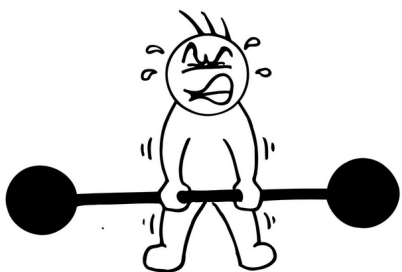


5—Weak Acids, Bases and their Salts



Name: _____

Date: _____

Section: _____

Objectives

- You will be able to define an acid and a base with the Arrhenius and Brønsted-Lowry definitions
- You will be able to predict the behavior of salts of strong acids and bases in water
- You will be able to predict the behavior of weak acids and bases in water
- You will be able to explain the pH scale of water in terms of $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$
- You will be able to explain the degree of ionization of a weak acid or weak base in terms of the dissociation constants, K_a and K_b .

Pre-Laboratory Requirements

- Read Chapter 18 in Silberberg, pp 776-814.
- Pre-Lab Questions (if required by your instructor)
- Laboratory Notebook—prepared before lab (if required by your instructor)

Safety Notes

- Eye protection must be worn at all times.
- Hydrochloric acid and sodium are caustic and should not come in contact with your skin or clothing. Wear gloves when handling these chemicals. A lab coat or lab apron is recommended.

Discussion

Strong Acids and Bases. Acids were known to the Romans who used the word *acidus* to describe foods that had a sour taste. It was later determined that the sour taste was caused by citric acid. All acids have a sour taste, which explains why good cooks keep a variety of acids available for use in preparing food.

Arrhenius defined acids as substances that form the hydronium ion, H_3O^+ , when placed in water. A more general definition of acids was later provided by Brønsted and Lowry, who defined acids as proton donors. The Arrhenius definition of a base is a substance that forms hydroxide ion when placed in water, and the Brønsted Lowry definition of a base is a “proton acceptor.”

The pH Scale. The hydronium ion concentration in water can vary over 14 orders of magnitude. A logarithmic scale has been adopted as a convenient way of expressing hydrogen ion concentration in whole numbers, which is more convenient than writing out negative exponential numbers. This is known as the pH scale, and based on the definition of pH:

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$$

The “p” scale has been expanded to include hydroxide, pOH, and acid dissociation constants, pK_a:

$$\text{pOH} = -\log_{10}[\text{OH}^-] \text{ and } \text{pK}_a = -\log_{10}K_a.$$

Neutral water has equal amounts of hydronium ion and hydroxide ion, and these concentrations are known to be 1×10^{-7} M at room temperature. The pH of neutral water is therefore 7.0 (as is the pOH of neutral water).

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}[1 \times 10^{-7}] = 7.0; \quad \text{pOH} = -\log_{10}[\text{OH}^-] = -\log_{10}[1 \times 10^{-7}] = 7.0$$

The dissociation constant for water, K_w, is 1×10^{-14} , and pK_w would therefore have a value of 14.0.

$$\text{pK}_w = -\log_{10}K_w = -\log_{10}[1 \times 10^{-14}] = 14.0 = \text{pH} + \text{pOH}$$

Acids differ in their ability to generate protons when placed in water. Some acids, HCl and HNO₃ for instance, ionize completely when placed in water, forming hydronium ion, H₃O⁺, and either Cl⁻ (in the case of HCl) or NO₃⁻ (in the case of HNO₃). A 0.10 M HCl solution would contain 0.10 moles of H₃O⁺ and 0.10 moles of Cl⁻. Acids that dissociate completely when placed in water are called strong acids. Because strong acids are completely ionized in water we would expect a 0.010 M solution of HNO₃ to have a pH of 2.0

$$\text{pH} = -\log_{10}[1 \times 10^{-2}] = 2.0$$

The pOH for this solution can also be calculated:

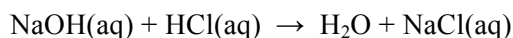
$$14.0 = \text{pH} + \text{pOH}; \quad 14.0 = 2.0 + \text{pOH}; \quad \text{pOH} = 12.0$$

Weak Acids and Bases. Weak acids and bases are only partially dissociated in water. The concentration of hydronium ions formed from weak acids is a function of the acid dissociation constant, K_a, for that substance. Compounds that have small acid dissociation constants generate less hydronium ion than substances that have larger dissociation constants. Acetic acid is a weak acid, and its K_a is 1.8×10^{-5} at room temperature. The acid dissociation constant for weak acids is an equilibrium constant, and all weak acids will therefore react with water to satisfy the K_a expression.

Degree of Ionization. The degree of ionization, α, is a measure of dissociation a substance undergoes when placed in water, and is defined as the ratio of ionized substance divided by the amount of substance initially placed in water. Strong acids or bases have α values of 1.0 (or very close to 1.0), and weak acids and bases have much smaller values. To calculate α you must know the initial concentration of weak acid or weak base in solution, and the equilibrium concentration of either H₃O⁺ or OH⁻. Consider a 0.010 M solution of acetic acid, HC₂H₃O₂, with a measured pH of 3.4. The degree of ionization, α, for this solution would be:

$$\begin{aligned} \text{HC}_2\text{H}_3\text{O}_2 &\rightleftharpoons \text{H}^+ + \text{C}_2\text{H}_3\text{O}_2^- \\ \text{pH} = 3.4; \quad [\text{H}_3\text{O}^+] &= [\text{C}_2\text{H}_3\text{O}_2^-] = 10^{-3.4} = 4.2 \times 10^{-4} \\ \alpha &= \frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = \frac{[4.2 \times 10^{-4}]}{[0.010]} = 0.042 \end{aligned}$$

