UNIT 8

When Chemicals Meet Water

Unit Overview

This unit addresses the physical and chemical properties of solutions. Although several topics addressed are common to all solution-phase systems, most of the examples are based specifically on aqueous systems since these are fundamental to life and common in inorganic chemistry. The formation of a solution is dependent upon the interactions between the solute, the substance that gets dissolved, and the solvent, the substance that does the dissolving. These interactions are heavily influenced by temperature and pressure. Pressure is particularly influential for gaseous solutes in liquid solvents. The physical properties of solutions differ from those of pure solvents and can reveal information about the identity and quantities of unknown components of a given solution.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Identify the solvent and solute(s) in different types of solutions.
2. Describe differences in solubility behaviors for solid, liquid, and gas solutes.
3. Calculate concentrations of solutions and understand the differences between different concentration units.
4. Understand how solutions' physical properties differ from those of pure solvents, and how concentration affects the extent of these differences.
5. Give examples of how solutions' physical behavior and solutes' chemical reactivity can be used to analyze the identity and amount of species in solutions.

Key Concepts and People

1. **Solutions:** Solutions are, by definition, homogeneous mixtures. Although we typically think of solutions as liquids, all phases of matter (in any combination) can potentially form solutions. Regardless of the physical state of solute and solvent, a solution can form only if the components can combine into a homogeneous mixture.
2. **Solubility**: In a solution, the substance present in the greatest quantity is referred to as the solvent; anything dissolved in the solvent is a solute. Structures of the components in the solution account for the observed "like dissolves like" rule.

3. **Solution Concentrations:** Concentrations of solutions (the amount of solute dissolved in a given volume of solvent) may be expressed in any of several ways. Concentration units, such as molarity, that express quantities in moles, simplify calculations for analysis.

4. **Analyzing Solutions:** Gravimetric and titrimetric analysis can reveal the contents of a solution. Gravimetric analysis is based on the collection of solid from a precipitation reaction; titration is based on the gradual addition of one component of a reaction to the other until the reaction goes just to completion.

5. **Gases and Solutions:** Gases dissolve in liquids and they can mix with other gases. The behavior of these solutions depends on partial pressures of the gases. Raoult's Law and Henry's Law apply to solutions with gases.

6. **Colligative Properties:** A solution's physical properties, such as boiling point, freezing point or vapor point depression, differ from those of a pure solvent. The effect is quantifiable based on the amount of solute, and also depends on whether the solutes are volatile, nonvolatile, or gaseous.

7. **Separation and Purification**: The physical and chemical properties of solutions are the basis for a number of purification and analytical techniques. Filtration, distillation, extraction, and chromatography all offer physical means of separating the components of mixtures, either for purification or analysis.

Video

The majority of chemical reactions happen in solutions—whether inside an espresso machine or in a human cell. For example, when we breathe, the nitrogen in the air dissolves in our blood. Henry's Law gives us the power to predict, prevent, and treat "the bends"—a life-threatening condition that can happen to SCUBA divers when nitrogen in the blood comes out of solution and forms gas bubbles. Solution chemistry provides tools to measure the concentration of components of solutions, like the CO_2 levels in ocean water. Knowing the concentrations of components in solutions can help determine the health of the world.

VIDEO CONTENT

Host Introduction
"What is a Solution?"

Dr. Adam Brunet, a chemistry professor at American International College, explains that solutions are homogenous mixtures and they are not limited to liquid mixtures. However, many chemical reactions take place in liquid solutions, especially aqueous ones.

Real World Application
"The Coffee Solution"

You might not think of coffee as a solution, but it is a solution containing hundreds, even thousands, of different solutes extracted from the coffee grounds. This segment gives an overview of how different brewing methods influence which compounds make it into the cup—and therefore how the coffee will taste. The different brewing methods capitalize on different parameters: temperature, time, and pressure, as well as the grain size of the coffee grounds.

Host Science Explanation
"The Greatest Solvent"

Dr. Adam Brunet dissolves table salt (NaCl) in water to show a simple model of the dissolution process. An animation at the molecular level illustrates how sodium and chloride ions are attracted to the opposite ends of the polar water molecule to form an aqueous solution.

Laboratory Demonstration
"Solids in Solution"

Harvard University Lecture Demonstrator Daniel Rosenberg attempts to generate a reaction by mixing the solids of two ionic compounds. He then mixes the aqueous solutions of the two compounds to generate a more successful reaction.

