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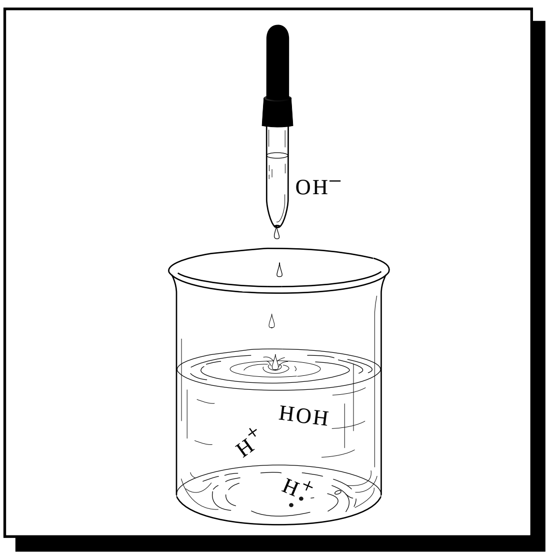

ACIDS AND BASES

ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Hydrangea is a shrub with clusters of showy flowers. The color of its flowers can change as the pH of the soil in which it grows changes. The flowers are red-pink under neutral to low soil acidity and blue under conditions of higher acidity. Hydrangeas contain a natural indicator, anthocyanin, which is responsive to the pH of its surroundings. Observations such as these about hydrangeas can constitute a central theme in understanding acids and bases.

It is useful to think of an acid as a substance that releases hydrogen ions (H^+) in water solution, as proposed by Arrhenius around 1885. Since bases have always been thought of as opposites of acids, Arrhenius also proposed that bases involve hydroxide ions (OH^-). The positive hydrogen ion, which is a proton, and the negative hydroxide ion together constitute water. A more useful definition of acids and bases was proposed by Brønsted and Lowry in 1923. This definition, which gives a more complete picture of acids and bases, treated an acid as a proton donor and a base as a proton acceptor. An acid-base reaction then is essentially a transfer of protons. During the past several decades other useful definitions of acids and bases have built upon these early concepts. Some of these—the Lewis definition, for example—may be beyond the scope of the high school general chemistry classroom and less useful to high school teachers.

Neither hydrogen ions nor hydroxide ions exist alone. They are found combined with other ions of the opposite charge. Hydrochloric acid, for example, produces hydrogen ions and negative chloride ions when in solution. Sodium hydroxide, a base, produces negative hydroxide ions and positive sodium ions when in solution. When solutions of hydrochloric acid (HCl) and sodium hydroxide (NaOH) are mixed, a product resulting from the combination of sodium ions and chloride ions (NaCl) can be separated from the solution. This kind of compound is called a salt. Thus, the products of a typical acid-base reaction are a salt and water. This acid-base reaction, called neutralization, can be represented by the following equation.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + HOH(l)$$

Acids differ from one another in strength. Bases differ from one another in a similar fashion. Chemists have found that some acids produce more hydrogen ions in solution than similar amounts of others. These acids (such as hydrochloric acid, HCl, and sulfuric acid, H_2SO_4) are strong acids, whereas those that release only a few percent of hydrogen ions (hydrofluoric acid, HF, and acetic acid, $HC_2H_3O_2$) are weak acids. Similarly, sodium hydroxide, NaOH, is a strong base and aqueous ammonia, $NH_3(aq)$, is a weak base. Strong and weak acids react with strong and weak bases in neutralization reactions to produce water and salts.

It is most helpful to the chemist to be able to express quantitatively the concentration of hydrogen ion in solution. We can do this with a term called pH. At room temperature, systems below pH 7 are acidic and systems above pH 7 are basic. A sample with a pH of seven is neutral. Special indicator solutions or indicator papers impregnated with certain dyes change colors at various pH values and allow us to estimate the pH of a solution. pH meters provide a more exact measurement of hydrogen ion concentration.

Acids and bases have real-life significance. The human body functions properly only when delicate acid-base balances are maintained; crops grow best in soil with the proper pH; substances released into the atmosphere as pollutants often form acid rain. In addition, many substances used in the home are acids or bases.

Place in the Curriculum

Acids and bases are an important topic in high school chemistry. This module can be used qualitatively to identify substances as acids or bases, in problem solving, or as an extension for related topics such as conjugate acid-base pairs, hydrolysis, and buffers.

Central Concepts

1. **Operational Definitions** Acids and bases are chemical species that exhibit distinctive sets of observable properties. Acids taste sour (like vinegar and lemon), cause blue litmus to turn red, liberate hydrogen gas when they react with certain metals (like iron, zinc, and aluminum), and neutralize bases. Bases taste bitter (like soap), feel slippery to the touch, cause red litmus to turn blue, and neutralize acids.
2. **Conceptual Definitions** Acids and bases can be defined conceptually to help account for what is happening on a microscopic level.

   Arrhenius concept An acid is a substance that, when dissolved in water, forms hydrogen ions (or protons, H^+). A base is a substance that, when dissolved in water, forms hydroxide ions (OH^-). The Arrhenius concept is limited in several ways. Hydroxide ion is singled out as the only source of base character. However, other species can display basic properties (*e.g.*, ammonia). The hydrogen ion cannot exist alone in water solution. It is chemically bonded to water to form what is conventionally written as the hydronium ion, H_3O^+. In fact, the hydronium ion is associated through hydrogen bonding with a variable number of water molecules (*e.g.*, $[H_9O_4]^+$, in which H_3O^+ is associated with three water molecules). Moreover, the Arrhenius concept does not cover acid-base reactions in nonaqueous solvents.

   Brønsted-Lowry concept An acid is a proton donor; a base is a proton acceptor. Any Arrhenius acid is also a Brønsted-Lowry acid, and any Arrhenius base is also a Brønsted-Lowry base. Hydrochloric acid (HCl) is an Arrhenius acid and is therefore, a Brønsted-Lowry acid. As a Brønsted-Lowry acid, HCl can donate a proton to water.

   $$HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq) \qquad (1)$$

   Magnesium hydroxide, $Mg(OH)_2$, is an Arrhenius base and therefore also a Brønsted-Lowry base. As a Brønsted-Lowry base, OH^- can accept a proton:

   $$OH^-(aq) + H_3O^+(aq) \rightleftharpoons 2H_2O(l) \qquad (2)$$

   Ammonia, NH_3, can also serve as a Brønsted-Lowry base, because it can accept a proton from water:

   $$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \qquad (3)$$

   Water can function both as a Brønsted-Lowry acid (Equation 3) and a Brønsted-Lowry base (Equation 1). The Brønsted-Lowry concept of acids and bases is more general than is the Arrhenius concept. In the Brønsted-Lowry concept a base accepts protons (OH^- is only one example). Brønsted-Lowry acids and bases can be ions (*e.g.*, OH^-) or molecules (*e.g.*, NH_3 and H_2O). Some species can act as either acids or bases, depending on the nature of the other reactant (*e.g.*, H_2O). Furthermore, Brønsted-Lowry acid-base reactions are not restricted to aqueous solution. For example, in the reaction between ethyl acetate and sodium ethoxide in ethyl alcohol, ethyl acetate is a Brønsted-Lowry acid and ethoxide ion is a Brønsted-Lowry base.

   $$\underset{\text{BL acid}}{CH_3COOC_2H_5} + \underset{\text{BL base}}{Na^+O^-C_2H_5} \rightarrow {}^-CH_2COOC_2H_5 + HOC_2H_5 + Na^+$$

3. When a suitable amount of base is added to an acid solution, the base and acid properties disappear and the acid is said to be neutralized. A neutralization reaction is the reaction of an acid and a base that results in an ionic compound and possibly water. The ionic compound is called a salt.

4. Salts are made up of a cation (other than H^+) and an anion (other than OH^- or oxide, O^{2-}). The formula of the salt indicates the acid and base that could react to prepare the salt. The cation (positive ion) is derived from the base; the anion (negative ion) is derived from the acid.

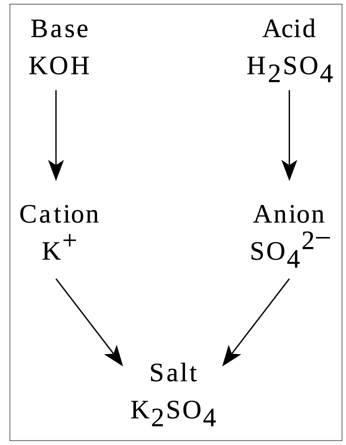

Figure 1. Formation of salt from acid and base.

5. A strong electrolyte is an electrolyte that exists in solution almost entirely as ions. Hydrochloric acid, HCl, is a strong electrolyte. A weak electrolyte is an electrolyte that dissolves in water to form an equilibrium between a molecular substance and a relatively small quantity of ions. Acetic acid and aqueous ammonia are examples of weak electrolytes.

6. In aqueous solution, H^+ (that is, H_3O^+) is the strongest stable acid, and OH^- is the strongest stable base.

7. pH is the negative logarithm of the molar hydrogen ion concentration; pOH is the negative logarithm of the molar hydroxide ion concentration. Mathematically,

$$pH = -\log [H^+] \quad \text{and} \quad pOH = -\log [OH^-]$$

8. $pH + pOH = 14.0$ (at 25 °C)

9. The relative acidity (or basicity) of a solution can be experimentally determined with an acid-base indicator. Indicators are either weak acids or weak bases that change colors over a short pH range. That point in a titration when a suitably chosen indicator changes color is called the *end point*. Some typical acid-base indicators, their acid and base colors, and the pH range for the color change are given in Figure 2.

Indicator	Acid color	Base color	pH range
Bromophenol blue	Yellow	Blue-violet	3.0 - 4.5
Methyl orange	Red	Yellow	3.1 - 4.4
Universal indicator	Red	Blue-violet	4.0 - 10.0
Litmus	Red	Blue	4.7 - 8.2
Methyl red	Red	Yellow	4.8 - 6.2
Bromthymol blue	Yellow	Blue	6.0 - 7.6
Phenol red	Yellow	Red	6.8 - 8.4
Phenolphthalein	Colorless	Red	8.2 - 10.0

Figure 2. Indicator colors and pH range.

10. A pH meter can give quantitative pH values to several decimal places. Usually, a specially designed combination electrode is dipped into the solution to be tested. An electrical potential is generated between the electrodes that depends on the pH. The electrical potential, in volts, is digitally displayed or read on a meter calibrated directly in pH.

Related Concepts

1. **Ions**
   a. An **ion** is an electrically charged atom or group of covalently bonded atoms—for example, sodium ion, Na^+, and carbonate ion, CO_3^{2-}.
   b. An **electrolyte** is a substance that, in water solution, is capable of conducting an electric current. An aqueous solution of sodium chloride conducts an electric current; therefore NaCl is an electrolyte. A **nonelectrolyte** is a substance that does not dissociate into ions in solution and therefore does not conduct an electric current. Sugar is a nonelectrolyte.
   c. **Ionization of covalent substances** is the breaking apart of a covalent substance into ions in aqueous solution. Hydrogen chloride ionizes completely in aqueous solution into hydrogen ions and chloride ions and is thus classified as a strong acid. Acetic acid only slightly ionizes in water; it is a weak acid.
   d. **Dissociation** is separation of an ionic substance into ions in an aqueous solution. For example, sodium chloride, NaCl, dissociates as follows:

$$NaCl(s) + H_2O(l) \rightarrow Na^+(aq) + Cl^-(aq)$$

2. **Solution concentration** The ratio of the quantity of solute to the quantity of solution (or solvent). It may be expressed as molal concentration, molar concentration, weight percent, volume percent, mole fraction, *etc.* The molarity or molar concentration (M) of a solution is the number of moles of solute per liter of solution. A solution of 20.0 g sodium hydroxide, NaOH, in one liter of solution has a concentration of 0.500 mol/L NaOH, or 0.500 M NaOH.
3. Classification and identity of some common acids and bases.

   Strong acids There are six common strong acids; molarities of these acids as they are commonly sold are indicated in Figure 3.

Formula	Name	Molarity	Weight %	Density (g/mL)
HI	Hydriodic acid	7.6	57.0	1.70
HBr	Hydrobromic acid	9.0	50.0	1.52
HCl	Hydrochloric acid (muriatic acid)	12.0	37.0	1.18
HNO_3	Nitric acid (aqua fortis)	15.0	70.0	1.35
$HClO_4$	Perchloric acid	11.6	70.0	1.67
H_2SO_4	Sulfuric acid (oil of vitriol)	18.2	97.0	1.84

Figure 3. Common strong acids.

Weak acids Unless otherwise in-formed, one can assume other common acids are weak. Figure 4 presents three examples:

Formula	Name	Molarity	Weight %	Density (g/mL)
$HC_2H_3O_2$ or CH_3COOH	Acetic acid (glacial acetic acid or vinegar)	17.4	99.7	1.05
HF	Hydrofluoric acid (acid of fluor)	31.3	53.0	1.18
H_3PO_4	Phosphoric acid	14.7	85.0	1.69

Figure 4. Some weak acids.

Strong bases The hydroxides of alkali and alkaline earth elements (except Be). Two common strong bases are shown in Figure 5.

Formula	Name
KOH	Potassium hydroxide (caustic potash)
NaOH	Sodium hydroxide (caustic soda or lye)

Figure 5. Some strong bases.

Weak bases Unless otherwise informed, one can assume other common bases are weak. One example is shown in Figure 6.

Formula	Name	Molarity	Weight %	Density (g/mL)
NH_3	Ammonia (spirit of hartshorn)	15.7	30.0	0.890

Figure 6. A weak base.

4. Equation writing and balancing
5. Stoichiometry
6. Exponents, powers of ten, logarithms

Related Skills

1. Reading volumetric devices
2. Using calculator for exponents, logarithms, arithmetic operations
3. Proper handling of potentially hazardous liquids and solids. Minimum amounts of acids and bases should be kept in the laboratory, preferably only those amounts needed for a given activity.

