UNIT 10

Acids and Bases

The Voyage of the Proton

Unit Overview

This unit introduces the key concepts of acids and bases. The acidity of a solution is a measurement of the concentration of hydrogen ions it contains, and the acidity has a great impact on the properties of the solution. Pure water contains a certain concentration of hydrogen ions. Dissolving acids in water raises the concentration of hydrogen ions, while dissolving bases in water lowers the concentration. Acidity is measured in pH units, and the acidity of a solution depends both on the concentration of acid dissolved and the strength (degree of dissociation) of the acid. Chemists use indicators to measure pH as well as titration to determine the concentration and strength of an acid.

Learning Objectives and Applicable Standards

Participants should be able to:

1. Explain the dissociation of water molecules into hydronium and hydroxide ions and explain the equilibrium constant (K_w) for that reaction.
2. Define acids and bases using the Arrhenius and Brønsted-Lowry (conjugate acid/base pairs) definitions of acids and bases.
3. Explain acid dissociation constants (K_a). Define the pK_a of an acid and understand its relationship to the acid's strength.
4. Write the chemical equation for neutralization reactions.
5. Calculate the pH of a solution.
6. Explain the titration curves of monoprotic acids.
7. Define acid/base **indicators** and explain their uses.
8. Define **buffer** and explain how one works.

Key Concepts and People

1. **Acids and Bases:** An Arrhenius acid is defined as a molecule that when dissolved in water, increases the concentration of H^+ ions. An Arrhenius base is a molecule that

dissolves in water to produce a hydroxide ion, thereby increasing the concentration of hydroxide ions in aqueous solution.

2. **The Dissociation of Water:** Water molecules exist in equilibrium with hydronium ions and hydroxide ions. Arrhenius acids and bases form water when they neutralize each other.

3. **The pH Scale:** The acidity of a solution is measured in pH. A solution's pH indicates the concentration of hydronium ions. The pH scale is a logarithmic scale running from zero to 14. Anything with a pH below seven is considered an acid, and anything with a pH above seven is considered a base. Solutions with a pH of seven are neutral.

4. **Strengths of Acids and Bases:** Acids and bases come in different strengths. The stronger the acid or base, the more it dissociates in aqueous solutions. The dissociation constant (K_a) reflects the extent of dissociation.

5. **Measuring pH:** pH meters use electrochemistry to measure the concentration of hydronium ions (H_3O^+) and thus can provide the pH of solutions. Indicators change color at certain pH levels, so chemists use them to estimate the pH of solutions

6. **Neutralization and Acid Base Titrations:** In general, when acids and bases react together, they neutralize each other, and the pH of the solution gets closer to seven. A titration curve charts the pH of a neutralization reaction as the acid and base slowly combine. Titration curves are used to determine the concentration of an unknown and to find pK_a values.

7. **Conjugates of Acids and Bases:** A Brønsted-Lowry acid is anything that can donate a proton (H^+), and a Brønsted-Lowry base is anything that can accept a proton. When the proton is dissociated from an acid, the remaining molecule is then called the conjugate base of the acid. Similarly, once a proton is picked up from a base, the molecule is now called a conjugate acid.

8. **Buffers:** Indicators and buffers are examples of shifting equilibria. Buffers stabilize pH levels at a certain value.

Video

Acids and bases are found all around us, and the currency of acid-base chemistry is the proton, or hydrogen ion. Acid-base chemistry is part of everyday life, from baking and the food we eat to the innumerable reactions that keep the human body alive. Acid-base chemistry is measured on the pH scale—the concentration of hydrogen ions in a solution. Buffers can control pH, whether used in the lab or in the acid-base components of human blood. The role of acids and bases will be shown in food—from the rise of a cake to the making of cheese. In the environment, acid rain plagues industrial portions of the world; the chemical nature of acid rain reactions and the environmental response and impact are part of acid-base chemistry.

VIDEO CONTENT

Host Introduction
"What is an Acid, What is a Base?"

Dr. Adam Brunet, a chemistry professor at American International College, explains that water is both an acid and a base. With the help of the Boston University synchronized swimming team, he explains the process in which water dissociates into acid (hydronium) and base (hydroxide). He also introduces the logarithmic pH scale.

Laboratory Demonstration
"Corrosive Acids and Bases"

Harvard University Lecture Demonstrator Daniel Rosenberg makes two solutions. One is a strong acid—HCl, and the other is a strong base—NaOH. Both solutions are highly corrosive and Daniel shows this by dissolving metal in each. He then neutralizes the solutions by pouring them together.

Host Science Explanation
"A Cheesy Weak Solution"

Dr. Adam Brunet demonstrates weak acids and weak bases by combining some milk and vinegar to make cheese. He does a quick lesson on partial dissociation, and defines weak acids as partially dissociated.

Real World Application
"Baked Reactions"

Joanne Chang of the famed Flour bakeries in Boston talks about the acid-base chemistry involved in baking. She demonstrates what can happen to a cake if a critical basic ingredient (baking powder) is left out.

Host Science Explanation
"Acidic Pond"

Dr. Adam Brunet travels to Little Pond in Vermont to discuss the effects of acid rain on a pond ecosystem. He describes the acidity of natural rainwater and how pollutants have created more acidic rain that is detrimental to Little Pond. He also introduces the idea that the rocks around the pond can act as a chemical buffer and help offset the effects of acid rain.

Laboratory Demonstration
"Buffered Lemonade"

Harvard University Lecture Demonstrator Daniel Rosenberg makes two solutions. One is pure water and the other is water with equal amounts of weak acid and weak base. When lemonade is introduced into the pure water solution, it instantly turns acidic from the citric acid in the lemonade. When the lemonade is added to the buffered solution with a weak acid and weak base, the pH doesn't change substantially. The buffered solution is able to take on the added acid without becoming acidic.

Real World Application
"Removing Acids from the Body"

Dr. Robert Stanton of the Joslin Diabetes Center in Boston explains how the human body makes acids during normal metabolic processes, and also how the body removes those acids. The lungs expire CO_2 to remove volatile (gaseous) acid while the kidney is responsible for the removal of non-volatile acids. In both processes the blood contains buffers that allow acids to be stored and moved throughout the body until they can be expelled.

Unit Text

Content Overview

Previous units have discussed the properties of solutions, the general properties of chemical reactions, and equilibrium. This unit examines acid-base reactions, which are equilibrium reactions that occur in aqueous solutions and involve the release of hydrogen ions by a kind of molecule called an acid. The text begins by defining acids and bases and then explains how the acidity of a solution is measured using the pH scale. It then describes what happens when acids and bases combine. Additionally, the unit defines the difference between a strong and weak acid and shows how titration analysis can be used to determine the amount of acid in a given solution. Finally, the text ends by discussing buffered solutions. This unit provides the foundational principles behind acid-base chemistry.

Sidebar Content

1. **Svante Arrhenius:** This Swedish chemist created a definition of acids and bases still used today.

2. **Calculating pH:** This sidebar describes the origins of the pH scale and how the pH of a solution is calculated.

3. **Cooking Fish With Citrus Juice – Ceviche:** The acids in citric juice denature the proteins in fish and make the fish appear cooked.

4. **The Dissociation Constant of Water:** The dissociation of water is an equilibrium reaction, with an equilibrium constant called K_w.

5. **Color Changing Flowers:** Many plants contain various forms of anthocyanins and respond to changing pH differently, so they can be used as indicators in a broad range of solutions of different acidities. For example, hydrangea flowers actually change color depending on the acidity of the soil they're growing in. In acidic soils they are blue and in basic soils they appear pink.

6. **Acid Rain:** The release of sulfur dioxide by burning coal leads to acid rain and widespread environmental damage.

7. **Amphoteric Compounds:** Amphoteric compounds are molecules that can act as an acid or a base depending on the situation.

8. **Resonance and Acidity:** Carboxylic acids are common in organic and biochemistry and they can form resonance structures and give up a proton easily.

9. **Acidosis and the Body:** Changes in the pH of blood can lead to problems such as rickets and the bowing of legs.

Interactives

Historical Timeline of Chemistry

This interactive illustrates how different discoveries build upon, disprove, or reinforce previous theories. This not only reinforces basic chemistry concepts, but also emphasizes the nature of science. Scientists mentioned in this unit are also listed on the timeline.

Chemistry of Running Interactive

In this interactive, students will see the effect of lactic acid build-up during strenuous exercise. With a decrease in pH, muscle enzymes fail to work properly and muscles don't contract well. The interactive also shows that blood acts as a buffer to prevent dangerous changes in pH during exercise.

During the Session

Before Facilitating this Unit

The "Acid and Base Safety" demonstration is a dramatic introduction to acid-base chemistry. However, it is also important that students realize that acids and bases are extremely common chemicals, and just because something is an acid or a base does not make it dangerous. The "Strong vs. Weak" demonstration introduces the concept of how both concentration and degree of dissociation influence pH, while the lab activity "Indicators and pH" reinforces these concepts and requires them to calculate pH using K_a values.

Tips and Suggestions

1. **Chemically speaking, acid strength (the K_a value) is different from concentration.** A highly concentrated solution of a weak acid may have a lower pH than a very dilute solution of a strong acid.

2. **Indicators and Buffers are applications of Le Chatelier's principle.** At this point in most chemistry courses, students have learned about chemical equilibrium.

It is helpful to refer them back to Le Chatelier's principle to help them understand how indicators and buffers work.

3. **Neutralization reactions always go to completion.** In advanced classes many problems ask students to calculate the pH of a solution made by mixing an acid and a base. Emphasize that no matter what the strength of the acid and base, they should assume the neutralization reaction goes to completion. Then they should determine what dominant chemical species remain and calculate the pH from there.

Starting the Session: Checking Prior Thinking

You might assign students a short writing assignment based on the following questions, and then spend some time discussing prior thinking. This will help elicit prior thinking and misconceptions. Many students will have heard the terms associated with acids and bases and have inaccurate notions about them.

1. What comes to mind when you hear the term **acid** and **base**?
2. Are acids and bases always dangerous?
3. How are acids and bases used in the home? In industry?
4. Where do you find acids and bases in "real life"?

Before/While Watching the Video

1. What is a hydronium ion? Is it a proton donor or proton acceptor?
2. What is a hydroxide ion? Is it a proton donor or proton acceptor?
3. What is dissociation?
4. How can you make a neutral solution?
5. What is the definition of a strong acid? What is the definition of a weak acid? Name an example of each.
6. What is acid rain?

Watch the Video

After Watching the Video

1. What is an acid?
2. What is a base?
3. What does the pH scale represent?
4. What is a buffer? How can a buffer be made?

5. Provide some examples of the roles acids and bases play in every day life.

Group Learning Activities

Buffer Activity

Objective

This activity is designed to help students understand how buffers work and what they are used for. Students compare how many drops it takes for a "buffered" solution to change pH compared to a solution that is not buffered. Universal indicator is a solution that contains multiple indicators. It can be used to estimate the pH of a solution as follows:

Color	red	orange	yellow	green	blue	indigo	violet
pH	4	5	6	7	8	9	10

List of Materials

Each group requires:

- An Erlenmeyer flask
- Universal indicator
- 10 mL 0.1 M HCl in dropper bottles
- 10 mL 0.1 M NaOH in dropper bottles
- 5 mL of 0.1 M Na_2HPO_4 in dropper bottles
- 5 mL 0.1 M NaH_2PO_4 in dropper bottles

Procedure (Part I)

1. In a small Erlenmeyer flask, pour approximately 10 mL of distilled water. Add four drops of universal indicator. The solution should be yellow to green (about pH 5 to 7).

2. Drop by drop, add 0.1 M HCl to the flask. Record how many drops it takes to decrease the pH by one unit.

3. Rinse out the flask and repeat the procedure with 0.1 M NaOH. Record how many drops it takes to increase the pH by one unit.

Procedure (Part II)

1. Using a pipette bulb and pipette, put 5 mL of 0.1 M NaH_2PO_4 solution into a small Erlenmeyer flask. To the same flask, add 5 mL of 0.1 M Na_2HPO_4. Add four drops of universal indicator. The solution should be green (about pH 7).

2. Drop by drop, add 0.1 M HCl to the flask. Record how many drops it takes to decrease the pH by one unit.

3. Rinse out the flask and repeat the procedure using 0.1 M NaOH instead of HCl. Record

how many drops it takes to increase the pH by one unit.

Discussion

The following chart can help students organize how they record their data.

Part I		**Part II**	
drops HCl	drops NaOH	drops HCl	drops NaOH

If the lab went well, you should have found that the solution in part II resisted the change in pH much more than the solution in part I. In other words, it should have taken many more drops of HCl or NaOH to change the pH in part II.

In part II, the solution contained the following weak acid equilibrium:

$$HA \ \square\ H^+ + A^-$$

A solution that contains this kind of equilibrium is called **buffered**. The following questions can help guide students thinking while doing this activity.

1. Compare part I and part II. Why was it so easy to change the pH in part I, but so hard in part II?
2. Would a solution that contained predominantly HA be an effective buffer? Explain.
3. Would a solution that contained predominantly A- be an effective buffer? Explain.
4. A similar solution to part II could be made from HCl and Cl^- ions. Would this be an effective buffer? Why or why not?

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, lab coat and gloves.

Universal indicator: Alcohol-based solution; flammable liquid.

Hydrochloric acid: Toxic by ingestion or inhalation; severely corrosive to skin and eyes.

Sodium hydroxide: Corrosive liquid; skin burns are possible; very dangerous to eyes; wear gloves.

Sodium phosphate, dibasic: Body tissue irritant. Avoid all body tissue contact.

Sodium phosphate, monobasic: Body tissue irritant. Avoid all body tissue contact.

Disposal

Check local regulations for the proper disposal of chemicals used in this activity. In general, the phosphate solutions may be disposed of in the sink, however, pH restrictions usually prevent the disposal of NaOH and HCl down the drain without proper neutralization.

Indicators and pH

Objective

Students will understand how indicators work based on shifts in equilibrium, and how to use indicators to estimate the pH of a solution.

List of Materials

- 0.02 M HCl
- 0.02 M $HC_2H_3O_2$
- Distilled water
- 0.02 M NH_3
- 0.02 M NaOH
- Turmeric extract (ground turmeric dissolved in ethanol)
- Methyl red
- Thymolphthalein
- Well plate

Procedure

1. Pre-lab Calculations**:** Calculate the pH of the following solutions. The K_a of $HC_2H_3O_2$ is 1.8×10^{-5}; the K_b of NH_3 is 1.8×10^{-5}.

0.02 M HCl	0.02 M $HC_2H_3O_2$	H_2O	0.02 M NH_3	0.02 M NaOH

2. Obtain a well plate and fill it with two drops of each solution as follows. Record the colors in each well.

	HCl	$HC_2H_3O_2$	H_2O	NH_3	NaOH
turmeric extract					
methyl red					
thymolphthalein					

Discussion

The following questions will help students think about this activity.

1. At approximately what pH do the indicators change color?
2. All substances that behave as indicators are weak acids. How do indicators work? Why do they change color at different pH levels? (Think about Le Chatelier's principle.)
3. If you had a wide range of different indicators, how could you use them to determine the pH of an unknown solution?

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, lab coat and gloves.

Hydrochloric acid: toxic by ingestion or inhalation and severely corrosive to skin and eyes.

Acetic acid: Corrosive liquid.

Ammonia: Both liquid and vapor are extremely irritating—especially to eyes. Dispense in a hood and be sure an eyewash is accessible. Moderately toxic by ingestion and inhalation. Serious respiratory hazard. LD50 350 mg/kg.

Sodium hydroxide: Corrosive liquid; skin burns are possible; very dangerous to eyes; wear gloves.

Ethanol: Dangerous fire risk; flammable; addition of denaturant makes the product poisonous-it cannot be made nonpoisonous. TLV 1880 mg/m^3.

Methyl red: Substance is not considered hazardous. However, not all health aspects of this substance have been thoroughly investigated.

Thymolphthalein: Alcohol solution; flammable liquid.

Disposal

Check local regulations for the proper disposal of chemicals. In general, the ammonia, ethanol, and methyl red solutions may be disposed of in the sink, however, pH restrictions usually prevent the disposal of NaOH and HCl down the drain without proper neutralization. Thymolphthalein must be disposed of by a licensed hazardous waste company.

Red Cabbage Indicator, the pH of Liquids, and Antacid Titration

Objective

The compound responsible for red cabbage's color is sensitive to changes in pH, and can be extracted in boiling water for use as a pH indicator. This homemade indicator changes from red/pink in acidic solutions to green in basic solutions. The acidity or basicity of various household products can be determined using the indicator; and, if desired, compared against the colors of the indicator in solutions of known pH. A rough titration analysis can also be performed between antacids and vinegar using the indicator to find the approximate endpoint.

List of Materials

- Red cabbage, chopped
- Water to boil cabbage, plus distilled water for dilution of solutions
- Stockpot, strainer and bowl/jar
- Filter paper, if desired, to saturate with the indicator and make pH papers
- 50-100 mL each of lemon juice, vinegar, bleach, ammonia, other kitchen/cleaning products, other beverages as desired.
- Baking soda
- Antacid tablets, 1-2 per group
- If desired: 1.0 M NaOH and 1.0 M HCl to generate solutions of known pH for color comparison.
- Disposable pipets, one per group for transfer of indicator plus more for dilution of HCl/NaOH
- Large test tubes or 50 mL beakers, one per liquid to be tested per group, plus 15 for a complete set of pH standards.

