

Acids, Bases, Salts, and Buffers

OBJECTIVE: The goal of this series of experiments is to investigate the characteristics of acidic and basic solutions. We will explore the neutralization of an acid with a base, measure the pH of a variety of common substances, demonstrate the color changes of various acid-base indicators, examine the hydrolysis reactions of salts, and analyze several buffer systems.

SKILLS: Observations, determining pH using pH paper, preparing buffer solutions

EQUIPMENT: pH paper

REFERENCES: *Chemistry: Matter and its Changes*, Brady & Senese, 5th ed, pp 606–673
Chemistry for Today: General, Organic, and Biochemistry, Seager et al, 5th ed, Experiment 12

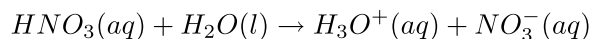
SAFETY AND DISPOSAL: Liquid samples and glassware used in this experiment present no special hazards. Disposal of all liquids in the lab sinks is appropriate; solids may be put in the trash baskets. Be careful to keep working areas clean.

INTRODUCTION: Acid/base properties of solutions are important in understanding a variety of biological and environmental systems. Arrhenius first proposed that acids dissociate to some extent in water to produce hydrogen ions (H^+) while bases liberate hydroxide ions (OH^-) in water. Brønsted and Lowry developed a more general definition of acids and bases. They defined acids as substances that donate H^+ ions (protons) and therefore are called **proton donors**. A Brønsted/Lowry base can accept H^+ ions (protons) and are defined as **proton acceptors**. Using these definitions, the autoionization of water can be written as:



In this reaction, one water molecule donates a proton and behaves as an acid while the other water molecule accepts the proton and behaves as a base.

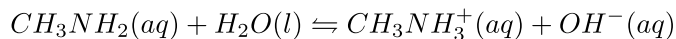
Nitric acid behaves as a typical Brønsted/Lowry acid in water solution:



The nitric acid is acting as a proton donor while the water acts as proton acceptor (base). We see that according to this reaction when nitric acid is added to water, hydronium ions (H_3O^+) are generated in solution. This causes the solution to contain more hydronium than hydroxide ions and thus the solution is acidic. Notice that because nitric acid is a strong acid this reaction is written with a single rightward facing arrow, indicating that this reaction goes completely to the

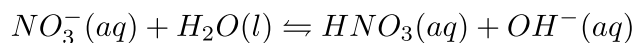
right. If we dissolve 1 mole of nitric acid in water, 1 mole of hydronium and 1 mole of nitrate ions are formed.

Similarly, we can write the reaction of a base (methylamine, CH_3NH_2) with water to show a base acting as a proton acceptor:

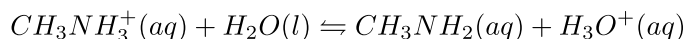


Notice that in this reaction, the double arrow is used to indicate equilibrium is established. Methylamine is a weak base and the ionization of a weak base does not occur completely. Part of the methylamine remains in solution as the unionized base. If we dissolve 1 mole of methylamine in water, less than 1 mole of each of the products are formed.

The anions from acids (such as NO_3^-) and the cations from bases (such as CH_3NH_3^+) can also behave as Brønsted/Lowry bases and acids, respectively. Thus, the nitrate ion can accept protons and behave as a base:

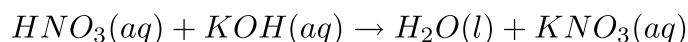


However, since HNO_3 give up protons readily (the conjugate acid to the nitrate ion is nitric acid, a **strong acid**), NO_3^- is not going to accept protons readily. Therefore, the nitrate ion does not react significantly as a Brønsted base. On the other hand, CH_3NH_3^+ can act as a Brønsted acid because its conjugate base is a weak base:



Since CH_3NH_2 does not readily accept protons (CH_3NH_2 is a weak base), CH_3NH_3^+ has the tendency to give up H^+ ions and acts as an acid. In general, the anions of weak acids behave as bases while the anions of strong acids do not. Similarly, the cations of weak bases behave as acids, while the cations of strong bases do not.

A. Neutralization Reactions: A neutralization reaction occurs when an acid and base are mixed together. The products of the reaction are water and a salt solution. For example, the neutralization of nitric acid with potassium hydroxide can be written as:



Since potassium nitrate is a water-soluble salt, the solution after the neutralization reaction contains potassium (K^+) and nitrate (NO_3^-) ions.