Performance Objectives

After completing their study of acids and bases, students should be able to:

1. define and classify acids and bases operationally and conceptually.
2. write neutralization equations, given the identities of reacting acids and bases, and complete relevant calculations based on these equations.
3. identify the salt product produced in an acid-base reaction. Given a salt, specify the acid and base from which it could be produced.
4. complete an acid-base titration (small-scale or standard).
5. determine whether a solution is acidic or basic using an indicator or a pH meter.
6. interpret pH values in terms of powers of 10.
7. distinguish between a strong acid (or base) and a weak acid (or base) operationally and conceptually.
8. present useful mental pictures of (a) ionization, (b) dilute *vs.* concentrated acid/base, (c) weak *vs.* strong acid/base, and (d) neutralization.
9. use a matrix to record and interpret experimental data.
10. provide a chemical explanation for acid rain.
11. express their opinions about the roles of acids and bases in the world.
12. explain how their study has changed their opinions about acids and bases and other substances.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Classifying Substances Based on Their Reactions

Introduction

Chemists find it convenient to classify the overwhelming number of known substances into categories that share common properties. In this laboratory activity you will become familiar with some chemical behaviors of certain types of substances that allow them to be classified into acids, bases, and salts.

Purpose

To develop a scheme for classifying different substances as acids, bases, or salts.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. The acids and bases used are corrosive. Follow your teacher's instructions in handling acids and bases and for waste disposal. All spills should be cleaned up immediately.
3 Accidental spills on the body should be flushed with water; notify your teacher.
4. Barium hydroxide is toxic.
5. Dispose of the chemicals as your teacher directs.

Part I. Procedure

1. Arrange a matrix within the well plate or among the test-tubes to match the pattern shown in Figure 7. Add a drop or two of each test solution.
2. Add a drop of phenolphthalein to a sample of each of the seven test solutions. Record your observations.
3. Add a drop of bromthymol blue to a different sample of each of the seven test solutions. Record your observations.
4. Add a drop of universal indicator to a different sample of each of the seven test solutions. Record your observations.
5. Clean a piece of magnesium ribbon with fine-grade sandpaper or steel wool until it is shiny. Cut the ribbon into seven pieces small enough to fit in the test-tubes or wells.
6. Add a piece of magnesium, Mg, to each of the seven test solutions. Record your observations.
7. Add a drop of calcium nitrate solution, $Ca(NO_3)_2(aq)$, to a different sample of each of seven test solutions. Record your observations.
8. Use the conductivity tester on the last sample of each of the seven test solutions. Rinse the conductivity tester in fresh distilled water after testing each sample. Record your observations.

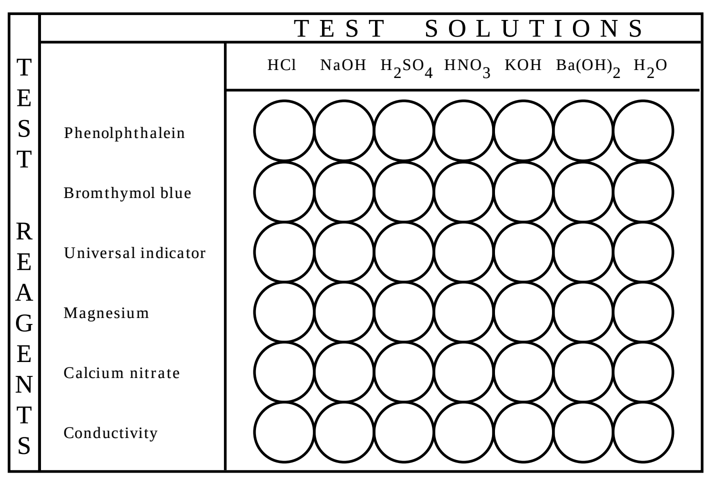

Figure 7. Reactant matrix.

Part I. Data Analysis

1. Use your observations to group the test solutions into as many categories as are needed to account for similarities and differences in chemical reactivity. Be prepared to justify your classification scheme.
2. For those solutions that indicated high conductivity with the tester, identify the ions present in solution.
3. Hydrochloric acid is a member of a class of compounds called acids. Sodium hydroxide is a member of a class of compounds called bases. What other substances in the group of seven tested should be called acids or bases?

Part I. Implications and Applications

1. What strategy could you use to identify an unknown substance as an acid or a base?
2. A drop of bromthymol blue placed in an unknown solution turns yellow. Is the solution acidic or basic?
3. Why do you think universal indicator is called an indicator?
4. What ionic species is common among the group of compounds you categorized as acids? As bases?

Part II. Procedure

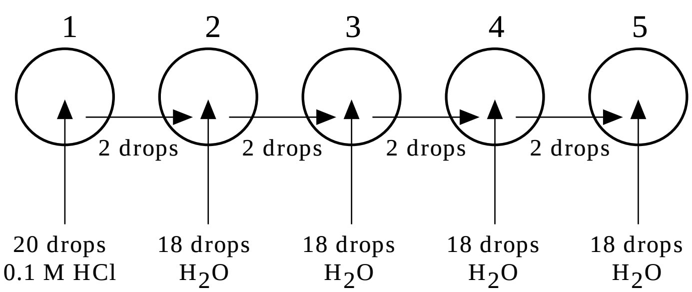

Figure 8. Serial dilution procedure.

1. Add 20 drops 0.1 M HCl to the first well or tube.
2. Remove two drops and place them in the second tube or well. Add 18 drops of water to the second tube or well, and mix.
3. Take two drops of the second solution and add them to the third well. Add 18 drops of water to the third well and mix.
4. Continue this "serial dilution" until you have five increasingly dilute solutions of hydrochloric acid, HCl, as shown in Figure 8.
5. Repeat Steps 1 through 4, starting with 0.1 M sodium hydroxide, NaOH, rather than hydrochloric acid.
6. Add one drop of universal indicator to each well or tube. Record your observations.

		Concentration	Observations
HCl Solutions (Most Concentrated)	1	0.1 M	
	2		
	3		
	4		
	5		
NaOH Solutions	5		
	4		
	3		
	2		
(Most Concentrated)	1	0.1 M	

Figure 9. Sample data table.

Part II. Data Analysis

1. Calculate the molar concentration (mol/L) of each dilution of hydrochloric acid, HCl, and sodium hydroxide, NaOH. Record these values in the chart above. If your teacher so instructs, show how you completed your calculation for the concentration of hydrochloric acid in the second well or test-tube.
2. Propose a convenient pattern to summarize your observations. (*HINT: Think about the colors of the spectrum.*)

Part III. Procedure

1. Add 3-4 drops of each commercial liquid sample to a separate well or test-tube. Label each location or draw a map so you can identify each sample.
2. For commercial solid samples, add an amount about the size of a pencil lead tip into a separate well or small test-tube. Add 3-4 drops of water to each.
3. Add a drop of universal indicator to each well or test-tube. Record your observations.
4. If your teacher so instructs, try any of the other tests you completed in *Part I* with additional samples of the commercial materials.
5. Thoroughly wash your hands before leaving the laboratory.

Part III. Data Analysis

1. Classify each commercial sample as acidic, basic, or neutral.
2. Match the particular concentration of hydrochloric acid or sodium hydroxide from *Part II* that each commercial sample most closely resembles in its acidic or basic properties.

Part III. Implications and Applications

1. How important is it for consumers to understand some fundamental ideas about acids and bases?
2. Is the acidity or basicity of a commercial product usually related to how it is used? Explain.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Classifying Substances Based on Their Reactions

Major Chemical Concepts

This laboratory activity is designed to develop operational definitions for acids, bases, and pH. It should be used as an introduction to acids and bases before students learn verbal or mathematical definitions of important terms. Conceptual definitions can be developed, if desired, during the post-laboratory discussion.

Expected Student Background

Students should be able to calculate the new concentrations of solutions after they are diluted and be familiar with standard laboratory procedures. Students should also be familiar with the electrical conductivity of ionic solutions.

Level

Appropriate for any basic or general student.

Time

This laboratory activity can be completed in two class periods. However, it can occupy over three periods if you use extended discussion to isolate, define, and solidify concept acquisition.

Safety

Read the *Safety Considerations* section in the *Student Version* . In addition, be sure that students treat commercial products as potentially hazardous; some are caustic or can lead to dangerous reaction products if mixed together. For example, if Vanish™ and household bleach are mixed together, poisonous chlorine gas is produced.

Materials (For 24 students working in pairs)

Part I. Materials

- 0.1 M Hydrochloric acid, HCl (2 mL conc HCl, diluted to 250 mL)
- 0.1 M Sodium hydroxide, NaOH (1 g NaOH, diluted to 250 mL)
- 0.1 M Sulfuric acid, H_2SO_4 (see *Advance Preparation*)
- 0.1 M Nitric acid, HNO_3 (1.5 mL conc HNO_3, diluted to 250 mL)
- 0.1 M Potassium hydroxide, KOH (1.5 g KOH, diluted to 250 mL)
- Saturated barium hydroxide solution, $Ba(OH)_2$ [15 g $Ba(OH)_2$, diluted to 250 mL]
- 1.0 M Calcium nitrate, $Ca(NO_3)_2$ [8 g $Ca(NO_3)_2$ diluted to 50 mL]
- Distilled water
- Magnesium ribbon, Mg, 10-cm length
- Bromthymol blue indicator (see *Advance Preparation*)
- Phenolphthalein indicator (see *Advance Preparation*)
- Universal indicator
- 12 Conductivity testers
- 12 Well-plates (96 wells) or 504 small test-tubes and racks
- 12 Small 25-cm^2 pieces of fine-grade sandpaper or steel wool
- 12 Scissors

Part II. Materials

12 Small dropper bottles of each of the following:

- 0.1 M Hydrochloric acid, HCl
- 0.1 M Sodium hydroxide, NaOH
- Distilled (or deionized) water
- Universal indicator
- 12 Well-plates or 120 small test-tubes (can clean ones used previously)
- 60 Toothpicks or small stirring rods

Part III. Materials

12 Small dropper bottles of each of the following:

- A number of commercial samples, *e.g.*, household ammonia (see *Advance Preparation*)
- Universal indicator
- Distilled (or deionized) water
- Well-plate, or small test-tube for each commercial product tested
- Conductivity tester (Directions for building a conductivity apparatus are given in the *Instrumentation* section of this module.)
- Scissors

Advance Preparation

Each student pair should have one dropper bottle of each of these:

- **0.1 M H_2SO_4**. Add 1.5 mL of concentrated (18 M) sulfuric acid, H_2SO_4, to enough water to make 250 mL of solution. *WARNING: Be sure to add acid slowly with continual stirring to several hundred milliliters of water, then make up to 250 mL of solution.*
- **Indicators** Methyl red, bromthymol blue, phenolphthalein, and universal indicator may be purchased as previously prepared solutions from appropriate commercial vendors. Alternatively, prepare methyl red by dissolving 0.02 g in 60 mL of ethanol and 40 mL water; bromthymol blue by dissolving 0.1 g in 16 mL 0.01 M NaOH and adding 234 mL water; phenolphthalein by dissolving 0.05 g in 50/50 ethanol/water. Feenamint™ and Ex-Lax™ are alternative sources of phenolphthalein. Each group should have one dropper bottle of each indicator.
- **A group of household or commercial samples spanning a wide range of pH** Suggestions, with approximate pH values in parentheses, include: liquid Vanish™ (pH 1), solid Vanish™ (pH 2), vinegar (pH 3), Sprite™ (pH 3), orange juice (pH 4), most vinyl cleaners (pH 4), Fluoriguard™ anti-cavity rinse (pH 5), Cepastat™ Mouth Spray (pH 5), most nail polish removers (pH 6), many shampoos (pH 6), many liquid detergents (pH 7 and above), borax solution (pH 9), window cleaner such as Windex™ (pH 10), Dow™ bathroom cleanser (pH 11), many dishwashing detergents (pH 12), Drano™ (pH 14).

Preparing a large number of solutions and filling numerous small dropper bottles with solutions is often a tedious and time-consuming task. Begin this process several days in advance of the activity. If you or your students have not already made conductivity testers (*Instrumentation* section) you should do so in advance of this activity. Be sure the electrodes are small enough to fit into the tubes or wells.

Pre-Laboratory Discussion

Little pre-laboratory discussion is needed with this activity other than a safety review and ensuring that students understand the instructions, including use of the conductivity tester. Make clear to students that they will be classifying materials based on their reactions and insure that they understand in general what "classifying" involves. *Do not* use terminology such as *acid, base, pH, etc.*; the purpose of this activity is to encourage students to invent these concepts from relevant chemical experiences.

Teacher-Student Interaction

The effectiveness of this activity will depend on your interaction with small groups as they work within parts of the procedure and with the entire class as they prepare to move from *Part I* to *Part II*, and from *Part II* to *Part III*.

Part I

Move from group to group to ensure that each group sets up the reaction matrix as instructed. If you use test-tubes rather than well plates, students may need to complete one reaction series and wash the tubes before they conduct the next reaction series. In either case, be sure students record their observations.

To prepare students for *Part II*, point out that most solutions were 0.1 M concentration. Invite students to speculate on the results if similar tests were conducted with solutions that were (a) more concentrated or (b) less concentrated. Consider recording the predictions of several students on the chalkboard before proceeding to *Part II*.

Part II

Circulate to ensure that students correctly complete the dilution series and record their observations.

When students have completed their observations, check on whether they need help in the dilution calculations.

When students have answered *Data Analysis*, Question 2, within their group, have them share their suggestions. When the gradation of colors associated with more and less acidity and less and more basicity is recognized by students, inform them that this is one way chemists identify the property called pH. Tell them that they will develop their understanding of pH as their acid-base study progresses. Universal indicator will continue to serve as a useful way to identify the pH of materials.

Point out that the idea of using universal indicator to determine pH will be used in the *Part III* of the activity. If students are to complete additional tests (in addition to universal indicator) provide suitable instructions.