Laboratory Demonstration
"The Ammonia Fountain"

Ammonia is a very water-soluble gas. A large quantity can be dissolved in a small volume of water. Therefore, if a few milliliters of water are introduced into a closed flask containing ammonia gas, the internal pressure will decrease as the gas dissolves into the water. This pressure decrease can be used to siphon water from an open container. The flask and the open container can be set up so that the water shoots upward into the flask, producing a well-contained fountain.

Host Science Explanation
"Gases in Solution"

Dr. Adam Brunet goes scuba diving and explains that solutes can be gases and can be soluble in liquids. The solubility of a gas is affected by the external partial pressure of the gas. This is why SCUBA divers must return to the surface slowly after being at depths where the partial pressures of the different gases are higher than at the surface. As the diver comes up to the surface, the partial pressures decrease and the excess gases are released from the blood solution, which can be painful or even fatal if not allowed to happen gradually.

Current Chemistry Research
"CO_2 in the Water"

An increase in carbon dioxide emissions not only affects the composition of our atmosphere, but it also affects our oceans. The concentration of carbon dioxide is of particular concern because aqueous carbon dioxide can form carbonic acid, which in turn lowers the pH of the water. A group of scientists at the Smithsonian Environmental Research Center use Henry's law to monitor the concentrations of carbon dioxide dissolved in the ocean at various locations. They are studying how different concentrations of carbon dioxide affect marine life by monitoring the growth and development of oysters at different carbon dioxide levels.

Unit Text

Content Overview

The majority of chemistry happens in solution, and aqueous solutions in particular are the basis for biological, environmental, and even some industrial processes. This unit discusses the formation of solutions, first in terms of the limits of solubility based on concentration, temperature, and (for gases) partial pressure at the surface. The limits of solubility are particularly influential in separation and analysis of mixtures.

The physical and chemical properties of solutions are addressed both qualitatively (*e.g.*, solutes lower the freezing point of the solution relative to that of the pure liquid) and quantitatively (*e.g.*, the concentration of a solution influences how pronounced the freezing-point effect will be). The use of mole ratios, as introduced in Unit 6, is applied to solution-phase reactions through mole-based concentration units.

Sidebar Content

1. **Homogenized Milk**: Homogenized milk isn't a solution; it's a colloid. The homogenization process forces the fats in milk to disperse evenly; otherwise the fat would float on the top.

2. **Parts Per Million and Parts Per Billion:** When there are very low concentrations of a substance in a solution or if the substance is unknown, scientists can express the quantity in "parts per million (ppm)" or "parts per billion (ppb)."

3. **A Practical Titration: Vitamin C in Grapefruit Juice:** Titration analysis can determine the amount of vitamin C in grapefruit juice.

4. **Partial Pressures and Space Suits:** Humans are adapted to a certain partial pressure of oxygen. Space suits are made so that astronauts can get the right amount of oxygen to breathe.

5. **Liquid Breathing:** The liquid perfluorodecalin can dissolve large amounts of oxygen. Researchers are conducting experiments to see if patients can breathe perfluorodecalin into the lungs in order to treat pulmonary or cardiac trauma.

6. **Osmotic Pressure and Cells:** Osmotic pressures have to do with the amount of dissolved solute between two solutions that are separated by a semipermeable membrane, such as a cell membrane.

Interactives

Historical Timeline of Chemistry

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

Chemistry of Running Interactive

All physiological reactions, regardless of how heavily they influence athletic performance, take place in aqueous solution. The concentration of electrolytes, hemoglobin, and oxyhemoglobin are key variables in the physiological responses to exercise.

During the Session

Before Facilitating this Unit

The video for this unit begins with a view of solution-phase chemistry far more practical than most people might think when they first think of chemistry. Encourage students to think about everyday examples of solution-phase reactions after they view the video and as they read the unit. The video highlights the behavior of gas-phase, acid-generating solutes in the segment on dissolved CO_2. This segment ties in closely with the demonstration on the effect of dry ice on the pH of a beaker of water, which can provide a nice introduction to acid-base chemistry.

Tips and Suggestions

1. **Titration and gravimetric analysis can be performed regardless of the stoichiometry of the reaction**. A successful analysis, therefore, must be treated as a stoichiometry problem with mole ratios as a conversion factor between the reacting species. Because dilution calculations often accompany titration problems, students often assume that the $M_1V_1 = M_2V_2$ relationship automatically holds for titrations as well, and ignore the stoichiometry even when the mole ratio is not 1:1.

