Oakland Schools Chemistry Resource Unit

Acid-Base Concepts

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<u>Acid-Base Concepts</u>

Content Statements:

C5.7

Acids and bases are important classes of chemicals that are recognized by easily observed properties in the laboratory. Acids and bases will neutralize each other. Acid formulas usually begin with hydrogen, and base formulas are a metal with a hydroxide ion. As the pH decreases, a solution becomes more acidic. A difference of one pH unit is a factor of 10 in hydrogen ion concentration.

C5.7x

Chemical reactions are classified according to the fundamental sub molecular changes that occur. Reactions that involve proton transfer are known as acid/base reactions.

Content Expectations:

C5.7A - Recognize formulas for common inorganic acids, carboxylic acids, and bases formed from families I and II.

C5.7B - Predict products of an acid-base neutralization.

C5.7C - Describe tests that can be used to distinguish an acid from a base.

C5.7D - Classify various solutions as acidic or basic, given their pH.

C5.7E - Explain why lakes with limestone or calcium carbonate experience less adverse effects from acid rain than lakes with granite beds.

C5.7f - Write balanced chemical equations for reactions between acids and bases and perform calculations with balanced equations.

C5.7g - Calculate the pH from the hydronium ion or hydroxide ion concentration.

C5.7h - Explain why sulfur oxides and nitrogen oxides contribute to acid rain.

C5.r7i - Identify the Bronsted-Lowry conjugate acid-base pairs in an equation.

Instructional Background Information:

Common acids and bases

Examples of inorganic acids are limited to HCl, HBr, and HI. Common oxyacids are limited to H_2SO_4 and HNO_3, carboxylic acids are limited to H_2CO_3 and CH_3COOH, and bases should be limited to hydroxides of alkali and alkaline earth metals. A Bronsted acid can donate a proton and a Bronsted base can accept a proton. For every Bronsted acid there is a conjugate base and vice versa.

Ex: $NH_3 + HF \leftrightarrow NH_4^+ + F^-$
Base acid acid base

EX: $CN^- + H_2O \leftrightarrow HCN + OH^-$
Base acid acid base

Neutralization

Various examples of neutralization can be used to illustrate that a salt and water are products of the reaction.

Example: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Salt – ionic compound w/ cation other than H^+ and anion other than OH^-; strong electrolyte

Testing for the presence of acids and bases

Indicators can be limited to litmus, phenolphthalein, and universal indicator (aqueous or paper). Use a color chart for universal indicators. Everyday examples of acidic or basic foods can be used to identify acids and bases due to their sour or bitter taste. Also, since acids react with metals to produce hydrogen gas, unknown samples may be tested by combining them with metals. pH testing should be used to classify solutions as acidic basic or neutral.

Summary of properties:

Acids:
Sour taste
Change color of indicators
React w/ metals to produce hydrogen
Conduct electricity
React w/ carbonates to produce CO_2

Bases:
Bitter taste
Feel slippery
Conduct electricity
Change color of indicators

Many household cleaners are acidic or basic. Examples of these are soaps, shampoos, window and toilet bowl cleaners, vinegar and drain cleaners. Foods such as soda, antacids, vinegar, and salad dressings are acidic or basic and many food processing techniques adhere to strict pH ranges. Indicators can be used to test swimming pools, and some fruits and vegetables may be used as indicators. Plants such as hydrangeas bloom blue in acidic soil and pink in alkaline soil.

Performing calculations with balanced equations

When calculating the pH of a solution given the hydrogen ion concentration $[H^+]$ (otherwise called hydronium ion concentration), or the hydroxide ion concentration of a strong acid or base, the following may be used in a conceptual course:

pH – A measure of acidity (derived to make a number scale consisting of numbers between zero and 14)

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

pH < 7 acidic
pH > 7 basic
pH = 7 neutral

<u>EX:</u> calculate the pH of a HNO_3 solution
Whose H^+ conc. is 0.76 M

$pH = -\log\ [H^+]$
$pH = -\log[0.76]$
$pH = 0.12$

<u>EX:</u> The pH of a juice is 3.33, calculate the H^+ conc.

$pH = -\log[H^+]$
$3.33 = -\log[H^+]$
$10^{-3.33} = [H^+]$
$4.68 \times 10^{-4}\ M = [H^+]$

$pOH = -\log[OH^-]$

$4 = pOH + pH$

$14 - pOH = pH$

The ion product constant of water may be used to find the pH of a solution. The value of the hydrogen ion concentration indicates the acidity and basicity of the solution. K_w (the ion product constant for water) is used to calculate either $[H^+]$ or $[OH^-]$ when the other is known. The pH can then be calculated using the $[H^+]$.

$Kw = [H^+]\ [OH^-]$

In pure water at 25ºC $[H^+]$ & $[OH^-]$ are equal to each other (1.0×10^{-7} M)

so $Kw = 1.0 \times 10^{-14}$ M

EX: Calculate the H^+ conc. in a solution whose OH^- conc. is 1.3 M

$Kw = [H^+][OH^-]$

$[H^+] = \frac{Kw}{[OH^-]}$

$= \frac{1.0 \times 10^{-14}}{1.3\ M}$

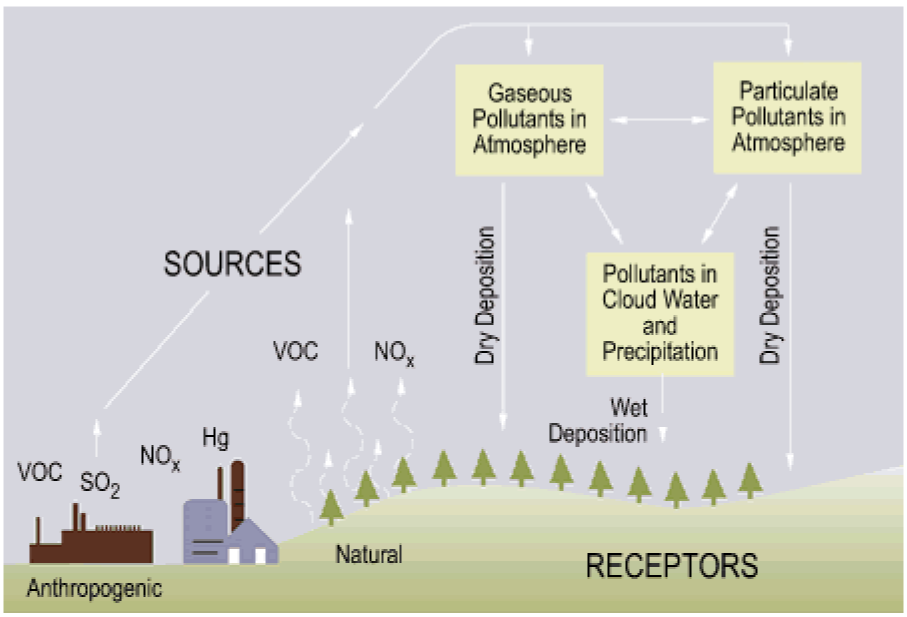
$= 7.7 \times 10^{-15}$

$pH = -\log[7.7 \times 10^{-15}]$

pH = 14

Acid Rain

"Acid rain" is a broad term referring to a mixture of wet and dry deposition (deposited material) from the atmosphere containing higher than normal amounts of nitric and sulfuric acids. The precursors, or chemical forerunners, of acid rain formation result from both natural sources, such as volcanoes and decaying vegetation, and man-made sources, primarily emissions of sulfur dioxide (SO_2) and nitrogen oxides (NO_x) resulting from fossil fuel combustion. In the United States, roughly 2/3 of all SO_2 and 1/4 of all NO_x come from electric power generation that relies on burning fossil fuels, like coal. Acid rain occurs when these gases react in the atmosphere with water, oxygen, and other chemicals to form various acidic compounds. The result is a mild solution of sulfuric acid and nitric acid. When sulfur dioxide and nitrogen oxides are released from power plants and other sources, prevailing winds blow these compounds across state and national borders, sometimes over hundreds of miles.



