - The equilibrium constant expression is

$$\frac{[H_3O^+][Acs^-]}{[HAcs]} = K_a$$

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - If we substitute the equilibrium concentrations and the K_a into the equilibrium constant expression, we get

$$\frac{x^2}{(0.0036 - x)} = 3.3 \times 10^{-4}$$

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A Problem To Consider

- **What is the pH** at 25 ^{o}C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 ^{o}C.
  - You can solve this equation exactly by using the quadratic formula.
  - Rearranging the preceding equation to put it in the form $ax^2 + bx + c = 0$, we get

$$x^2 + (3.3 \times 10^{-4})x - (1.2 \times 10^{-6}) = 0$$

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A Problem To Consider

- **What is the pH** at 25 ^{o}C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 ^{o}C.
  - Now substitute into the quadratic formula.

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

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A Problem To Consider

- **What is the pH** at 25 ^{o}C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 ^{o}C.
  - Now substitute into the quadratic formula.

$$x = \frac{-(3.3 \times 10^{-4}) \pm \sqrt{(3.3 \times 10^{-4})^2 - 4(1.2 \times 10^{-6})}}{2}$$

  - The lower sign in ± gives a negative root which we can ignore.

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A Problem To Consider

- **What is the pH** at 25 °C of a solution obtained by dissolving 0.325 g of acetylsalicylic acid (aspirin), $HC_9H_7O_4$, in 0.500 L of water? The acid is monoprotic and K_a=3.3 x 10^{-4} at 25 °C.
  - Taking the upper sign, we get

$$x = [H_3O^+] = 9.4 \times 10^{-4}$$

  - Now we can calculate the pH.

$$pH = -\log(9.4 \times 10^{-4}) = 3.03$$

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Polyprotic Acids

- Some acids have two or more protons (hydrogen ions) to donate in aqueous solution. These are referred to as **polyprotic acids**.
  - Sulfuric acid, for example, can lose two protons in aqueous solution.

$$H_2SO_4(aq) + H_2O(l) \rightarrow H_3O^+(aq) + HSO_4^-(aq)$$

$$HSO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_4^{2-}(aq)$$

  - The first proton is lost completely followed by a weak ionization of the hydrogen sulfate ion, HSO_4^-.

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Polyprotic Acids

- Some acids have two or more protons (hydrogen ions) to donate in aqueous solution. These are referred to as **polyprotic acids**.
  - For a weak diprotic acid like carbonic acid, H_2CO_3, two simultaneous equilibria must be considered.

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$$

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Polyprotic Acids

- Some acids have two or more protons (hydrogen ions) to donate in aqueous solution. These are referred to as **polyprotic acids**.
  - Each equilibrium has an associated acid-ionization constant.

For the loss of the **first** proton

$$K_{a1} = \frac{[H_3O^+][HCO_3^-]}{[H_2CO_3]} = 4.3 \times 10^{-7}$$

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Polyprotic Acids

- Some acids have two or more protons (hydrogen ions) to donate in aqueous solution. These are referred to as **polyprotic acids**.
  - Each equilibrium has an associated acid-ionization constant.

For the loss of the **second** proton

$$K_{a2} = \frac{[H_3O^+][CO_3^{2-}]}{[HCO_3^-]} = 4.8 \times 10^{-11}$$

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Polyprotic Acids

- Some acids have two or more protons (hydrogen ions) to donate in aqueous solution. These are referred to as **polyprotic acids**.
  - In general, the second ionization constant, K_{a2}, for a polyprotic acid is smaller than the first ionization constant, K_{a1}.
  - In the case of a triprotic acid, such as H_3PO_4, the third ionization constant, K_{a3}, is smaller than the second one, K_{a2}.

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Polyprotic Acids

- Some acids have two or more protons (hydrogen ions) to donate in aqueous solution. These are referred to as **polyprotic acids**.
  - When several equilibria occur at once, it might appear complicated to calculate equilibrium compositions.
  - However, reasonable assumptions can be made that simplify these calculations as we show in the next example.

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - For diprotic acids, K_{a2} is so much smaller than K_{a1} that the smaller amount of hydronium ion produced in the second reaction **can be neglected**.

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - The pH can be determined by simply solving the equilibrium problem posed by the first ionization.

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.

  - If we abbreviate the formula for ascorbic acid as H_2Asc, then the first ionization is:

$$H_2Asc(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HAsc^-(aq)$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.

$$H_2Asc(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HAsc^-(aq)$$

	H_2Asc	H_3O^+	$HAsc^-$
Starting	0.10	0	0
Change	-x	+x	+x
Equilibrium	0.10-x	x	x

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.

  - The equilibrium constant expression is

$$\frac{[H_3O^+][HAsc^-]}{[H_2Asc]} = K_{a1}$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?
  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - Substituting into the equilibrium expression

$$\frac{x^2}{(0.10 - x)} = 7.9 \times 10^{-5} \cong \frac{x^2}{0.10}$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?
  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - Assuming that x is much smaller than 0.10, you get

$$x^2 \cong (7.9 \times 10^{-5}) \times (0.10)$$

$$x \cong 2.8 \times 10^{-3} = 0.0028$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?
  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - The hydronium ion concentration is 0.0028 M, so

$$pH = -\log(0.0028) = 2.55$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.

  - The ascorbate ion, Asc^{2-}, which we will call "y", is produced only in the second ionization of H_2Asc.

$$HAsc^{-}(aq) + H_2O(l) \rightleftharpoons H_3O^{+}(aq) + Asc^{2-}(aq)$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.

  - Assume the starting concentrations for $HAsc^-$ and H_3O^+ to be those from the first equilibrium.

$$HAsc^{-}(aq) + H_2O(l) \rightleftharpoons H_3O^{+}(aq) + Asc^{2-}(aq)$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.

$$HAsc^{-}(aq) + H_2O(l) \rightleftharpoons H_3O^{+}(aq) + Asc^{2-}(aq)$$

	$HAsc^-$	H_3O^+	Asc^{2-}
Starting	0.0028	0.0028	0
Change	-y	+y	+y
Equilibrium	0.0028-y	0.0028+y	y

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$? The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - The equilibrium constant expression is

$$\frac{[H_3O^+][Asc^{2-}]}{[HAsc^-]} = K_{a2}$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$? The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - Substituting into the equilibrium expression

$$\frac{(0.0028 + y)(y)}{(0.0028 - y)} = 1.6 \times 10^{-12}$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$? The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.
  - Assuming y is much smaller than 0.0028, the equation simplifies to

$$\frac{(0.0028)(y)}{(0.0028)} \approx 1.6 \times 10^{-12}$$

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A Problem To Consider

- Ascorbic acid (vitamin C) is a diprotic acid, $H_2C_6H_6O_6$. **What is the pH** of a 0.10 M solution? What is the concentration of the ascorbate ion, $C_6H_6O_6^{2-}$?

  The acid ionization constants are $K_{a1} = 7.9 \times 10^{-5}$ and $K_{a2} = 1.6 \times 10^{-12}$.

  - Hence, $y \cong [Asc^{2-}] = 1.6 \times 10^{-12}$
  - The concentration of the ascorbate ion equals K_{a2}.

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Operational Skills

- Determining K_a (or K_b) from the solution pH
- Calculating the concentration of a species in a weak acid solution using K_a
- Calculating concentrations of species in a solution of a diprotic acid

Time for a few review questions

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