2. **The colligative properties are dependent upon the number of particles in solution**, **regardless of their identity**. Therefore, ionic solutes will have a net effect proportional to the number of moles of ions in solution, not just the molarity or molality of the compound overall. Similarly, weak electrolytes (such as acetic acid) will have a net effect from both the ionized portion and any dissolved but un-ionized solute.

Starting the Session: Checking Prior Thinking

You might assign students a short writing assignment based on the following questions, and then spend some time discussing prior thinking. This will help elicit prior thinking and misconceptions.

1. What is a solution?

2. What are some examples of solutions?

3. What is the difference between a pure substance and a solution?

4. Suppose you have two mugs of hot water and put a tea bag into each mug. If you remove the tea bag from one mug after 30 seconds and leave the tea bag in the other mug for several minutes, what will be the same about the two solutions? What will be different? Explain.

5. Suppose you dissolve 10 g of sugar in 100 mL of water in one container, and another 10 g of sugar in 500 mL of water in a different container. Which one would taste sweeter? Why?

6. If you dissolved 50 g of sugar in 500 mL of water, to which of the two solutions above is this solution the most similar? Why?

Before Watching the Video

Students should be given the following questions to consider while watching:

1. What is a solute?

2. What is a solvent?

3. Do you think coffee is a pure substance or a solution? Explain your answer.

4. How do you think liquid solvents are able to dissolve solids?

5. Explain what happens when a SCUBA diver gets the bends.

Watch the Video

After Watching the Video

Use these additional questions as follow-up, either as a group discussion or as short writing assignments.

1. What is a solution?

2. Describe what happens when salt is dissolved in water.

3. Provide an example of how Henry's Law is applied in a real world situation.

4. Explain how the ammonia fountain demonstration works.

Group Learning Activities

Unknown Identification By Precipitation

Objective

Metal cations are sometimes classified according to their "analytical group," based on how sensitive they are to precipitation with various anions. By analyzing how solutions of known cations behave when combined with different anions, students can determine the identity of an unknown compound. This activity illustrates both the principle of precipitation and the use of precipitation as an analytical technique.

List of Materials

- Plastic page protector sheets or overhead transparencies, one per group
- 10 mL each of 1.0 M solutions of $AgNO_3$, $Zn(NO_3)_2$, KNO_3, $Sn(NO_3)_2$, $Ca(NO_3)_2$, NaOH, Na_2SO_4, Na_2CO_3, and NaCl
- 10 mL each of 1.0 M solutions of $AgNO_3$, $Zn(NO_3)_2$, KNO_3, $Sn(NO_3)_2$, and $Ca(NO_3)_2$, labeled as unknowns
- Disposable pipets, one per solution per group

Set Up

Prepare solutions to be used in the activity.

Procedure

1. On a sheet of paper, draw a grid with rows labeled NaOH, Na_2SO_4, Na_2CO_3, and NaCl; and columns labeled $AgNO_3$, $Zn(NO_3)_2$, KNO_3, $Sn(NO_3)_2$, $Ca(NO_3)_2$, unknown. Place the paper under the transparency sheet or in the page protector.

2. Using a disposable pipet, put 1-2 drops of each solution down its column, in each of the rows.

3. Using a disposable pipet, add 1-2 drops of each solution across its row, in each column.

4. Make a note of whether a combination of solutions produces a precipitate, a color change, or bubbles.

5. Your unknown will be one of the five compounds in the columns. Based on your observations, identify your unknown.

Discussion

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

1. Which compound was your unknown? Explain your reasoning.

2. Lack of a precipitate can be as important an observation as formation of a precipitate. Explain.

3. Would the reaction of an acid and base, such as HCl and NaOH, have been observable here? Why or why not? If not, what could you do to make the reaction noticeable?

4. This experiment was run on the microscale in order to minimize waste, but doing so limits certain types of observations. What types of observations do you think might be easier if this experiment were performed on a larger scale and why?

Hazards
It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Prepare solutions under a fume hood. Wear eye protection, gloves, and lab coat while preparing the solutions and performing the experiment. Wash hands after performing the experiment.

Disposal
Page protectors can be wiped down with a damp paper towel, and the paper towels disposed in the trash. Check local regulations regarding the disposal of the chemicals.