Terms and Concepts

Acid Rain	Acid/Base reaction	Acidic
Alkaline	Basic	Bronsted-Lowry
Carboxyl group	Hydrogen Ion	Hydronum Ion
Hydroxide Ion	K_w	Neutral
Neutralize	pH	

Instructional Resources

Environmental Protection Agency
http://www.epa.gov/acidrain/what/index.html

Raymond Chang. General Chemistry: The Essential Concepts. New York, McGraw Hill. 2006.

Process Oriented Guided Inquiry Learning
www.pogil.org

Michigan Department of Education
www.michigan.gov/mde HSSCE Companion Document

Flinn Scientific
www.flinnsci.com

Acid-Base Concepts

Activity #1 – Intro to Acids and Bases

Questions to be investigated

How can we identify an unknown solution as an acid or base?

Objectives

To identify unknown solutions as acids or bases using pH values.

Recognize formulas for common inorganic acids, carboxylic acids, and bases formed from families I and II.

Predict products of an acid-base neutralization.

Materials

8 small beakers or cups
8 unknown solutions such as lemon juice, bleach, liquid soap, ammonia, soda water, tap water, etc.
pH paper

Safety Concerns

Some unknowns can be dangerous. Wear goggles. Congo red indicator can be a skin and body irritant. Wear gloves when doing the demo.

Real-World Connections

Many students are not aware that the household substances they use are acids or bases. This lab allows them to examine the different substances and determine whether they are acidic, basic, or neutral.

Teacher Notes

Prior to doing this lab, students may observe the indicator sponge demo. This is a discrepant event demo to stimulate a discussion and have students develop a hypothesis about what is happening with the sponge.

Indicator Sponge Demo:

Materials for Demo

Cellulose sponge	Congo red indicator
Distilled or deionized water	Sodium hydroxide, saturated solution, 100 mL
Hydrochloric acid 1.0 M, 100 mL	Beaker (1000 mL)
Red food coloring	Blue food coloring

Make the indicator sponge by taking a clean cellulose sponge and submerging it into a prepared congo red dye solution. When the sponge is soaked in the congo red, it becomes permanently bonded to the cellulose fibers. The sponge turns blue in acidic solutions and goes back to red in basic solutions.

1. Add 100mL of 1 M HCl to a 1000 mL (or larger) beaker. Fill the beaker about ¾ full with tap water.
2. Add enough red food coloring to the solution to turn it a deep red color.
3. Add 100 mL of 1 M NaOH solution to a 1000 mL beaker. Fill the beaker about ¾ full with tap water
4. Add enough blue food coloring to the basic solution in the beaker until it is a deep blue color.
5. When the sponge is red, rinse it with tap water and squeeze it out.
6. Slowly place one end of the red sponge into the red acid solution and remove it. The sponge should turn blue. Place the entire sponge into the acid solution and then rinse the sponge with tap water to prove that the sponge is actually blue, not just saturated with the acid solution.
7. Squeeze the excess acid out of the sponge, and place the now blue sponge into the blue NaOH solution. The sponge should turn red back to red. Repeat the rinsing process and lead a discussion with the students about what they think happened. Use this discussion to lead in to the Acid & Base pH Lab.

Procedure / Description of Lesson

Acid & Base pH Lab

Name(s) ____

Date ____________________________ Period _________

Q: Which household solutions are acids and which are bases?

Measuring pH indicates whether a solution is an acid or a base. A pH less than 7 indicates an acid and the lower the pH, the stronger the concentration of the acid. A pH more than 7 indicates a base and the higher the pH, the stronger the concentration of the base. Acids and bases are chemical opposites but if a solution's pH is 7, it is neutral which means it is neither an acid nor a base.

Try the experiment below and discover the pH of solutions of many common household items.

Materials

8 plastic cups (or beakers) each containing an unknown
8 pieces of pH paper and a pH color scale
Lab sheet with data table (on back or next sheet)

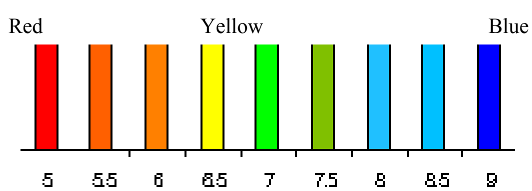
Safety

- **Be sure to wear your goggles at all times!**
- **DO NOT TASTE ANY OF THE SOLUTIONS!**
- **Smell only by "waving" some of the fumes toward your nose.**

Procedure

Determine whether a household solution is an acid or a base and rank the relative strength of the solutions according to pH.

1. Take turns testing the solutions in the numbered plastic cups by dippping a pH strip into the liquid for about 2 secs. Take the strip out and "read" the pH within 10 sec. by comparing the color of the strip to the pH color chart.
2. Record the pH on your data chart for each of the numbered solutions.
3. Now determine if the solutions were acids (pH from 0 - 6), neutral (pH 7) or bases (pH from 8 - 14)
4. Record this information on your data chart beside each solution.
5. Clean up your lab area completely. Get the names of the solutions from your teacher and go to your seats to complete the rest of the lab.



Data Table:

Solution #	pH # from test	Acid or Base?	Solution Name
#1			
#2			
#3			
#4			
#5			
#6			
#7			
#8			

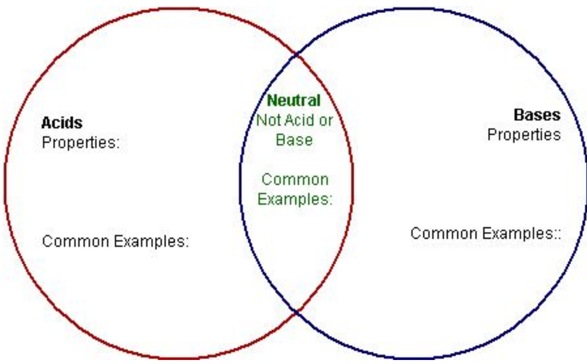
Summary Report:

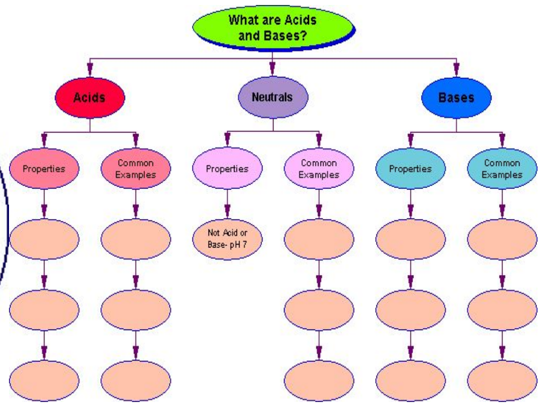
Choose and complete 1, 2, or 3. You must do 4 with your partner. You may need to re-read the text book to find the properties of acidic, basic, and neutral solutions.

1. Make a Venn diagram and fill in the names of the items that were in each of the solutions in the correct section in the diagram and write the properties in the correct section. (See example below.)
2. Write a brief report about acids, bases, and neutrals in paragraph form and include examples from the testing lab.
3. Draw a concept map or word web comparing and contrasting acids and bases and add descriptive words and examples. (See examples on the next page.)
4. Create a continuum (similar to a time line) with the solution substances listed along it in order of their pH. Show the pH and label the continuum to show where acids, bases, and neutral occur along with descriptive phrases.

