1. The Relative Strengths of Some Acids

Introduction

Acids and bases, two closely connected types of substances, are part of our everyday lives. They can be found in foods, soft drinks, medicines, and cleaning products.

The relationship between acids and bases occurs because these substances react readily with each other. An acid-base reaction is a competition for protons (Ebbing/Gammon, Chapter 16). The extent of the reaction depends on the strengths of the acid and the base. A strong acid reacts with a base more completely than a weaker acid reacts with the same base. This difference enables us to measure the relative strengths of acids (and bases).

Purpose

You will estimate or measure the pH of various solutions, using indicators, pH paper, and (optionally) a pH meter.

Relative strengths of acids

How can you determine the relative strengths of two acids? As an example, let us consider how you would determine the relative strengths of HF (hydrofluoric acid) and HCN (hydrocyanic acid). You would prepare isomolar aqueous solutions of these acids and examine the extent of their ionizations in water. The equations that describe these reactions are

$$HF(aq) + H_2O(l) \rightleftharpoons F^-(aq) + H_3O^+(aq)$$

$$HCN(aq) + H_2O(l) \rightleftharpoons CN^-(aq) + H_3O^+(aq)$$

The ionization of the stronger acid will be greater than the ionization of the weaker acid. As a result, the stronger acid will provide more H_3O^+ ions than the weaker acid. You would find in this case that $[H_3O^+]$ is greater in the solution of HF. You would then know that HF is a stronger acid than HCN.

The relative strengths of the acids in Table 1.1 were determined by this method. The table also shows the relative strengths of the conjugate bases of the acids. When you compare the strengths of any two acids in the table, you will note that the stronger acid has the weaker conjugate base.

Two special notes

The aluminum cation, Al^{3+}, exists in aqueous solution as the hydrated ion, $Al(H_2O)_6{}^{3+}$. Table 1.1 shows that this ion is an acid. Why? The water molecules are bonded to the metal through lone pairs of electrons on the oxygen atoms. The positively charged aluminum ion draws electrons from the oxygen atoms, which, in turn, draw electrons from the O–H bonds and weaken them. As a result, the water molecules tend to be acidic:

$$Al(H_2O)_6{}^{3+}(aq) + H_2O(l) \rightleftharpoons Al(H_2O)_5OH^{2+}(aq) + H_3O^+(aq)$$

Table 1.1 Relative Strengths of Some Acids and Bases

	Acid	Base	
Strongest acids ↓	$HClO_4$	ClO_4^-	Weakest bases ↑
	H_2SO_4	HSO_4^-	
	HI	I^-	
	HBr	Br^-	
	HCl	Cl^-	
	HNO_3	NO_3^-	
	H_3O^+	H_2O	
	HSO_4^-	SO_4^{2-}	
	H_2SO_3	HSO_3^-	
	H_3PO_4	$H_2PO_4^-$	
	HNO_2	NO_2^-	
	HF	F^-	
	$HC_2H_3O_2$	$C_2H_3O_2^-$	
	$Al(H_2O)_6^{3+}$	$Al(H_2O)_5OH^{2-}$	
	H_2CO_3	HCO_3^-	
	H_2S	HS^-	
	$HClO$	ClO^-	
	$HBrO$	BrO^-	
	NH_4^+	NH_3	
	HCN	CN^-	
	HCO_3^-	CO_3^{2-}	
	H_2O_2	HO_2^-	
	HS^-	S^{2-}	
	H_2O	OH^-	
	NH_3	NH_2^-	
Weakest acids	OH^-	O^{2-}	Strongest bases

You will also note that Table 1.1 indicates that the ammonium ion, NH_4^+, is an acid. The acidity of this ion is a result of the reaction shown in the following chemical equation:

$$NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3 + H_3O^+(aq)$$

pH

The concentration of H_3O^+ ions in a solution is usually determined by measuring the pH of the solution. The pH is defined as

$$pH = -\log[H_3O^+]$$

(Ebbing/Gammon, Section 16.8). If the pH of the solution is low, the ionization equilibrium favors the right side of the chemical equation because $[H_3O^+]$ is large. However, if the pH is high, the equilibrium favors the left side because $[H_3O^+]$ is small. When two acids are compared, the stronger acid yields a solution with the lower pH.

