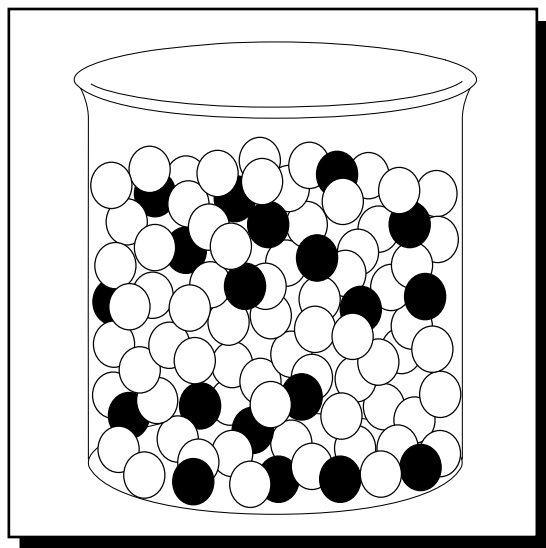


A SourceBook Module

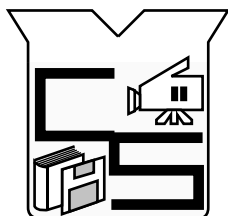
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*ChemSource Project Principal Investigator:
Mary Virginia Orna, OSU
Department of Chemistry
College of New Rochelle
New Rochelle, NY 10805
Phone: (914) 654-5302
FAX: (914) 654-5387*



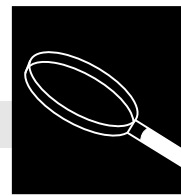
SOLUTIONS



ChemSource

*Instructional Resources for Preservice and
Inservice Chemistry Teachers*

Topic Overview



CONTENT IN A NUTSHELL

Solutions make up a large, important part of chemistry. Usually solutions are considered in terms of the ability of liquids to dissolve solids, other liquids, or gases, forming a homogeneous mixture. However, solutions can be solids in solids, liquids in solids, gases in solids, solids in liquids, liquids in liquids, gases in liquids, and gases in gases. The major component of a solution is termed the solvent and the other component(s) is(are) termed the solute(s). In some cases these terms are arbitrary.

A somewhat arbitrary distinction is often made between homogeneous solutions and colloids, which are solid, liquid, or gaseous particles (made up of particles large enough to diffract or scatter light—the Tyndall effect) dispersed in solid, liquid, or gaseous media. The dispersed particles are called colloids. The thermal motion (Brownian movement) of such small particles is sufficient to keep them from settling out in the earth's gravitational field.

“Like dissolves like” is an expression that qualitatively expresses the experimental observations that polar and ionic substances dissolve in polar solvents and nonpolar substances dissolve in nonpolar solvents—*other things being equal*. A solvation model that considers interactions between polar solvent molecules and polar and ionic solutes is helpful in considering the solvation properties of water and other polar solvents. However, some aqueous solutions, and most solutions involving nonpolar solvents, depend on the randomness (entropy increase) obtained in the solution process.

Solution concentrations (*e.g.*, mol/L) should be distinguished from amounts (*e.g.*, moles). Several methods for expressing solution concentrations are important—reaction stoichiometry in solution is normally expressed in terms of the molarity (M) of the reacting species, whereas colligative properties are compared using either molality (*m*) or mole fractions (*X*), and impurities are often quoted in parts per million (ppm) or parts per billion (ppb). Relative concentrations are often expressed as dilute or concentrated, or by the terms unsaturated, saturated or supersaturated.

Solution components can be detected and evaluated by a variety of methods and can often be separated by distillation, crystallization, or chromatography. Colligative properties are useful in determining the nature of solutes in solutions. These properties together with conductivity have allowed scientists to quantify electrolytes (ionic solutes) in solutions. Solute in water are often categorized as strong electrolytes, weak electrolytes, and non-electrolytes, reflecting the conductivity of the substance when dissolved.

PLACE IN THE CURRICULUM

This module should follow a discussion of the states of matter, changes of state, moles, and stoichiometry. It provides a good vehicle for either introducing or augmenting a discussion of solution stoichiometry.

CENTRAL CONCEPTS

1. The designations solute and solvent can be usefully made in many types of solutions (*Demonstration 1*).
2. Various types of solutions are possible, based on whether the intermixed species are solids, liquids, or gases (*Demonstration 1*).

3. A solution is a homogeneous mixture of two or more substances where the solvent and solute particles are of typical molecular dimensions (*i.e.*, < 1 nm). A mixture made up of larger particles is considered a colloidal suspension, which is observable via the Tyndall effect (*Demonstration 2*).
4. Like dissolves like, in general. Nonpolar solvents tend to dissolve nonpolar substances and polar solvents tend to dissolve polar and ionic substances.
5. A microscale model of the solution process within a liquid solvent can be constructed based on the principle of solvation.
6. Concentration (*e.g.*, mol/L) should be distinguished from amount (*i.e.*, moles).
7. Common quantitative methods for expressing solution concentration include percent by mass, percent by volume, molality (m), molarity (M), mole fraction (X), and parts per million (ppm).
8. Stoichiometry in solutions is conveniently based on molar concentrations (M) of the reacting species (*Laboratory Activity 1*).
9. Relative concentrations of a solute/solvent system can often be expressed by the terms dilute and concentrated, or by the terms unsaturated, saturated, and supersaturated.
10. A solution can be separated into its components through processes such as distillation, crystallization, and chromatography (*Demonstration 3*).
11. Colligative properties of solutions include vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure (*Laboratory Activity 2*).
12. Solutes in water are often categorized as either electrolytes (strong, if completely ionized in water or weak, if only partially ionized) or non-electrolytes (nonionized).
13. Generally, gases become *less* soluble in liquid solvents as the temperature is increased. On the other hand, most (but not all) solids and liquids become *more* soluble in liquid solvents as the temperature is increased.

1. The Mole
2. Gases
3. Condensed States
4. Stoichiometry

RELATED CONCEPTS

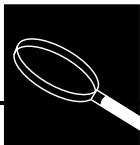
1. The ability to measure mass, volume, and temperature.
2. The ability to write and balance chemical equations.
3. The ability to interpret graphs.

RELATED SKILLS

After completing their study of solutions, students should be able to:

1. distinguish between solutes, solvents, and solutions.
2. list the multitude of solution types that exist.
3. discuss the difference between solutions and colloids.

PERFORMANCE OBJECTIVES



4. apply the generality that polar solvents tend to dissolve polar substances and ions, whereas nonpolar solvents tend to dissolve nonpolar species.
5. explain the distinction between concentrations and amounts.
6. calculate the concentration of solutions using molarity and molality. (Calculations using the other units are optional.)
7. define: dilute, concentrated, unsaturated, saturated, and supersaturated.
8. explain how distillation, crystallization and chromatography are used to separate solutions into their components.
9. discuss colligative properties including vapor pressure lowering, boiling point elevation, freezing point depression, and osmotic pressure.
10. describe the differences between electrolytes (both weak and strong) and non-electrolytes.
11. explain the temperature dependence of gas solubility in liquid solutions.
12. explain the temperature dependence of liquid and solid solubility in liquid solutions.



Concept/Skills Development

Activity 1: Limiting Reactant for a Chemical Reaction in Solution

Introduction

Many chemical reactions are conducted by combining solutions of reagents that will react rather than by reacting the pure materials. This strategy is particularly useful when the reactants are pure solids. Also, when synthesizing a compound it is common to have an excess of one of the reactants present in order to assure complete reaction of the second reactant. In such a reaction, the reactant present in excess is called the excess reactant, and the reactant that limits the extent of the reaction is called the limiting reactant (see *Stoichiometry* module). It is important to know which solution placed into the reaction vessel is providing the limiting reactant in order to know the maximum amount of product (theoretical yield) that can be obtained.

Purpose

To determine the limiting reactant in a chemical reaction and the stoichiometric ratio in the balanced equation.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. If you should come into contact with any reagents you should immediately wash the affected area with large amounts of water.
3. Be sure to operate the centrifuge according to the procedure given by your teacher. Be sure to place a second test-tube containing the same amount of material in a test-tube holder *opposite* the holder containing the sample. Do not attempt to stop the centrifuge by making contact with the spinning head. If a test-tube should break while in the centrifuge, notify your teacher at once—do not attempt to clean the centrifuge without direction from your teacher and only after the centrifuge has been disconnected from the electrical outlet.
4. Be careful not to contact the hot surfaces of the oven when placing objects into or removing objects from the oven. Also, objects being removed from the oven are very hot and should be handled with tongs or other appropriate utensils as demonstrated by your teacher.
5. Do not pour any substances from this experiment down the drain. Dispose of them as indicated by your teacher.

Procedure

1. Obtain four 13 x 100-mm test-tubes or other tubes usable as centrifuge tubes, two 50-mL beakers, two 250-mL beakers, one stirring rod, three Beral™ pipets, and a wash bottle filled with distilled water.
2. Place 10 mL 1.0 M BaCl₂ solution into a 50-mL beaker and label the beaker. Place 15 mL 0.50 M Na₂SO₄ solution into another 50-mL beaker and label the beaker.

LABORATORY ACTIVITY: STUDENT VERSION





3. Take clean and dry test-tubes and label the test-tubes 1, 2, 3, and 4 using a pencil. Measure and record the mass of each of the test-tubes to the nearest 0.01 g.
4. Use one of the Beral™ pipets to measure 2.0 mL 1.0 M BaCl_2 solution from your beaker into each of the four numbered test-tubes. *Be careful when filling the pipet to avoid air bubbles in the liquid column in the pipet. Use one of the 250-mL beakers as a test-tube rack. Make sure that the pipet used with the BaCl_2 solution stays in the beaker containing the BaCl_2 solution after you have finished using it.*
5. Use a second Beral™ pipet to measure volumes of 0.50 M Na_2SO_4 solution indicated below into the test-tubes. Expel the Na_2SO_4 solution into the test-tubes in a fast stream to promote mixing of the solutions and make sure that the tip of the pipet does not go into the solutions in the test-tubes.
6. Place 2.0 mL 0.50 M Na_2SO_4 into test-tube number 1, 3.0 mL into test-tube 2, 4.0 mL into test-tube 3, and 5.0 mL into test-tube 4. *Make sure that the Beral™ pipet used for the Na_2SO_4 solution stays in the beaker containing the Na_2SO_4 solution after you have finished using it.*
7. Stir the solutions with the stirring rod; rinse the stirring rod with a stream of distilled water from your wash bottle each time it is withdrawn from a test-tube making sure that the washings go into the test-tube from which you are withdrawing the rod.
8. Using water from your wash bottle fill all the test-tubes to the same height. The liquid should be about 2 cm from the top of the test-tubes. Do not overfill! Place the test-tubes into the centrifuge using two of them to counterbalance the other two and spin them for about one minute. Remove the test-tubes from the centrifuge. *The solid BaSO_4 should be at the bottom of the test-tube, and the supernatant liquid (the liquid above the precipitated solid) should be clear, or slightly cloudy. If the liquid is very cloudy, check with your teacher to determine if you should continue to centrifuge the tubes. Your teacher may also recommend additional steps to aid the settling of the precipitate.*
9. Use the third Beral™ pipet to remove the supernatant liquid from above the precipitates and discard the liquid into the 250-mL beaker not being used as a rack. *Be careful not to remove any of the settled precipitate from the test-tubes.*
10. Fill each test-tube about 2/3 full with distilled water from your wash bottle. Stir each solution with the stirring rod so that the settled precipitate is dispersed into the liquid. Use the wash bottle to rinse the stirring rod as you did in Step 7. Now add distilled water from your wash bottle to the test-tubes to make the level of liquid the same in each test-tube, but at least 2 cm from the top. Centrifuge the tubes as you did in Step 8. *Once again, the solid should be settled and the supernatant liquid clear.* Remove the supernatant liquid from the test-tubes with the third Beral™ pipet taking care not to remove any of the solids.
11. Place the test-tubes containing the wet solids into a 250-mL beaker and place the beaker into the drying oven. *Be sure it has an ID mark, such as your initials, on it.*
12. When the solids have dried remove them from the oven. *CAUTION: The beaker is **hot** and should be handled with tongs or other approved handling equipment.* Place the beaker on a wire gauze on your desk and allow the beaker, the test-tubes and the solids to cool. Measure and record the mass of each of the test-tubes to the nearest 0.01 g.

13. Dispose of the test-tube contents, the excess reagents, and waste solutions as directed by your teacher. (Do not pour the chemicals down the drain!) Wash all glassware and the pipets and return all items to the appropriate place.
14. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. Make a table with the following column headings (use the long dimension of your paper for the top of the table): [1] Test-tube number, [2] Mass tube (g), [3] mL 1.0 M BaCl₂, [4] mL 0.50 M Na₂SO₄, [5] Mass tube + BaSO₄ (g), [6] Mass BaSO₄ obtained (g), [7] Moles BaSO₄ obtained, [8] Moles BaCl₂ added, [9] Moles Na₂SO₄ added, [10] Moles BaSO₄ expected if all BaCl₂ reacted, [11] Moles BaSO₄ expected if all Na₂SO₄ reacted, and [12] Limiting reactant.
2. Fill in columns [1] through [5] from information from your data sheet.
3. Calculate the mass and moles BaSO₄ obtained in each test-tube and record the values in columns [6] and [7], respectively.
4. Calculate the moles BaCl₂ and Na₂SO₄ added to each test-tube and record the values in columns [8] and [9], respectively.
5. Write a balanced equation for the reaction that occurred and use information from columns [8] and [9] to calculate the moles BaSO₄ expected if (a) all the BaCl₂ added had reacted and (b) all the Na₂SO₄ reacted for each of the test-tubes, and record the values in columns [10] and [11], respectively.
6. Compare the values in columns [7], [10] and [11] and decide which of the reactant solutions, BaCl₂ or Na₂SO₄, contained the limiting reactant and indicate your choice in column [12].

Implications and Applications

1. Does the solution containing the lower concentration of reactant always contain the limiting reactant? How do the experimental results verify your conclusion?
2. Does the smaller volume of solution always contain the limiting reactant? How do the experimental results verify your conclusion?
3. What pieces of information do you need before calculating the maximum amount of product to be expected from a reaction where the reactants are in solution?
4. Corn sweetener is a sweetening agent obtained from corn. A recipe calls for 8.0 ounces of pure corn sweetener. You go to the supermarket and find that there are two brands of corn sweetener both in solution form. Brand A comes in a 16 ounce bottle, is 40% corn sweetener by mass and costs \$ 0.59. Brand B comes in a 24-ounce bottle, is 45% corn sweetener by mass, and costs \$ 0.95. Which brand costs less per ounce of pure corn sweetener? How many ounces of the less expensive brand should be added to the recipe in order to get the 8.0 ounces of pure corn sweetener?
5. Barium ion in solution is highly toxic if ingested. Why can large quantities of BaSO₄ be ingested by individuals undergoing gastrointestinal X-rays without any apparent ill effects?



**LABORATORY
ACTIVITY:
TEACHER
NOTES**

***Activity 1: Limiting Reagent for a Chemical
Reaction in Solution***

Major Chemical Concept

The major chemical concept is the stoichiometry involving solutions and the determination of the limiting reactant in a chemical reaction.

Level

Honor students.

Expected Student Background

Students should be capable of accurately weighing items on whatever balance is being used. A balance such as a Centigram™ balance should suffice for this experiment. The correct method of using a centrifuge should be demonstrated (or reviewed if used before). Related concepts include balancing equations, molarity, and stoichiometric calculations and the use of limiting reactant in such calculations.

Time

Because of a bottleneck that can occur with using the centrifuge, and the need to allow about 20 min for drying precipitates in the drying oven and 5-10 min needed for cooling, about 80 min is required to complete this experiment. The experiment can be conveniently broken into two parts with Steps 1 through 10 being conducted one day and the experiment being completed the second day. If you intend to complete the experiment in one day you should plan to conduct the pre-laboratory discussion the day before the experiment. Teacher/technician preparation time for the experiment is about 45 min, mainly devoted to making the two required solutions and calibrating the Beral™ pipets.

Safety

Read the *Safety Considerations* in the *Student Version*. Barium ion in solution, or in a soluble compound, is *highly toxic if ingested*; therefore, students should wash their hands thoroughly before leaving the laboratory. Barium compounds must *not* be discarded down the drain. There should be a waste container for these salts and solutions. After all students have completed the activity, you can convert the sample to *very* insoluble barium sulfate by adding excess sodium sulfate. The barium sulfate waste can be filtered, dried, and stored for proper disposal.

Safety precautions related to using a centrifuge should be stressed—be sure to place a second tube containing the same amount of material in a tube holder *opposite* the holder containing the sample. Because students may be in a hurry at the time when they are using the centrifuge, stress that the centrifuge should not be manually stopped by the insertion of fingers or other objects. If safety considerations don't impress students you may point out that quickly stopping a centrifuge tends to bring the solid back up into the liquid, thereby undoing what centrifuging has just accomplished. Also warn students to keep ties and long hair away from the spinning centrifuge. Occasionally a test-tube breaks while being centrifuged. This usually does not result in flying glass but students should be given instructions about cleanup procedures.

Probably the most common accident involves finger burns when placing glassware into or removing it from the drying oven or handling glassware that has not sufficiently cooled. Proper techniques for handling hot glassware should be reviewed before the experiment.

Materials (For 24 students working in pairs)

Balances, Centigram™ or equivalent
Centrifuge
Drying oven

Nonconsumables

48 Test-tubes, 13 x 100-mm or larger to fit available centrifuge
24 Beakers, 50-mL
24 Beakers, 250-mL
36 Plastic Beral™ pipets, stems calibrated to 1.0 mL
12 Stirring rods
12 Wash bottles filled with distilled water

Consumables

1.0 M Barium chloride, BaCl₂·2H₂O, 200 mL (48.8 g BaCl₂·2H₂O per 200 mL)
0.50 M Sodium sulfate, Na₂SO₄·10H₂O, 250 mL (40.3 g Na₂SO₄·10H₂O per 250 mL)

Advance Preparation

Solutions should be prepared as above. If the Na₂SO₄ in your stockroom is in the anhydrous form, 17.8 g per 250 mL yields the desired 0.50 M solution.

All glassware should be clean and dry. If students have to clean and dry the glassware, particularly the test-tubes, before beginning they will not have sufficient time to complete the activity. If the plastic Beral™ pipets are to be reused they should also be carefully washed to avoid cross-contamination during reuse and all visible traces of moisture removed.

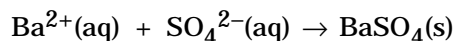
The Beral™ pipet stems may be calibrated for 1 mL by filling a 10-mL graduated cylinder to the 10 mL mark with distilled water and drawing sufficient water into the pipet to bring the water in the cylinder to the 9 mL mark. The water level in the pipet stem should then be marked with a waterproof felt marker.

The drying oven should be turned on at the beginning of the class so that it comes to drying temperature (120-130 °C) by the time it is to be used. The thermostat settings should be determined the day before the laboratory.

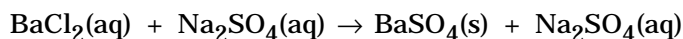
Pre-Laboratory Discussion

The pre-laboratory discussion should take place the day before the activity is to be performed if it is to be completed in one 80-min period.

1. Review safety considerations and disposal methods involved with the handling of the chemical substances being used.
2. Emphasize the necessity not to mix the Beral™ pipets during the activity—one should be used for the barium chloride solution and another used for the sodium sulfate solution, and the third for removal of the supernatant liquid.
3. Make sure that students understand the difference between *amount* and *concentration* of a solution.
4. Review mole and molarity calculations as well as reaction stoichiometry calculations. The net-ionic equation for the reaction in this experiment is:



The overall reaction is:





Therefore, one mole of BaCl_2 reacts with one mole of Na_2SO_4 . You should be able to elicit this information from your students if they have studied chemical stoichiometry.

- Review the procedure used and safety considerations in operating the centrifuge and the drying oven.
- Explain that students can use the calibration marks on the beakers to measure the 10- and 15-mL volumes needed in Step 2 as only approximate volumes are necessary.
- Because the test-tubes are to be put in the oven, labeling in Step 3 should *not* be made with a wax marking crayon or with paper labels.

Teacher-Student Interaction

One problem that arises is the fineness of the barium sulfate precipitate. Even after extended centrifuging the precipitate may not settle; some precipitate may adhere to the sides of the test-tube. The adherence to the test-tube is not a problem, since the precipitate will remain in the test-tube and be weighed. The precipitate that does not settle is a very small fraction of the total precipitate and, if removed with the supernatant liquid, will cause a small percent error. However, you may want to tell students to add a small quantity of *soap* (not detergent) solution by dipping the stirrer in a soap solution, mixing the solution with the stirrer, and then recentrifuging. This will usually aid the settling of the precipitate.

Monitor student use of the centrifuge. Any large vibration usually signals that the centrifuge is not balanced. When in a hurry, students tend to disregard the admonition against using fingers or other objects to stop the centrifuge.

Monitor the use of the drying oven. Students may need help in removing the hot beaker and test-tubes. It is best if all students place their beakers into the oven within minutes of each other. This strategy permits the oven to heat up to the required drying temperature and remain there, speeding drying.

Circulate around the room while students are doing their weighings. It is important that the weighings be done properly and masses recorded to 0.01 g.

Anticipated Student Results

1	2	3	4	5	6	7	8	9	10	11	12
T.T. No.	T.T. Mass (g)	Reactant Vol. 1.0 M BaCl_2 (mL)	0.50 M Na_2SO_4 (mL)	Mass of T.T. + BaSO_4 (g)	Mass of BaSO_4 Obtained (g) (5-2)	Moles of BaSO_4 Obtained $\times 10^4$ *	Moles BaCl_2 Added $\times 10^4$	Moles Na_2SO_4 Added $\times 10^4$	Moles BaSO_4 to Expect if all BaCl_2 Na_2SO_4 Reacted $\times 10^4$		Limiting Reactant
1	7.75	2.0	2.0	7.95	0.20	8.6	20	10	20	10	Na_2SO_4
2	8.66	2.0	3.0	8.97	0.31	13	20	15	20	15	Na_2SO_4
3	7.55	2.0	4.0	7.99	0.44	19	20	20	20	20	Either
4	7.50	2.0	5.0	7.95	0.45	19	20	25	20	25	BaCl_2

* The actual yield should be close to the theoretical yield. The factor 10^4 means that the quantities in columns 7-11 have been multiplied by that amount to yield numbers in the tens range. The actual quantity is obtained by reversing the process and dividing by 10^4 ; e.g., $8.6 = 8.6 \times 10^{-4}$ (in Column 7; see *Introductory Module* for use of exponents).

Answers to Implications and Applications

1. No. Sodium sulfate was the solution of lower concentration, but the results in Tube 4 show that barium chloride was the limiting reactant.
2. No. There were 2.0 mL barium chloride solution and 3.0 mL sodium sulfate solution in Tube 2 and sodium sulfate was the limiting reactant.
3. You must know the stoichiometric ratio from the balanced equation, the concentrations of the reactants and the volumes of solutions mixed together in order to determine the limiting reactant and consequently the maximum amount of product to be expected.
4. Brand A costs \$ 0.092 per ounce of corn sweetener. Brand B costs \$ 0.088 per ounce of corn sweetener. Brand B (17.8 ounces) has to be used to obtain 8.0 ounces of corn sweetener.
5. Barium sulfate is so insoluble that a negligible amount of barium ions go into solution. The barium sulfate ingested in “liquid” during a gastrointestinal X-ray is really a slurry of barium sulfate in water similar to the slurry obtained in this activity when the barium sulfate precipitate is washed with distilled water.

Possible Extensions

Depending upon the students' ability, the topics covered prior to the *Solutions* module and the time allotted, the following topics might be discussed:

1. The overwhelming number of chemical processes that take place in solution.
2. Introduction to the use of stoichiometric ratios in quantitative analysis.
3. The need to introduce excessive concentrations of reactants to shift equilibrium.

Assessing Laboratory Learning

1. Laboratory Practical. The use of the centrifuge may be tested as part of a laboratory practical. Also, an “unknown” concentration of one of the solutions in this experiment may be presented and students asked how they would go about determining the concentration of that solution—or asked to actually do it.
2. Written Examination. Problems involving the use of solutions and limiting reactant determination could be presented. Students could be asked to construct a “pictures in the mind” illustration of the limiting reactant for the $\text{BaCl}_2\text{-Na}_2\text{SO}_4$ reaction similar to the $\text{H}_2\text{-O}_2$ reaction illustrated in the *Stoichiometry* module.



**LABORATORY
ACTIVITY:
STUDENT
VERSION**

**Activity 2: The Freezing Point Depression
of Lauric Acid**

Introduction

Dissolved solute particles disrupt the crystallization process and lower the freezing point of the solvent. The amount the freezing point is depressed is directly proportional to the number of moles of particles dissolved.



Purpose

To determine the freezing point of lauric acid and to observe the effect a solute has on the freezing point.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Proper laboratory safety procedures must be followed.

3. The burners heat metal and glass, so proper care must be taken not to burn yourself. Hot and cool glass look alike.
4. Any substances spilled on you should be immediately washed off with large amounts of water and the incident reported to your teacher.

Procedure

Part I: The Freezing Point of Lauric Acid

1. Prepare one hot water bath (65 °C) and one warm water bath (30 °C). Set up two ring stands and place 400 mL water into each 600-mL beaker; using the burner, heat one bath to 65 °C and warm the other bath to 30 °C.
2. Weigh the test-tube (25 x 200-mm) empty; fill the test-tube to a height of about 5 cm with solid lauric acid and weigh the test-tube and lauric acid.
3. Clamp the test-tube with the lauric acid in the hot water bath and allow the acid to melt.
4. When the lauric acid has just melted, remove the test-tube from the water bath and place the thermometer and stirrer into the lauric acid.
5. Place the test-tube containing the melted lauric acid, thermometer and stirrer into the 30 °C water bath and clamp the test-tube in place. Stir the lauric acid and record the temperature every 30 seconds until the temperature of the lauric acid has dropped to 40 °C. *CAUTION: Do not attempt to remove or even move the thermometer or stirrer from the solidified lauric acid!*

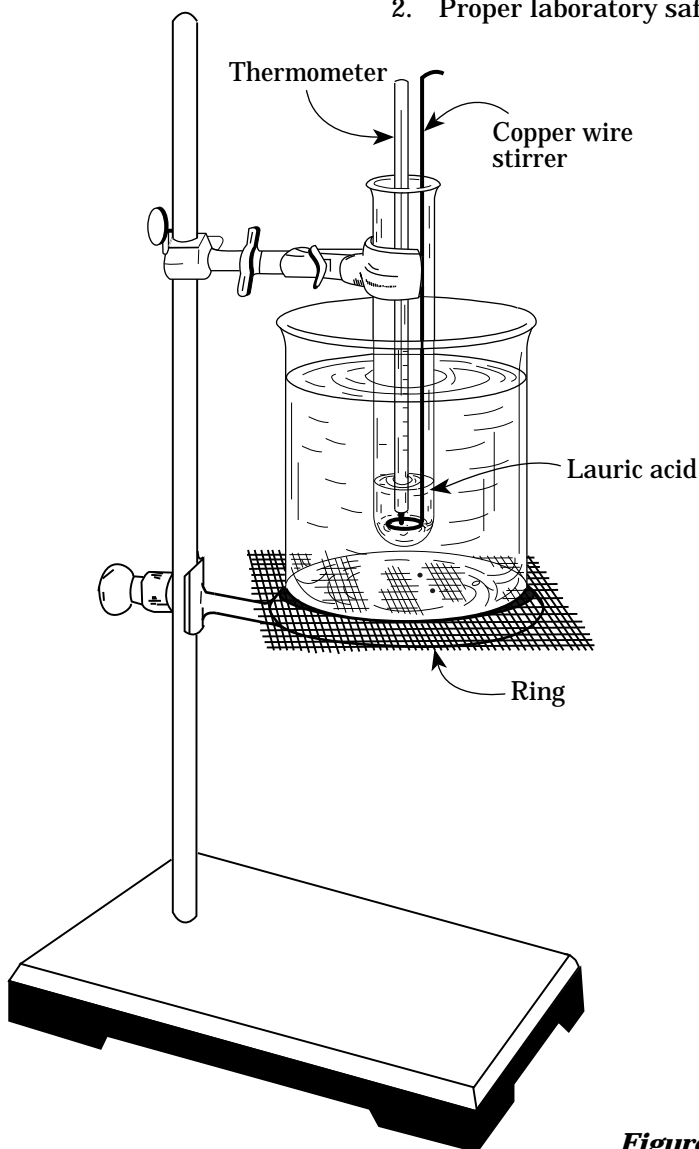


Figure 1. Elevated-temperature freezing-point apparatus.

Part II: The Freezing Point Depression of Lauric Acid

6. Weigh 2.5 g benzoic acid.
7. Remelt the lauric acid as you did in Step 3 above, and as soon as the lauric acid is melted, dissolve the benzoic acid in the lauric acid. Take care when adding the benzoic acid to make sure that all of it adds to the lauric acid, and that none of it sticks to the sides of the test-tube. You will need to stir the mixture vigorously with your stirrer in order to get the benzoic acid to dissolve completely.
8. Place the test-tube containing the melted lauric acid, benzoic acid, thermometer and stirrer into the 30 °C water bath and clamp the test-tube in place. Stir the mixture and record the temperature every 30 seconds until the temperature of the mixture has dropped to 36 °C.
9. Melt the mixture in the hot water bath and carefully remove the thermometer and stirrer from the mixture. Quickly wipe the thermometer and stirrer with a few layers of paper towels before the mixture has had a chance to solidify onto the thermometer.
10. Dispose of the paper towels and the test-tube contents as directed by your teacher, and return all materials to their proper places.
11. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. Graph the temperature *vs.* time for each trial.
2. What is the freezing point of lauric acid? Explain your reasoning.
3. What is the freezing point of the mixture of lauric acid and benzoic acid? Explain your reasoning.
4. How many moles of benzoic acid were used?
5. Calculate the molality (mol of benzoic acid/kg of lauric acid) of the solution.
6. Calculate the change in the freezing point temperature in going from the pure solvent to the solution.
7. Calculate the freezing point depression constant (K_f) and compare this with class results. $K_f = \text{change in freezing point divided by molality of solution}$. Remember that molality is moles of solute per 1000 g of solvent.

Implications and Applications

1. Is the freezing point the same as the melting point for pure lauric acid? Why?
2. Explain what is happening to the energy removed from the lauric acid in each part of the graph for the cooling of the pure lauric acid.
3. What is happening, on the molecular level, during the time when the graph of the pure lauric acid is a flat, straight line?
4. Why is the graph of the pure lauric acid different from the graph of the mixture of acids?
5. What is happening, on the molecular level, during the time when the graph of the lauric acid plus benzoic acid shows a minimal temperature change with time?



**LABORATORY
ACTIVITY:
TEACHER
NOTES**

***Activity 2: The Freezing Point Depression
of Lauric Acid***

Major Chemical Concept

The major chemical concept is the freezing point depression due to a dissolved solute.

Level

General high school chemistry

Expected Student Background

The major prerequisite for this activity is the ability to read and follow instructions, make and record observations, plot graphs, and safely handle hazardous materials.

Related concepts necessary for understanding this exercise include: atomic structure, formula writing, equations, and concentration calculations which are a direct part of the exercise. Students should know and understand the concept of colligative properties.

Time

The laboratory can be completed in two 45-min periods with *Part I* completed on the first day and *Part II* on another day *without dismantling the apparatus*. Teacher preparation time is about one hour. A great timesaving step is to supply the hot water baths at the proper temperature. The hot water from the tap may be hot enough, though further heating may be necessary in some schools.

Safety

Read the *Safety Considerations* in the *Student Version*. Students should not heat the test-tube of lauric acid directly in the flame as this can cause it to vaporize. It is also flammable. Care should be exercised in using the thermometer, particularly if mercury thermometers are used. Have sulfur or other mercury clean-up materials ready and immediately available in case of accidental breakage if mercury thermometers are used.

Materials (For 24 students working in pairs)

At least one balance with 0.1 g sensitivity should be available to the class.

Nonconsumables

- 24 Ring stands with iron rings and wire gauze
- 24 Thermometer clamps
- 12 Test-tube clamps
- 24 Beakers, 600-mL
- 24 Thermometers
- 12 Burners
- 12 Test-tubes, 25 x 200-mm
- 24 Copper heavy gauge wires with loops for stirring

Consumables

- Lauric acid, 300 g
- Benzoic acid, 30 g
- Paper towels

Advance Preparation

It is important that you do the laboratory yourself beforehand. Time may be saved by pre-filling the test-tubes with about 25 g of lauric acid and having the hot and warm water baths ready. [NOTE: The mass of the lauric acid in the test-tubes must be determined to the nearest gram.]

Pre-Laboratory Discussion

Make sure students understand the safety instructions. It is important that you demonstrate the proper placement of the thermometer, clamp and stirrer. Caution students not to attempt to remove or even move the thermometer or stirrer from the solidified lauric acid. A brief discussion of the reason for freezing point depression of solutions as explained in *Answers to Implication and Applications* (Question 4) would be quite useful, too.

Teacher-Student Interaction

One of the most common problems encountered in this procedure is the fact that the molten acid will warm the 30 °C water so that it no longer cools, or the reverse, that warming of the solid acid will be hampered by the warm water cooling off. If you have sufficient water in the beakers, this problem is avoided. During the session, make sure students do not use the thermometers as stirrers, and that they do not attempt to remove the thermometers after the lauric acid has solidified.

Anticipated Student Results

1. The melting point of lauric acid is 44 °C.
2. The average freezing point depression observed is around 2°C. This translates into a K_f of about 2.5 °C/molal.

Answers to Data Analysis

1. The graphs should consist of a downward slope followed by a flat plateau and then another downward slope after the lauric acid has solidified. In the case of the solution of benzoic acid in lauric acid, the plateau following the initial downward slope is not perfectly level (as discussed in *Answers to Implications and Applications*, Question 4).
2. The freezing point of lauric acid is 44 °C, based on the plateau observed at this temperature on the graph.
3. The freezing point of the lauric acid/benzoic acid mixture is 42 °C, based on the start of the plateau observed on the graph at this temperature.
4. Moles of benzoic acid used: $\frac{2.5 \text{ g}}{122 \text{ g mol}^{-1}} = 0.020 \text{ mol benzoic acid}$
5. Molality of the solution: $\frac{0.020 \text{ mol benzoic acid}}{0.025 \text{ kg lauric acid}} = 0.80 \text{ m}$
6. Freezing point depression: $44 \text{ °C} - 42 \text{ °C} = 2 \text{ °C}$.
7. $K_f = \frac{2 \text{ °C}}{0.80 \text{ m}} = 2.5 \text{ °C/m}$



Answers to Implications and Applications

1. The freezing point and the melting point of all pure substances are the same. (Naturally, superheating or supercooling can provide apparent differences.) There cannot be a “forbidden region” where a substance exists as neither a liquid nor a solid.
2. As we cool the liquid lauric acid, the energy we remove initially comes from the slowing down of the molecules as the temperature drops. The plateau is the point where the energy is lost in the phase change. The lauric acid molecules are forming a solid (crystallizing) and releasing energy as they do so. During the phase change no temperature changes are observed. As we cool the solid lauric acid further the energy we remove comes from the slowing down of the molecules of the solid and the temperature continues to drop.
3. This is the point of the phase change; *i.e.*, the molecules become an ordered solid. Although the kinetic energy of the molecules remains constant, as the liquid solidifies, the potential energy decreases while the molecules become part of an ordered state. (In the opposite direction, when melting occurs, the added energy is absorbed as potential energy as the molecules move into the liquid phase.)
4. The graph of the pure lauric acid shows an abrupt change to a level plateau at the freezing point. In the solution, the lauric acid solidifies, though the molecules of benzoic acid partially disrupt the process and cause solidification to happen at a lower temperature. When the lauric acid solidifies, the benzoic acid tends to stay in the liquid phase and further decrease the freezing point somewhat (*i.e.*, the higher molality of the remaining liquid phase has an even lower freezing point). Thus the plateau is not flat in the solution cooling curve.
5. The lauric acid is crystallizing into an ordered solid state as in the case of the pure acid, though the benzoic acid eventually also precipitates when its solubility in lauric acid is exceeded.

Extensions

1. Repeat the procedure with a larger amount (*e.g.*, 5.0 g) of benzoic acid.
2. Repeat the procedure with a soluble ionic solute to see the particle relationship that exists. A feasible system is water and CaCl_2 if you can obtain dry ice from the grocery store to use as a coolant. A dry ice/acetone slurry should be an efficient cooling bath.

Assessing Laboratory Learning

1. Laboratory Practical. A laboratory exercise requiring the determination of the freezing point depression of some other material may be used as assessment. Water is one such solvent.
2. Written Examination. Questions about the molecular level changes and questions requiring calculation of freezing point depression may be used.

Reference

This activity is an extension of a phase-transition experiment by Wilbraham *et al.* (1987).

CAUTION: Use appropriate safety guidelines in performing demonstrations.

DEMONSTRATIONS

Demonstration 1: Types of Solutions

Purpose

To demonstrate types of solutions.

Safety

See individual demonstration cautions and use generally accepted procedures for safety as outlined in the safety section of the *ChemSource* materials.

Procedures

Gas in gas: Gently warm I_2 crystals in a beaker over a water bath to observe I_2 vapor in the air. The beaker should be covered with a watch glass or evaporating dish. (*CAUTION: Iodine vapors are corrosive to eyes, skin, respiratory tract, and other materials.*)

Gas in liquid: Deaerate well-aerated water by warming (bubbles of air should be apparent) or uncap soda pop bottle previously shaken lightly (CO_2 will be evolved).

Gas in solid: Previously freeze well-aerated water. In class, melt the ice and warm further to deaerate. (*NOTE: Bubbles in the ice are not dissolved gases.*)

Liquid in liquid: Mix ethanol with water. (*CAUTION: Ethanol dehydrates the skin. Use with care. No flames should be present. Avoid breathing solvent vapors.*) The sum of the two volumes is not the same as the volume of the solution and can be shown by careful volume measurement. (100 mL ethanol + 100 mL water \approx 190 mL solution. The solute fills in the void volume (holes) of the solvent.)

Liquid in solid: The substance in a disposable diaper can dissolve water. The powder can be removed by cutting the diaper into 10 cm² pieces. Place the pieces in a plastic bag, close the bag, and shake out the powder. Sprinkle about one-half teaspoon of the powder onto one-half cup water. Observe. (A gel-like solid should form.) Practice this since the amount of water that can be absorbed may vary with the brand of diaper. (This demonstration is adapted from the *Polymer* module. See *Demonstration 3* in the *Polymer* module for another example of this type of solution.)

Solids in liquids (best known): Examples include solutions of sugar or salt (NaCl) in water. Anhydrous $CuSO_4$ (white) dissolved in water (colorless) gives a blue solution. I_2 in different solvents gives different colors (violet in cyclohexane, yellow in ethyl acetate, brown in water).

Solids in solids: Using a soldering iron compare the ease of melting of a typical Sn-Pb alloy (solder) vs. Sn metal and/or Pb metal. (*CAUTION: Beware of metal vapors from the solders. Alternatively, compare the properties of brass [Cu-Zn alloy] vs. the metallic elements Cu and Zn.*)

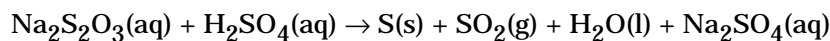
Demonstration 2: Colloids, The Tyndall Effect and More

Introduction

The Tyndall effect is usually given as a definitive test to distinguish between a true solution and a colloid. The Tyndall effect involves the scattering of a beam of light as the light passes through a medium having particles of colloidal size. Since particles such as molecules of sugar or sodium ions or chloride ions in solution are too small to scatter light, a beam of light passing through such a solution is not scattered. However, the protein molecules in milk are of colloidal



size and consequently a drop of milk mixed into water will cause a light beam traversing the solution to be scattered. To demonstrate the difference, the two systems described above are usually employed. The milk in water changes color as more milk is added (bluish to yellow to red). However, a single system where the particles go from “solution size” to “colloidal size” provides a more dramatic demonstration and as the particles continue to grow additional optical effects may be demonstrated. One such system is the production of sulfur by the reaction between sodium thiosulfate and sulfuric acid.



Here, the particles of sulfur grow from solution size to colloidal size and finally begin to precipitate. As this phenomenon occurs, it is possible to demonstrate the Tyndall effect and to examine some of the characteristics of the scattered and transmitted light. These characteristics may then be related to other phenomena, such as red sunsets.

Safety

Some students are very allergic to sulfur dioxide (SO_2), which is generated during this demonstration. You should offer a warning and excuse students if they have a known sensitivity to sulfur dioxide (or to sulfites in food, which may signal such a sensitivity).

Option A

Materials

Square or rectangular battery jar or a small fish tank
Stirring rod (long glass rod to stir contents of the battery jar)
Parallel beam light source or a flashlight
Piece of frosted glass or a sheet of white paper stapled to a frame
Ring stand and clamp to hold the glass or paper in a vertical position
Prewighed sample of $\text{Na}_2\text{S}_2\text{O}_3$ or $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ to make a 0.01 *m* solution when dissolved in the water in the battery jar (1.6 g $\text{Na}_2\text{S}_2\text{O}_3$ per L water or 2.5 g $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$ per L water)
10 mL Concentrated sulfuric acid, H_2SO_4 per L of water in the battery jar

Optional:

One or more pieces (sheets) Polaroid material
Cardboard silhouette of a flying duck (about the size of half the diameter of the parallel beam) and hung like a mobile from a fine piece of string
Recording of “Canadian Sunset” and appropriate player

Directions

The demonstration should be conducted in a well ventilated area because of the production of $\text{SO}_2(\text{g})$ whose odor can be mildly detected while the demonstration is in progress.

Set up the light source, battery jar, and frosted glass as illustrated in Figure 2. The battery jar should be filled with sufficient depth of water so that the entire diameter of the light beam passes through the liquid. (The amount of water should be predetermined in order to have a preweighed sample of sodium thiosulfate and a premeasured volume of concentrated sulfuric acid prepared.) Also, the water should be placed into the battery jar a few minutes prior to beginning the demonstration to permit air bubbles to escape.

Turn on the light source and darken the room. If you will be using the Polaroid sheets, pass these out to students seated directly in front of the demonstration. Ask students whether they can see the beam traversing the

water. If using the Polaroid sheets, ask whether rotating them has any effect. *[They should not be able to see the beam but the white round disk of the transmitted beam hitting the frosted glass should be visible. Nothing different should be seen using the Polaroid sheet.]* Point out that the transmitted beam striking the frosted glass is colorless (white). Add the preweighed sample of sodium thiosulfate to the water and stir until it is dissolved. Wait for any bubbles to leave the solution. Repeat the above question(s). *[The reply should be the same.]* Carefully and with stirring, add the premeasured concentrated sulfuric acid to the sodium thiosulfate solution. Wait about 15 seconds for bubbles to clear and repeat the above question(s). *[The reply should be the same.]* Ask if there is any evidence of a chemical reaction? *[There should be no such evidence.]* Ask students to tell you when they notice any change. *[Using the concentrations recommended and a solution temperature of about 20°C, it should take about two minutes for the Tyndall beam to start to appear and another two minutes for the particles to become large enough and concentrated enough for the percent of transmitted light to go to zero. The reaction rate appears to be first order in thiosulfate concentration so if you want to slow things down, decrease the concentration of sodium thiosulfate.]* As soon as the Tyndall beam becomes visible ask students to compare the color of the scattered light to that of the transmitted light striking the screen. Also ask students using the Polaroid sheets whether rotating them has any effect on the intensity of the scattered light. *[The scattered beam is bluish, whereas the transmitted light starts to turn yellow. As the Tyndall beam becomes more visible the scattered beam loses a little of its bluish color as the green and yellow portion of the spectrum is scattered and the transmitted light striking the screen goes from yellow to red. This observation occurs because the wavelengths of the light scattered are directly proportional to the size of the particle scattering the light. Also, the scattered light is polarized.]* If you have the silhouette of a “flying duck” you can add a dramatic closing touch to the demonstration. Start the recording of “Canadian Sunset” and hang the silhouette over the frosted glass so that it intercepts the disk of the reddening transmitted beam before the beam strikes the glass.

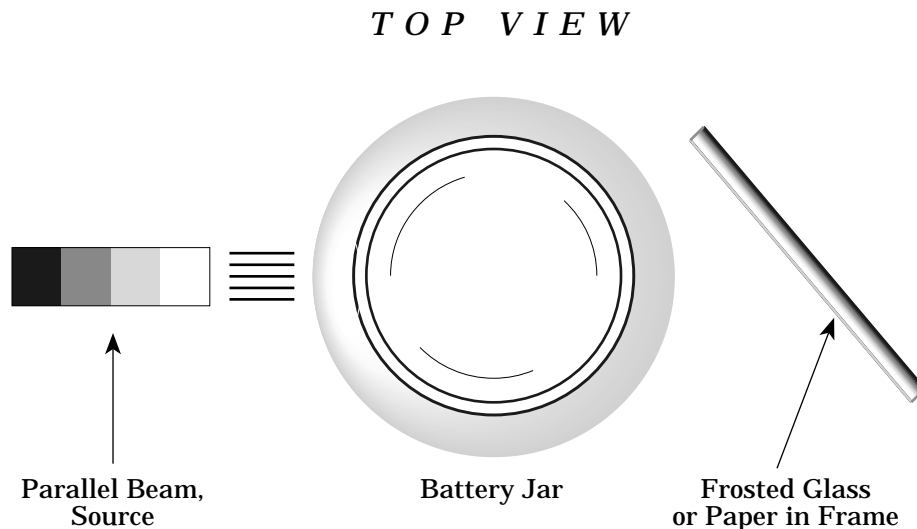


Figure 2. Apparatus for Tyndall Effect demonstration.



Dispose of the battery jar solution as soon as the demonstration is ended to minimize the amount of SO_2 entering the atmosphere.

After summarizing for students that the Tyndall beam was not noted until the particle size had become that of a colloid and that no beam had been visible when there was a true solution, you might want to extend the discussion and demonstration to other areas. If the day happens to be sunny with a relatively clear sky you might ask students to use the Polaroid sheets to observe the light coming from the blue sky at right angles to the direction of the sun. (CAUTION: Make sure students do not look toward the sun since this is very dangerous.) [The scattered blue light is polarized.]

Questions

You may follow the demonstration with the following questions:

1. Why does the sun appear exceptionally red when it sets behind a city? [The blue end of the visible spectrum is scattered by the dust and aerosols in the air over the city and the red end of the spectrum is transmitted.]
2. Why are fog lights usually yellow and not white? [The yellow light being of relatively long wavelength is transmitted through the colloidal size fog particles thereby permitting the driver to see, whereas the blue component of white light is scattered back to the driver thereby obscuring vision.]
3. Why not use red fog lights? [Red light is usually a signal of danger and approaching motorists may interpret it as such. Also, the human eye is much more sensitive to yellow light than red light.]
4. Why do things look clearer through rose (pink) colored glasses? [The pink glass filters out blue light that is scattered by aerosols in the environment. Since this blue light does not reach our eyes, the clarity of objects is increased.]

Option B

Materials

Overhead projector
Medicine dropper
Beaker of water
Milk, 10 mL

Directions

A simpler and quicker demonstration involves a beaker of water on the stage of an overhead projector and a small amount of milk in a medicine dropper. As the milk is slowly added to the water, the color of the water becomes bluish, while the circle of light thrown up onto the screen gradually becomes yellow then orange and then red. The questions suggested above (Option A) are also appropriate for this option.

Demonstration 3: Paper Chromatography

Introduction

Chromatography is used to separate the components of a mixture, operating on the basis of selective adsorption. The solvent moves through the porous material called the adsorbent. The various parts of the mixture are attracted to the adsorbent differently. The more strongly attracted parts do not move as far as the more weakly attracted parts. It is possible to calculate an **R_f value**, which is defined as the distance traveled by the solute divided by the distance traveled by the solvent. The purpose of this demonstration is to separate the pigments in a food color by paper chromatography.

Safety

Proper laboratory safety procedures must be followed. Alcohols are flammable; no open flames are allowed. Any substances spilled should be immediately washed off with large amounts of water.

Materials

- 2 Petri dishes (less than 11 cm diameter)
- Capillary tube (thin-walled, open-ended melting point tube or Pasteur pipette)
- Filter paper, 11 cm
- Ethanol, 95% or absolute (Solvent 1)
- 0.1% Sodium chloride, NaCl (0.1 g NaCl or table salt per 100 mL water; Solvent 2)
- Food coloring—Green typically works best, but try more than one color
- Scissors
- Distilled water

Procedure

1. Cut two pieces of filter paper as shown in Figure 3. Fold the tongue down so that it will just touch the surface of the Petri dish.
2. Using the capillary tube to transfer the food dye to the paper, make a small spot behind the folded tongue with one of the samples of food coloring. When the spot is dry, touch the capillary tube to the paper again to increase the amount of pigment on the paper. Do not increase the size of the spot; just try to make it more concentrated.
3. Prepare the other sample in the same way.
4. Place enough of one of the solvents to cover the bottom of one of the Petri dishes. Repeat with the other solvent in another Petri dish.
5. Place one piece of filter paper over each dish with the tongue just touching the solvent.
6. After the solvent is well absorbed (*i.e.*, traveled along) onto the paper without completely moving to the edge, remove the filter paper and mark the outer edge of the solvent line using a pencil. Do not allow the solvent front to reach the edge of the filter paper. When the papers have dried, mark the edges of each pigment using a pencil.
7. Calculate the R_f value for each spot in each solvent.

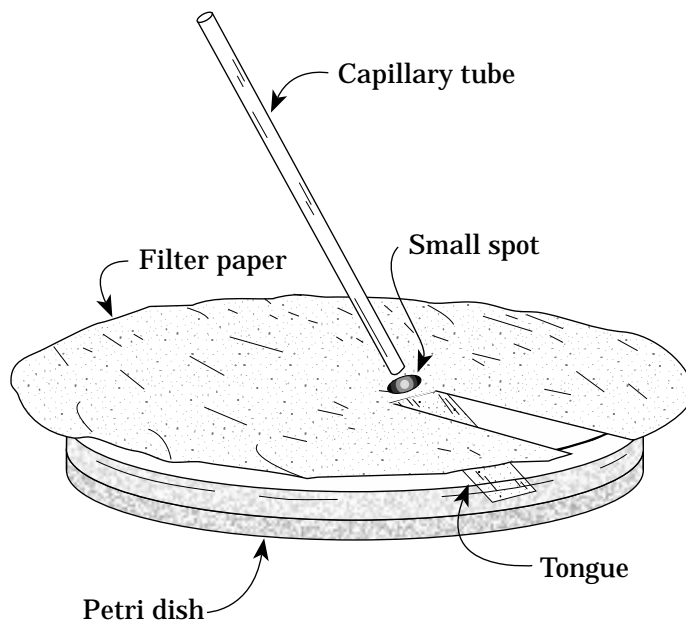


Figure 3. Set-up for paper chromatography demonstration.

$$R_f = \frac{\text{distance traveled by solute}}{\text{distance traveled by solvent}}$$

(NOTE: R_f values are always ≤ 1 , and that a smaller R_f value corresponds to a more tightly held solute.)

8. If time permits redo with other colors.



Data Analysis

1. How far did each solvent travel?
2. How many pigments were observed?
3. How far did each pigment travel from the original spot?
4. Calculate the R_f value of each major pigment.
5. Which color had the largest R_f value?
6. Was the same order of elution (as determined by the order of the colors observed) for each solvent?

Extensions

1. Many inks and dyes can be separated in this way (see *Demonstration 4*).
2. Likewise you can try many other solvents, *e.g.*, nonpolar organic solvents. Mixed solvents (*e.g.*, aqueous ammonia and ethanol) can also be used.
3. This demonstration can also be developed into a full laboratory experiment.
4. Chromatographic paper used in chromatographic developing tanks can also be used but is unnecessary. Strips or cylinders of rectangular chromatographic paper in small glass jars with covers or beakers with Petri dish covers can also be used. The one directional movement allows the development of several spots at once.
5. For a detailed experiment using paper chromatography with food dyes, see Markow, P. G. (1988). *J. Chem. Ed.*, 65, 899-900. This paper also references six other versions of this experiment or demonstration for teaching the basics of chromatography.

Demonstration 4: Separation of the Components of a Mixture by Column Chromatography

Introduction

The purpose of this activity is to separate the components of a mixture by column chromatography, a very common method for separating a mixture. It is based on the fact that different materials are attracted to adsorbents with varying strengths. These attractions are based on the types of forces of attraction; very polar solutes are strongly attracted to some column media while nonpolar solutes are not. The solvent also plays an important role in the proper operation of the column. If the solvent attracts the solutes too strongly, all of the solutes travel close to the solvent front (high but similar R_f values) and good separation is not obtained.

Safety

Proper laboratory safety procedures must be followed. Any substances spilled should be immediately washed off with large amounts of water.

Materials

Support stand and clamps
Chromatography column*
Cellulose packing*
Glass wool*
Dyes to be separated*
Distilled water should also be available.

*These materials are contained in a kit available from Sargent-Welch; stock number S-18818-10. The refill materials are also available; stock number S-18818-20.

Procedure

It is important that you do the demonstration prior to doing it before a group. It will require several hours to set up and perform.

1. Support the column with clamps on a support stand. Be sure that the column is vertical and that there is sufficient room to place a collecting flask under it.
2. Place a small quantity of glass wool in the bottom of the column.
3. Close the stopcock and fill the column with distilled water (solvent).
4. While tapping the column (use your thumb or the eraser head on a pencil) sprinkle the cellulose into the column slowly. The purpose of the tapping is to help the packing material settle uniformly in the solvent. It will take about 15-20 min to fill the column.
5. Once the column is filled, place a plug of glass wool on the top. The purpose of this plug is to protect the surface of the packing while you add sample and solvent. Open the stopcock allowing solvent to run out of the column until the top level of the solvent is even with the top of the upper glass wool plug, and then close the stopcock.
6. Dissolve a small amount of the solute in 10-15 mL water; use a dropper to place it carefully onto the column.
7. Allow the solvent to run out until the level of the pigment is even with the top of the upper glass wool plug.
8. Carefully fill the column with solvent, then allow the column to run. *CAUTION: You must maintain the solvent above the top of the glass wool plug atop the column. Do not allow the column to run dry.*
9. Once separation is evident you can show students and explain the system.

Data Analysis

1. What evidence do you have that separation occurred?
2. Which colored substances moved the fastest? The slowest?

Extensions

1. If you continue running the column, and collect the different fractions in separate beakers you can demonstrate how complete separation is attained.
2. Try other samples and solvent systems.

Suggestions for Other Demonstrations

CAUTION: Use appropriate safety guidelines in performing these activities.

Dilution: CEPUP (Chemical Education for Public Understanding Program, Lawrence Hall of Science, University of California, Berkeley) has a semiquantitative dilution experiment in which dilution is continued on a multi-well strip or plate until the solution appears colorless.

Supersaturation: *Doing Chemistry* DMEX C22 (Videodisc side C, Frame 20792, and Teacher's Manual, p. C22-48) provides details for the classical sodium acetate hydrate supersaturation demonstration. (See also *The Chemistry of Rocks, Minerals, and Gems* module.)

Distillation: ChemCom (Chemistry in the Community) has a distillation activity using ethylene glycol and water that can be done either as a laboratory exercise or as a demonstration.



Conductivity of Solutions: *Doing Chemistry* DMEX C26 (Videodisc Side C, Frame 25883, and Teacher's Manual, p. C26-62) provides details for a small-scale electrical conductivity demonstration. However, a simpler small-scale electrical conductivity meter can be constructed from a battery and a low-voltage light bulb. Either type is safer than using 120-volt electricity as the electrical circuit, although the use of direct current has the disadvantage that electrochemical deposition can ruin the electrodes if used for long time periods (see *Periodicity* module).

Ice Cream: A fun activity is to study colligative properties through the making of ice cream. The students should undoubtedly perform the activity in the kitchens of the school's Home Economics facilities or at home rather than in the chemistry laboratory. Use a thermometer to show the depressed freezing point of the ice-salt bath, *etc.*

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. Distinguish between the solute and the solvent in each solution in the laboratory and demonstration activities of this unit. *[NOTE: The selection is arbitrary in some cases.]*
2. Why do colloidal particles remain dispersed throughout the solvent instead of falling to the bottom of the container? *[See Content in a Nutshell.]*
3. Why is molarity (M) a convenient unit for solution reactions? *[Solutions used in reactions are usually measured by volume. The stoichiometry of reactions is based on moles. Thus, moles of solute per liter of solution is the most convenient unit.]*
4. Why are so many different concentration units used? Consider the types of situations when each might be preferred. *[Molarity is discussed in the answer to Question 3. Molality is useful for colligative property studies to keep the amount of liquid solvent constant; mole fraction is useful for colligative properties of gases; % by mass for components that are easily weighed; % mass to volume for solids in liquids; % by volume and volume:volume for liquids in liquids; and ppm for trace amounts of solids in liquid solvents.]*
5. Differentiate the properties that would allow the use of distillation for the separation of solution components from those that would allow the use of crystallization. *[Distillation would be necessary if both components were liquids, whereas crystallization is satisfactory for solid components.]*
6. Why are the colligative properties of aqueous solutions affected more by a given number of moles of an electrolyte than by the same number of moles of a nonelectrolyte? *[Colligative properties depend on the number of particles and electrolytes have more particles per mole as a result of ionization or dissociation.]*

Counterintuitive Examples and Discrepant Events

1. When a dog urinates on grass, urea (a fertilizer) is added to the soil in high concentration. Why does the grass die? *[The osmotic imbalance dehydrates the cells in the grass, and it dies.]*
2. Deep-sea divers use helium/oxygen gas solutions rather than air for breathing because of the problem (the bends) that occurs when nitrogen is used. Why? *[Because of the high solubility of nitrogen in water at higher pressures. Helium is much less soluble.]*

Language of Chemistry

Concentration Terms

molarity (molar concentration) (M) moles of solute per liter of solution.

molality (molal concentration) (m) moles of solute per 1000 grams (1 kg) of solvent.

mole fraction (X_1) moles of component 1 divided by the total moles of all of the solution components (including the solvent).

percent by mass mass of solute divided by the total mass of all of the solution components (including the solvent).

percent mass to volume mass of the solute relative to the volume of solvent.

percent by volume volume of the solute divided by the total volume of the solution.

volume:volume the volume ratios of the solution components.

parts per million (ppm) number of mass units of the solute of interest relative to one million mass units of solution; *e.g.*, micrograms per gram or milligrams per kilogram.

NOTE: The last five terms are not on a particle or mole basis for either solute or solvent!

- Molality is moles per 1000 *g* of *solvent* whereas molarity is moles per 1000 *mL* of *solution*.
- Mole fraction is the only concentration unit that is consistent from solvent to solvent.

solution types

A **solution** is a homogeneous mixture. Although solutions are often thought of in terms of liquid solvents, solutions can be solids in solids (alloys, such as brass—zinc in copper), liquids in solids (mercury in zinc—referred to as zinc amalgam), gases in solids (hydrogen in palladium), solids in liquids (salts in water), liquids in liquids (ethylene glycol or ethanol in water), gases in liquids (air in water), and gases in gases (air). Although liquids and solids in gases are theoretically possible, if the particles are small enough to provide a *homogeneous* material in the gas, the material is effectively a gas itself.

colloids

Colloidal particles, which must be large enough to diffract or scatter light, can be either large single molecules (egg albumin, a protein of molar mass about 43,000, or hemoglobin, the oxygen-carrier in your blood of molar mass about 68,000) or clusters of normally insoluble materials (colloidal gold in water or dust in air).

Pattern Recognition

One of the most important aspects of this unit is solving solution concentration problems. Inasmuch as most textbooks have many solution concentration problems, none have been included other than those included with the laboratory and demonstration activities. Instilling the molarity concept is most important for reaction stoichiometry; however, selected problems using the other common concentration units (see *Language of Chemistry*) should also be included.

TIPS FOR THE TEACHER



Common Student Misconceptions

1. **“The solution and melting processes are the same.”**

This misconception is illustrated by the conversation:

Warning: “Are you going out? It’s raining very hard!”

Response: “Don’t worry, I won’t *melt*.”

2. **“It doesn’t matter whether a liter of solution or 1000 g of solvent are used in calculations.”**

Because molarity and molality are very similar in dilute aqueous solutions, students often assume that the above statement is true. To avoid this confusion, use solvents with densities other than approximately 1.0 g/mL and check aqueous solution calculations very carefully to insure that the right units are being used.

Decision Making

Challenge students to consider the practical methods for desalination of sea water. Processes that have been considered include reverse osmosis, ion exchange, freezing, nuclear-powered distillation, and solar vaporization and recondensation. Have them evaluate the relative costs, volume of water per unit of space required, and other factors that they feel are important to finding the best solution to the long term water shortage problem that exists in many parts of the world.

HISTORY: ON THE HUMAN SIDE

1. When Arrhenius developed the concept of electrolytes in solution, he was still a student. The idea that ions were moving about in water was considered so preposterous by his doctoral committee that he just barely passed his doctoral examination (see *Acids and Bases* module).
2. Solutions have played an important part in human history for many centuries. It is believed that wealthy Roman citizens met early deaths because of lead poisoning from the lead pipes they used. Other Romans, lacking the luxury of piped water, were not affected by the problem.
3. Shortly after the introduction of detergents containing sulfonate groups, lakes and streams developed huge amounts of foam. The synthetic detergents did not break down. A switch to phosphate detergents solved the foam problem, but led to eutrophication (see definition in *Links and Connection* section) of lakes and streams instead. The reformulation with minimal phosphate content has minimized that situation considerably though further research is continuing to develop even better detergents.

HUMOR: ON THE FUN SIDE

1. “Mathematicians have problems, but chemists have solutions.” [Students could make bumper stickers.]
2. $\frac{2 \text{ NaCl(aq)}}{C_7}$ [*Saline, saline, over the seven C’s*]
3. Then there are the witches from William Shakespeare’s *Macbeth*:
“Double, double, toil and trouble,
Fire burn and cauldron bubble.”
Even the great bard of the English language was into making solutions!

4. A modern young man thought he'd boobed
When the milk he had ordered came tubed
Instead of "One liter"
The package read (neater)
"Contents are one dm cubed."

(CHEM 13 NEWS, September 1978, p. 25)

5. a. **Question:** How do you make antifreeze?
Answer: Hide her scarf and mittens.
b. Solute is what you give a general.

(CHEM 13 NEWS, December 1976, p. 1123)

6. **Another message on a bumper sticker:** Chemists have solutions...

7. Word Search (see Appendix for master copy)

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

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

- Moles of solute per liter of solution.
- Homogeneous mixture of two or more substances that are at the smallest levels of their states of subdivision.
- Solution of copper and zinc.
- Solution of copper and tin.
- Moles of solute per 1000 grams of solvent.
- Type of unstable solution that has a higher concentration of solute than normal at a given temperature.
- Extent to which a substance dissolves in a fixed volume or mass of a solvent at a given temperature.
- Adjective describing a solution in which water is the solvent.
- That solution component into which the solutes are considered to have dissolved; the component that is present as the continuous phase.
- Ratio of moles of one component divided by the total moles of all solution components, including the solvent. (two words)



Answers: 1. MOLARITY 2. SOLUTION 3. BRASS 4. BRONZE 5. MOLALITY
6. SUPERSATURATED 7. SOLUBILITY 8. AQUEOUS 9. SOLVENT
10. MOLE FRACTION

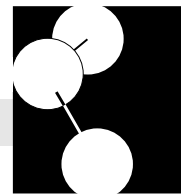
8. See cartoons at end of module.

MEDIA

1. The World of Chemistry videotape "Number 10: Water" has some good materials on solutions. It shows water dissolving a salt, discusses the polarity of water, discusses miscibility-immiscibility, water-soluble polymers, *etc.* 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. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
 - a. *Animated Demonstrations*, by Philip Pavlik. Vol. V B, No. 2, for IBM PS/2. PC-compatible computers.
 - b. *The Acid-Base Package*, by Richard Ramette. Vol. II B. No. 2, for IBM PS/2. PC-compatible computers.
 - c. *Acid-Base Package*, by Richard Ramette and Jon L. Holmes. Vol. IV C. No. 2, for the Apple Macintosh.
 - d. *Equilibrium Calculator*, by Robert D. Allendoerfer. Vol. VI B, No. 1. for IBM PS/2. PC-compatible computers.
 - e. *Equilibrium Calculator*, by Robert D. Allendoerfer, Vol. I D. No. 1, for Windows running on IBM PS/2, PC-compatible computers.
3. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
 - a. For the Apple II computer: AP 502, AP 506.
 - b. For the Apple Macintosh: MC 50 1.
 - c. For IBM PCs and PC-compatibles: PC 2501
4. Videodiscs published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue, Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
 - a. "Changes of State in Bromine," "The Hydrologic Cycle," "Acid Rain and Limestone" and "Lime Water and Carbon Dioxide," four chapters on *The World of Chemistry: Selected Demonstrations and Animations: Disc II* (double sided, 60 min.), Special Issue 4.
 - b. "A Solution with Water," "Alcohol Dissolves in Water," and "Water and Oil," three chapters on *The World of Chemistry: Selected Demonstrations and Animations: Disc I* (double sided, 60 min.), Special Issue 3.

5. Additional software available from QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
- J & S Solutions* (J & S; Apple-II and IBM-PC) Naming of compounds, calculating ion concentration, molarity problems, calorie problems, and the formation of precipitates.
 - Solutions and Near Solutions* College Board Achievement Test in Chemistry Series (Intellectual Software, Queue; Apple-II, IBM-PC, and Macintosh) Properties of solutions, molarity, molality, normality, colloidal suspensions, Brownian motion, Tyndall effect, hydrates, and efflorescent and deliquescent substances.
 - Solution Equilibrium, Colligative Properties, and Acids and Bases.* College Board Achievement Test in Chemistry Series (Intellectual Software, Queue; Apple-II, IBM-PC, and Macintosh) Vapor pressure depression, boiling point elevation, freezing point depression, osmotic pressure, and other topics.
 - FM Solutions* (Focus Media; Apple-II) Types of solutions, solubility factors and curves, and solution problems.
 - SEI Solutions* (Sliwa Enterprises; Apple-II and IBM-PC) Types of solutions and chemical equilibria.
 - Topic 11: Solutions.* General Chemistry 1A Series for AP or College Chemistry (Knowledge Factory; Apple-II) Solution terminology, concentration, colligative properties, electrolytes and non-electrolytes, separation methods, and diffusion.
6. *Solutions: Ionic and Molecular*, 23 min. video available from Coronet/NTI Film and Video, 108 Wilmot Road, Deerfield, IL 60015; (800) 621-2131; (708) 940-3640 (FAX).

Links/Connections



WITHIN CHEMISTRY

Solutions and solution concentrations are very important in solubility and precipitation, chemical reactions, stoichiometry, chemical equilibria, acids and bases, electrochemistry, separations, qualitative analysis, and rates of reactions. See the respective *SourceBook* modules.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Biological sciences require a thorough understanding of aqueous solutions and concentrations. Health-related sciences—as pharmacy, nursing, medicine, and dentistry—make extensive use of solutions. Solutions are also of concern in chemical engineering and related fields. Because both air and natural waters are solutions, environmental scientists are considerably involved with solutions.

TO THE CONTEMPORARY WORLD

Personal

Consider the concentrations of key ingredients of various items in your home. Look at a box of breakfast cereal that is vitamin fortified and note the concentration units that are used. Check other food and beverage items as well. If you have a bottle of tincture of iodine in your medicine chest, note the units of concentration that are used. Check other over-the-counter drug items; *e.g.*, 0.09% saline, 3% hydrogen peroxide. How about shampoos? Cleaning solutions? Adhesives?

Community

Both air and natural water supplies are solutions. Water supplies typically have many ions and polar molecules dissolved in them. Consider the water supply system in your community. Is your water considered a “hard-water” supply? Why is it called hard water? How can you test for the hardness? (See *Biogeochemical Cycles* module.) How is the level (concentration) of chlorine in your water determined? Find out by contacting the appropriate “water chemist” in your community. Are there pollutants in your water supply? How are they measured? Or are they? What about air pollution in your community? Acid rain? More detailed information can be found in the *Chemistry in the Community (ChemCom)* units on water and the atmosphere.

Other ecological situations can be considered; *e.g.*, DDT (an organic pesticide once in very wide use), accumulates in the fatty tissues of animals. Determine how this fact fits with the classic expression: “like dissolves like.” Also water passing through a landfill or garbage dump dissolves many substances that can eventually contaminate water supplies. These problems involving solutions can provide interesting class discussions, especially when probed to find answers (solutions!) to these problems.

Many workers in your community use solutions in their work, whether they know it or not. Have students think of persons who deal with solutions. A few hints follow: Gasoline has additives in addition to its mixed hydrocarbon content. Antifreeze is an ethylene glycol/water solution with added rust inhibitors. Hair stylists use “permanent” wave solutions and shampoos, which may be either true solutions or colloids. Many food and beverage solutions can be obtained in stores and restaurants. [The brine (aqueous NaCl solution) used in pickles provides a good example of solution control. If the NaCl concentration in the brine is too high, the pickles will shrivel up.] The use of solutions in the medical field was briefly noted above. Divers use a helium/oxygen gas mixture for breathing in order to avoid “the bends.” (Helium is not as soluble in water as is nitrogen.)

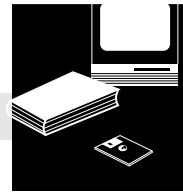
The World

The oceans provide the most abundant aqueous solution on earth. The seas contain 97.6% of the water on this planet. The water content of the ice caps and glaciers (1.9%), ground water (0.5%), rivers and lakes (0.02%), soil moisture (0.01%), and the atmosphere (0.0001%) is significantly lower. Although we normally consider ocean water as salt water, ions other than Na^+ and Cl^- are also quite prevalent as are dissolved gases and organic materials. A consideration of the components of the ocean water is an interesting exercise. See the article on "Oceans and Seas" in the *Encyclopaedia Britannica* or the *Chemistry of Seawater* module for details.

Naturally, the atmosphere is another important solution composed of nitrogen, oxygen, argon, carbon dioxide, water vapor, smaller amounts of sulfur oxides and nitrogen oxides, *etc.* Colloidal dust particles exist in both water and air supplies.

Another important solution example is the *eutrophication* of fresh-water lakes and streams, *i.e.*, the overgrowth of algae due to excessive nutrients (nitrates and phosphates, in particular) in the water from fertilizer runoff from homes, farms, golf courses, *etc.* and from detergents (phosphates).

References



Module developed by R. D. Archer, W. G. Cumming, and A. M. Rennert, the New England team.

Brooks, D. W. (Producer). (1989). *Doing chemistry* [Videodiscs, computer program, and supporting written materials]. Washington, DC: American Chemical Society.

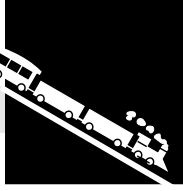
Lippincott, W. T. (Principal Investigator) and American Chemical Society team. (1988). *ChemCom: Chemistry in the community*. Dubuque, IA: Kendall/Hunt. Concept of solutions is discussed in "Investigating the Cause of the Fish Kill."

Markow, P. G. (1988). The ideal solvent for paper chromatography of food dyes. *Journal of Chemical Education*, 65, 899-900.

The New Encyclopædia Britannica. (1986). Chicago, IL: Encyclopedia Britannica. The articles on "Solutions" and on "Oceans and Seas" are relevant to this module.

Wilbraham, A. C., Staley, D. C., Simpson, C. J., and Matta, M. S. (1987). *Chemistry laboratory manual*. Menlo Park, CA: Addison Wesley.

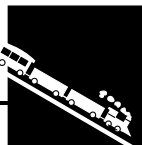
Laboratory Activity 2 is an extension of a phase-transition experiment in this manual.



Appendix

- **Transparencies**
 1. Solute Concentration Units
 2. Aqueous Salt Solution
 3. Word Search

- **Humor**



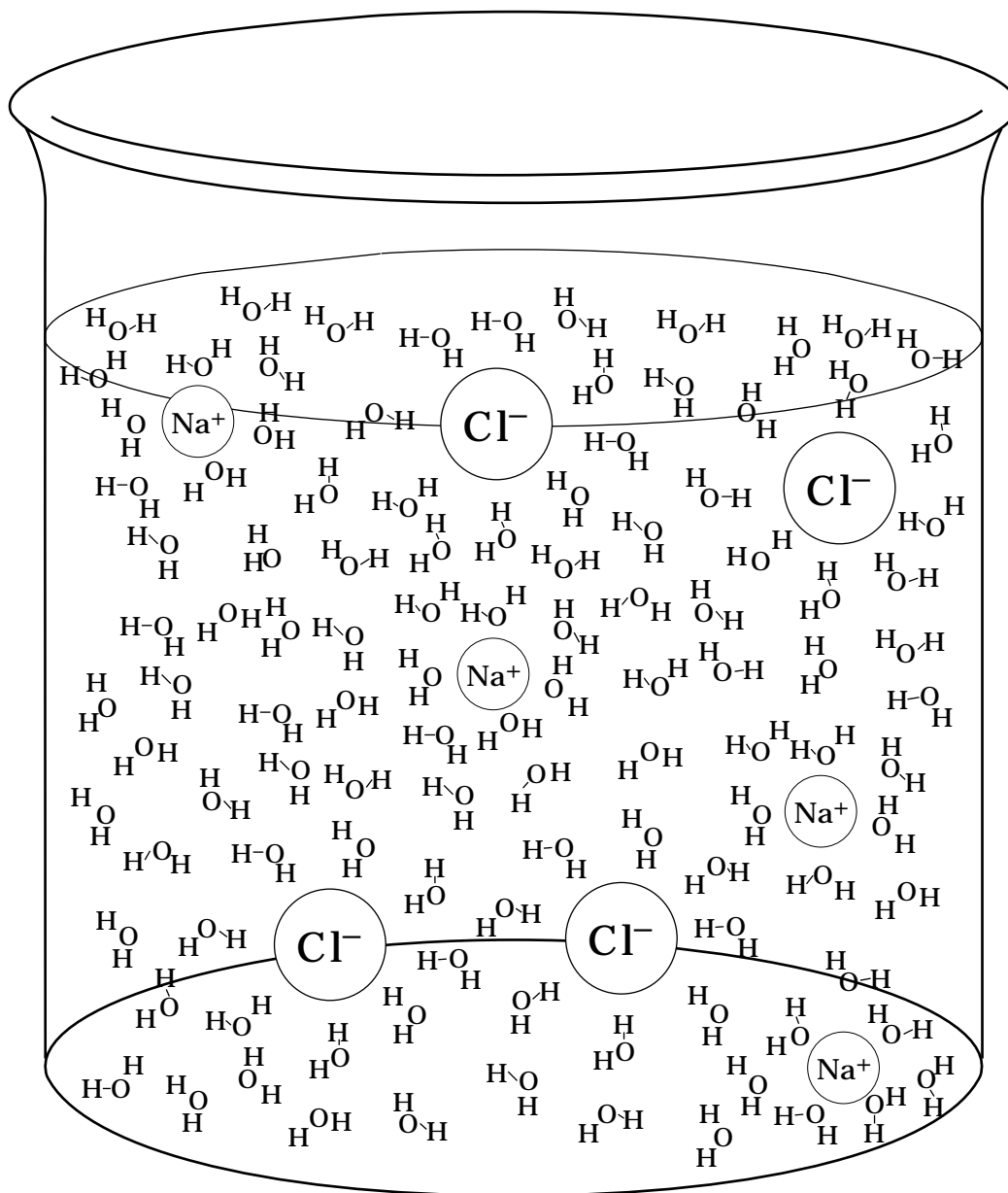
Solute Concentration Units

Molarity = Moles of solute/liter of solution

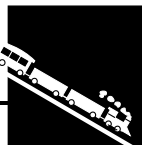
Molality = Moles of solute/1000 g of solvent

Mole fraction = Moles of solute/moles of solution*

*including moles of solvent

Aqueous Salt (NaCl) Solution

*Note H end of water nearest Cl⁻ and O side of water nearest Na⁺.

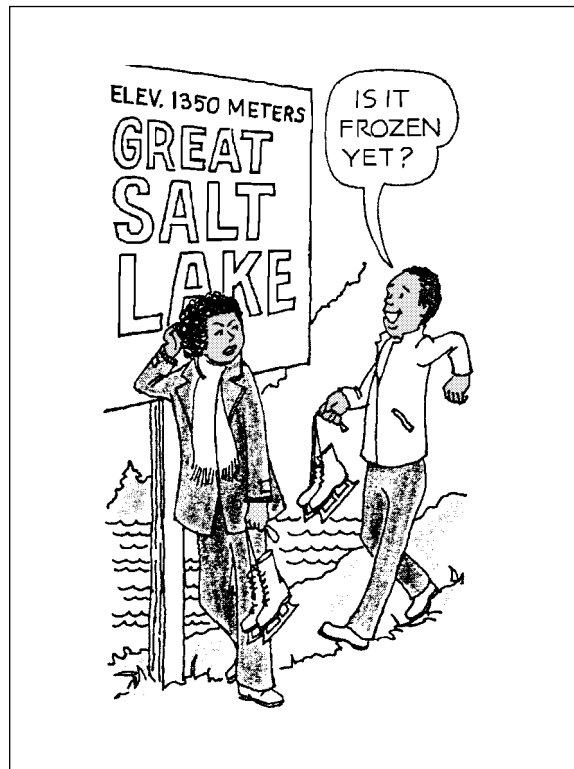


Word Search

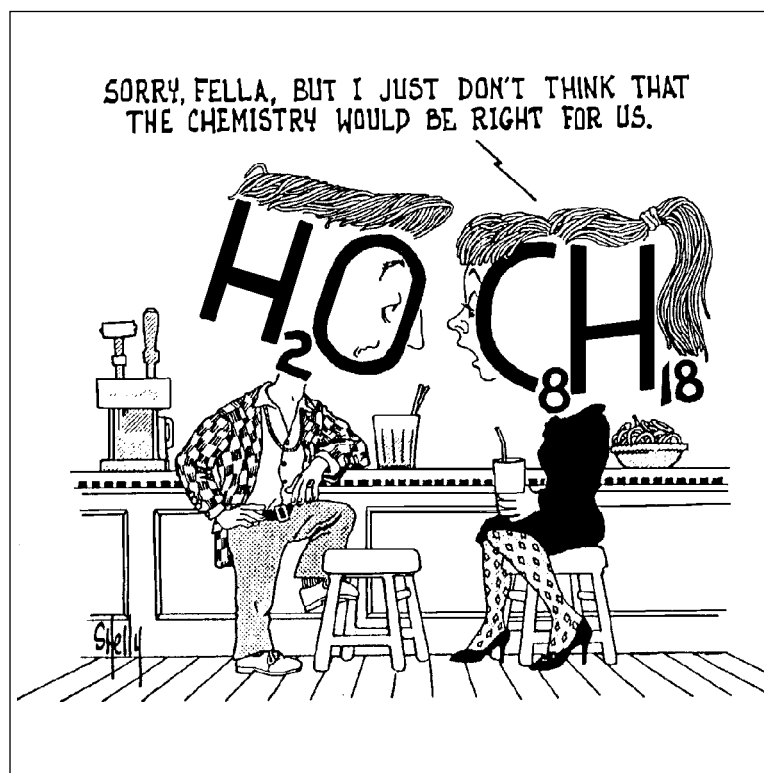
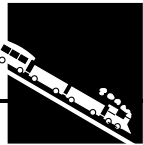
M E F P P T F T B O P J G F T B B
O K E Y C H N N O I T U L O S R J
L C R T R E Q E U W Q T R A O A D
E W U I V W G M Y P H Z F N F G S
F A G L K N N T T C G C Z K A I Z
R Q O I S C I V I P T E F M Y G S
A S Y B S L W I R J J W R D S U P
C C V U A Z L I A R H V D L O U Q
T L G L R K D K L Q Q T T E G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

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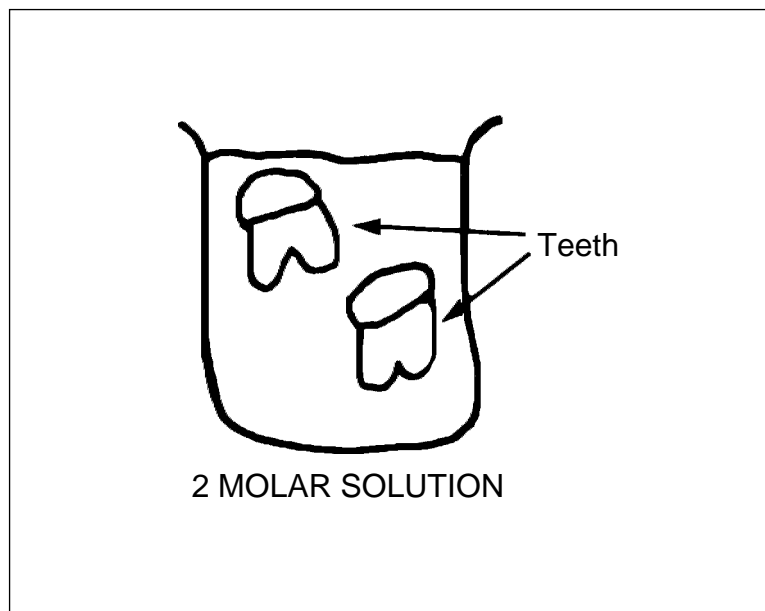
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8. Adjective describing a solution in which water is the solvent.
9. That solution component into which the solutes are considered to have dissolved; the component that is present as the continuous phase.
10. Ratio of moles of one component divided by the total moles of all solution components, including the solvent. (two words)



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