B. & C. pH and Indicators: The concentration of hydronium ions in a solution is often expressed in terms of the pH scale developed by Sørensen:

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

Thus the acidity of a solution can be characterized by pH, a function of the concentration of hydronium ions in water solutions. Neutral solutions have a pH of 7, acids have pH less than 7, while bases have pH greater than 7. Lower pH

means greater acidity, and vice versa. pH can be measured by using pH paper or a pH meter. Both are sensitive to the hydrogen ion concentration, but differ somewhat in sensitivity. pH paper is impregnated with indicators, substances that appear different colors depending on the pH of a solution. Indicators are often used to visually determine the end point of a titration and need to be selected in order to change color in the correct pH range.

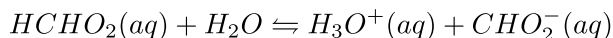
D. Hydrolysis: As discussed earlier, the presence of anions and cations of weak acids and bases can impact the acidity/alkalinity of a solution. If a soluble salt contains the cation of a strong base and the anion of a weak acid, the solution will be basic, with a pH greater than 7. Conversely, if a soluble salt contains the cation of a weak base and the anion of a strong acid, the pH of the solution will be less than 7, indicating an acidic solution.

For example, potassium nitrite (KNO_2) contains a cation of a strong base (KOH) and the anion of a weak acid (HNO_2 , nitrous acid). The nitrite ion (NO_2^-) is a significant Brønsted base and reacts with water in a **hydrolysis reaction**:

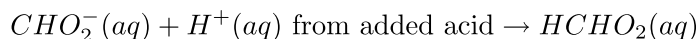


The hydrolysis reaction above produces hydroxide ions, thus the solution of the salt potassium nitrite will be basic.

E. Buffers: Buffers are solutions designed to maintain a relatively constant pH when an acid or base is added. This can be achieved through the presence in solution of a substance with the ability to donate protons and another substance with the ability to accept protons. A buffer typically consists of a weak acid and its conjugate base in solution. For example a solution of formic acid (HCHO_2) and sodium formate (NaCHO_2) can act as a buffer. In this case formic acid is the weak acid (proton donor) and the formate ion from the sodium formate acts as the conjugate base (proton acceptor).



If an acid is added to this buffer, the extra H^+ from the added acid can react with the formate ions in solution to minimize the change in pH:



Similarly, any added base will introduce OH^- ions to the solution, which can react with the formic acid to produce more formate ions, thus minimizing the change in pH. As long as both the acid and conjugate base are present in solution, the buffer will be able to withstand changes in pH with the addition of an acid or base. When one of the two species is mostly depleted, the buffer has reached its buffer capacity and the pH will begin to change significantly if more acid or base is added.

EXPERIMENTAL PROCEDURE

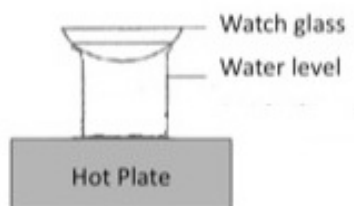
This experiment consists of five parts (A–E), which can be completed in any order. Follow the directions of your instructor as to which part you are to begin with. Solutions for each part are located in the fume hoods while solids are located next to the balances. The bottles and jars for each part are color coded, so refer to the posted guide when completing each part.

Important Notes:

1. You will need boiled distilled water for several portions of the experiment. Put about 300 mL of distilled water in your largest beaker and boil for about 5 minutes. Set this aside to cool and use as needed later.
2. Be sure to thoroughly clean your glassware/spot plates between each part of the experiment. Contamination from previous solutions can impact your measurements.

A. Acid-Base Neutralization

Begin by measuring 2.0 mL of 1.0 M sodium hydroxide (NaOH) in a clean, dry 10 mL graduated cylinder and transfer the 2.0 mL to a test tube. Rinse and dry your graduated cylinder and then put 2.0 mL of 1.0 M hydrochloric acid (HCl) into another test tube. Next, pour the contents of one of the test tubes into the other test tube and mix the contents well by repeatedly blowing air bubbles through it using one of your plastic droppers. Pour the resulting neutralized solution into a clean watch glass and evaporate to dryness over a boiling water bath as shown in the figure below.



When the evaporation is complete, record the appearance of the solid that forms on your evaporating dish. Take approximately half of the solid and dissolve it in about 20 drops of distilled water in a clean test tube. Test the dissolved solid for the presence of chloride ions (Cl^-) by adding 2 drops of 0.1 M silver nitrate solution (AgNO_3). The formation of a white, cloudy solution indicates the presence of chloride ions. Your instructor will demonstrate the flame test for the presence of sodium for the entire class.