Part III

As you initially circulate among students, ensure that they use only small samples of household materials and that they do not mix the materials with each other. Encourage groups to share household materials if only limited samples are on hand.

Have students share and compare information. Encourage them to use language such as "has the same color as 0.001 M NaOH" as well as "is an acid" or "is a base."

Both items in *Part III. Implications and Applications* are open-ended; encourage students to evaluate their colleagues' answers. In dealing with the second question, encourage students to relate their hypotheses to their original observations and correlate answers with the pH chart found in *Tips for the Teacher* (Item 8).

Anticipated Student Results

All acids should cause each indicator to exhibit a characteristic color; the bases should cause a different color to appear. Because water can act as both a weak acid and a weak base, its indicator's color will be intermediate to the colors shown by more concentrated and stronger acids and bases.

Magnesium metal will react with acids to produce gas bubbles, but should not react with bases. All acids and bases should be electrically conductive; water is not. Bases used in this activity should produce a precipitate (cloudiness) with calcium nitrate; acids should not.

If you are using universal indicator as recommended, the tested solutions should grade from red through yellow and green to blue and violet, following the visible light spectrum from most acidic (red) to most basic (violet).

Answers to Data Analysis

Part I

1. When analyzing data, student groups should compare classification schemes and then discuss any differences noted. After discussion, acids should be grouped together due to their common indicator colors, reactivity with magnesium, and failure to react with calcium nitrate. Bases should be grouped together based on indicator color, failure to react with magnesium, and reaction with calcium nitrate. Water should be identified as belonging to neither group. At this point, *and not before this point*, introduce the terms "Acid," "Base," and "Neutral." These terms, as used by chemists, refer to the properties (and others not yet encountered) just observed by the students.
2. The ions in solution are: HCl(aq) contains H^+(aq) and Cl^-(aq), NaOH(aq) contains Na^+(aq) and OH^-(aq), H_2SO_4(aq) contains H^+(aq) and HSO_4^-(aq) [remember HSO_4^-(aq) is a weak acid] and some SO_4^{2-}(aq), HNO_3(aq) contains H^+(aq) and NO_3^-(aq), KOH(aq) contains K^+(aq) and OH^-(aq), and $Ba(OH)_2$(aq) contains Ba^{2+}(aq) and OH^-(aq).
3. HCl(aq), H_2SO_4(aq), and HNO_3(aq) are acids; NaOH(aq), KOH(aq), and $Ba(OH)_2$(aq) are bases.

Part II

1. For the HCl solutions: Solution 1 is 0.1 M, 2 is 0.01 M, 3 is 0.001 M, 4 is 0.0001M, 5 is 0.00001 M. For the NaOH solutions: Solution 5 is 0.00001 M, 4 is 0.0001 M, 3 is 0.001 M, 2 is 0.01 M, 1 is 0.1 M.
2. Follow the spectrum from red to violet as the solutions become less acidic and more basic.

Part III

1. The results depend on which household materials are tested. For general results see the pH chart in *Tips for the Teacher* (Item 8).

Post-Laboratory Discussion

The suggested class development outlined above is intended to introduce the concepts of acid, base, pH, and possibly indicators from phenomena students directly observe in the laboratory. Thus the post-laboratory discussion can employ a form of "stand-and-deliver" strategy in which you call on students to define main concepts and illustrate definitions with examples from reactions just studied. Other students can be invited to evaluate or extend previous comments by their colleagues.

Students should also be asked occasionally to give examples of acid-base phenomena that were not observed in the laboratory activity to see whether they can extend the concepts.

Answers to Implications and Applications

Part I

1. Use one of the three indicators to test for an acid or a base.
2. Bromthymol blue changes to yellow in an acid solution or in a solution of pH less than 6 (see Figure 2).
3. In solutions of different concentrations of acids or bases, the universal indicator has different colors.
4. All acids seem to have H^{+}(aq), and all bases seem to have OH^{-}(aq).

Part III

1. Many household products display acidic or basic properties.
2. Most commercial cleansing products are basic. Most fruits tend to be acidic.

Possible Extension

A good group activity is to conduct an acid-base scavenger hunt. Send students out with a well plate or test-tubes and universal indicator to identify five additional acids and five additional bases in their everyday environment. They can complete this as a take-home activity, if you choose.

Assessing Laboratory Learning

1. **Laboratory practical.** Give individuals or groups of students one or more unknowns. Have students categorize the samples as acidic and/or basic, indicating which solution(s) from *Part II* had comparable pH values.
2. **Demonstration examination.** Same as laboratory practical, but you demonstrate the tests. Carry out a series of reactions, describing at each step what you are adding and allowing students to observe the results. Then have students categorize and/or identify the pH of the unknown substance.
3. **Written examination.** Here are some possible items:
   a. Three identical samples of a clear, colorless liquid are tested: Magnesium is added to Sample 1; no reaction is observed. To Sample 2 is added calcium nitrate; the liquid turns cloudy. Universal indicator is added to Sample 3; the indicator turns bluish-purple. Is the unknown liquid acidic, basic, or neutral? *[Basic]*
   b. Universal indicator can be made to turn various colors, depending with what it is mixed. What property determines the color shown by universal indicator? *[Whether the substance being tested is an acid or a base.]*
   c. In what ways are some materials commonly available in grocery stores similar to materials commonly used in a chemistry laboratory? *[Some materials are acids (vinegar, 7-Up, orange juice); some are bases (baking soda, milk of magnesia).]*

Suggestions for Other Laboratory Activities

These suggested activities are in outline format. Details for starred (*) activities are found in Shakhashiri, *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (see *References*). Suggested demonstrations in the section that follows could also be used as laboratory activities. *CAUTION: Use appropriate safety guidelines in performing these activities, since only brief procedural steps are described here.*

1. **Properties of acids and bases*** Arrange eight beakers in two sets of four. Add 20 mL each of the following to separate beakers: water, 2 M HCl, 2 M H_2SO_4, and 2 M $HC_2H_3O_2$. Test each solution with litmus paper. Add several magnesium turnings, Mg, to each beaker in one set of the four liquid samples. To each beaker in the second set of four, add 10 drops of litmus solution. Pour 25 mL 6 M NaOH in each beaker. Add magnesium turnings to each beaker and compare the results with the previous set.
2. **Define acid, base, and pH operationally** This objective can be accomplished with cabbage juice and/or universal indicator in household materials, hydrochloric acid, and sodium hydroxide at various concentrations, as outlined: Decant the purple liquid from canned red cabbage. Add 5 mL of the extracted juice to (1) 125 mL vinegar, (2) 125 mL household ammonia, (3) 1 teaspoon baking soda in 125 mL distilled water, (4) 125 mL colorless, carbonated beverage, (5) 5 mL of laundry detergent in 125 mL distilled water, (6) 125 mL milk, (7) 125 mL 0.1 M HCl, and (8) 125 mL 0.1 M NaOH. Record the color of each mixture. Decide on the indicator color of red cabbage juice in acidic and basic solutions. Repeat with universal indicator.
3. **Neutralization** Titrate 20 mL 0.1 M HCl with 0.1 M NaOH using phenolphthalein as indicator. Repeat with 0.1 M H_2SO_4. Determine which titration requires more base. Explain results using balanced equation for neutralization of each acid by NaOH.
4. **Titration of an antacid*** Weigh an antacid tablet accurately to nearest 0.01 g. Place tablet in an Erlenmeyer flask and add 100 mL of 0.15 M HCl to dissolve tablet. Add 5 drops of bromophenol blue indicator solution. If the indicator is blue, add HCl until it turns yellow. Record volume of HCl added. Titrate the sample with 0.1 M NaOH until the solution just turns blue. Calculate the mass of HCl solution consumed per gram of antacid tablet.
5. **Distinction between acid strength and acid concentration*** Separate beakers are filled with 40-mL samples of 0.1 M HCl, H_2SO_4, and $HC_2H_3O_2$. Ten drops of phenolphthalein are added to each. The pH of each solution is determined with a pH meter or with pH test paper. 0.1 M NaOH is slowly added to each beaker of acid until the indicator just turns from colorless to pink. Compare the volumes of NaOH needed for each acid. Explain the results.
6. **Strong vs. Weak Acids*** Add 25 mL of 6 M, 1 M, 0.1 M HCl (strong acid); 6 M, 1 M, 0.1 M $HC_2H_3O_2$ (weak acid) to separate beakers. Measure pH. Check conductivity of solutions and rank. Add Mg turnings to each beaker. Rank by reactivity.
7. **Acid and Base Indicators*** Mix 10 mL indicator solution with 50 mL colorless solutions having pH values ranging from 1 to 13. For indicator solutions that change color below pH 7 use bromophenol blue or methyl orange. For indicator solutions that change color near pH 7 use litmus or phenol red. For indicator solutions that change color above pH 7 use phenolphthalein. Compare colors in solutions of various pH.

8. **pH of Familiar Products*** Prepare standards for comparison by adding 10 drops of universal indicator to 50 mL of standard solutions representing pH values of 1, 3, 5, 7, 9, and 11. Add 10 drops of universal indicator to 50 mL of familiar products (fruit juice, vinegar, carbonated colorless beverage, milk, mouthwash, ammonia, bleach). Dissolve 5 mL (one teaspoon) of any of the following in 100 mL water: laundry detergent, hand soap, shampoo, drain cleaner, baking soda, antacid, aspirin. Add 10 drops of universal indicator to 25 mL of each. Compare colors for pH. Use a pH meter to determine accurately.
9. **Concept of pH** Serial dilution of an acid; measure pH with (1) pH paper, (2) pH meter, (3) various indicators, and/or (4) with assistance of a computer interface.

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Metal-Acid Reactions

This demonstration illustrates that most active metals replace hydrogen from an acid.

Materials

6 M Hydrochloric acid, HCl
Small samples of any four metals (Cu, Mg, Al, Zn, Fe, or Sn)
Slant-top overhead apparatus or Petri dishes (or 50-mL beakers or titer plates)

Safety

6 M HCl is corrosive and must be handled with care. Goggles should be worn.

Directions

Place vessels on an overhead projector. Add 6 M HCl to each vessel, followed by the sequential addition of the metal samples. Observe; ask students to explain the results.

Demonstration 2: Nature of Indicators

Many common foods have natural indicators.

Materials

Natural indicators Red cabbage juice, grape juice, or fillings from berry fruit pies
Vinegar
Baking soda (sodium bicarbonate), $NaHCO_3$
6 Petri dishes, slant test-tube apparatus, or 50-mL beakers

Safety

Most of the materials are very weak acids or bases, but goggles should still be worn.

Directions

Place natural indicators in each of three dishes, test-tubes, or beakers. You might want to dilute the fruit fillings with a little distilled water and use the resulting mixture as the sample. Add vinegar to one solution, add bakingsoda to another, and use the third as a control. Have students make observations and state their conclusions.

Demonstration 3: Indicators and the pH Scale

Different indicators change colors at different pH values.

Materials

Universal indicator solution
Bromthymol blue indicator solution
Phenol red indicator solution
Other indicator solutions if desired, or solutions of known pH: 1, 4, 7, 10

Safety

Goggles should be worn even though these solutions are weak acids and bases.

Directions

Line up twelve Petri dishes (or small beakers) on the overhead projector as shown in Figure 10.

O O O O
O O O O
O O O O

Figure 10. Arrangement of Petri dishes.

Place a separate solution of pH 1, 4, 7, and 10 in each row. Add universal indicator to each Petri dish in the first column, bromthymol blue to each dish in the second column, and phenol red to each dish in the third column. Add one to three drops (or the amount of indicator needed to produce an observable color) to each Petri dish. Either invite students to make their own observations and interpretations or, alternatively, interpret the results for them.

Demonstration 4: Conductivity

Acids and bases conduct an electric current.

Materials

Conductivity apparatus (see *Instrumentation*)
Beaker, 250-mL
10 Beakers, 50-mL
0.01 M or 0.1 M Sulfuric acid, H_2SO_4
0.01 M or 0.1 M Barium hydroxide, $Ba(OH)_2$
Phenolphthalein solution, 1%
0.1 M Hydrochloric acid, HCl
0.1 M Acetic acid, $HC_2H_3O_2$
0.1 M Aqueous ammonia, NH_3
0.1 M Sodium hydroxide, NaOH
Glacial acetic acid, $HC_2H_3O_2$
Selected household products (*e.g.*, Drano™, vinegar, *etc.*)

Safety

Many commercial products can be corrosive to the skin or damage clothing. Apron and goggles should be worn.

Conductivity Test Directions

Test the 0.1 M solutions and selected household products for their relative conductivities (do not test H_2SO_4 or $Ba(OH)_2$; save these for the conductivity titration). Use a clean beaker for each solution. Test glacial acetic acid last in this manner: test its conductivity, then double the volume by adding distilled water and test the conductivity again. Continue in this fashion until adding water no longer changes the conductivity of the acetic acid.

Conductivity Titration Directions

For chemical and electrical reasons this activity should be done only as a teacher demonstration. Set up clean burets using a ring stand and buret clamp. Fill one buret with H_2SO_4 solution and the other with $Ba(OH)_2$ solution. It is not necessary to fill each buret completely, but at least 40 mL of each solution should be present. Add 30-40 mL of H_2SO_4 from the buret to the 250-mL beaker.

Add two to three drops of phenolphthalein indicator solution. Stir. (Ifavailable, a magnetic stirrer will be helpful.) Test the conductivity. (It is best to use a conductivity tester with a 60-W light bulb for this and to leave the electrodes immersed in the liquid.) Slowly add $Ba(OH)_2$ solution from the other buret with stirring. Note the conductivity as the base solution is added. Stop the $Ba(OH)_2$ addition as close as possible to the neutrality point. The light bulb should glow more dimly until it completely stops glowing. This is the neutrality point. The phenolphthalein color should become visible at the same point.

After students have time to observe and record their observations, add additional $Ba(OH)_2$ solution. The light bulb should begin to glow increasingly brightly and the phenolphthalein color should become more pronounced. Ask students to propose explanations for the behavior of the conductivity light bulb.

Demonstration 5: pH and Environmental Gases

Dissolved gases often produce acidic or basic solutions.

Materials

5 Flasks, conical or round, 250-mL
Universal indicator solution
Gas sources for CO_2, NH_3, and SO_2
Dry ice, if available

Safety

Dry ice can cause severe frostbite and should be handled with gloves or tongs. Other gas sources need to be handled carefully. Use the fume hood. Goggles should be worn.

Directions

Add 100 mL of distilled water to each flask. Add universal indicator solution to produce an intense color. Bubble each individual gas, one at a time, into the distilled water in the separate flasks. Add a piece of dry ice to the fourth flask. The fifth flask can be left with distilled water as a control. Have students observe and interpret the results.

Demonstration 6: pH Measurement

pH can be measured in several convenient ways.

Materials

Narrow and broad range pH paper
pH pen or pH meter
Solutions with various pH values

Safety

Become familiar with the operating instructions for the pH measuring device.

Directions

If a pH meter is available and used, show students not only how to use the meter but also how to calibrate the instrument at a pH close to the value to be measured. Adapt this demonstration to your needs and available pH measuring devices. Indicators may also be used to supplement the pH paper or meter observations.

Demonstration 7: Silent Demo—Reactivity and Concentration

As the concentration of an acid decreases, reactivity with metals decreases also.

Materials

6 M Hydrochloric acid, HCl
Three samples (about 4 cm) of magnesium metal ribbon
An overhead projector or a light box

Safety

6 M HCl is corrosive and must be handled with care. Goggles should be worn.

Directions

Inform the students that this will be a silent demonstration for both the teacher and students. Information can, however, be written on the board. Place one piece of Mg ribbon into 50 mL HCl in a 250-mL beaker. With a large-face timer or stop-watch, indicate the time needed for the Mg to react completely on the board. Show the students a second beaker with 50 mL water; then pour the acid into the water beaker. Again note the time required for a complete reaction. The teacher can continue the dilutions as long as there is an observable reaction. The discussion should contain other examples where concentration is a big factor in reaction rates and toxicity. Hopefully the students will begin to understand the importance of developing a quantitative method of expressing concentration, which will follow in the lesson.

Suggestions for Other Demonstrations

Bo, Sk, Su, and Sq refer to monographs by Borgford, Shakhashiri, Summerlin, and Sarquis listed in *References*. *CAUTION: Check each reference for appropriate safety guidelines before performing these demonstrations.*

Natural indicators Extractions of colored dyes from plants that can function as acid-base indicators. *[Sk v3 50; Sq CF 27; Bo CA 92]*

Properties of acids and bases Demonstrates classical properties of acids and bases. *[Sk v3 58; Su v2 171]*

pH of familiar products pH measurement of household chemicals using indicator solutions or a pH meter reveals that many substances are acids and bases. *[Sk v3 65; Bo CA 133]*

Disappearing ink A blue “ink” spot on white cloth is caused to fade and disappear. *[Su v2 174; Sq CF 325]*

Invisible painting A pattern is painted on a sheet of paper with phenolphthalein (colorless), and the pattern is developed when the paper is sprayed with sodium hydroxide solution. *[Sk v3 47]*

Cylinders of color Create color changes in solutions containing several different acid-base indicators. *[Sk v3 114 (uses dry ice); Sq CF 127 (uses Alka Seltzer™), Su v1 38]*

Acid-Base nature of common gases The properties of several gases are revealed: carbon dioxide in breath *[Bo CA 15]*, sulfur dioxide *[Su v2 165]*; and ammonia *[Sq CF 59]*.

Sealed-bag reactions with acids and bases A reaction occurs among three substances mixed in a sealed plastic bag *[Sq CF 345]* or between two substances in a sealed plastic bag *[Sk v3 100]*.

Acid-base titration/conductivity Conductometric titrations of barium hydroxide and sulfuric acid. *[Sk v3 152]*

Making hydrogen gas from an acid and a base Hydrogen gas is collected in a balloon attached to each of two flasks. One flask contains aluminum foil and hydrochloric acid; the other, aluminum foil and sodium hydroxide. *[Su, v2, 33]*

Acid-base indicators: A voice-activated chemical reaction Allow students to speak into a flask containing bromthymol indicator. After several tries, the solution will turn yellow. *[Su v1 40]*

Miscellaneous applications Measuring acid rain *[Su CA 191]*, measuring soil pH *[Bo CA 194]*, what's in baking powder *[Bo CA 223]*, carbonates in rocks, minerals, and eggshells *[Bo CA 17; Sq CF 305]*.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What are common observable properties of acids and bases? *[Acids taste sour, cause blue litmus to turn red, liberate hydrogen gas when they react with certain metals, and neutralize bases. Bases taste bitter, feel slippery, cause red litmus to turn blue, and neutralize acids.]*
2. Identify acids and bases you have encountered in and around your home. *[Baking soda (sodium bicarbonate, $NaHCO_3$, base), drain cleaner (sodium hydroxide, NaOH, base), vinegar (acetic acid, $HC_2H_3O_2$, acid), citrus fruits (citric acid, acid), battery acid (sulfuric acid, H_2SO_4, acid), antacids (*e.g., *Milk of Magnesia, $Mg(OH)_2$, base), household ammonia (aqueous ammonia, NH_3, base)].*
3. How does the concentration of hydrogen ion vary as pH changes? *[As the pH increases, the concentration of H^+ decreases.]*
4. From what acid and base could we synthesize sodium nitrate, $NaNO_3$? *[The anion of the salt comes from the base and the cation comes from the acid. For $NaNO_3$, the sodium ion, Na^+, is supplied by sodium hydroxide, NaOH, and the nitrate ion, NO_3^-, is supplied by nitric acid, HNO_3.]*
5. What are some problems associated with "acid rain"? *[Acid rain is a dilute solution of sulfuric acid and nitric acid. Because acids are corrosive, acid rain can cause damage to humans, metal and brick buildings, marble and limestone structures, and the environment (kill trees and crops, change pH of soil,* etc. *See "Checking Out Acids and Bases" in* ChemMatters, *April 1983.]*
6. Why do some shampoo bottles say "acid-balanced"? *[These shampoos claim to maintain the normal acid pH of hair and contain an acid, usually citric acid, to counteract the alkalinity of the detergents. See "Checking Out Acids and Bases" in* ChemMatters, *April, 1983.]*
7. What do antacids do? Are there any problems with taking too much? *[Antacids neutralize excess stomach acid brought on by stress, overeating, or disease. Some antacids contain sodium ion (Alka Seltzer™), which can aggravate hypertension. Milk of Magnesia, $Mg(OH)_2$, in large doses has a laxative effect. See "Checking Out Acids and Bases" in* ChemMatters, *April 1983.]*

Counterintuitive Examples and Discrepant Events

1. Diluting a weak electrolyte ($HC_2H_3O_2$) with water increases the electrical conductivity (see *Suggested Laboratory Activity* , Strong *vs.* Weak Acids).
2. Carbon dioxide bubbled through limewater causes a precipitate to form. Continued bubbling causes the precipitate to disappear. [See Underground sculpture. (1984). *ChemMatters,1* (2), 10-11.]
3. Some active metals react with either acids or bases to produce hydrogen gas. (For example, aluminum will react with either hydrochloric acid or sodium hydroxide, releasing hydrogen gas. See *Suggestions for Other Demonstrations* , Making hydrogen gas from an acid and a base.)
4. The same amount of hydrogen gas will be produced when a sample of an active metal is added to equal volumes of concentrated acid solution and dilute acid solution (*if* the metal is the limiting reactant).
5. Bicarbonate salts can be used to neutralize either an acid or a base. This can be illustrated by the following equations:

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$$

$$HCO_3^-(aq) + OH^-(aq) \rightarrow H_2O(l) + CO_3^{2-}(aq)$$

Metaphors and Analogies

1. $[H^+]$ *vs.* pH: A see-saw relationship (see *Pictures in the Mind*); one goes up (increases) while the other goes down (decreases).
2. A proton shifting from an acid to a base can be likened to a baseball being thrown from a pitcher (the acid) to a catcher (the base).
3. Universal indicator color changes follow the colors in the rainbow as the pH moves from 2 to 10. The name ROY G BIV helps keep the colors straight: Red, Orange, Yellow, Green, Blue, Indigo, Violet. A pH of 7 produces a yellow-green hue.

Pictures in the Mind

1. **Ionization** Graphical pictorial representation of the behavior of acids of different strengths in aqueous solution.

   a. A strong acid is completely dissociated.

   b. In a weak acid only a small fraction of molecules is dissociated in solution.

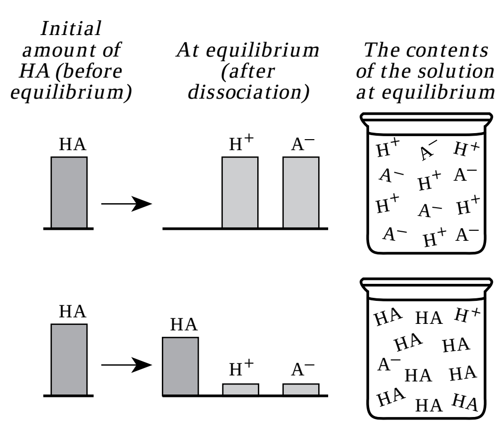

Figure 11. Picturing ionization of strong acids (upper) and weak acids (lower).

2. Pictures of the composition of a concentrated-strong acid, a dilute-strong acid, concentrated-weak acid, and a dilute-weak acid.

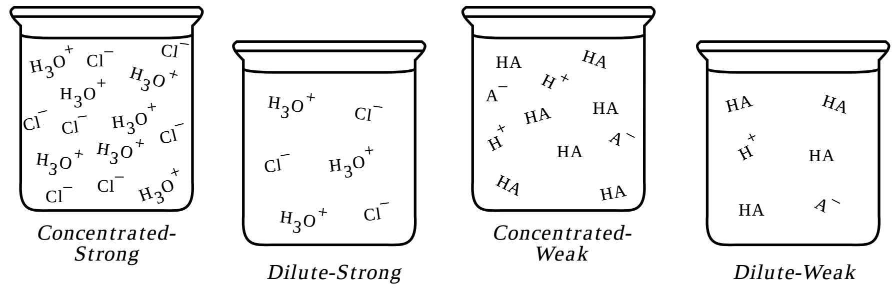

Figure 12. Picturing acid strength and concentration.

3. pH rectangle

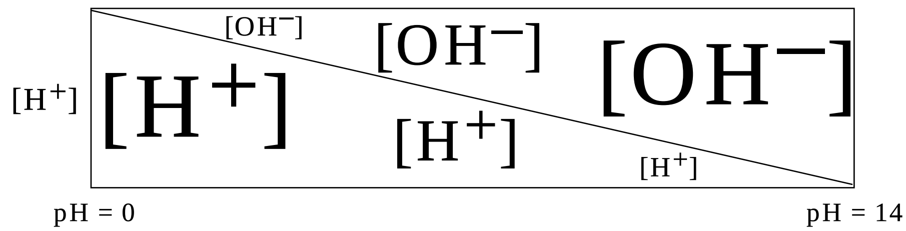

Figure 13. The pH rectangle.

4. $[H^+]$ versus $[OH^-]$: A see-saw

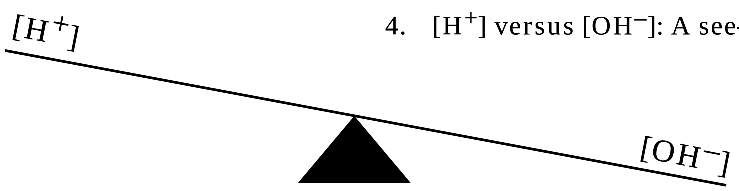

Figure 14, See-saw demonstrates inverse relationship of $[H^+]$ and $[OH^-]$.

5. pH versus $[H^+]$: a see-saw

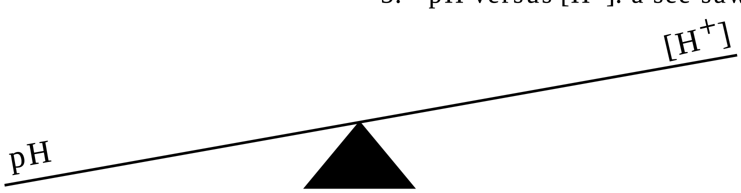

Figure 15. A pH/$[H^+]$ see-saw.

6. Bronsted-Lowry "Umbrella"

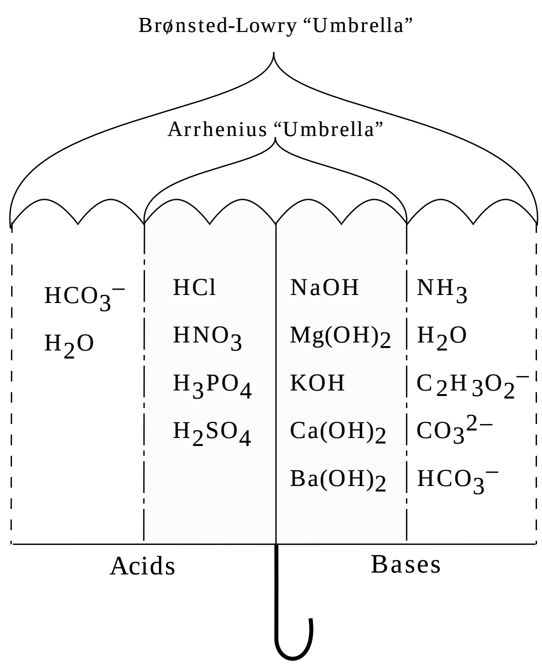

Figure 16. Umbrellas of acid-base definitions.

Language of Chemistry

TIPS FOR THE TEACHER

Brønsted-Lowry acid proton donor.

Brønsted-Lowry base proton acceptor.

concentration amount of solute in some specified amount of solution.

end point point in a titration when the indicator changes color.

indicator weak acid or base that changes colors over a narrow pH range.

Lewis acid electron pair acceptor.

Lewis base electron pair donor.

molarity concentration term defined as the number of moles of solute per liter of solution.

neutralization reaction between an acid and a base in which the acid and base properties disappear.

strength percentage of ionized molecules of acid (or base).

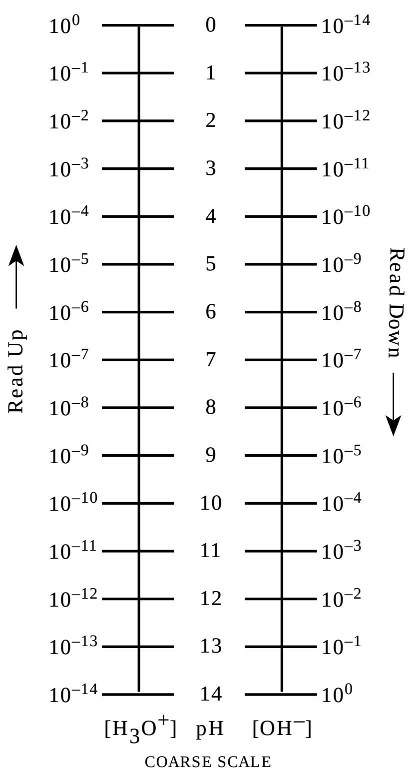

Figure 17. Scales for pH-concentration relationships.

1. The word acid comes from the Latin *acidus*, for sour. One property of acids is their sour taste, such as that experienced with yogurt, pickles, and lemons.
2. The terms strong and weak are not synonymous with concentrated and dilute, respectively (see *Common Misconceptions*).
3. For consecutive integer values of pH, H^+ concentrations differ by a factor of 10. For example, solutions of pH = 1 ($[H^+] = 10^{-1}$ M) and pH = 2 ($[H^+] = 10^{-2}$ M) have a H^+ concentration ratio of 10, that is, $10^{-1}/10^{-2} = 10$.
4. Solutions with high hydrogen ion concentration have a low pH. A solution with pH = 1 has a $[H^+] = 10^{-1}$ M; a solution with pH = 6 has a $[H^+] = 10^{-6}$ M. Figure 17 provides convenient conversion scales between pH and concentration.
5. pH-pOH square. Equations over or next to arrows are conversion equations. Given one corner, the other can be calculated using the conversion equation.

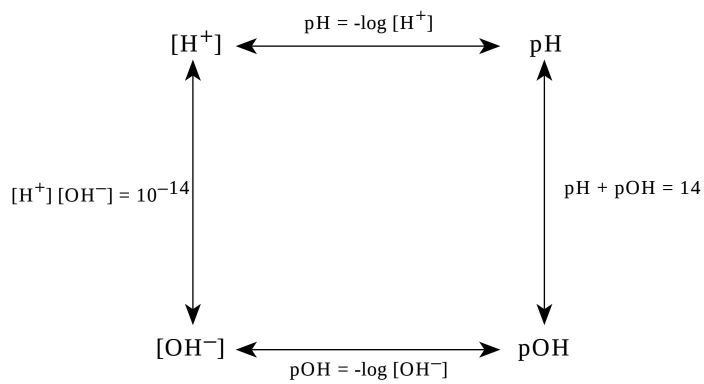

Figure 18. pH-pOH square.

6. A pH rhyme from Kolb, D. (1979). *Journal of Chemical Education*, *56*, 53:

For coffee it's 5, for tomatoes it's 4;
While household ammonia is 11 or more.
It's 7 for water, if in a pure state,
But rain water is 6, and sea water is 8.
It's basic at 10, quite acidic at 2,
And well above 7 when litmus turns blue.
Some find it a puzzlement. Doubtless their fog
Has something to do with that negative log.

7.

Solution	pH
Battery acid	0
Stomach acid	1.4 - 1.8
Lemon juice	2.1
Vinegar	2.9
Soda water	3
Wine	3.5
Tomato juice	4
Black coffee	5
Urine, sour milk	6
Rain water	6.5
Pure water (25 °C)	7
Blood	7.35 - 7.45
Sea water	8
Soaps, shampoos	8 - 9
Detergents	9 - 10
Milk of magnesia	10
Household ammonia	11.9
Liquid bleach	12
Household lye	14

Figure 19. Approximate pH of some common solutions.

8. Remember the color litmus becomes in acidic and basic solutions:

Blue Base
reD aciD

or

aci
reD

ase
Blue

9. Generally, concentrations of acidic and basic solutions are expressed as moles per liter (M). Commercial acids and bases are sold as percent by mass. For example, concentrated hydrochloric acid is 37% by mass. The concentration of sulfuric acid in rain water is measured in parts per million (ppm). (See *Introductory* module for conversion table.)

10. Practice writing balanced equations for neutralization reactions.
11. Review safety precautions when handling acids and bases.

Common Student Misconceptions

1. **"Substances containing H are acidic; substances containing OH are basic."**

   Many substances that contain H are not acids and many substances that contain OH are not bases. Table sugar (sucrose), $C_6H_{12}O_6$, contains H and OH; however, when dissolved in water it dissolves as intact molecules and does not ionize to produce any H^+ or OH^- ions, as shown by the equation: $C_6H_{12}O_6(s) + H_2O(l) \rightarrow C_6H_{12}O_6(aq)$. Since sugar molecules do not lose any H^+ or OH^- ions, sugar is not considered an acid. Alcohols have a characteristic hydroxyl group, OH, covalently bonded to the rest of the molecule. When alcohols dissolve in water they also dissolve molecularly and do not ionize. Since no OH^- ions are released, alcohols are not bases—they actually are weak acids! Another approach is to consider the following substances containing OH groups: NaOH, CH_3OH, CH_3COOH, and $SO_2(OH)_2$. The first is a strong base in water, the second (methanol) has neither acidic or basic properties in water, the third (acetic acid) is a weak acid in water, and the latter (H_2SO_4) is a strong acid in water. The concept of bonding (see *Chemical Bonding* module) is useful with this approach.

2. **"When a 'proton donor' acid reacts, the nucleus of an atom loses a proton."**

   When we speak of acids as proton donors we are talking about the single proton in a hydrogen ion, H^+, not a proton from the nucleus of some other atom. Although chemists commonly refer to acids and bases as proton donors and proton acceptors, it is important to realize that H^+ is being moved from one chemical species to another, not a proton from one nucleus to another. Furthermore, if nuclear protons *were* transferred, then elements would change their identities in acid-base reactions, since the atomic numbers of the atoms involved would be altered. Another effective way to deal with this confusion is to have students draw a "structure" for the hydrogen atom. Attention should be drawn to what is in the hydrogen nucleus, and students ask, "If we remove an electron from a hydrogen atom, what is left?" It is then fairly easy for students to understand that a hydrogen ion is a bare proton, but not the proton drawn from the nucleus of just any atom.

3. **"Strength and concentration mean the same thing."**

   These two terms are often confused by students. Concentration refers to the amount of solute in some specified amount of solution. Strength refers to the percent of molecules that ionize and form ions in solution. Hydrochloric acid is considered a strong acid because in aqueous solutions nearly all the molecules ionize to produce H^+ and Cl^- ions. Acetic acid is considered a weak acid because only about 1% of the molecules normally ionize. It is interesting to note that the percent of acid molecules that normally ionize in a weak acid actually increases as the acid concentration is decreased through dilution with water. Thus, acetic acid actually becomes slightly stronger as the solution is diluted.

4. **"Concentration and amount mean the same thing."**

Some students have difficulty distinguishing the concepts of amount and concentration. By international convention (SI), the *amount of a substance* is measured in *moles*. Solution concentration is frequently expressed as Molarity (M), which is a ratio of moles of solute to liters of solution. Thus, simply stating Molarity does not give amount of substance; the volume must also be known. The amount of acid, in moles, is found from the relationship:

Moles = Molarity X Volume (in liters)

Thus, the amount of solute in these solutions: 1.0 L of 2.0 M
2.0 L of 1.0 M
0.50 L of 4.0 M

is 2.0 mol in each. It is apparent that the concentrations, 2.0, 1.0, and 4.0 M, are not identical. Concentration is not, then, the same as amount. This is emphatically illustrated by discussing the relative effects of 1 drop of concentrated sulfuric acid (0.05 mL of 18 M H_2SO_4) and 9 L of 0.1 M sulfuric acid. While both contain 0.9 moles of acid molecules, a single drop of the former solution causes a painful burn, while the latter is harmless. Demonstrate this using a paper towel rather than skin! The concentrated acid eats through the paper, leaving a carbonized patch; but dropping a towel into a plastic bucket of the diluted acid (1 drop in 9 L) causes no change.

You may choose to illustrate concentration *vs.* amount using the following visual system, which does not refer to acid-base concepts, but uses food coloring (see *SourceView:* Problem Solving, also). Place 3-4 drops of food coloring in a 100 to 250-mL graduated cylinder filled about a third full with water. Stir. Set this on an illuminated overhead projector stage. Add water, having students observe from top and side. Ask if any more dye was added. Point out that the side view shows the dilution effect while the top view (color constant) indicates the total amount of dye substance. For an explanation of how color and concentration are related, read on.

The color varies according to Beer's Law.

A = *abc* where,

A = Absorption
a = Molar absorptivity (a constant for a given colored species)
b = Light path length or solution "thickness," and
c = Concentration of colored species

As water was added, *a* remained constant, *c* decreased, cross section *b* remained constant. Thus from the side, A = *abc* decreased. From above, however, *b* increased as much as *c* decreased so A = *abc* was constant.

5. **"Ammonium hydroxide is the name of the solution formed when ammonia gas dissolves in water."**

Ammonia gas does readily dissolve in water. After it has dissolved, the equilibrium represented by the following expression is eventually attained.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq).$$

Sometimes this solution is called ammonium hydroxide. However, the equilibrium lies far to the left; few ammonium ions, NH_4^+, or hydroxide ions, OH^-, are present. The compound NH_4OH has not been isolated. Water solutions of ammonia are weak electrolytes. Aqueous ammonia is a better name. Ammonia is a weak base.

6. **"A neutral solution always has a pH equal to 7."**

   The percent of water molecules that ionizes depends a great deal on temperature. Only at 25 ^{o}C does $[H^+] \times [OH^-] = 10^{-14}$, giving a pH of 7 at neutrality. As temperature increases so does the percent of water molecules that ionize. Thus, at higher temperatures the pH of neutrality decreases. At body temperature, 37 °C, the pH of neutral water is 6.81.

7. **"A neutralization reaction produces a neutral solution."**

   A neutralization reaction is one in which acid and base properties disappear when an acid solution is added to a base solution. A solution of an ionic compound (a salt) results. This solution may or may not be "neutral' (that is, have a pH of 7; see the Hydrolysis section under *Extensions*).

8. **"A pH = 0 means that no acid or base is present."**

   A solution with a pH = 0 would be *very* acidic—in fact, the solution would have a hydrogen ion concentration of 1 M. [pH = –log(1) = 0]

9. **"The compound HNaO would be considered an acid."**

   Students often think that the way a compound's formula is written determines whether the compound is an acid or a base. Because of the way that Na, O, and H interact, they produce a compound commonly known as sodium hydroxide. No matter how we write the formula (HNaO or NaOH), the same basic compound is produced. A similar confusion can arise when acetic acid is written as $HC_2H_3O_2$ or CH_3COOH. In both cases an acidic compound is represented. The difference is that the first is written to correspond to the way most mineral acids are written, while the second gives a better indication of its actual organic carboxylic structure. It's not how we write the formulas, but how the atoms in the structure actually interact that determines the properties of compounds.

10. **"The end point and neutralization point in a titration are the same."**

    Not so—neutralization involves consuming all acid and base; end point depends on indicator color change.

11. **"A concentrated acid is a pure substance."**

    Acids are usually used in aqueous solutions and are mixtures, not pure substances.

12. **"Concentrated acids are much more dangerous than concentrated bases."**

    Most beginning chemistry students have some knowledge of the dangers of concentrated acids. (Battery acid, horror movies, *etc.*) These same students have little or no knowledge of bases and their properties. Students should be reminded that most "bowl cleaners" are concentrated bases because they are good at dissolving organic material. When students realize they are made up of organic material, they begin to build a healthy respect for bases.

13. **"Only pH values of 1-14 exist."**

    The definition of pH ($-\log[H^+]$) permits values of pH outside the 1-14 range, although they are seldom encountered. Substituting molarity of commercial concentrated hydrochloric acid (12 M) in the definition equation gives a pH = –1.1; for 18 M sulfuric acid, the pH = –1.3. Similarly, the 8 M sodium hydroxide solution used to make soap from animal fat has $[H^+] = 1.25 \times 10^{-15}$ (that is, $10^{14}/8$), and a corresponding pH = 14.9.

    Solutions of such extreme concentrations are seldom used, except as concentrates for making the dilute solutions of pH 1-14.

14. **"For sulfuric acid, to say that dissociation is complete means the acid has lost both protons."**

When we say that dissociation of sulfuric acid is complete, we are not saying that the species in solution are two hydrogen ions and one sulfate ion for each molecule dissolved. The fact is, sulfuric acid is "strong" only in regard to removal of the first hydrogen ion.

HISTORY: ON THE HUMAN SIDE

1. **Svante August Arrhenius** taught himself to read at age three and graduated from high school as the youngest and brightest in his class. As a student at the University of Uppsala he studied how electricity passed through solutions. He proposed that the resulting electric current from an aqueous sodium chloride solution was carried not by ordinary sodium and chlorine atoms, but by atoms carrying an electric charge—sodium ions and chloride ions. These ideas, which Arrhenius presented in 1884 in his Ph.D thesis, met with resistance. He was awarded the lowest possible passing grade by his examiners.

Fortunately, Van't Hoff and Ostwald, two influential physical chemists, were impressed and encouraged him to continue work in physical chemistry. In 1887, Arrhenius proposed that the characteristic properties of acids in water solution are the properties of hydrogen ion and those of bases, the properties of hydroxide ion. When J. J. Thomson discovered the electron in the 1890s, the idea of "ions" that Arrehenius proposed in his thesis suddenly became credible. In 1903, for the same thesis that had barely earned him a passing grade in his Ph.D examination, Arrhenius was awarded the Nobel Prize in chemistry. In 1959 Sweden issued a postage stamp on the occasion of his birth centenary.

2. **Johannes Brønsted** was born in Denmark in 1879. When he was 13 his father died; he was to follow in his father's profession of civil engineer. However, he was interested in chemistry and studied it in college. In 1908 he earned his doctorate and was selected as a new professor of chemistry at the University of Copenhagen. While studying how acids and bases catalyzed reactions in 1921, he clarified what acids and bases were. Since the properties of acids and bases were opposites, Brønsted felt it made more sense to supply definitions that were opposites. In 1923 he suggested that if acids were species that gave up a proton, bases are species that accept a proton. At the same time, Thomas Lowry, a British chemist, independently proposed the same definition. This definition represented a greater flexibility that extended acid-base notions into areas in which the old view was inadequate. For examples, ions could be classified as acids or bases and acid-base reactions could occur without water being present.

HUMOR: ON THE FUN SIDE

1. Confucius: "Chemist who falls in acid is absorbed in his work!" (Chem 13 News, January 1975, p. 838.)
2. Rhyme:

   Johnny was a chemist
   Now Johnny is no more
   For what he thought was H two O
   Was H two S O four!

3. Student responses on exams:
   a. A strong base is a salt that produces a strong acid when reacted with water.
   b. A strong acid is one that will not make a light burn bright.
   c. Filter paper is an indicator.
4. Sign on a bumper sticker: *Chemistry is pHun!*
5. Rhyme:

   Mamie often wondered why
   acids trouble alkali.
   Mamie in a manner placid
   fed the cat some boric acid;

   whereupon the cat grew frantic
   executing many an antic.
   Concluded Mamie from its pace,
   "Pussy cat must be a base."

   (*Chem 13 News, November 1988, p. 13.*)
6. Word Search (see *Appendix* for master copy)

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Z S N M A T K B Z R U O S T M P Y
X T S Q R A D H V K Q K Y O S E W
R P H T D R B U X E R I L D U R H
E U Y O R U M X J O S A I Q H B Y
T R D F H E P G N H R W N A V W D
T M R Z D V N O Q I O P O Y K M R
I R O C V M D G T R T T I X D O O
B I G V F N J Y T P A V W L X D X
A U E A O X Y N W H C C R Z J S I
M Y N T K A T H T N I O P D N E D
G Y O P V T E I D Z D F E S I K E
M R O Y D I A V C M N G J T H S J
P Q Z N E U T R A L I Z E I N I R
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   Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

   1. Weak acids or bases that change color over a narrow pH range.
   2. To eliminate the acidic or basic character of a solution.
   3. Term used to denote the degree of ionization of an acid or base.
   4. The point in a titration when the indicator changes color (2 words).
   5. Concentration of a solution defined in terms of moles of solute per liter of solution.
   6. Ion most often associated with an acid.
   7. Ion most often associated with a base.
   8. Taste most often associated with acids.
   9. Taste most often associated with bases.
   10. Brønsted-Lowry acid (2 words).

   Answers: 1. INDICATORS 2. NEUTRALIZE 3. STRENGTH 4. END POINT 5. MOLARITY 6. HYDROGEN 7. HYDROXIDE 8. SOUR 9. BITTER 10. PROTON DONOR
7. See relevant cartoons at end of module.

MEDIA

1. *The World of Chemistry* videotape "Number 16: The Proton in Chemistry," World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

2. CHEM Study films/videos available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.

   a. A*cid-Base Indicators*

   b. *Nitric Acid*

3. Moore, J. (Ed.). (1988). Acid-base indicator simulation for Apple II microcomputer. *Journal of Chemical Education, 65,* 798. Computer series.

4. Atkinson, G., Doadt, E., and Rell, C. (1986). Master variable diagrams for acid-base systems for an IBM personal computer. *Journal of Chemical Education, 63,* 841.

5. Software published by *JCE: Software*, a publication of the Journal of Chemical Education, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. *The Acid-Base Package*, by Richard Ramette. Vol. II B, Bo. 2, for IBM PS/2 PC-compatible computers.

   b. *Acid-Base Package*, by Richard Ramette and Jon L. Holmes. Vol. IV C, No. 2, for the Apple Macintosh.

6. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   a. For the Apple II computer: AP 501

   b. For IBM PCs and PC-compatibles: PC 2701, PC 3901

7. Videodisc published by *JCE: Software,* a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   "Acid Rain and Limestone" and "Lime Water and Carbon Dioxide," two chapters on *The World of Chemistry: Selected Demonstrations and Animations* : Disc II (double sided, 60 min.), Special Issue 4.

8. *Acids, Bases and Salts,* a 20 min. video available from Coronet/MTI Film and Video, 108 Wilmot Road, Deerfield, IL, 60015; (800) 621-2131; (708) 940-3640 (FAX).

INSTRUMENTATION

1. Vernier's Voltage Plotter with pH electrode (Vernier Software Science Supply, 2920-T.S.W. 89th Street, Portland, OR 97225).
2. Nester model and Corning model pH meters (Flinn Scientific, Inc. Chemical Catalog/ Reference Manual, 131 Flinn Street, P. O. Box 219, Batavia, IL 60510-0219).
3. Corning Conductivity Meter (Flinn Scientific).
4. Conductivity Apparatus: Russo, T. (1986). A low-cost conductivity apparatus. *J. Chem. Ed.*, *63*, 981; Gadek, F. (1987). Easily made electronic device for conductivity experiments. *J. Chem. Ed.*, *64,* 628; Gadek, F. (1987). A commercially available device for conductivity experiments. *J. Chem. Ed.*, *64*, 281; and Vitz, E. (1987). Conductivity of solutions apparatus. *J. Chem. Ed.*, *64*, 550.
5. Oyster-Shell Type pH Meter. Extech, 150 Bear Road, Waltham, MA 02154, 617-890-7440. This instrument costs less than $200. It has two calibration adjustments to give results comparable to more expensive pH meters.

Links/Connections

WITHIN CHEMISTRY

1. **Periodicity**

Metal oxides usually form basic solutions in water; aqueous solutions of nonmetal oxides are acidic.

$$\underset{\text{Metal oxide}}{K_2O} + H_2O \rightarrow 2KOH$$

$$\underset{\text{Nonmetal oxide}}{SO_3} + H_2O \rightarrow H_2SO_4$$

The acidity of the hydrogen halides increases in the sequence

$$HF < HCl < HBr < HI$$

The basicity of Period 3 hydroxides decreases in the sequence

$$NaOH > Mg(OH)_2 > Al(OH)_3$$

The acidity of Period 3 oxyacids increases in the sequence

$$H_3PO_4 < H_2SO_4 < HClO_4$$

2. **Solutions/solubility**

Most acids tend to be water soluble. NaOH and KOH are quite soluble in water. $Ca(OH)_2$ is moderately water soluble. Acidic solutions dissolve some metals (Fe, Zn, and Al), releasing hydrogen gas. Zinc and aluminum will dissolve in base. Drano™ consists of solid NaOH and Al turnings. When water is added to the mixture, hydrogen gas is evolved.

$$2Al(s) + 2NaOH(aq) + 2H_2O(l) \rightarrow 2NaAlO_2(s) + 3H_2(g)$$

3. **Equilibrium; Ionization Constant**

The dissociation of a strong acid or base is complete. The ionization constant expression is a measure of the equilibrium for the ionization of a weak acid or base. For example, the equilibrium established when the weak acid, acetic acid, ionizes,

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

is given by the ionization constant expression:

$$K_a = K[H_2O] = \frac{[H_3O^+]\,[C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Water has the important characteristic of being able to act as either an acid or base. This tendency allows two water molecules to react accordingly

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

The equilibrium constant expression for the process is

$$K_w = K\,[H_2O]\,[H_2O] = [H_3O^+]\,[OH^-]$$

K_w, the ion product constant for water, has a value of 1.0×10^{-14} at 25 °C and 2.4×10^{-14} at 37 °C (body temperature). This is significant because it means that although at 25 °C a neutral solution has a pH = 7, at 37 °C, the pH of a neutral solution is less than 7 (6.81).

4. **Electrochemistry**

Electrolysis of water produces an acidic solution at the anode and a basic solution at the cathode. When an electrical current is passed into graphite electrodes immersed in a sodium sulfate solution, hydrogen gas and hydroxide ions are produced at the cathode:

$$2H_2O(l) + 2\,e^- \rightarrow H_2(g) + 2OH^-(aq)$$

At the anode, oxygen gas, and hydrogen ions are formed:

$$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$

The OH^- and H^+ ions combine to reform water. The overall reaction is

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

5. **Stoichiometry**

See the *Stoichiometry* module.

6. **Bonding**

Hydrogen halides have a polar covalent bond. $H:\ddot{\underset{\cdot\cdot}{F}}:$ $\quad \overset{\delta+}{H}-\overset{\delta-}{F}$

In oxyacids, hydrogen is bonded to oxygen. The central atom becomes very important in determining acidity.

H_2SO_4 is H-O-SO_2-O-H and $HClO_4$ is H-O-ClO_3

Chlorine is more electronegative than sulfur and is able to weaken the O-H bond, increasing the acidity of $HClO_4$ over that of H_2SO_4.

7. **Thermochemistry**

Thermochemistry is a study of the quantity of heat absorbed or evolved by chemical reactions. The neutralization of a strong acid and strong base, represented by the equation

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) + \text{Heat}$$

is highly exothermic. Acid or base spills on the body should not be neutralized since the affected area could actually experience a heat burn due to high heat of neutralization. Instead, dilute the acid or base by flushing with lots of water.

Between Chemistry and Other Disciplines

1. **Environmental Science**

Acid rain is rain with a pH less than 5.6. Normal rain water has a pH of 5.6 due to dissolved carbon dioxide, CO_2. Acid rain is primarily a dilute mixture of sulfuric acid and nitric acid that is formed when sulfur and nitrogen oxides (nonmetal oxides) dissolve in rain or snow. Acid rain destroys lakes, kills forests and crumbles buildings and statues.

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Acid mine water is due to sulfuric acid produced from iron sulfide (pyrite, FeS_2) in coal. Acid mine water is associated with strip mining, produced when pyrite deposits are exposed to the atmosphere.

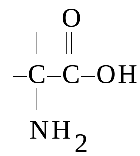
$$2FeS_2(s) + 7\,O_2(g) + 2H_2O(l) \rightarrow 2FeSO_4 + 2H_2SO_4$$

$Ca(OH)_2$ and $Al(OH)_3$ are important in water purification. During the settling stage, $Ca(OH)_2$ and $Al_2(SO_4)_3$ are added to produce $Al(OH)_3$, a sticky, gelatinous precipitate that settles out slowly, carrying suspended dirt particles and bacteria with it.

2. **Biochemistry**

Gastric (stomach) acid is essentially hydrochloric acid with a pH of about 1.6-1.8. It activates protein-digesting enzymes, which are responsible for protein digestion in the stomach.

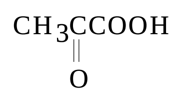
Proteins are polymers of amino acids. Amino acids have two functional groups, the basic amino group, NH_2, and the acidic carboxyl group, COOH:



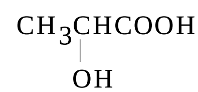
About 20 different amino acids make up the proteins in the body.

DNA (deoxyribonucleic acid) and RNA (ribonucleic acid) are complex molecules whose biological activity is dependent on the composition of organic nitrogen bases present in their structures. DNA is the essential hereditary material of all living cells. RNA controls the production of proteins, the building blocks of living cells.

Physical activity is a series of muscle contractions, each requiring energy. Muscles store energy as glycogen. This glycogen is the storage form of glucose in the blood arising from carbohydrate in the diet (candy bar, fruit, potatoes, pancake syrup, *etc.*). When muscles contract, the glycogen is converted to pyruvic acid and, if sufficient oxygen is present, to CO_2 and H_2O. This is aerobic exercise or aerobics. If sufficient oxygen is unavailable, pyruvic acid is converted to lactic acid. This is anaerobic exercise. If the buildup of lactic acid continues, the muscles are less responsive to contraction, and the person experiences some fatigue and even pain. After exercising, more oxygen becomes available to convert the lactic acid to pyruvic acid, and the latter is converted to CO_2, H_2O, and energy. Sprinting and weight lifting are typical anaerobic activities; a marathon race is largely aerobic. Warm-up activities before exercise help to prevent rapid buildup of lactic acid in the muscle tissues.



lactic acid

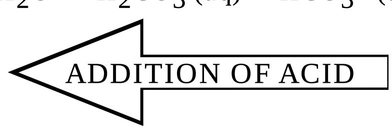


pyruvic acid

To insure necessary biological functions, body fluids are maintained within a narrow pH range. For example, protein molecules within cells can be altered if the pH changes too much in either direction of the optimum pH. This alteration can change the usual properties of the protein such that the protein would not be able to carry out its normal function.