Set Up

1. Before the session, chop the cabbage, cover with boiling water, and allow to come to room temperature. Strain the liquid.

2. If desired, soak strips of filter paper in the liquid and allow to dry for pH paper.

Procedure: Making a set of pH standards:

1. Using a graduated disposable pipet, transfer 1 mL of 1.0 M HCl or NaOH into a test tube or beaker and add 9 mL of distilled water. Mix well.

2. Transfer 1.0 mL of this diluted solution into another test tube, and add 9mL of distilled water. Mix well.

3. Repeat this procedure five more times, then repeat entire procedure for the other compound. Each tenfold dilution will produce a solution with a pH approximately one unit different from the previous test tube. 1.0 M HCl (pH 0) can be diluted down to pH ~7, as can 1.0 M NaOH (pH 14).

4. Label each test tube with the approximate pH, and place a few drops of the indicator in each tube to show the color for that pH.

Procedure: Testing the pH of household liquids or beverages:

1. If testing baking soda or antacids, grind antacid tablets or clumps of baking soda into a fine powder and dissolve in a few milliliters of distilled water.

2. Pour a few milliliters of each liquid into a clean test tube or beaker.

3. Add a few drops of indicator to each liquid, and record the color and the approximate pH (or whether the liquid is acidic or basic).

4. If using strips of pH paper, a drop of each liquid may be placed on a strip of the paper instead of filling a test tube and adding indicator.

Procedure: "Titrating" an antacid tablet

1. Grind an antacid tablet into a fine powder and dissolve in water.

2. Add a few drops of indicator to the beaker or test tube.

3. Using a graduated disposable pipet, add vinegar a few drops at a time and stir.

4. Add vinegar a drop at a time and record the volume added until the indicator turns pink/red (showing the presence of excess acid).

Discussion

These questions can help guide students thinking during and after the activity:

1. Why did you make a pH standard scale? What information does that scale provide? Calculate the concentration of acid and base for each dilution. How does this reflect a logarithmic scale?

2. Which household liquids were acidic? Which ones were basic? Which ones appeared close to a pH of 7? How did your results compare with your expectations?

3. What information does the volume of vinegar needed to titrate the antacid tell you?

4. Vinegar is diluted acetic acid (CH_3COOH, one ionizable proton). If the main component

of antacid tablets is $CaCO_3$, write the balanced equation for the reaction.

5. Vinegar is typically 5% w/v, meaning that every 100 mL of solution contains 5 g of acetic acid. Look at the mass of $CaCO_3$ per tablet as listed on the antacid bottle, and use your equation above to determine how many milliliters of vinegar should be required to react completely with one tablet. How well does this calculation agree with your experimental results?

6. All substances that behave as indicators are weak acids. How do indicators work? Why do they change color at different pH levels? (Think about Le Chatelier's principle.)

Hazards
It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear eye protection and gloves when preparing and performing this experiment. DO NOT mix ammonia or acids with bleach.

Disposal
Beverages, vinegar, and antacids may be poured down the sink. Check local regulations for proper disposal of chemicals and household cleaning products. As mentioned in the hazards, do not mix household cleaning chemicals. In general, pH restrictions usually prevent the disposal of NaOH and HCl down the drain without proper neutralization.

In-Class Chemical Demonstrations

Acid and Base Safety

Crack a raw egg into a petri dish on an overhead projector. Explain that human skin and eyes contain a lot of protein, just like the egg white. Add a few drops of 6M HCl to the egg white; the acid will "cook" the egg white and make it opaque. Add some drops of baking soda solution to demonstrate that the damage done by the acid is irreversible. Put a few drops of 6M NaOH solution on another region of the egg white to show that bases cause just as much damage.

Hazards
It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them as well. Wear proper protective equipment at all times: lab coat, chemical apron, splash goggles, and gloves.

Hydrochloric acid: Toxic by ingestion or inhalation; severely corrosive to skin and eyes,

Sodium hydroxide: Corrosive liquid; skin burns are possible; very dangerous to eyes; wear gloves.

Disposal
Check local regulations for proper disposal of chemicals. In general pH restrictions usually

prevent the disposal of NaOH and HCl down the drain without proper neutralization.

Strong vs. Weak Acids

Using litmus paper, indicators, or a pH meter, measure the pH of four solutions: 0.02 M HCl, 0.02 M $HC_2H_3O_2$, 0.02 M NH_3, and 0.02 M NaOH. Emphasize that the concentrations of all the solutions are the same, and ask the students to explain why the weak acid and weak base do not produce such large changes in pH.