In-Class Chemical Demonstrations

Colligative Properties and Henry's Law Observed: Instant Ice

Objective
A solution of carbon dioxide and water has a lower freezing point than pure water. When a sealed bottle of carbonated water is placed in an ice-salt bath, the carbon dioxide in the water prevents the carbonated water from freezing. However, once the bottle is removed from the ice bath and opened, carbon dioxide is released and the water inside the bottle instantly freezes.

Materials
- Ice bucket
- Rock Salt
- Bottle of unopened carbonated water
- Bottle of unopened regular water
- Thermometer

Procedure
1. Add rock salt to a large ice bucket.

2. Submerge the unopened bottle of carbonated water and the unopened bottle of regular water into the ice bath. Monitor the temperature of the ice bath so that it goes below 0°C. Allow both bottles to sit in ice bath for about 10-20 minutes and ask students to predict what will happen to both bottles of water.

3. Remove both bottles of water after enough time has passed to allow the temperature of the water inside the bottles to match the temperature of the ice-salt bath. Ask

students to observe the bottles. It may be easier for students to see the inside of the bottle when the labels are removed, however, be sure to label which bottle contains carbonated water and which bottle contains regular water.

4. Ask students to predict what will happen when you open both bottles of water.
5. Open both bottles of water. The carbonated water will instantly freeze once the carbon dioxide leaves the water solution.

Discussion
These questions can help guide students thinking during and after this demonstration:

1. Using the principles of Henry's Law explain how water can be carbonated.
2. What happens when you open a regular bottle of soda water? What is the hiss you hear?
3. Why does the non-carbonated water freeze in the bucket, while the carbonated water remains a liquid?
4. Why does the carbonated water instantly freeze when it is opened after being submerged in the ice bath?
5. Why did we add salt to the ice in the ice bath?

Hazards
If using a glass container of water, when the water freezes it may expand enough to break the glass. Be careful when handling; wear eye protection and protective gloves.

Disposal
There are no special disposal considerations.

Carbon Dioxide Acid/Base Chemistry

Objective
As illustrated in the video, even small amounts of carbon dioxide dissolved in water can generate enough acid to produce a noticeable change in the pH of the resulting solution. This demo uses dry ice as a source of large amounts of CO_2. Visual indicators and/or pH meters may be used to illustrate the change in pH from the introduction of dry ice into a beaker of water.

List of Materials
- Beaker(s) of distilled water
- Dry ice chunks
- NaOH solution to adjust initial pH of water if necessary
- Acid/base indicator(s) (Alizarin, bromothymol blue, and bromcresol purple change color in the pH 6-7 range and will likely require the least adjustment of the initial pH; phenolphthalein becomes colored only around pH 8 and so will require addition of more

NaOH in order to show any effect from addition of dry ice; universal indicator is colored at every pH, but the easiest pH changes to see are likely the blue-green-yellow from pH 9 down to pH 6. Red cabbage indicator may also be used.)

- pH meter if desired
- HCl solution to demonstrate "control" sample(s) for indicator, if desired

Procedure

1. Set up one or more beakers of distilled water containing just enough NaOH to bring the desired indicator(s) to the color corresponding to the upper bound of their color transition.

2. If using a pH meter, demonstrate that the pH of the solution(s) is at or above 7.

3. Add a chunk of dry ice to the beaker. As the dry ice dissolves into the solution, the pH of the solution changes due to the formation of carbonic acid. The indicator(s) will change color accordingly.

4. If using a pH meter, demonstrate that the pH of the solutions after the addition of the dry ice is lower than the initial pH.

5. If desired, a parallel demonstration with the addition of HCl instead of dry ice may be run to demonstrate the effect of the addition of acid to each indicator used.

Discussion

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

1. What compound is dry ice?

2. In the video, you learned about the acidification of bodies of water from higher partial pressures of CO_2 in the atmosphere. Do you think that the dry ice can make the CO_2 concentration in water higher than the partial pressure of the atmosphere can? Explain.

3. The lowering of the pH of water by CO_2 actually requires two steps. In the first, a water molecule combines with a CO_2 molecule to form H_2CO_3. In the second, a hydrogen ion is released from H_2CO_3, which in turn lowers the pH of the solution. Which of these two reactions is the more favorable one? Draw Lewis structures of both compounds to show which bonds are the hardest and easiest to break.

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear gloves, eye protection and a lab coat when handling acids and bases. Dry ice must be handled with tongs or gloves in order to prevent frostbite.

Disposal

Check local regulations for proper disposal.

Going Deeper (In-Class Discussion or Reflection)

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

1. Gases become less soluble in liquids as the partial pressure of the gas on the liquid decreases. When a gas becomes insoluble in a liquid, it bubbles out. What happens when a solid becomes insoluble in a liquid? What about when a liquid becomes insoluble in another liquid?
2. If you boil water to sterilize it before putting it into an aquarium, why is it important to let the water sit at room temperature for several hours before putting it into the tank?
3. Hand-crank ice cream makers use a sleeve filled with water, ice, and rock salt to cool the ice cream mixture. Why does this work?
4. Ethanol and water can be separated by simple distillation only up to 95% pure ethanol. Why is the remainder of the water so difficult to remove?
5. A solution contains one or more of the following cations: [Ag^+, Ba^{2+}, Cu^{2+}, Ca^{2+}]. When an excess of sodium chloride is added to the solution, a precipitate forms. This precipitate is filtered off and sodium sulfate is added to the solution. No precipitate forms. Sodium sulfide is then added to the solution, which produces a precipitate. The precipitate is filtered off and sodium phosphate is added to the solution, which also produces a precipitate. Which ions were present in the original solution?

Before the Next Unit

Students should read the Unit 8 text if they haven't already done so. They may be assigned one or more reading assignments from the list below, or if you choose to have them use the course materials outside of class, they can watch the Unit 9 video and/or read the Unit 9 text as an assignment before the next unit.

References and Additional Resources

Thompson, Stephen. *Chemtrek : Small-Scale Experiments for General Chemistry*. Boston: Allyn and Bacon, 1990.

Atkins, P.W. *Reactions: The Private Life of Atoms*. New York: Oxford University Press, 2011.

For Professional Development

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

Further Reading & Reflection Questions

Valanides, Nicos. "Primary Student Teachers' Understanding of the Particulate Nature of Matter and its Transformations During Dissolving." *Chemistry Education: Research and Practice in Europe*. 1:2 (2000): 249-262. Accessed August 2, 2013. http://www.uoi.gr/cerp/2000_May/pdf/33-06valanides.pdf

1. Have you ever held any of the same misconceptions as these pre-service teachers? If so, do you remember how you addressed them? What concepts do you find particularly challenging?

2. Have you encountered students with these misconceptions? Are there any misconceptions that are more prevalent than others? How have you addressed those misconceptions in the past?

3. Do you agree with the implications of this study that the author presents in the paper? If so, which ones and why? If not, why not?

Calik, Muammer and Alipasa Ayas. "A Cross-Age Study on the Understanding of Chemical Solutions and their Components." *International Education Journal*. 6:1 (2005): 30-41. Accessed August 13, 2013. http://ojs-prod.library.usyd.edu.au/index.php/IEJ/article/viewFile/6792/7434#page=34

1. Were you surprised by the results of the study? Why or why not? Have you encountered students with similar misconceptions as those uncovered in this study? In your experiences, are there any misconceptions that are more prevalent than others? Which ones? How have you addressed those misconceptions in the past?

2. Do you agree with the implications of this study that the authors present in the paper? If so, which ones and why? If not, why not? Do you foresee any challenges in trying to implement any of these recommendations?

3. Does this paper influence how you will approach teaching solution chemistry? If so, how? If not, why not?

Rickey, Dawn and Angelica M. Stacy. "The Role of Metacognition in Learning Chemistry." *Journal of Chemical Education*. 77:7 (2000): 915-919. Accessed August 2, 2013. http://www.physics.emory.edu/Faculty/weeks/journal/rickey-jce00.pdf

1. Did this paper convince you of the importance of metacognition in learning chemistry and solving chemistry problems? Why or why not?
2. How could you use concept maps/ Concept Tests, POE tasks, or MORE Thinking Frames in your chemistry lessons?
3. Does this paper change the way you will approach teaching chemistry concepts? If so, how? If not, why not?

Professional Development Assignments

1. After reading the papers above and reflecting on the questions presented, develop a lesson plan designed to teach material presented in this unit.
2. Using a group activity or classroom demonstration presented in this course guide, show how you would implement it into your classroom. Where would it fit into your curriculum or standards? How would you assess student learning?