Assessment Ideas:

Sample Venn Diagram:	Sample Concept Map:





Acid-Base Concepts
Activity #2 – Indicators

Questions to be investigated

How can household indicators be made and are they acidic or basic?

Objectives

Describe tests that can be used to distinguish an acid from a base. Classify various solutions as acidic or basic, given their pH. (C5.7C)(C5.7D)

Teacher Notes

Students will use different fruits and vegetables and create acid-base indicators. Students must make the solutions and then prove that they work as indicators. You may want to have them prepare the solutions on their own at home, and then bring them in to do the lab.

Teacher tips and answers:

The following tips may help you to troubleshoot problems that students may have with their labs:

- Tomato: The solution remains the same color in acid, but turns light green in base. Some students may question whether this truly means it's an indicator solution – point out that any difference in color between acid and base makes it a good indicator.
- Turnip: Turnips typically contain only a small amount of red coloring – as a result, it may be difficult to get a solution concentrated enough to see color changes. In NaOH, the indicator will turn green, in HCl it will turn pink.
- Blackberry: An easy fruit to make an indicator from. Green in base, pink in acid.
- Lime: It may be necessary to grind this with a mortar and pestle to get enough color to make a good solution. Colorless in acid, yellow in base.
- Red grapes: Students have found that the skins of the grapes are good for making indicators, while the flesh yields a basically worthless sticky juice. Olive green in base, no color change in acid.
- Red pepper: Difficult to make an indicator from, since the color seems bound tightly to the skin. Students found that pounding the pepper with a hammer was the best way to break the color loose. No color change in acid, very slightly yellow in base.
- Red apple: Yellow in base, pink in acid.
- Eggplant: Eggplants were a real mess, since they tend to oxidize quickly and turn brown. Additionally, the skin has the color tightly bound into it, so smashing with a mortar and pestle is about the only way to get any appreciable quantity of color out of them. Yellow in base, pinkish in acid.

- Radish: In NaOH, the indicator will turn green, in HCl it will turn orange/red.
- Red onion: Very easy. Yellow in base, pink in acid.
- Yellow pepper: Similar difficulties to red peppers, yielding a very faint indicator. Colorless in acid, yellow in base.
- Habanero pepper: Similar difficulties to red peppers, yielding a very faint indicator. Colorless in acid, purple in base. Caution: habanero peppers are EXTREMELY spicy, and great care should be taken to ensure that students keep their hands away from eyes, nose, and mouth until they have washed them several times in soap and water. It may be a good idea to have them wear gloves.

Materials

3 small dropper bottles containing 0.1 M NaOH
3 small dropper bottles containing 0.1 M HCl
6 Bunsen burners
6 ring stands with rings
6 wire gauzes
6-250 mL beakers
1 L distilled water (though tap water may be used)
2 kitchen knives (knives with serrated edges seem to hold up best over the long haul)
3 mortars and pestles
3 pair scissors
6 dropper bottles for students to place their final extracts into
Various fruits and vegetables with colored skins. (It's best if each lab group has a unique food.)
6 watch glasses
Filter paper and funnels for removing the pulp from each extract.

Class time required:

One-half to one hour.

Safety Concerns

1. Goggles should be worn by all students during the entire lab.

2. Care should be exercised before letting students use knives to cut up vegetables. I would recommend that only high school students be permitted to use knives, and even then only if you believe they are mature enough to handle them. If you're not sure, DO NOT USE KNIVES!

3. The boiling extracts tend to splash somewhat. Make sure all possible safety precautions involving the use of Bunsen burners and boiling liquids are used.

4. The small quantities of HCl and NaOH used in this lab are still enough to cause injury, especially to the eyes. Make sure students are aware of this danger!

5. Habanero peppers, should you choose to use them, are extremely hot and will cause painful (but harmless) chemical burns on the skin and mucous membranes. It is not known whether these burns are harmful to the eyes, but all possible care should be exercised with these peppers. The use of gloves is recommended – even then, everything used to handle these peppers should be cleaned before students are permitted to touch with their bare hands.

Sources

Modified from www.chemfiesta.com Indicator Lab

Procedure/Description of Lesson

Indicator Lab

Many naturally colored compounds can behave as acid-base indicators. In this lab, you'll be given a choice of several different fruits and vegetables and given the chance to make an indicator of your own.

For full credit, you must do the following:

- Provide a complete set of instructions for making the indicator solution from the raw materials.
- Provide a sample of your indicator solution in a form that would be usable by others.
- Demonstrate that the indicator solution truly works for acids and bases. This should be done in the form of observations of what the indicator looks like in each.

Assessment Ideas

This lab may be used for a lab report since the students are designing there own lab. The following things might be required as part of the report:

1. A purpose and background section that includes a description of the procedure as well as background information that includes the definition of an indicator.
2. A procedure that shows the steps to the experiment. The steps should be repeatable.
3. Data tables showing the results of testing each acid and base with each indicator the group made.

Acid-Base Concepts
Activity #3 – Acids & Bases

Questions to be investigated

What are Bronsted-Lowry acids and bases and their conjugates?

Objectives

(C5.r7i) Identify the Bronsted-Lowry conjugate acid-base pairs in an equation.

Sources

ChemQuests from www.ChemistryInquiry.com with permission from Jason Neil

Procedure / Description of Lesson

Students should work in small groups (2-3) in order to read and answer the questions in the ChemQuest. Discuss answers following activity. This is an inquiry based activity because the students are given the information one piece at a time and then apply the concepts to scenarios at the end of the assignment.

ChemQuest 48

Intro to Acids and Bases

Name: ______________________

Date: ________________

Hour: _____

Information: Definitions of Acids and Bases

Arrhenius definitions

1) acid: substance that when dissolved in water increases $[H^+]$; (note: H^+ exists bonded to water as the hydronium ion, H_3O^+, so $[H^+]$ and $[H_3O^+]$ are equivalent expressions)
2) base: substance that when dissolved in water increases $[OH^-]$

Bronsted-Lowry definitions

1) acid: substance that donates a proton, H^+, in a reaction
2) base: substance that accepts a proton, H^+, in a reaction

Table 1: Equilibrium constants (at 25°C) for some acid-base equilibrium reactions.

Reaction	K_c
1. $C_2H_3O_2^- + H_2O \leftrightarrow HC_2H_3O_2 + OH^-$	1.07×10^{-11}
2. $HCN + SO_4^{2-} + H_2O \leftrightarrow HSO_4^- + CN^- + H_2O$	4.9×10^{-11}
3. $HC_2H_3O_2 + H_2O \leftrightarrow H_3O^+ + C_2H_3O_2^-$	3.09×10^{-7}
4. $H_2CO_3 + H_2O \leftrightarrow H_3O^+ + HCO_3^-$	7.82×10^{-9}
5. $HCl + H_2O \leftrightarrow H_3O^+ + Cl^-$	2.0×10^4

Critical Thinking Questions

1. List all of the reactants in Table 1 that are Arrhenius acids.

2. List all of the reactants in Table 1 that are Arrhenius bases.

3. List all of the reactants in Table 1 that are Bronsted-Lowry acids.

4. List all of the reactants in Table 1 that are Bronsted-Lowry bases.

5. Is it possible for an ion to act as a base? Explain.

6. Can a substance act as both an acid and a base under different conditions? Explain.

7. Is this statement true: "All substances that are Arrhenius acids are also Bronsted-Lowry acids"? Explain.

8. If a substance is a Bronsted-Lowry acid, can we conclude that the substance is also an Arrhenius acid? Explain.

9. a) What is the strongest acid among the reactants in reactions 3-5 inTable 1? Explain.

   b) What is the weakest acid among the reactants in reactions 3-5 inTable 1? Explain.