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them as well.

Disposal

Check local regulations for proper disposal of chemicals.

Going Deeper (In-Class Discussion or Reflection)

Instructors should allow up to 30 minutes for discussion at the end of the session, or students can use the time to reflect on one or more of these questions in journals.

1. Why do we use logarithmic units like the pH scale? Can you name another common logarithmic unit?
2. Would it be better to neutralize an acid spill with a weak base like baking soda or a strong base like sodium hydroxide? Why?
3. Buffers have limitations; if you add enough acid to a buffer, the pH will eventually change. Why? What determines how much acid the buffer can handle?
4. If we combine equal numbers of moles of a strong acid and a weak base, the pH of the resulting solution is neutral. If we combine equal numbers of moles of a *weak* acid and a weak base, the pH of the resulting solution is basic. Why?
5. Why do indicator solutions change color gradually, instead of having a sharp cutoff?

Before the Next Unit

Students should read the Unit 10 text if they haven't already done so. They may be assigned one or more activities from the list below, or if you choose to have them use the course materials outside of class, they can watch the Unit 11 video and/or read the Unit 11 text as an assignment before the next unit.

References and Additional Resources

Phet Interactive Simulations. "pH Scale." University of Colorado at Boulder. Accessed May 22, 2013. http://phet.colorado.edu/en/simulation/ph-scale

Phet Interactive Simulations. "Acid-Base Solutions" University of Colorado at Boulder. Accessed May 22, 2013. http://phet.colorado.edu/en/simulation/acid-base-solutions

For Professional Development

In addition to watching the videos, reading the text, and going through the activities listed in the course guide, participants taking this course for professional development should read the following papers and answer the corresponding reflection questions. Participants should then complete the accompanying professional development assignments.

Further Reading and Reflection Questions

Sheppard, Keith. "High School Students' Understanding of Titrations and Related Acid-Base Phenomena." *Chemistry Education Research and Practice*.7: 1 (2006): 32-45. Accessed August 13, 2013. http://www.rsc.org/images/Sheppardpaper_tcm18-46455.pdf

1. Were you surprised by the students' ideas about pH and neutralization that this study uncovered? Why or why not? What did you find most surprising and why?
2. Have you encountered students with similar understandings/ misunderstandings about pH and/or neutralization? How have you addressed those ideas in the past?
3. Were you surprised by the students' explanations of the sections of the titration curve? Why or why not? What did you find most surprising and why?
4. Do you agree with any of the author's conclusions presented in this paper? Why or why not?
5. Has this study influenced how you will approach teaching acid-base chemistry? If so, how? If not, why not?

Dreschler, Michal and Hans-Jurgen Schmidt. "Textbooks' and Teachers' Understanding of Acid-Base Models Used in Chemistry Teaching. *Chemistry Education Research and Practice*. 6:1 (2005): 19-35. Accessed August 13, 2013. http://www.rsc.org/images/p2_drechsler_tcm18-31134.pdf

1. How have you traditionally presented acid-base models when teaching about acid-base reactions? Do you find that students tend to prefer one model over the other? Have you encountered students that are confused by when to invoke different models of acids and bases? How could you reconcile that confusion?
2. What are the similarities and differences between the presentations of acid base chem-

istry in the textbooks that were analyzed in this study compared to the textbooks you use? Do the textbooks you use influence how you present this material? If so, how?

3. Do you agree with the authors' recommendations for how to improve teaching of acid-base topics in chemistry? If so, which ones and why? If not, why not? Do you foresee any difficulties or challenges to implementing these recommendations?

4. Has this paper influenced how you will approach teaching acid-base chemistry? If so, how? If not, why not?

Sisovic, Dragica and Snezanna Bojovic. "Approaching the Concepts of Acids and Bases by Cooperative Learning." *Chemistry Education: Research and Practice in Europe*. 1:2 (2000): 263-2745. Accessed August 13, 2013. http://www.uoi.gr/cerp/2000_May/pdf/34-07sisovic.pdf

1. Were you surprised by the results of this study? Why or why not? What was particularly surprising to you?

2. Does this study influence how you will approach teaching acid-base chemistry or chemistry in general? If so, how? If not, why not?

Professional Development Assignments

1. After reading the papers above and reflecting on the questions presented develop a lesson plan designed to teach material presented in this unit.

2. Using a group activity or classroom demonstration presented in this course guide, show how you would implement it into your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?