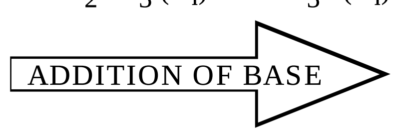
The H_2CO_3/HCO_3^- buffer is the major buffer in blood. This buffer counteracts either added acid or bases. Acidic by-products (H^+) of metabolic cycles are neutralized by $HCO_3^-(aq)$ forming $H_2CO_3(aq)$. Excess $H_2CO_3(aq)$ is removed from the body as $CO_2(g)$ in the lungs.

$$CO_2\,(g) + H_2O \rightleftharpoons H_2CO_3\,(aq) \rightleftharpoons HCO_3^-\,(aq) + H^+\,(aq)$$



Excess base is neutralized by $H_2CO_3(aq)$ forming $HCO_3^-(aq)$. To compensate for the diminished $H_2CO_3(aq)$, $CO_2\,(g)$ dissolves in water to form carbonic acid.

$$CO_2\,(g) + H_2O \rightleftharpoons H_2CO_3\,(aq) \rightleftharpoons HCO_3^-\,(aq) + H^+\,(aq)$$



3. **Geochemistry**

Magnesium, Mg, is the lightest structural metal in common use. It is obtained from sea water. Lime, CaO, is added to sea water to precipitate magnesium hydroxide, $Mg(OH)_2$. Magnesium hydroxide is then filtered and neutralized with hydrochloric acid. After evaporation of the water, the resulting $MgCl_2$ is melted and electrolyzed to give magnesium metal.

After recovery from coal and petroleum or even sulfur mines, sulfur is converted to sulfuric acid in four steps: (1) the sulfur is burned in air to sulfur dioxide, SO_2; (2) sulfur dioxide is passed over a hot platinum surface where it is converted to sulfur trioxide, SO_3; (3) sulfur trioxide is passed into sulfuric acid where pyrosulfuric acid, $H_2S_2O_7$, is formed, and (4) the latter is diluted with water to give sulfuric acid, H_2SO_4. Sulfuric acid is the number one chemical substance in terms of total mass produced by U.S. industry (see *Industrial Inorganic Chemistry* module).

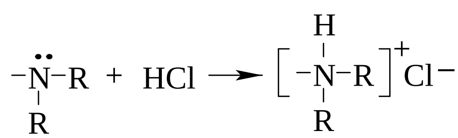
Sodium hydroxide is prepared by electrolysis of a concentrated aqueous sodium chloride solution.

$$2NaCl(aq) + 2H_2O(l) \rightarrow 2NaOH(aq) + Cl_2(g) + H_2(g)$$

To the Contemporary World

1. **Drugs**

Lidocaine and novocaine are bases that are used by the medical profession as anesthetics. Because of their limited solubility in water (the typical injection solvent), the basic amino group in these molecules is converted to their hydrochloride salt form, which is water soluble. Furthermore, as bases these drugs would cause localized changes in pH at the injection site. Thus, the use of salt solutions minimizes this localized irritation.



Antacids are basic compounds that decrease the amount of hydrochloric acid in the stomach. Examples are milk of magnesia (magnesium hydroxide), $Mg(OH)_2$, and Rolaids™ (sodium dihydroxy aluminum carbonate, $NaAl(OH)_2CO_3$). Sodium bicarbonate (baking soda), $NaHCO_3$, is the antacid in Alka Seltzer™. When Alka Seltzer™ is placed in water, bicarbonate ions react with hydronium ions from the acid producing the familiar fizz:

$$HCO_3^-(aq) + H_3O^+(aq) \rightarrow CO_2(g) + 2H_2O(l)$$

2. **Foods**

Acids and bases are used in cooking. Baking soda is sodium bicarbonate, $NaHCO_3$. Baking powder is a mixture of sodium bicarbonate, $NaHCO_3$, and an acid salt, potassium hydrogen tartrate, $KHC_4H_4O_6$. When bread bakes, these ingredients react to produce carbon dioxide and steam that expands and produces an airy loaf.

$$NaHCO_3(s) + KHC_4H_4O_6(s) \rightarrow H_2O(g) + CO_2(g) + NaKC_4H_4O_6(s)$$

Vinegar (acetic acid, $HC_2H_3O_2$) is a component of salad dressings. Citric acid is a component of fruit juices. Salts of acids (sodium benzoate and sodium propionate) are food additives that preserve foods from microorganisms.

3. **Agriculture**

The carbon in decaying organic matter uses oxygen to form carbon dioxide. This increased concentration of carbon dioxide causes the ground water flowing through to become acidic. Acidic soils are sour soils.

$$CO_2(g) + H_2O(l) \rightarrow H^+(aq) + HCO_3^-(aq)$$

Limestone (calcium carbonate, $CaCO_3$) is applied to acid soils to produce a pH that is slightly basic. Such a soil is called a sweet soil.

$$CO_3^{2-}(aq) + H^+(aq) \rightarrow HCO_3^-(aq)$$

4. **Insects**

Some species of millipedes store the compound mandelonitrile. When attacked by a predator, the millipede combines an enzyme with the mandelonitrile and discharges a toxic repellant, hydrogen cyanide, HCN.

$$C_6H_5CH(OH)CN(s) \rightarrow C_6H_5CHO(l) + HCN(g)$$

The whip scorpion, which is only a few centimeters long, can accurately direct a spray many times that distance at a predator. The spray contain a mixture of three organic carboxylic acids:

HCOOH	CH_3COOH	$CH_3(CH_2)_6COOH$
Formic acid	Acetic acid	Caprylic acid

Although acetic acid is the actual irritant, its effectiveness is enhanced by the presence of caprylic acid. Caprylic acid permits penetration of the waxy coating found on many recipients.

Formicidae ants defend themselves by discharging a spray of formic acid.

5. **Community Resources**

Try to plan field trips to a water or sewage treatment plant, a food processing plant, or dairy in your area. Invite knowledgeable people in your area associated with one of these to your classroom to discuss the chemistry involved in their particular plant.

6. **Acids and Foods**

Lactic acid, sauerkraut, hot BBQ sauce—all find their way to our tables. The relationship between acids and foods is a rich one to explore. (Idea contributed by Angie Matamoros.)

Extensions

Conjugate Acid-Base Pairs

The Brønsted-Lowry definition of an acid is a species that donates protons, and a base is a species that accepts protons. These definitions have a valuable application in interpreting experimental facts. The dissociation of an acid (or a base) is an equilibrium reaction. For example, in the dissociation of acetic acid,

$$HC_2H_3O_2(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

acetic acid donates a proton to water. Acetic acid is a Brønsted acid. Water, which accepts the proton, is a base.

In the reverse reaction (which is proceeding at the same rate), hydronium ion donates a proton to acetate ion. Hydronium ion is an acid. Acetate ion is a base because it accepts a proton.

Acetic acid by losing a proton forms a base, acetate ion. Acetate ion, by accepting a proton, can form an acid—acetic acid. This relationship can be represented as follows:

$$\underset{\text{Acid}}{HC_2H_3O_2} \underset{\text{Gains } H^+}{\overset{\text{Loses } H^+}{\rightleftharpoons}} \underset{\text{Base}}{C_2H_3O_2^-}$$

This acid-base pair, formed from each other by the loss or gain of a proton represents a conjugate acid-base pair. Conjugate means joined in pairs, or coupled.

Similarly, water (a base) accepts a proton to form the acid hydronium ion, which in turn, forms water by loss of a proton.

$$\underset{\text{Base}}{H_2O} \underset{\text{Loses H+}}{\overset{\text{Gains } H^+}{\rightleftharpoons}} \underset{\text{Acid}}{H_3O^+}$$

Hydronium ion and water are a second conjugate acid-base pair in the acetic acid dissociation.

Hydrolysis

When a salt dissolves in water, it completely dissociates to produce cations and anions. Either one or both of these ions may then react with water. This reaction is called hydrolysis. Hydrolysis of a salt can affect the pH of the solution. The resulting pH depends on the nature of the salt dissolved in water (see Figure 20).

Type of salt derived from	pH	Example
Strong acid-strong base	7	NaCl, KNO_3
Weak acid-strong base	>7	$NaC_2H_3O_2$, KCN
Strong acid-weak base	<7	NH_4Cl, $AlCl_3$
Weak acid-weak base	Depends on salt	$NH_4C_2H_3O_2$

Figure 20. pH effects by hydrolyzing typical salts.

Sodium acetate forms a basic solution because the acetate ion reacts with water to form hydroxide ion:

$$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + \underset{\text{Base}}{OH^-}$$

Similarly, ammonium chloride forms an acidic solution because only the cation hydrolyzes:

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + \underset{\text{Acid}}{H_3O^+}$$

Suggested Demonstration

Materials

Universal indicator

Five 0.1 M solutions selected from: sodium bicarbonate, $NaHCO_3$; ammonium chloride, NH_4Cl; sodium bisulfate, $NaHSO_4$; sodium chloride, NaCl; sodium carbonate, Na_2CO_3; ammonium sulfate, $(NH_4)_2SO_4$; aluminum nitrate, $Al(NO_3)_3$; sodium nitrate, $NaNO_3$; ammonium acetate, $NH_4C_2H_3O_2$; and potassium phosphate, K_3PO_4. Select solutions that will give acidic, basic, and at least one neutral solution.

5 Petri dishes or small beakers

Directions

Add solutions to vessels placed on the overhead projector. Add 1-3 drops of universal indicator solution to each container. Alternatively, large test-tubes in a test-tube rack may be used with a diffuse light source behind them.

Buffers

A buffer is a system that maintains a nearly constant pH in a solution when relatively small amounts of acid or base are added to the solution. A buffer is a mixture of a weak acid and its conjugate base, or a weak base and its conjugate acid.

A buffer can be prepared by mixing a weak acid (for example, carbonic acid, H_2CO_3) with one of its salts (for example, sodium bicarbonate, $NaHCO_3$), since the anion (HCO_3^-) of the acid is the conjugate base. This buffer is very important in maintaining a nearly constant pH of the blood. A weak base (NH_3) mixed with one of its salts (NH_4Cl) can also function as a buffer.

Suggested Demonstration

Materials

3 M or 6 M Hydrochloric acid, HCl (25 or 50 mL 12 M HCl diluted to 100 mL with distilled H_2O)

3 M or 6 M Sodium hydroxide, NaOH (12 g or 24 g NaOH(s) dissolved in 100 mL with distilled H_2O)

9 Large (about 180-mm) test-tubes

Test-tube rack with illuminated lightbox or white background

Universal indicator solution

Distilled water

Buffer Solution A Prepare by mixing equal volumes of 4 M acetic acid, $HC_2H_3O_2$ (24 mL 17 M $HC_2H_3O_2$ diluted to 100 mL with distilled H_2O) and 4 M sodium acetate, $NaC_2H_3O_2$ (32.8 g $NaC_2H_3O_2$ dissolved in 100 mL distilled H_2O).

Buffer Solution B Prepare by mixing equal volumes of 4 M aqueous ammonia, NH_3 (27 mL 15 M NH_3 diluted to 100 mL with distilled H_2O) and 4 M ammonium chloride, NH_4Cl (21.4 g NH_4Cl dissolved in 100 mL distilled H_2O).

Directions

Set up a 3 x 3 matrix of test-tubes.

HCl	Control	NaOH	
O	O	O	**Buffer A**
O	O	O	**Water**
O	O	O	**Buffer B**

Add Buffer Solution A (acetic acid buffer) to three test-tubes. Place distilled water into the second tube of each set. Then add Buffer Solution B (ammoniabuffer) to the last three test-tubes. Add a drop or two of universal indicator solution to each tube; stir with a stirring rod until a brightly colored solution results. To the first of each of the three test-tube sets, add a dropperful of 6 M HCl; to the right hand member of each of the three test-tube sets, add a dropperful of 6 M NaOH. Stir the solutions. Have students record observed changes. To illustrate the capacity of the buffer, slowly add 6 M NaOH to the third test-tube in the acid buffer set (Set A) and 6 M HCl to the third test-tube in the basic buffer set (Set B) until pronounced color changes are noted.

Optional

Dissolve an Alka-Seltzer™ tablet in water in each of four 100-mL beakers on an overhead projector. Add universal indicator solution to each. Reserve the first beaker as a control. Add 1 M HCl dropwise to the second beaker with stirring; observe. Then add 1 M NaOH dropwise to the third beaker with stirring; observe. Finally, add distilled water dropwise to the fourth beaker with stirring; observe.

Lewis Acids and Bases

The Brønsted-Lowry concept of an acid and a base was limited to electron pair sharing to the proton (H^+). A new concept of acids and bases was proposed in the early 1920s by G. N. Lewis, a professor of chemistry at the University of California at Berkeley. In the Lewis proposal, an acid is an electron-pair acceptor ("take two from you"), a base is an electron-pair donor ("have pair, will share"), and an acid-base reaction involves a base sharing an electron pair with an acid.

Lewis extended the Brønsted-Lowry definition of an acid and a base. Lewis acids and bases are not dependent upon the proton (H^+) or the hydroxide ion. In fact, the Lewis concept revolutionized the theory and practice of acid catalysis in organic chemistry (see last equation in Figure 21, formation of a carbocation).

In a Lewis acid-base reaction a coordinate covalent bond is formed between an acid and base. Examples are in the equations shown in Figure 21:

Acid + Base

$$F_3B + :NH_3 \rightarrow F_3B:NH_3$$

$$Ag^+ + 2:NH_3 \rightarrow [H_3N:Ag:NH_3]^+$$

$$AlCl_3 + (CH_3)_3C\text{-}Cl \rightarrow [Cl:AlCl_3]^- + (CH_3)_3C^+$$

Figure 21. Examples of Lewis acid-base reactions.