10. Consider Reaction 3.
    a) What substance is formed (by the acid) after the acid loses a proton?

    b) Is this substance an acid or a base? (Hint: look at reaction 1.)

11. Drawing a conclusion from question 10, what can be said about a substance after it loses a proton? Is the substance formed acidic or basic?

Information: Conjugate Acid-Base Pairs

After an acid loses a proton in a reaction, the substance formed behaves like a base. Verify this by examining Reactions 3 and 1 in Table 1. Notice from reaction 3 that $HC_2H_3O_2$ is an acid. After it loses a proton it becomes the acetate ion, $C_2H_3O_2^-$. The acetate ion is a base, as seen in reaction 1; there is a special name for this base: it is a conjugate base. So, $C_2H_3O_2^-$ is the conjugate base of $HC_2H_3O_2$. Similarly, HSO_4^- is the conjugate acid of SO_4^{2-}. Verify this by examining Reaction 2.

Critical Thinking Questions

12. Describe how a conjugate base is formed.

13. How is a conjugate acid formed?

14. For each of the acids below, write the reaction of the acid with water and circle the formula of the conjugate base in your reaction.
    a) H_2SO_4

    b) HCO_3^-

    c) HF

15. For each of the bases below, write the reaction of the base with water and circle the formula of the conjugate acid in your reaction.
    a) NH_3

    b) OH^-

    c) NO_3^-

Acid-Base Concepts

Activity #4– A Simple Titration

Questions to be investigated

How can you perform a simple acid/base titration using an everyday problem ("flu virus")?

Objectives

C5.7D - Classify various solutions as acidic or basic, given their pH.

Teacher Notes

This lab should be prepared ahead of time. All students will be given a small amount of base with the exception of one student who is given a small amount of acid. They record their results after mixing their acid or base with two other students. The acid is considered to be "the virus" and it can be traced through the class using an indicator.

Real World Connection

pH measurements are used in many different situations such as swimming pool test kits, biological processes where correct pH dictates correct enzyme function, aquaria, hot tubs etc.

Sources

Simple titration lab used with permission from Craig Riesen, Clarenceville High School.

Procedure/Description of Lesson

Students are given a scenario to demonstrate a titration to determine who passed "an illness" from one person to another.

Purpose

To perform a simple acid/base titration using an everyday problem ("flu virus").

Discussion

Acids are found in many foods and taste sour. Vinegar, carbonated beverages, digestive juices in the stomach, car batteries, and fruit (especially citrus) all contain acids in some form. Acids react with metals to form gas in a single replacement reaction. Bases, on the other hand, have a bitter taste and feel slippery because the base reacts with the fat in our finger to form soap. Examples include drain cleaner, alum, cocoa, soap and deodorant. One should not deal with acids or bases without a solid understanding of what they are and how strong they are.

How do we know whether a substance is acidic or basic? How can we tell how strong an acid or base is? You may have seen a lifeguard using a plastic kit to scoop some pool water. Then, they place a few drops of chemicals into vials and shake it. What are they doing? They are testing the pH of the pool water. pH indicates acidity or basicity of solutions as well as acid or base strength. A pH scale was developed, ranging from pH of 1 (strongest acid) to pH 7 (neutral) to pH 14 (strongest base). Acids have pH's less than 7 and Bases have pH's greater than 7.

The lifeguard uses the color change of the pool water reacting with chemicals to determine if the water is safe to swim in. The chemicals are called indicators. Indicators are used to colorfully (and qualitatively) show whether a solution is acidic or basic without having to do extensive quantitative measurements. Normally, an indicator is specific for a particular pH range and will turn color as the pH of a solution within that pH range varies. Flower color is often determined by the pH of the soil the plant is potted in because the flowers contain natural indicators.

Some indicators are general and will turn one color in an acid, another color in a neutral solution and a third color in a base. They may not indicate the exact pH of an acid or base, however. Litmus paper is such an indicator. Litmus will turn red in an acid and blue in a base. Other indicators are pH specific. Hydrion paper and other brands of pH paper, for instance, indicate acidity or basicity in a narrow range of pH by changing colors for each range of pH. Some indicators are specific for bases and others for acids. For example, Phenolphthalein turns red in a base. Methyl red is used to indicate an acid.

Materials

Lab Apron	Safety Goggles	Bromthymol Blue or Phenol Red
Clear Plastic Cups	$NaOH_{(s)}$	concentrated HCl
1 ml Pipette (0.01 ml)	5 - 10 ml Pipettes	10 ml Grad Cylinder
100 ml Grad Cylinder	50 ml Beaker	250 ml Beaker

Procedure

1. The teacher will prepare <u>1 L</u> of a <u>0.001 M</u> solution of <u>NaOH</u> $_{(aq)}$ from solid NaOH crystal, using quantitative analysis [molarity].
2. The teacher will prepare make <u>100 ml</u> of a <u>0.1 M</u> solution of <u>HCl</u> $_{(aq)}$ from concentrated $HCl_{(aq)}$ [$M_1V_1 = M_2V_2$].
3. The two solutions will be set aside and labeled "acid" and "base".
4. Test each solution with Litmus paper. Record your results in the calculations.
5. Measure 5 ml of the acid solution and place this in a plastic cup.
6. Measure 5 ml of the base solution for all other plastic cups (one for each student).
7. The teacher will have an assigned area to place all the cups in order to redistribute the cups to students. The teacher must know who has the acid solution in order for this activity to be effective. (*Students should not know who has the acid solution*.)
8. List each student on the calculations and data sheet (#1-27).
9. Each student will be given a plastic cup containing a solution.

10. Find a partner and pour the contents of your cup into his/hers. GENTLY pour the contents back and forth a couple of times, and pour half of the mixture back into your partners' cup.
11. WRITE DOWN the initials or name of the person you "contacted" as 'Person 1' on the calculations and data sheet for YOUR row.
12. Circulate around the room and repeat this process (procedures 9-11) with TWO other students ONLY.
13. Each student should have made three (3) "contacts" total.
14. Add <u>1 drop</u> of the indicator solution (bromthymol blue) to your plastic cup and GENTLY swirl to mix the indicator with the solution.
15. Observe any color changes in your mixtures. A yellow solution indicates you caught the virus! A bluish-green color indicates you remained healthy.

Calculations and Data

1. Complete the chart on your Calculations and Data sheet by placing your name in the left hand column along with the initials of all other participants (# 1-27).
2. The three people whom you "mixed" with are "Persons #1 – 3" in the rows of the table. Place their names in the columns under "Person 1," "Person 2" and "Person 3."
3. The teacher will then lead a class discussion of who got "infected" by the "flu virus.
4. Can you figure out the one person who first carried the flu virus? If so, how?
5. What happens to Litmus paper in each of the two solutions that were made?

Conclusions and Questions

1. Name a disease that can be spread in the same kind of way as observed in this activity. Explain your answer.
2. Why did the color change in your solutions when mixed?
3. What was the general pH of the solution that was NOT "infected" with the "flu virus?"
4. What was used to show the acidity or basicity of the solution?
5. Generally speaking, what causes something to be an acid? A base?
6. What can be used to determine the pH of a solution within a certain pH range?
7. Give several examples of both acids and bases.

Calculations and Data

1. Chart of student results. Name "Person 1", "Person 2" and "Person 3". Then, place an "X" in the box that showed a color change after we added the indicator.

Student	**Person 1**	**Person 2**	**Person 3**
1			
2			
3			
4			
5			
6			
7			
8			
9			
10			
11			
12			
13			
14			
15			
16			
17			
18			
19			
20			
21			
22			
23			
24			
25			
26			
27			

4. Can you figure out the one person who first carried the flu virus? If so, how*? [The teacher will unveil the "infected" person eventually*.]

Conclusions and Questions

1. Name a disease that can be spread in the same kind of way as observed in this activity. Explain your answer.

2. Why did the color change in your solutions when mixed?

3. What was the general pH of the solution that was NOT “infected” with the “flu virus?”

4. What was used to show the acidity or basicity of the solution?

5. Generally speaking, what causes something to be an acid? A base?

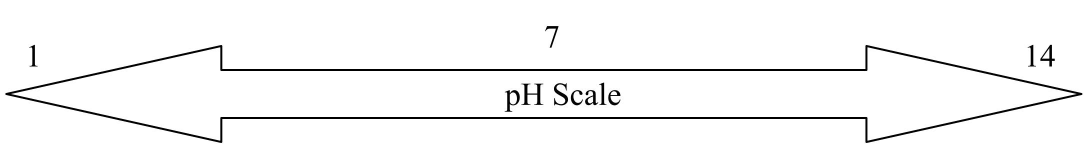
6. What can be used to determine the pH of a solution within a certain pH range?

7. Give several examples of both acids and bases.

   Acids →

   Bases →

8. Label the pH scale using the terms: acidic, neutral and basic; excess H+ ions, excess OH- ions



Solution Preparation of 1 L of 0.001 M NaOH

[For the accelerated students]

- Add **0.4 g** of NaOH solid **to 100 ml** of distilled water
- Then, add **1 ml** of that 1 M solution **to 1 liter** of distilled water

1. Find molar mass of NaOH = 40 g / mole

   Na = 23 g/mole
   O = 16 g/mole
   H = 1 g/mole

2. Theoretically, we can add 40 g of NaOH to enough distilled water to make1 liter of a 1 M solution. How do we get a 0.001 M solution of NaOH?

   Molarity = moles / liter

   0.001 M = X moles / 1 L

   X = 0.001 moles of NaOH x 40 g / mole = **0.04 grams** of NaOH added to **1 Liter** of distilled water

2. Since 0.04 grams is hard to measure, we can use a dilution:
   mass1 / volume1 = mass2 / volume2.

   0.04 grams / 1 liter = **0.4 grams / 10 liters** = 4.0 grams / 100 liters

   - but we do not want to use 10 liters of distilled water either, so ....
   - use **<u>0.4 grams</u>** of NaOH in **<u>100.0 ml</u>** of distilled water

     0.4 g x 40 g/mole = 0.1 mole of NaOH
     0.1 mole / 0.1 L = 1 M solution of NaOH ... but we want a 0.001 M solution

   - Use $M_1 V_1 = M_2 V_2$

   [*The molarity of the first solution times its volume = the molarity of the second solution times its volume*]

     (1 M)(V) = (0.001 M)(1 L)

     - In other words, how much of the 1 M NaOH solution do we need to add to 1 liter of distilled water to get the 0.001 M NaOH solution
     - We need 0.001 liters or 1 ml of 1 M NaOH $_{(aq)}$ → simply **<u>add 1 ml</u>** of the 1 M NaOH **to <u>1 liter</u>** of distilled water to get your desired 0.001 M solution

- -

10 ml of a 0.1 M solution of HCl $_{(aq)}$

- Add **0.8 ml** of conc. HCl **to 100 ml** of distilled water & use 5 ml each test

$M_1 V_1 = M_2 V_2$

- Concentrated HCl $_{(aq)}$ = 12 M   10 ml = 0.01 liters

$M_1 V_1 = M_2 V_2$ (12 M)(**V**) = (0.1 M)(0.01 L)

V = 0.00008 L = 0.08 ml → 0.08 ml / 10 ml = X ml / 100 ml

Add 0.0008 L or **0.8 ml** of concentrated HCl $_{(aq)}$ to **100 ml** of distilled water ... use 5 ml

Answers to Conclusions and Questions

1. Name a disease that can be spread in the same kind of way as observed in this activity. Explain your answer.
   Epidemics of various contagious diseases [flus, colds, bacterial & viral infections, measles, chicken pox, etc.] form by people contacting each other via coughing, sneezing, touching.

2. Why did the color change in your solutions when mixed?
   We placed an indicator that was sensitive to acid. The "flu virus" was the acid which was added to a base. The indicator (Phenol Red) would change color in an acid.

3. What was the general pH of the solution that was NOT "infected" with the "flu virus?"
   Every student but one had a base solution to begin with. Therefore, the pH is greater than 7.

4. What was used to show the acidity or basicity of the solution?
   We used two different indicators: Litmus paper and Bromthymol Blue or Phenol Red. Litmus turns red in an acid and blue in a base. Phenol Red was reddish purple in the base, but turned yellow in an acid. Bromthymol Blue is blue in a basic solution and yellow in an acidic solution.

5. Generally speaking, what causes something to be an acid? A base?
   An acid possesses and excess of Hydrogen ions (H^+ ions). A base contains an excess of hydroxide ions (OH^-).

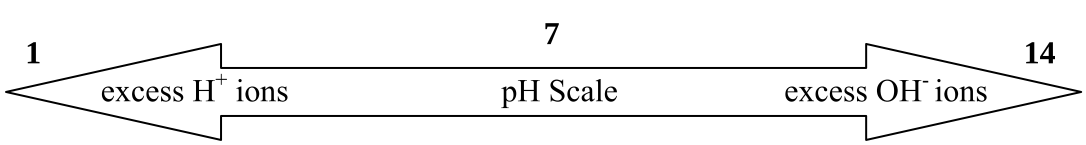
6. What can be used to determine the pH of a solution within a certain pH range?
   Hydrion paper is used to determine the pH of a solution within a certain pH range. A pH meter also works.

7. Give several examples of both acids and bases.
   Acids → stomach digestive juice, soda pop, fruit juice, vitamins, milk (lactic acid)

   Bases → alum, drain cleaner, cocoa, soap, deodorant

8\.

acidic **neutral** **basic**



Acid-Base Concepts

Activity #5 – Acids & Bases

Questions to be investigated

How can we find out the concentration of an acid or base using a titration?

Objectives

Students will use a titration to perform a neutralization reaction. (C5.7B)

Materials

Goggles, Aprons	1 - 100 ml Burette	Burette Clamp
1 M NaOH	1 M HCl	1 M H_2SO_4
1 M K(OH)	Distilled water	50 ml Beaker
250 ml beaker	Bromthymol Blue	Hydrion (pH) paper
Masking tape	Funnel	Ring stand

.

Safety Concerns

Safety goggles should be worn at all times. Acids and bases can be irritating to eyes and skin. Burettes are expensive and should be used with care.

Real-World Connections

Many companies, hospitals, universities etc. use acid base titrations to analyze the presence of compounds in various consumer products.

Sources

The website below is an online virtual titration lab that may be used to practice prior to doing the lab in class.

http://lrs.ed.uiuc.edu/students/mihyewon/chemlab_instruction.html

Acid and Base lab used with permission from Craig Riesen, Clarenceville High School.

Procedure / Description of Lesson

Purpose:

To investigate Acid-Base Titration using various indicators.

Discussion:

Chemists used **acid-base titration** to determine the composition of acids and bases in solution. Acid-base titration is used in nearly every laboratory, including universities, power plants, candy factories, hospitals, poison & emergency medical treatment facilities and even breweries in order to analyze mixtures of compounds in solutions. It is one of the most precise methods of chemical analysis. To produce "Coke" or "Pepsi" or a "Snickers Bar" requires consistent precision in terms of the ingredients used.

In this lab we will use **neutralization reactions** to understand acid-base theory. Acids and bases ionize readily, making them reactive compounds. When a strong acid and a strong base are mixed together, however, the reactivity is neutralized because salt and water are produced. Salt and water are non-reactive compounds. Neutralization is defined by the following equation:

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

$$HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H(OH)_{(l)}$$

Notice that water H(OH) is produced as the H^+ ion from the acid combines with the OH^- ion from the base. Salt (NaCl) is produced as the Na^+ ion from the acid combines with the Cl^- ion from the base. During this particular neutralization the following Net Ionic Equation took place:

$$H^+ + OH^- \rightarrow H(OH)_{(l)} \qquad \text{Spectator Ions: } Na^+, Cl^-$$

Acid-base titrations use indicators. An indicator is selected on the basis of color changes they produce at or near the point at which equal quantities of an acid and a base have been added (this is called the equivalence point). When titrating, however, the end point is often a more important pH value. The end point shows the point at which a change in the indicator color reveals that neutralization is complete. At the end point, one drop of an acid or a base can produce a color change.

In order to titrate, an equation is used to determine stoichiometric quantities of acids and bases:

$$n_a M_a V_a = n_b M_b V_b$$

n_a → mole factor of the acid
M_a → molarity of the acid
V_a → volume of the acid

n_b → mole factor of the acid
M_b → molarity of the base
V_b → volume of the base

Let's consider an example of titration using a strong acid, HCl, and a strong base, NaOH. If we know the molarity and volume of the acid used, we can predict the volume of a certain molarity of base needed to neutralize it. This type of reaction is used, for example, when something splashes in your eye at home. It begins to burn because it is an acid. You take the bottle of "stuff" that splashed in your eye and rush to the hospital. The emergency staff tells you to quickly and thoroughly flush your eye with water and sends the "stuff" to the lab. The lab does a titration to determine if the "stuff" is an acid or base and then to find a substance to neutralize your eye.

$$HCl + NaOH \rightarrow NaCl + H(OH)$$

1) The mole ratio between the acid, HCl, and the base, NaOH, is 1:1 in the balanced equation as shown.
2) If we use 10 ml of acid and know that the molarity of the acid and the base is 1 M, we can determine how much base is needed to neutralize the acid using:

$$n_a M_a V_a = n_b M_b V_b$$

$$(1)\ (1\text{ M})\ (10\text{ ml}) = (1)\ (1\text{M})\ (?\text{ ml})$$

3) 10 ml of the base, NaOH, is needed to neutralize 10 ml of the acid, HCl.

Materials

Goggles, Aprons	1 - 100 ml Berets	Buried Clamp
1 M NaOH	1 M HCl	1 M H_2SO_4
1 M K(OH)	Distilled water	50 ml Beaker
250 ml beaker	Bromthymol Blue	Hydrion (pH) paper
Masking tape	Funnel	Ring Stand

Procedures

1. Obtain goggles and aprons.
2. All titrations will be done next to a sink on the back counter. Do not set up on the lab tables.
3. Use stoichiometry to calculate how to prepare 250 ml of each acid and base solution using the molarity equation under the guidance of the teacher. We need 2 – 100 ml burettes for each solution (8 total).
4. Set up the ring stand and buret clamp next to a sink.
5. Clean one 100 ml buret by filling with distilled water (*from a 100 ml beaker*) and allowing the water to go through the stopcock. Use a 100 ml beaker to capture the water in the buret. Practice measuring out EXACTLY 10 ml of water from the buret into the beaker.
6. Repeat the filling, rinsing and practice measuring using the buret 2-3 times.
7. Each group will be assigned ONE acid OR ONE base to actually prepare.
8. Take the buret out of the clamp and hold it near the sink.
9. Place a clean, dry funnel in the buret.
10. The TEACHER will pour the acid or base into the respective buret, setting the meniscus of the acid or base solution at exactly the 0 milliliter calibration in the buret.
11. Place the buret back into the buret clamp on the ring stand.
12. Rinse the 250 ml beaker thoroughly with tap water and put it in the lab station.
13. Label which acid or base solution is in the buret. There will be 2 burettes of the same solution at each sink.
14. We will do reactions in sets as shown in the table below:

Acid Used	Base Used
HCl	NaOH
HCl	KOH
H_2SO_4	NaOH
H_2SO_4	KOH

15. Complete #1 - #2 on the Calculations and Data section concerning pH, balanced chemical equations, and predicting how much acid or base is needed for each solution tested.
16. Use pH or Hydrion paper to determine the pH of the acid or base that you used*.* *Record this in the data table "pH of the Acids & Bases" on your Calculations and Data Sheet*.

17. Start with your groups assigned solution (acid or base); add EXACTLY 5 ml of acid to a clean, dry 50 ml beaker.
18. Add 2 drops of Bromthymol Blue indicator to the acid in the 50 ml beaker.
19. If you start with an acid, go to a base buret and vice versa. At first, add exactly 1 ml of the base and swirl the beaker GENTLY. Add another milliliter of solution and gently swirl the beaker. Look for color changes. As soon as any kind of color change occurs, go SLOWLY.
20. KEEP TRACK OF EXACTLY HOW MUCH BASE YOU ADD TO THE ACID.
21. Look for a green color that begins to remain in the beaker (*neutralization has taken place*).
22. Use pH or Hydrion paper to determine the pH of the neutralized solution.
23. Rinse the 50 ml beaker thoroughly with distilled water and dry it out.
24. Repeat procedures #17-23 for all the acid-base combinations from the table.
25. When finished, turn on the sink water and let it run. Then, pour the acid or base CAREFULLY into the sink. AVOID getting splashed!

Calculations and Data

1. If 10 ml of the acid are used for each titration, predict (*based on your balanced equation*) how much base is needed to bring about a neutralization reaction.

Acid Used (10 ml)	Base Used	Predict Amount of Base Needed	Actual Amount of Base Used
HCl	NaOH		
HCl	KOH		
HCl	$Ca(OH)_2$		
H_2SO_4	NaOH		
H_2SO_4	KOH		
H_2SO_4	$Ca(OH)_2$		

2. Write balanced chemical equation for each of the acid-base reactions that occurred during titration (*see the table on page 2 for reactants*).
3. Fill out the data table “pH of the Acids & Bases” and “Actual Amount of Base Used.”

Conclusions and Questions

1. How did you predictions compare with the actual amount of base needed to neutralize the acids? Explain any error that may have occurred.

2. What is the difference between “neutralization” and “dilution”?

3. One mixes 2 M HCl with 20 ml of 0.5 M $Ca(OH)_2$.
   a. Write a balanced neutralization reaction, the net ionic reaction, and list the spectator ions.
   b. Determine how much acid is needed in this reaction.

Name ______________________________ Acid-Base Titration Lab
Date _________ Hour _________

1. If 10 ml of the acid are used for each titration, predict (*based on your balanced equation*) the amount of base needed to bring about a neutralization reaction. We assume that acids/bases are strong, but not all of them are. Therefore, one must consider the K_a or K_b value.