Concept of the experiment

You will estimate or measure the pH of isomolar solutions of HCl, H_3PO_4, $HC_2H_3O_2$, NaH_2PO_4, $Al(NO_3)_3$, $Zn(NO_3)_2$, and NH_4NO_3. Many of these substances can be found in Table 1.1. You will be able to judge the strengths of these acids from the pH values of their solutions. You will also have the opportunity to estimate or measure the pH of several common substances: vinegar, carbonated water, tap water, and distilled water.

Procedure

Getting started

1. Obtain 4 large test tubes. Use a graduated cylinder to place 4 mL of distilled water in each test tube. Mark the height of the meniscus in each tube with a marking pencil. Pour out the water and dry the test tubes.
2. Obtain a small (about 1/2-inch) strip of wide-range pH paper for each solution that you will test. Handle these pieces as little as possible. Place them on a clean paper towel.
3. Obtain a small (about 1/2-inch) strip of each narrow-range pH paper that is available. Handle these pieces as little as possible. Place them on a clean paper towel. Mark the pH range for each strip on the paper towel. Replace these strips as needed.
4. If you are going to use a pH meter, obtain instructions for using it.

Doing the experiment with indicators and pH paper

1. Complete Steps 2 through 6 for each solution whose pH is to be estimated.
2. Make sure your test tubes are clean and dry.
3. Obtain 4 mL of the solution to be tested in each of the test tubes.
4. Using a clean, dry stirring rod, place 1 drop of the solution from one of the test tubes on a piece of wide-range pH paper. Do not contaminate the other unused strips. Using the result that you obtain, select the appropriate short-range paper and use it in the same manner. Record your estimate.
5. Use 3 drops of a different indicator solution for each test tube. The indicators are thymol blue, methyl orange, methyl red, and bromthymol blue (see the figure in Appendix: Indicators, pH Paper, and pH Meters for the pH range for these indicators).
6. Estimate the pH by comparing your results with the figure in Appendix: Indicators, pH Paper, and pH Meters. In some instances, you may be able to estimate to the nearest 0.1 pH unit, whereas in others you may be able to estimate only to the nearest 0.5 pH unit. Record your estimate.

Measuring the pH with a pH meter (optional)

1. Calibrate the pH meter with solutions of known pH according to the methods prescribed by your laboratory instructor.
2. Measure, and record to the nearest 0.01 pH unit, the pH of each solution.

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Date:	....	Student name:	....
Course:	....	Team members:	....
Section:	....		....
Instructor:	....		....

Prelaboratory assignment

1. Provide definitions for the following terms:

   a. Brønsted-Lowry acid

   b. Brønsted-Lowry base

   c. pH

   d. Acidic solution

   e. Basic solution

   f. Neutral solution

   g. Indicator

   h. pH paper

i. pH meter

2. a. A solution causes a yellow color with thymol blue, a yellow color with methyl orange, an orange color with methyl red, and a yellow color with bromthymol blue. The pH must lie between .... and .... The estimated pH is .... ± ....

b. A solution causes a yellow color with thymol blue, yellow colors with both methyl orange and methyl red, and a green color with bromthymol blue. The pH must lie between .... and .... The estimated pH is .... ± ....

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Results

Solution	pH from pH Paper	pH from Indicators	pH from pH Meter

Questions

1. a. Calculate the concentration of H_3O^+ ions in the most acidic solution that you examined. Use your most precise pH measurement.

b. Calculate the concentration of H_3O^+ ions in the least acidic solution that you examined. Use your most precise pH measurement.

2. a. Arrange the acids that you used in order of *decreasing* acidity. Use your most precise pH measurements. Exclude the common household substances.

   b. How does your arrangement compare with the first column in Table 1.1, insofar as a comparison can be made? Comment.

   c. Comment on the relative acidities of H_3PO_4 and the $H_2PO_4^-$ ion.

3. Give the chemical reactions that show why solutions of $Al(NO_3)_3$, $Zn(NO_3)_2$, and NH_4NO_3 have the pH values that you found.

4. Arrange the common household substances according to *decreasing* acidity.