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Appendix

- **Transparency Masters**
  1. Acids: Concentration *vs.* Strength
  2. Acid-Base Definitions
  3. pH and Concentration
  4. Word Search
- **Humor**

Acids: Concentration *vs.* Strength

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<tr><th></th><th>CONCENTRATED</th><th>DILUTE</th></tr>
<tr><td>STRONG</td><td>
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<tr><td>WEAK</td><td>
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</table>

Acid-Base Definitions

	Arrhenius	Brønsted-Lowry	Lewis
Acid	Provider of H^+ in HOH	Proton donor	Electron pair acceptor
Base	Provider of OH^- in HOH	Proton acceptor	Electron pair donor
Neutralization	Formation of water	Proton transfer	Coordinate covalent bond formation
Equation	$H^+ + OH^- \rightarrow HOH$	$HA + B \rightarrow BH^+ + A^-$	$A + :B \rightarrow A:B$
Limitation	Water solution only	Proton transfer Reactions only	Generalized theory

pH and Concentration (25°C)

	$[H^+]$	pH	$[OH^-]$	pOH
Acidic Solution	$>10^{-7}$	<7	$<10^{-7}$	>7
Neutral Solution	10^{-7}	7	10^{-7}	7
Basic Solution	$<10^{-7}$	>7	$>10^{-7}$	<7

Word Search

Z	S	N	M	A	T	K	B	Z	R	U	O	S	T	M	P	Y
X	T	S	Q	R	A	D	H	V	K	Q	K	Y	O	S	E	W
R	P	H	T	D	R	B	U	X	E	R	I	L	D	U	R	H
E	U	Y	O	R	U	M	X	J	O	S	A	I	Q	H	B	Y
T	R	D	F	H	E	P	G	N	H	R	W	N	A	V	W	D
T	M	R	Z	D	V	N	O	Q	I	O	P	O	Y	K	M	R
I	R	O	C	V	M	D	G	T	R	T	T	I	X	D	O	O
B	I	G	V	F	N	J	Y	T	P	A	V	W	L	X	D	X
A	U	E	A	O	X	Y	N	W	H	C	C	R	Z	J	S	I
M	Y	N	T	K	A	T	H	T	N	I	O	P	D	N	E	D
G	Y	O	P	V	T	E	I	D	Z	D	F	E	S	I	K	E
M	R	O	Y	D	I	A	V	C	M	N	G	J	T	H	S	J
P	Q	Z	N	E	U	T	R	A	L	I	Z	E	I	N	I	R

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Weak acids or bases that change color over a narrow pH range.
2. To eliminate the acidic or basic character of a solution.
3. In a solution, the percentage of ionized molecules of acid or base.
4. The point in a titration when the indicator changes color (2 words).
5. Concentration of a solution defined in terms of moles of solute per liter of solution.
6. Ion most often associated with an acid.
7. Ion most often associated with a base.
8. Taste most often associated with acids.
9. Taste most often associated with bases.
10. Brønsted-Lowry acid (2 words).

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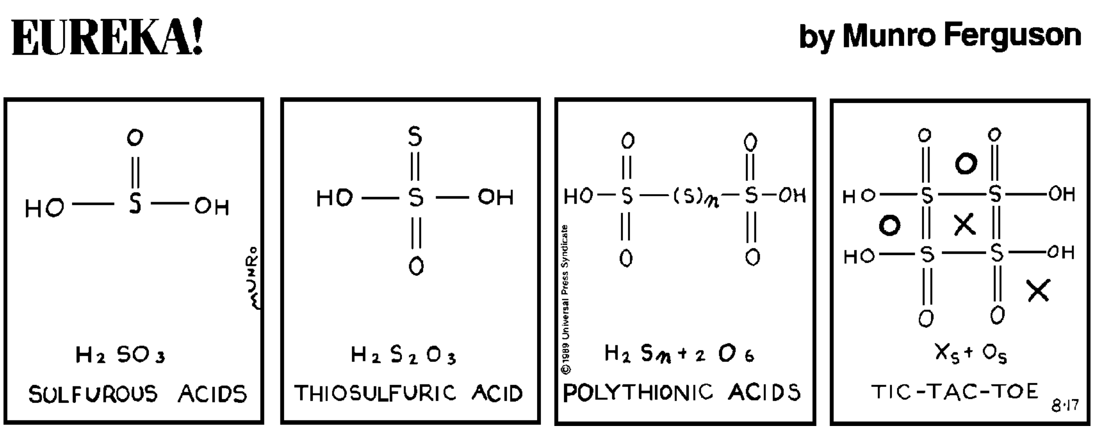

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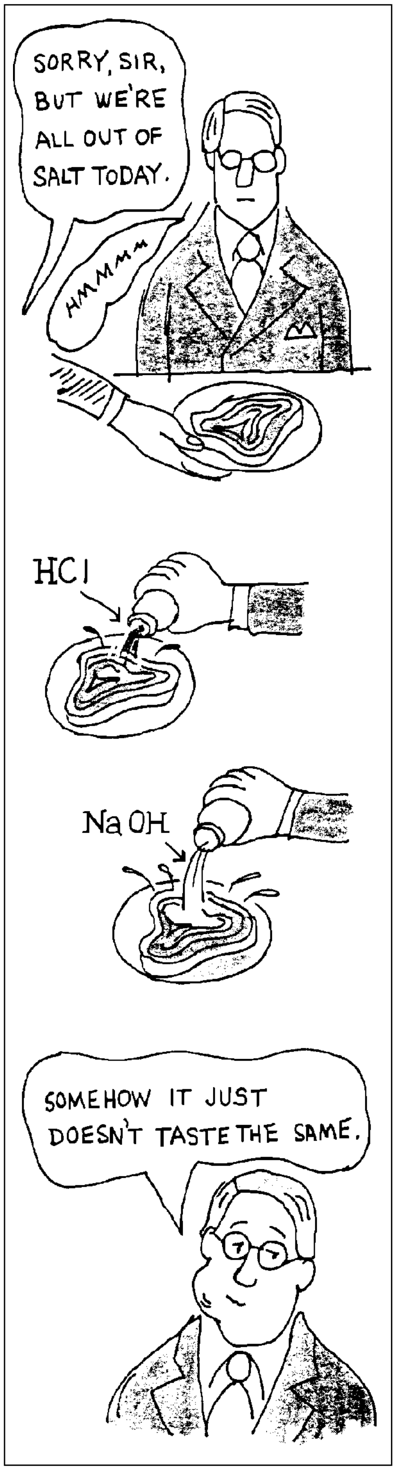

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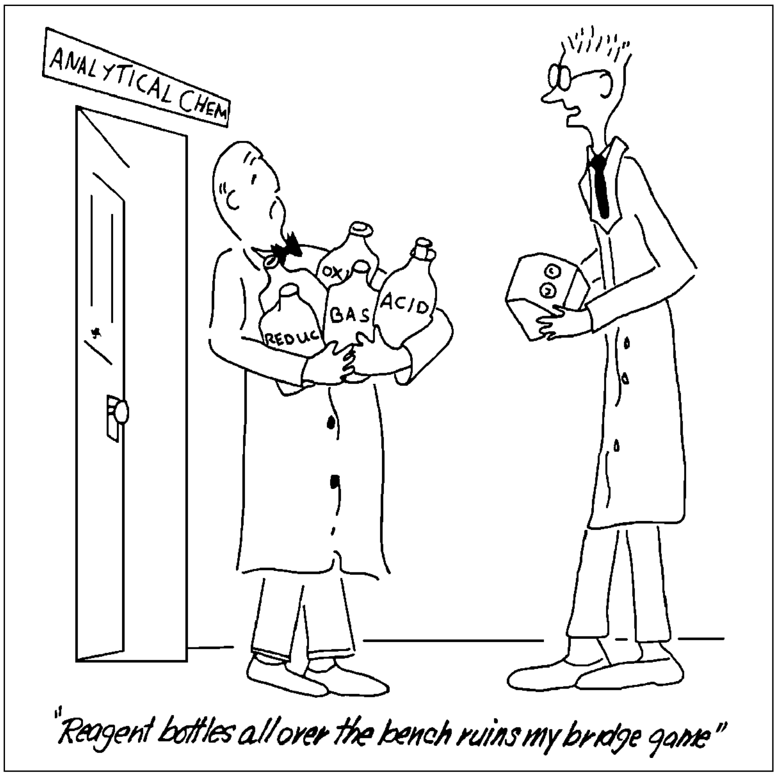

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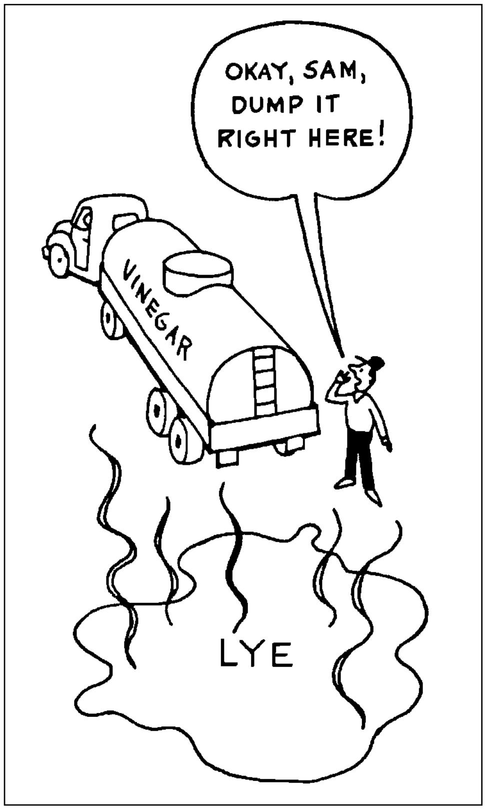

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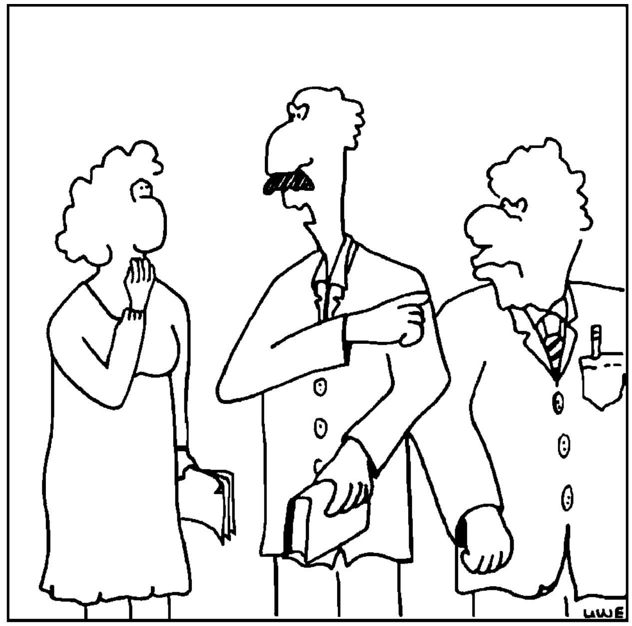

WITH HIS ACID TONGUE AND MY CAUSTIC WIT, WORKING IN THE SAME LAB COULD END IN ASSAULT.

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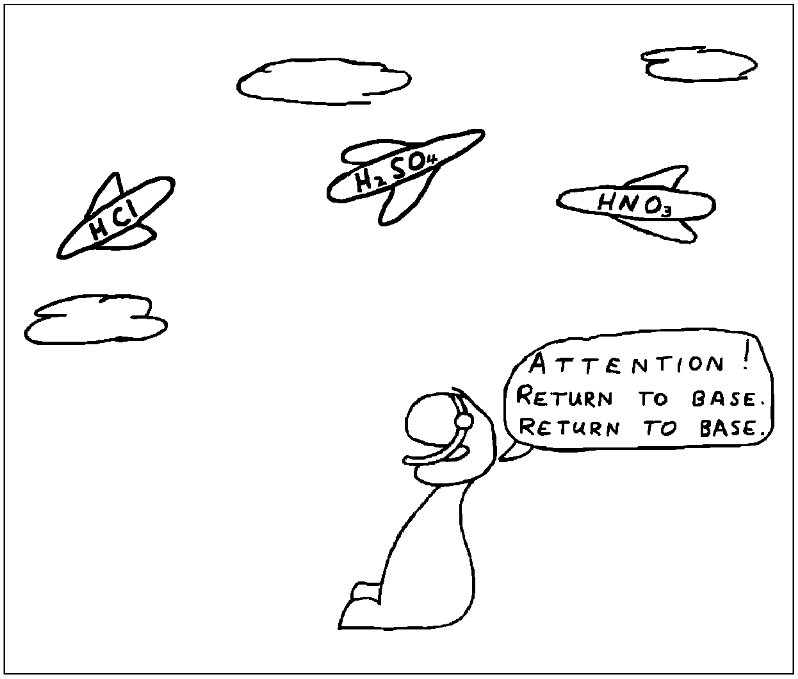

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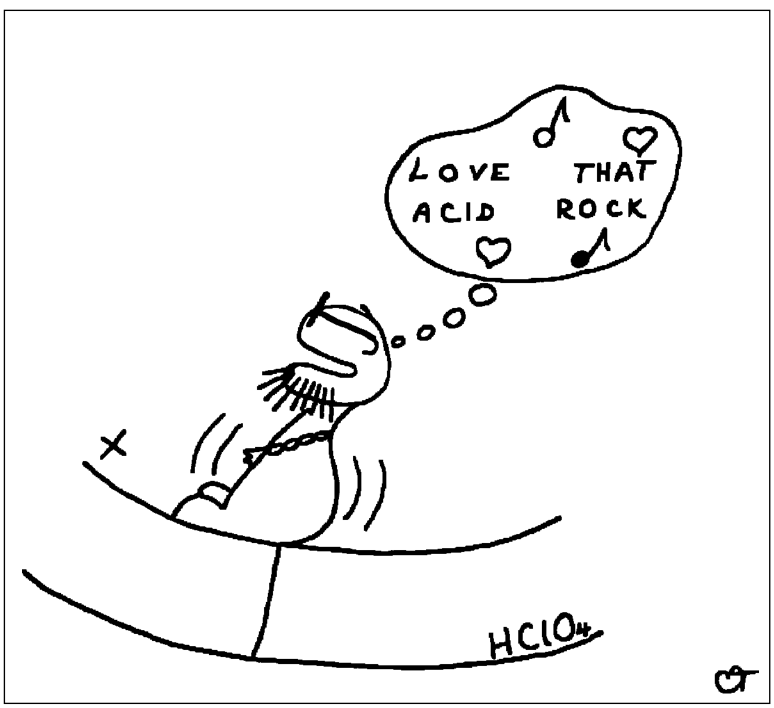

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