Acid Used (10 ml)	K_a or K_b	**Base Used**	K_a or K_b	**Predict Amount of Base Needed**	**Actual Amount of Base Used**
HCl	large	NaOH	>>>1		
HCl		KOH	>>>1		
H_2SO_4		NaOH	>>>1		
H_2SO_4		KOH	>>>1		
HCl		$Ca(OH)_2$	<<1		
H_2SO_4		$Ca(OH)_2$	<<1		

2. Write balanced chemical equation for each of the acid-base reactions that occurred during titration.

___ HCl + ___ NaOH $\rightarrow$

___ HCl + ___ KOH $\rightarrow$

___ H_2SO_4 + ___ NaOH $\rightarrow$

___ H_2SO_4 + ___ KOH $\rightarrow$

___ HCl + ___ $Ca(OH)_2$ $\rightarrow$

___ H_2SO_4 + ___ $Ca(OH)_2$ $\rightarrow$

3. pH of the Acids & Bases

Acid	**pH**		**Base**	**pH**		**pH of neutralized solution**
HCl			NaOH			
HCl			KOH			
H_2SO_4			NaOH			
H_2SO_4			KOH			

4. Complete the last column of the table above “Actual Amount of Base Used” as you proceeded through the procedures.

Conclusions and Questions

1. How did you predictions compare with the actual amount of base needed to neutralize the acids? Explain any error that may have occurred.

2. What is the difference between "neutralization" and "dilution"?

3. One mixes 2 M HCl with 200 ml of 0.5 M $Ca(OH)_2$.

   a. Write a balanced neutralization reaction, the net ionic reaction, and list the spectator ions.

   b. Explain why the equilibrium constant is important in determining how much acid/base is required to neutralize another solution.

   c. Determine how much acid is needed in this reaction. Show Work.

4. What type of chemical reaction is an acid-base titration [S, D, SR, DD]?

5. Give two examples in daily life where titrations are used. Be specific.

6. Write the general formula for a neutralization reaction:

7. Based on Table 2, is calcium hydroxide as strong a base as sodium hydroxide? Explain your answer.

8. How did you know that a chemical reaction (titration) had taken place in this lab?

9. Based on this lab, if you had 100 ml of hydrochloric acid, how much sodium hydroxide would be needed to neutralize it [less than 100 ml, 100 ml, greater than 100 ml]? Explain.

10. Based on this lab, if you had 100 ml of sulfuric acid, how much sodium hydroxide would be needed to neutralize it? Explain.

11. Acids and Bases are sometimes referred to as "electrolytes." Why? [Look it up!]

12. A strong acid or base is said to "ionize completely" in water. Use hydrochloric acid and sodium hydroxide to explain what this would mean in a neutralization reaction.

10\. Aluminum hydroxide is a weak base found in deodorants. In your own words and using question 9 for help, how would you describe the ionic character of aluminum hydroxide?

Teacher Answer Key

4. pH of the Acids & Bases

Acid	pH	Base	pH
HCl	1-2	NaOH	13-14
		KOH	13-14
		$Ca(OH)_2$	11
H_2SO_4	1-2	NaOH	
		KOH	
		$Ca(OH)_2$	

7-11. If 6 drops of the acid are used for each titration, record how many drops of base is needed to bring about a neutralization reaction.

Acid (10 ml)	Base	Actual Amount of Base Used	pH of neutralized solution
HCl	NaOH	10 ml	7
HCl	KOH	10 ml	7
HCl	$Ca(OH)_2$	100 ml	7
H_2SO_4	NaOH	20 ml	7
H_2SO_4	KOH	20 ml	7
H_2SO_4	$Ca(OH)_2$	200 ml	7

If K_a or K_b is large, it greatly favors the forward reaction (ionization). If small, it favors the reverse reaction.

Calcium hydroxide is a weak base. Therefore, one needs a lot of it to neutralize the strong acids.

- Write balanced chemical equations for each of the acid-base:

$1\ H^+Cl^- \quad + \quad 1\ Na^+(OH)^- \rightarrow 1\ Na^+Cl^- \quad + \quad 1\ H^+(OH)^-$

$1\ H^+Cl^- \quad + \quad 1\ K^{+1}(OH)^- \rightarrow 1\ K^{+1}Cl^- \quad + \quad 1\ H^+(OH)^-$

$2\ H^+Cl^- \quad + \quad 1\ Ca^{+2}(OH)^-_2 \rightarrow 1\ Ca^{+2}Cl_2^- \quad + \quad 2\ H^+(OH)^-$

$1\ H^+_2(SO_4)^{-2} \quad + \quad 2\ Na^+(OH)^- \rightarrow Na^+_2(SO_4)^{-2} \quad + \quad 2\ H^+(OH)^-$

$1\ H^+_2(SO_4)^{-2} \quad + \quad 2\ K^{+1}(OH)^- \rightarrow K^+_2(SO_4)^{-2} \quad + \quad 2\ H^+(OH)^-$

$1\ H^+_2(SO_4)^{-2} \quad + \quad 1\ Ca^{+2}(OH)^-_2 \rightarrow Ca^+(SO_4)^{-2} \quad + \quad 2\ H^+(OH)^-$

Answers to Conclusions and Questions

1. What kind of chemical reaction is an acid-base titration?

   Double displacement (double replacement)

2. Give two examples in daily life where titrations are used.
   a. **producing any drink, many food combinations**
   b. **drugs, chemicals used in hospitals, etc.**

3. Write the general formula for a neutralization reaction:

   Acid + Base → Salt + Water

4. Based on Table 2, is calcium hydroxide as strong a base as sodium hydroxide? Explain your answer.

   Explain why the equilibrium constant is important in determining how much acid/base is required to neutralize another solution.

   Sodium hydroxide is a stronger base than calcium hydroxide. More drops of calcium hydroxide were needed to neutralize the acids. The K_b of Calcium hydroxide is much less than 0. Therefore, it is a very weak base.

5. How did you know that a chemical reaction (titration) had taken place in this lab?

   There was a color change based on the "indicator" Bromthymol Blue

6. Based on this lab, if you had 100 ml of hydrochloric acid, how much sodium hydroxide would be needed to neutralize it?

   100 ml ... since the neutralization reaction is a 1:1 ratio of moles (see the balanced equation). Plus using the # of drops needed (6 drops of acid to 6 drops of base)

7. Based on this lab, if you had 100 ml of sulfuric acid, how much sodium hydroxide would be needed to neutralize it?

   200 ml ... since the neutralization reaction is a 1:2 ratio of moles (see the balanced equation). Plus using the # of drops needed (6 drops of acid to 12 drops of base)

8. Acids and Bases are sometimes referred to as “electrolytes.” Why? How does the concept of electrolyte relate to the strength of an acid or a base? [Look it up!]

Electrolytes are formed as substances form ions in solution. The more ions that form, the stronger the acid or the base. Hydrochloric acid and sulfuric acid are all strong acids while and sodium hydroxide is a strong base, meaning they ionize almost completely in water.

9. A strong acid or base is said to “ionize completely” in water. Use hydrochloric acid and sodium hydroxide to explain what this would mean in a neutralization reaction.

Since hydrochloric acid and sodium hydroxide are a strong acid and a strong base, they ionize completely. This means that HCl breaks apart into H^+ ions and Cl^- ions The NaOH breaks apart into Na^+ ions and OH^- ions. There would no longer be HCl or NaOH compounds in the water, only the H^+ , Cl^- , Na^+ , and OH^- ions.

10. Aluminum hydroxide is a weak base found in deodorants. In your own words and using question 9 for help, how would you describe the ionic character of aluminum hydroxide?

Aluminum hydroxide would form fewer ions in solution than a strong base. Therefore, Al^{+3} ions and OH^- ions will be present in solution, but so would $Al(OH)_{3\ (s)}$ compounds. In a neutralization reaction, aluminum hydroxide would offer less ions to react with an acid than a strong base like sodium hydroxide (NaOH) would.

Acid-Base Concepts
Activity #6 – Acid Rain Experiments

Questions to be investigated

What is acid rain and what are the problems associated with it.

Objectives

The students will first do an activity that examines the effects of a buffer followed by designing a series of experiments that test whether the presence of soil in water will influence the whether the pH of water will change when an acid or base is added. (C5.7h) (C5.7E)

Materials

distilled water
Buffered Solution: add 1/2 tsp baking soda to 1 liter distilled water
Acid Rain Solution: add 4 ml 1M H_2SO_4 to 2 liters distilled water.
pH meter, test kit, pH paper, or Universal Indicator Solution
beakers or clear plastic cups (200-ml size, two per student or group
25-ml graduated cylinders (one per student or group of students)
10-ml pipette (one per student or group of students)
safety goggles
gloves
optional: alkalinity test kit

Real-World Connections

Students will learn what causes acid rain and the aesthetic and natural problems associated with it.

Sources

Acid Rain Experiment is from Cornell University, Ithaca, NY 14853
EnvInquiry@cornell.edu http://ei.cornell.edu

This exercise was written by Nancy Trautmann (Cornell University) under NSF Award #9454428.

Procedure / Description of Lesson

Acid Rain Experiments

Cornell University, Ithaca, NY 14853 EnvInquiry@cornell.edu http://ei.cornell.edu

Background

Acid precipitation is defined to have a pH lower than 5.6. In New York, the average pH of rainfall is 4.0-4.5 and individual storms as low as 3.0 are not unusual.

Freshwater lakes commonly are slightly basic. pH in the range of 6.5 to 8.2 is optimal for most organisms, and below 5.0 is lethal to many fish species. The susceptibility of lakes to changes in pH varies depending on how well buffered they are. Measured as alkalinity, the buffering capacity of water is a function primarily of the concentration of carbonate (CO_3^{-2}) and bicarbonate (HCO_3^{-1}) ions.

In areas with limestone ($CaCO_3$) bedrock, surface waters have high concentrations of carbonate and bicarbonate and therefore are able to resist change in pH. The pH of a well-buffered lake does not change dramatically following a storm or snowmelt period because the acidity becomes neutralized by these ions.

In regions where the bedrock is granite, the soils and surface waters are naturally low in alkalinity. One such region is the Adirondack Mountains, where approximately 20% of the lakes are too acidic to support fish life.

One approach to restoring acidic lakes is to add lime to the lake itself, to the influent streams, or to the watershed land. This can be simulated in the laboratory using baking soda ($NaHCO_3$), horticultural lime, or a stomach antacid such as Tums, which is made up of $CaCO_3$.

Experimental Problem #1 (Standard Problem)

What is a buffer? What happens to the pH of a buffered solution when acid or base is added, and how does this compare to an unbuffered solution?

Introduction

Discussion of acid rain:

- What is acid rain?
- What are the causes?
- What regions are most affected?
- Why are some lakes more susceptible to acid rain than are others?

Materials

distilled water
Buffered Solution: add 1/2 tsp baking soda to 1 liter distilled water
Acid Rain Solution: add 4 ml 1M H_2SO_4 to 2 liters distilled water.
pH meter, test kit, pH paper, or Universal Indicator Solution
beakers or clear plastic cups (200-ml size, two per student or group
25-ml graduated cylinders (one per student or group of students)
10-ml pipette (one per student or group of students)
safety goggles
gloves
optional: alkalinity test kit

Procedure

1. Put 25 ml distilled water into one beaker and 25 ml Buffered Solution into another.
2. Add 6 drops Universal Indicator Solution to each beaker. {Note the color differences between the two solutions. The distilled water is slightly acidic because of dissolved carbon dioxide; the baking soda solution is slightly basic.}
3. Using a pipette, carefully add Acid Rain Solution drop by drop to the beaker containing distilled water, swirling after each addition until the color stabilizes. How much do you need to add to make the solution turn pink and stay that color, indicating that it is acidic?
4. Using a 25-ml graduated cylinder, carefully add Acid Rain Solution to the cup containing the buffered solution. Add a few ml at a time, swirling and observing the color changes. How much do you need to add to make the solution turn a stable pink?
5. Optional: Measure the alkalinity of distilled water and of Buffered Solution, then relate the alkalinity measurements to the amounts of acid needed to cause a pH change in the two solutions.

Discussion

1. Why were there differences in the amount of acid needed to change the pH of these two solutions?
2. What is a buffer?
3. • How does this relate to lakes?

Acid Rain Experiments Environmental Inquiry http://ei.cornell.edu 3

Experimental Problem #2 (Original Problem)

Materials:

Same as for Experiment Problem #1, plus:

- assorted soil samples
- coffee filters
- rubber bands

Procedure:

Design an experiment or set of experiments that use the materials provided to address one or more of these questions:

1. Does soil change the pH of water that drains through it?
2. Can soil reduce the acidity of water draining into a lake?
3. Do some types of soil buffer drainage water better than do other soil types?
4. Can the buffering capacity of soils be depleted?
5. Does the buffering capacity of a soil relate to the alkalinity of water that has drained through it?
6. What components of soil provide its buffering capacity?

Suggested Outline for Lab Write-up

I. Prediction Statement

II. Lab Log

What you did and when, problems that arose and how you addressed them.

III. Interpretation

A summary of your data, preferably in graphs.

IV. Tips for Future Experimenters

Recommendations for improving upon or extending the research, addressed to students taking this class next year.

Source: Acid Rain Experiments Environmental Inquiry http://ei.cornell.edu 4

Assessment Items

Traditional Items

1. Which of these two lakes is more likely to experience large drops in pH due to acid precipitation? Why?

	Mirror Lake (limestone bottom)	Loon Lake (granite bottom)
pH:	7.5	6.0
temp.:	6oC	4oC
alkalinity	50 ppm	200 ppm
dissolved oxygen	13 ppm	7 ppm

2. If you steadily add acid to a well-buffered solution, you would expect the pH to:
   (a) Increase immediately.
   (b) Initially remain constant, then begin dropping.
   (c) Decrease gradually but steadily.
   (d) Stay constant for a while, then begin rising.
3. Universal Indicator Solution is used to indicate the pH of liquids: it turns red in acidic solutions, green in neutral solutions, and purple in basic solutions. When Universal Indicator Solution is added to a sample of water from Mystery Lake, the water turns green. Adding some weak acid turns the solution red, but after mixing it returns to green. The most likely reason:
   (a) Mixing increases dissolved oxygen and therefore affects the pH.
   (b) Buffers in the water neutralize the acid.
   (c) The acid dissolves organic matter in the water, releasing compounds which cause the color change.
   (d) pH changes over time in all solutions.
4. Acid precipitation causes greater changes in lake acidity in the Adirondacks than in other parts of New York State because of differences in:
   (a) The type of bedrock and soil.
   (b) The acidity of the precipitation.
   (c) The size and shape of the lakes.
   (d) The lower fish populations.

Authentic Assessment Items

Part I.

Suppose you are the lake manager for an exclusive fishing club. You have read news accounts about acid precipitation, and you are worried about its possible affects on your fish populations. Design a study to determine whether acid precipitation is likely to cause problems in your club's lakes. What will you study, and why?