B. pH Measurement

For this portion of the experiment you will use pH paper to measure the pH of the substances listed in the table below. A single strip of pH paper can be used to test several substances. Dispense the substances into your clean spot plate. Be sure to record the location of each substance in your notebook. If the samples are solid, dissolve about 0.1 g in 10–15 drops of boiled distilled water. If the samples are liquid, dispense about 5 drops into your spot plate.

Dip a clean glass stir rod into the liquid being tested and then touch the rod to the pH paper. Compare the color of the spot on the paper to the chart that is located at your station. Be sure to clean your glass stir rod between each test. Rinsing with distilled water and drying with a paper towel is appropriate. One strip of pH paper of about 5 cm in length should be enough for all tests done in this part.

NOTE: Test the household ammonia in a clean test tube by itself. The vapors from the ammonia will cause the paper to change color and give incorrect readings for other solutions. Therefore, be sure to test the ammonia by itself in a clean test tube. Ammonium hydroxide is another name for aqueous ammonia or simply ammonia.

Substances for pH measurement in Part B		
Lemon Juice	Ammonia (see note above)	Baking soda
Orange juice	Laundry detergent	Aspirin
Milk	Dishwashing detergent	Buffered aspirin

C. Acid-Base Indicators

Clean and dry your spot plate before beginning this part of the experiment. Place 10 drops of 0.05 M hydrochloric acid into five wells of your spot plate. Repeat this procedure in a different set of five wells using 10 drops of 0.05 M sodium hydroxide. For each indicator listed below, add 1 drop of the indicator to one of your HCl samples and 1 drop to one of your NaOH samples. Be sure to record which solution and indicator is being added to each well. Observe the color of the resulting solution for each case and record your observations in your notebook.

Indicators for Part C
Methyl red
Bromcresol green
Phenolphthalein
Methyl orange
Methyl violet

D. Behavior of Salts in Solution: Hydrolysis

For this part of the experiment you will need 5 clean test tubes. The test tubes do not need to be dry. First, place approximately 0.1 g samples of the following salts into separate clean test tubes:

- Sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$)
- Sodium carbonate (Na_2CO_3)
- Ammonium chloride (NH_4Cl)
- Sodium chloride (NaCl)

Next, add approximately 5 mL of boiled distilled water to each sample and agitate until the solid dissolves. Finally, measure the pH of each solution using pH paper. To measure the pH, dip a clean glass stir rod into each test tube and then touch the rod to the pH paper. Be sure to clean the stir rod between each measurement. For comparison, measure the pH of the boiled distilled water.

E. Buffers

For this part of the experiment you will be preparing 4 solutions, A through D. For each solution, mix the required amounts of the stock solutions in clean test tubes.

Solution A	20 drops of 0.1 M sodium acetate solution ($\text{NaC}_2\text{H}_3\text{O}_2$) 20 drops 0.1 M acetic acid solution ($\text{HC}_2\text{H}_3\text{O}_2$)
Solution B	20 drops of 0.1 M aqueous ammonia solution (NH_4OH) 20 drops of 0.1 M ammonium chloride (NH_4Cl)
Solution C	20 drops of 0.1 M sodium acetate solution ($\text{NaC}_2\text{H}_3\text{O}_2$) 20 drops boiled distilled water
Solution D	20 drops 0.1 M sodium chloride solution (NaCl) 20 drops boiled distilled water

Be sure to mix each solution well by using a clean dropper to bubble air through it. Next, pour half of each solution into another clean test tube so you have two equal samples of each solution. Use pH paper to measure and record the initial pH of one sample of each solution. For comparison, also measure the pH of a sample of boiled distilled water.

To test the impact of adding a strong acid, add 8 drops of 0.05 M hydrochloric acid (HCl) to one sample of each solution and to a 20-drop sample of boiled distilled water. Mix each well and then record the resulting pH's. Repeat this procedure to test the effect of adding a strong base. To the other sample of each solution add 8 drops of 0.05 M sodium hydroxide (NaOH), mix well, and record the resulting pH. Measure the pH of a distilled water sample plus NaOH as well.

Notebook Template: You should record all of your observations and measured pH's in your notebook. It is advisable to set of a series of tables to help you organize your data. When using your spot plate, be sure to record what you are adding to each well. Each well in the spot plate is numbered. When using test tubes, use some method of labeling in order to record what solutions you have added to each test tube. It is also a good idea to try and write the associated chemical reactions when appropriate.