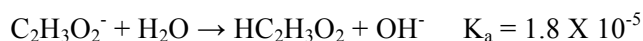
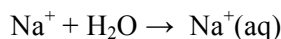
Salts. Salts are the product of neutralization reactions. Sodium chloride forms when equal amounts of NaOH combines with HCl as shown in the following equation:



All salts will ionize completely when placed in aqueous solution. One mole of sodium chloride will generate one mole of sodium ions and one mole of chloride ions when placed in solution because sodium chloride is the product of a neutralization of a strong acid with a strong base, i.e., both the anion (Cl⁻ in this case) and the cation (Na⁺ in this example) are always 100% ionized.

Salts from weak acids and weak bases will also ionize completely when placed in aqueous solution. However, weak acids and weak bases will react with neutral water to satisfy their dissociation constant (a condition that must always be met when these substances are placed in water). The need of weak acids and bases to satisfy their dissociation constant equilibrium requirement disrupts the pH balance of neutral water, resulting in an excess of hydroxide ion (in the case of a weak acid) or an excess of hydronium ion (in the case of a weak base).

For example, consider the compound sodium acetate (NaC₂H₃O₂). When sodium acetate is placed in water it dissolves completely, just like the sodium chloride. Since the sodium ion was the cation from a strong base (NaOH), the sodium cation will remain 100% ionized. The acetate ion (C₂H₃O₂⁻) on the other hand, reacts with neutral water to satisfy its acid dissociation constant K_a.



Because of the reaction of acetate ion with water, the resulting solution of sodium acetate would be expected to have an *excess of hydroxide ion, and a pH that is greater than 7.0*. Salts from a weak base and a strong acid also react with neutral water, but they would be expected to generate an *excess of hydronium ions, lowering the pH to below 7.0*.

Procedure

We will measure the pH of solutions that have been previously prepared using a pH meter. To perform the measurements follow the simple procedure below.

1. Rinse your test tube with DI water.
2. Place a small amount of your solution into the rinsed test tube.
3. Submerge the pH meter into your solution. Make sure to avoid contact with the sides of the test tube to ensure an accurate reading.
4. Repeat steps 1-3 for each sample. When using a pH meter you must avoid contamination between solutions by washing the end of the meter each time it is used.

Part 1-A. pH of Salt Solutions

Salt Solution	pH	[H ₃ O ⁺]	[OH ⁻]
1.0 M NaCl	_____	_____	_____
1.0 M NaC ₂ H ₃ O ₂	_____	_____	_____
1.0 M NaHSO ₄	_____	_____	_____
1.0 M Na ₂ CO ₃	_____	_____	_____
1.0 M NH ₄ Cl	_____	_____	_____
1.0 M NH ₄ C ₂ H ₃ O ₂	_____	_____	_____

Part 1-B.

Use the recorded pH values of these solutions to write chemical equations for the reaction of the salts with water. For acidic solutions write a chemical equation that shows the production of H_3O^+ , and if the solution is basic write a chemical equation that shows the production of hydroxide. Do not include spectator ions in your equations.

Solute	Chemical Equations
$\text{NaC}_2\text{H}_3\text{O}_2$	_____
NaHSO_4	_____
Na_2CO_3	_____
NH_4Cl	_____

Part 1-C.

For solutions that are acidic, write an expression for the acid dissociation constant, K_a , for that salt, and for solutions that are basic, write an expression for the base ionization constant, K_b . Use your pH measurements from Part A and the K_a or K_b expressions to calculate values for K_a and K_b .

Solute	Expression for K_a or K_b	Value of K_a or K_b
$\text{NaC}_2\text{H}_3\text{O}_2$	_____	_____
NaHSO_4	_____	_____
Na_2CO_3	_____	_____
NH_4Cl	_____	_____

Part 2-A

Percent Ionization of Strong and Weak Acids and Bases

Measure and record the pH values of hydrochloric acid, acetic acid and aqueous ammonia where the concentrations are 1.00M, 0.100 M and 0.010 M. Use the same procedure for determining pH that was used with the salt solutions in Part 1-A. Calculate the $[H_3O^+]$ and $[OH^-]$ concentrations in each of the nine solutions and enter the values into the table.

		1.00 M	0.10 M	0.010 M
HC ₂ H ₃ O ₂	pH	_____	_____	_____
	$[H_3O^+]$	_____	_____	_____
	$[OH^-]$	_____	_____	_____
NH ₃ (aq)	pH	_____	_____	_____
	$[H_3O^+]$	_____	_____	_____
	$[OH^-]$	_____	_____	_____

Calculate the acid dissociation constant, K_a , and the degree of ionization, α , for HCl and HC₂H₃O₂ from the data in the table above. Then calculate the base dissociation constant, K_b , and the degree of ionization, α , for aqueous ammonia.

		1.00 M	0.10 M	0.010 M
HC ₂ H ₃ O ₂	K_a	_____	_____	_____
	α	_____	_____	_____
NH ₃ (aq)	K_b	_____	_____	_____
	α	_____	_____	_____

Show Your Calculations Here: