



COURSE GUIDE

Chemistry

Challenges and Solutions

**A 13-part multimedia course
in introductory chemistry**

Produced by the Harvard-Smithsonian Center for Astrophysics

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Chemistry

Challenges and Solutions

ABOUT THE COURSE

Course Overview

Matter makes up the world around us, and by studying the properties, composition, and behavior of matter, chemists can manipulate it to help improve daily life. From developing drugs to fight cancer to designing systems to help reduce carbon dioxide emissions, chemistry is at the forefront of solving everyday problems.

By covering the content of a general introductory chemistry course, this course is designed to reveal how chemistry is a discipline that is integral to our understanding of how the world operates. The course itself consists of 13 units, which include text, interactives, and 13 half hour videos. The written text provides the backbone to understanding fundamental chemical principles, while the videos put a human face to chemistry by providing a bird's eye view of the chemistry content discussed in the text and exploring real world applications. The interactives reinforce content discussed in both the text and videos.

Course Guide Components

This guide is designed to help teachers use the course material to teach an introductory level chemistry course. Below is a description of each component of the course guide.

Unit Overview

The unit overview provides a brief look at what the whole unit covers. The goals of the unit are stated here.

Learning Objectives and Applicable Standards

This is a list that explains what the participants of this course should be able to do and understand as a result of completing the unit.

Key Concepts and People

This section provides a list of the major concepts and people discussed in the unit.

Video Overview

The video overview summarizes the key concepts discussed in the video. Generally, we recommend that students watch the video before reading the text.

About the Course

Video Segments Descriptions

The videos are split into segments, which include real world applications, history of chemistry, laboratory demonstrations, chemistry concepts explained by the host, and current chemistry research. The video segment descriptions explain what is covered in each segment of the half hour show. Some teachers may wish to watch certain segments of the videos that relate specifically to sections of the text or other activities or demonstrations described in the guide.

Text Content Overview

This provides an overview of the concepts covered in the written text and shows how the text is organized.

Interactives

There are three interactives for this course: *Historical Timeline of Chemistry*, *Control a Haber-Bosch Ammonia Plant*, and *Chemistry of Running*. The interactives section of the course guide describes how each interactive relates to the content discussed in the specific unit. A separate lesson plan and worksheet are available online for the Control a Haber-Bosch Ammonia Plant Interactive.

Tips and Suggestions for Teaching this Unit

This section points out specific concepts from the unit that are typically difficult for students to understand and provides advice for addressing these difficulties. It also highlights common misconceptions students may have about the related materials and ways to address those misconceptions.

Checking Prior Thinking

The “checking prior thinking” section provides general questions to see what students think they know about the material. The goal is to elicit prior thinking, including misconceptions.

Before Watching the Video

This is a list of questions that should be given to students to answer while they watch the video. The questions are designed to help students think about chemistry concepts covered in the video.

After Watching the Video

This section provides a list of questions to ask students after they have watched the videos in order to help reinforce concepts covered in the video. It may also be helpful to revisit the pre-video questions or the “checking prior thinking” questions to see if students’ thinking has changed.

Group Activities

The course guide outlines some group activities that are designed to complement material presented in the unit. Each activity has an objective describing the learning goals of the activity, as well as sections that include: a list of materials for the activity, procedure, discussion questions, safety hazards, and disposal considerations.

Going Deeper

This is a list of questions designed to help students think critically about the topics discussed in the unit. These are more open-ended questions that are good for both a group discussion and a journal reflection.

Before the Next Unit

This section describes how students can prepare for the next unit.

For Professional Development

This section is for teachers who are taking this course for professional development credit. It provides links to readings from educational research literature as well as questions to help teachers reflect on the readings.

About the Contributors

COURSE DEVELOPER

Christopher Morse

Dr. Christopher Morse is a Lecturer in Chemistry at Olin College. Dr. Morse studied at Dartmouth College before earning his doctoral degree in inorganic chemistry at the Massachusetts Institute of Technology in the lab of Alan Davison where he was a National Science Foundation Predoctoral Fellow. Before coming to Olin College, Dr. Morse was a faculty member in the chemistry department at Tufts University, where his courses covered both graduate and undergraduate curricula. At Tufts, he successfully ran the Summer Institute on College Teaching for seven years and served as the Graduate Training Coordinator, with the responsibility for the pedagogical training of the graduate students, especially those interested in careers in academia. At Olin, Dr. Morse teaches courses in general and organic chemistry. He recently co-authored a textbook for a course about art, art history, and art preservation from a chemical perspective. Additionally, Dr. Morse is the science editor at Sporcle.com, a quiz site where he coordinates and creates study guides and chemistry quizzes for students.

COURSE GUIDE WRITERS

Karen Atkinson

Dr. Karen Atkinson is an associate professor of chemistry at Bunker Hill Community College (BHCC). Dr. Atkinson earned her B.A. in chemistry and medieval/renaissance studies from Wellesley College and her doctoral degree in chemistry from Northeastern University. She has prior teaching experience at Northeastern's University College (now the College of Professional Studies) and Boston College. Dr. Atkinson serves as the liaison for BHCC's Science and

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Engineering Department for various outside research opportunities for BHCC students, including Boston University, Massachusetts Institute of Technology, and Wellesley College. She also developed (and teaches) the department's science writing elective course.

Adam Brunet

Dr. Adam Brunet is an assistant professor of chemistry at American International College. Dr. Brunet obtained degrees in biology and biochemistry from American International College, a private, liberal arts college in Springfield, Massachusetts, then went on to Princeton University for his doctoral degree in chemistry and to University of Massachusetts for a master's degree in business administration. He returned to American International College in 2008 to teach, and currently teaches organic chemistry and biochemistry.

Louisa Morrison

Louisa Morrison is a chemistry teacher at Wellesley High School in Massachusetts. Ms. Morrison studied chemistry at Barnard College and obtained her M.S. in biochemistry and toxicology from Vanderbilt University Medical Center and an M.A. in Teaching from Belmont University. She worked in biochemical research at Woods Hole Oceanographic Institution and Vanderbilt University. She began teaching high school chemistry in the Nashville Public Schools from 2009 to 2011, and is currently teaching at Wellesley High School, a public school in Wellesley, Massachusetts.

Thomas van Geel

Dr. Thomas van Geel has taught chemistry at Wellesley (Massachusetts) High School since 2005. Dr. van Geel holds an M.D., and prior to coming to WHS, he was a teaching fellow at Harvard University.

Jennifer Weeks

Jennifer Weeks worked for fifteen years as a Congressional staffer and public policy analyst before becoming a freelance science writer. She graduated from Williams College and earned master's degrees at the University of North Carolina (political science) and Harvard University (environmental policy). As a freelance journalist, Ms. Weeks has written for more than 50 newspapers, magazines, and web sites, and worked on diverse projects for nonprofit and government clients. She is a contributing writer for CQ Researcher magazine, published by Congressional Quarterly Press, and belongs to the Society of Environmental Journalists and the National Association of Science Writers. She has covered topics including energy, climate change, organic food, agriculture, land and wildlife conservation, waste management, and earth science, in formats ranging from news reports to service articles, detailed features, and book reviews.

UNIT I

Matter and the Rise of Atomic Theory

The Art of the Meticulous

Unit Overview

Since the first time early humans lit a fire and cooked food, people have been manipulating the matter around them. That's what chemistry is – the manipulation of matter – and it has been performed since long before the word “chemistry” ever existed. Every person does chemistry every day, and not just the chemistry that is going on inside the human body.

Ancient people became adept at manipulating materials with the goal of creating products to improve their lives. For example, they fired clay to make ceramics and mixed herbs to create medicines. Eventually, humans tried to create models and theories to explain why certain chemical manipulations worked. This is the point at which chemistry became “a science.”

Our goal in this first unit is to show how chemistry evolved from manipulating materials for practical purposes to Dalton's first atomic theory of matter. It is a story that will start with Democritus around 400 BCE and go to the dawn of the 1800s, when modern scientific methods transformed the way people approached doing chemistry.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Trace the origins of chemistry back to prehistoric times.
2. Understand the scientific method and the use of models in chemistry.
3. Define chemistry broadly as the accumulated knowledge, skills, methods, procedures, and theoretical framework that people use to manipulate matter to serve human needs.
4. Trace how phlogiston and other now-discredited chemical theories came to be replaced by a more scientific model of matter based on particles.
5. Show how the work in the late 1700s and early 1800s set the stage for the chemical revolution that followed.

Key Concepts and People

1. **Origins of Chemistry:** Chemistry started as a set of practices that were passed from generation to generation, with gradual improvement over time but without a detailed theory to guide progress. Chemistry became a formal discipline when researchers began to analyze chemical processes in meticulous detail in order to gain an understanding of their observations.
2. **Chemistry in the Ancient World:** Metallurgy is chemistry's oldest ancestor. It gradually expanded, starting with metals that appear in nature as pure elements, like copper and gold, and later moving to metals that required processing, like iron.
3. **Arab Contributions:** Jābir ibn Hayyān and other Arabs began pioneering the chemical sciences in the 7th century when Europe was in the Dark Ages.
4. **Alchemy:** Nicolas Flamel and Paracelsus developed instruments, techniques, and knowledge of chemical substances and early medical chemistry, which contributed to the rapid advance of chemistry in the 18th century.
5. **Antoine Lavoisier and Other Advances:** Antoine Lavoisier disproved phlogiston theory and demonstrated that oxygen is a substance that combines with metals and can be released again as a gas. This led other researchers, including Joseph Proust and Claude Berthollet, to further define matter as small, indivisible particles that combine in different ways to make all the materials we know.
6. **Proust's Law of Definite Proportions:** All elements combine to make compounds in the same ratios, indicating that there must be a fundamental component that is replicated over and over.
7. **Dalton's Atomic Theory:** John Dalton postulated that matter is made up of small indivisible units called atoms. He then built on the work of Proust and Berthollet to theorize that atoms could only combine together in simple whole number ratios.

Video

This program traces the story of how humans have always practiced chemistry; how, over time, it developed from a practical discipline into a science. Today, we keep up the chemists' tradition to refine and purify substances. A current, real-life application of the "art of the meticulous" is the refining and purification of pure silicon from a common material—silica sand—for advanced electronics, such as cell phones and solar cells.

VIDEO CONTENT

Host Introduction

After a general introduction to the field of chemistry, the host, Dr. Chris Morse, a professor of chemistry at Olin College, introduces chemistry as a discipline and sets the stage for a silicon theme. He walks down a beach and explains that silicon is a component of sand and it is the cornerstone of modern electronic technologies.

History of Chemistry

“The Origins of Chemistry”

This historical overview of chemistry begins with the Greek philosopher Democritus’s theory of “atomos.” Other Greek philosophers like Aristotle came to believe that all matter was made of four elements: air, earth, fire, and water. Harvard Historian, Dr. Ahmed Ragab, explains Jābir ibn Hayyān’s important contributions to early chemistry.

Host Science Explanation

“Separating the Beach”

In this demonstration Dr. Chris Morse separates out sand, salt, and water from a mixture of sand and saltwater. First he passes the mixture through a filter to separate the solids from the liquids. Then, using a distillation apparatus, he is able to separate the salt from the water in the seawater.

Real World Application

“Sand to Sun”

Solar World uses silicon to manufacture solar photovoltaic panels in Oregon. This segment explains the chemistry behind this process.

History of Chemistry

“The History of Oxygen”

Dr. Michal Meyer of the Chemical Heritage Foundation discusses Lavoisier’s experiments that allowed him to disprove phlogiston theory and isolate and measure the elements oxygen, nitrogen, and hydrogen. Lavoisier was further able to demonstrate that water was not an element, but rather a compound of hydrogen and oxygen.

Laboratory Demonstration

“Making Water the Hard Way”

Daniel Rosenberg, Lecture Demonstrator at Harvard University, demonstrates that combining two parts hydrogen gas and one part oxygen gas makes water.

Current Chemistry Research

“Atomic Manipulation”

Dr. Tonio Buonassisi, a professor at the Massachusetts Institute of Technology, studies how to make solar panels more efficient. He and his team are researching the next generation of photovoltaics by examining the purity of the silicon and how it is formed and processed.

Unit Text

Content Overview

The text of this first unit parallels the video, but includes more detail and less emphasis on specific real-world examples. The unit begins by setting up the scientific method. It uses chemistry to help explain the nature of science, specifically that scientific investigations are conducted through careful observation, measurement, and experimentation designed to explain why something happens. Chemistry is a science that requires meticulous attention to detail and relies on careful measurements. This unit emphasizes the difference between precision and accuracy.

The text then provides an overview of 2,000 years of models developed to explain the chemical changes that occur in nature as well as through human activities. Failed models, such as phlogiston theory, are shown to be part of the process that led to the development of a better, more robust science of chemistry.

The unit culminates in Dalton's theory, which explained that atoms were tiny particles that make up all matter. The rules predicted by his theories guided the renaissance of chemistry research during the 1800s.

Sidebar Content

An Official Language for Chemistry: In 1787 Antoine Lavoisier published *Methode de Nomenclature (Essay on Chemical Nomenclature)*, a book that proposed a system for naming and classifying chemicals.

Interactives

Historical Timeline of Chemistry

The Historical Timeline of Chemistry is directly related to Unit I. 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.

During the Session

Before Facilitating this Unit

Generally, we recommend watching the videos first before reading the written units. Some of the video segments directly relate to the suggested group activities or in-class demonstrations. For example, the "History of Oxygen" video segment complements the group exercise "Debating Phlogiston vs. Oxygen" and the "Democritus' Air Displacement Demonstration." Similarly the "Separating out the Beach" video segment relates to the separation exercise presented in this guide.

Tips and Suggestions

1. **Find out ahead of time what your students know about this topic.** Many times, students will hold some of the same misconceptions that early scientific thinkers did.
2. **Some students are confused about the nature of science.** The content and the history in this unit provide a good opportunity to discuss what science is in general.
3. **Some students don't see the relevance of studying the history of science.** To counter this, emphasize that many "facts" that we are convinced of today may turn out to be wrong by future generations. Use this as a way to reinforce the nature of science.
4. **In the West, we tend to discount the scientific contributions made by people of other cultures.** Jabir and the other Arab scientists are excellent talking points for analysis of the factors that might lead toward Eurocentrism and offer the opportunity to bring a multicultural perspective to an introductory chemistry class.

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. Many chemistry courses begin with a unit on measurement. This historical introduction to how chemists' tools, measurements, and procedures gradually improved can be the launching point for a detailed look at measurement in chemistry labs today.

1. What is chemistry?
2. What is the difference between craftsmanship and science?
3. What makes an experiment "scientific"?
4. What is the scientific method?

Before Watching the Video

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

1. When was chemistry invented?
2. What were the first practical uses of chemistry?
3. What major historical events provided the context for the transformation of chemistry into a science?
4. Has the definition of chemistry changed over time? If so, how?

5. How did the development of tools contribute to chemistry as a science?

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. Revisit your answer to the question, “What is chemistry?” Do you wish to change your answer now?
2. What happens when two different theories are created for explaining the same thing?
3. What did each of the following scientists do that was different than their predecessors?
 - a. Jabir
 - b. Lavoisier
 - c. Dalton
4. How does chemistry research advance today? What are the most significant differences between Lavoisier’s laboratory practices and today’s? What are the similarities?

Group Learning Activities

Debating Phlogiston vs. Oxygen

Objective

This activity helps students think about how scientists form hypotheses based on their observations.

Students can take sides and debate the question of phlogiston vs. oxygen theories. Ask them to form groups and work for 5-10 minutes to come up with a convincing argument for their side. Then have a debate where members of each group present their position, followed by a counterargument from the other group. What evidence can they produce to support each argument?

Separation Exercise

Objective

This activity is designed to help students think about how chemists use properties of chemicals and tools to purify substances and separate mixtures. Chemists separate out substances in a mixture by taking advantage of the chemical properties of the different components within a mixture. In this activity, participants are given a mixture of objects and some tools and are asked

to separate the mixture. Students will need to strategize how the tools available can best exploit the different properties of the objects in order to separate them all out.

List of materials

- Small bucket or large beaker (one per group)
- Sand (handful per group)
- Object that floats in water: woodchips, cork, Styrofoam ball, etc. (handful per group—choose only one type)
- Magnetized object: paperclips, screws, nails, etc. (handful per group—choose only one type)
- Lightweight object that is easily blown: feathers, pieces of paper/ confetti (handful per group—choose only one type)
- Heavy object that is not magnetized: buttons, marbles, coins, construction blocks

NOTE: It is important to only have one type of object from each of the groups listed above. For example you want to have either woodchips or cork, not both. You don't want to have multiple objects that have the same properties.

- Magnets (one per group)
- Sieves (one per group)
- Large beaker of water (one per group)
- Trays or pans
- Water

Setup

1. For each group participating in the activity combine sand, floating objects, magnetized objects, lightweight objects, and heavier objects into a small bucket or large beaker. Make sure items are well mixed.
2. Provide each group with a magnet, a sieve, and a large beaker of water. Trays or pans can be used to keep the work area tidy.

Procedure

1. Separate the different components of the mixture. You may use the tools provided to help you sort.
2. Record your methods and reasons for separating the objects the way you did.

Discussion

It is likely that not all the groups will separate the components in the same way. Lead a discussion to compare the different ways groups chose to separate out their mixtures.

What are some of the advantages of doing it one way or another? For example, did any groups

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use the water to float some of the objects? Did they use the water before or after they separated out the sand? Did any group spread out the mixture on a surface and blow the feathers or paper away from the rest of the mixture? The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal.

1. If you were to do this exercise again, how would you separate the objects? Why?
2. How does this exercise mirror how chemists separate different substances?
3. How do tools help advance separation techniques? What would you do differently if you did not have a sieve or if you did not have a magnet?
4. How do tools help advance chemistry as a science?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

In-Class Chemical Demonstrations

Democritus' Air Displacement Demonstration

Objective

This simple demonstration shows that air is matter – it takes up space. In this demonstration a covered glass funnel is inserted into a tank of water. The water does not go up the funnel since air is trapped inside. Because air is made of matter, it takes up space and will prevent other matter from entering that space. When the finger covering the hole of the funnel is removed, the air can escape allowing the water to travel up the funnel.

List of Materials

- Clear glass funnel
- Clear glass tank of water

Procedure

1. Place your finger over the end of an inverted clear glass funnel.
2. Ask the students what will happen when the open end of the funnel is placed face down into a tank of water.

3. Place the funnel into the water and then, after students have predicted the result, release your finger from the short end and let the water seek its own level.

Discussion

These questions can help guide students to think about what they learned from observing this demonstration.

1. What did you observe?
2. How did what you observe compare to what you thought would happen?
3. Why did the water not go up the funnel when a finger covered the hole?
4. What happened when the finger was released from the hole? Why?

Hazards

Glass is slippery when wet. Be careful not to break the glass funnel.

Disposal

There are no special disposal considerations.

Going Deeper (In-Class Discussion or Reflection)

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

1. Why do you think the phlogiston theory was so popular? What aspects of it appeal to “common sense”?
2. Why do you think humans have named the eras of civilization after chemical innovations (Stone Age, Bronze Age, Iron Age, etc.)?
3. Lavoisier was executed by guillotine following the French Revolution. What might the impact have been on history if his wife, Marie Anne Paulze Lavoisier, had not worked to publicize their results?
4. What do you think science is?
5. How does discovering the properties of the materials around you and developing tools based on that help advance chemistry?

Before the Next Unit

Learners should read the Unit 1 text if they haven't already done so. They may wish to read one or more of the 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 2 video and/or read the Unit 2 text as an assignment before the next session.

References and Additional Resources

Djerassi, Carl and Roald Hoffman. *Oxygen*. Wiley, February 2001.

Salzberg, Hugh. *From Caveman to Chemist: Circumstances and Achievements* (American Chemical Society, 1991).

Jackson, Joe. *A World on Fire: A Heretic, an Aristocrat, and the Race to Discover Oxygen*. (Viking, 2005).

Lavoisier, Antoine. *Elements of Chemistry In a New Systemic Order, Containing all the Modern Discoveries*. Project Gutenberg e-book. Released December 28, 2009.

<http://www.gutenberg.org/ebooks/30775>

Levore, Trevor H. *Transforming Matter: A History of Chemistry from Alchemy to the Buckyball*, (Johns Hopkins University Press, 2011).

U.S. National Institute of Health, National Library of Medicine, "Paracelsus and the Medical Revolution of the Renaissance: A 500th Anniversary Celebration,"

<http://www.nlm.nih.gov/exhibition/paracelsus/index.html>.

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 and Reflection Questions

Horton, Christopher. Student Alternative Conceptions in Chemistry. *California Journal of Science Education*, 7(2), 2007. Accessed June 25, 2013.

modeling.asu.edu/modeling/Chem-AltConceptions3-09.doc

1. Do you remember ever holding any of the misconceptions discussed in this article? If so, do you remember how you came to reconcile your way of thinking?
2. How can you use this information about misconceptions to inform how you teach?

McComas, William F. "The Principle Elements of the Nature of Science: Dispelling the Myths." In *The Nature of Science in Science Education*, ed. W.F. McComas, 53-70.

Netherlands: Kluwer Academic Publishers, 1998. Accessed June 25, 2013.

<http://www.torjah.org/articles/McComas-ScienceMytrhs-1998.pdf>

1. How does the information in this unit help clarify any of the myths discussed in the reading?

Do any of the materials in this unit have the potential to perpetuate any of the myths mentioned in this paper? What are ways that you could address these issues in the classroom?

Professional Development Assignments

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

UNIT 2

The Behavior of Particles

Unit Overview

Fundamentally, chemistry is the science of interacting particles. In this unit, students will learn about the variables that control the physical state of matter: pressure, volume, and temperature. Students will be able to distinguish among the phases of matter and explain the properties of solids, liquids, and gases based on their kinetic energy. The text and video provide real life applications of phase changes and a context for interpreting phase diagrams.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Describe phase changes based on kinetic energy.
2. Relate pressure, volume, and temperature.
3. Interpret phase diagrams.

Key Concepts and People

1. **Classifying the States of Matter:** Today, scientists understand that melting, freezing, evaporating, and other changes of state do not alter the identity of a substance, just the distance between the molecules. Gases were the final phase studied because early theories of matter defined “air” as one of the four elements, rather than a mixture of many substances including oxygen, carbon dioxide, and nitrogen.
2. **Measuring Temperature:** Fahrenheit, Celsius, and Kelvin are all temperature scales named after the people who discovered them. Reliable temperature scales allowed for scientists around the world to compare their results, which provided a reproducible standard for studying phase changes.
3. **Measuring Pressure:** Pressure is defined as a force per area, and can be visualized as gas particles colliding with the walls of their container.
4. **Early Gas Chemistry:** Robert Boyle, an English chemist, discovered that pressure and volume are inversely proportional; at a constant temperature, decreasing volume corresponds with increasing pressure. Joseph Black, Joseph Priestly, and Henry Cavendish all made valuable contributions to the understanding of gases.

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5. **Formulating the Gas Laws:** Robert Boyle, Jacques Charles, Joseph Louis Gay-Lussac, and Amedeo Avogadro lend their names to the gas laws that show the relationship between pressure, volume, moles, and temperature. The Ideal Gas Law is the combination of all of the work by Boyle, Charles, Gay-Lussac, and Avogadro. The Ideal Gas Law establishes that one mole of any gas at standard temperature and pressure has a volume of 22.4 L.
6. **Phase Diagrams:** The physical state of a substance depends on the pressure and temperature of the surroundings. Phase diagrams demonstrate this relationship based on the substance.

Video

This program explores the phases of matter—solids, liquids, and gases—and how particles in a given phase interact with each other. Phase diagrams explain at what temperature and pressure a given substance will be in a solid, liquid, or gas phase. Practical problems, like how to safely store enough hydrogen gas to power an automobile, are solved by understanding the different behaviors of solids, liquids, and gases. Understanding the relationships between temperature, pressure, and volume eventually led to the Ideal Gas Law, which provides the platform for examining the conditions under which matter can form a supercritical fluid. Researchers are investigating underground sequestration of supercritical carbon dioxide to mitigate the environmental impact of burning fossil fuels.

VIDEO CONTENT

Host Science Explanation

“Phase Diagram”

The host of this show, Dr. Mala Radhakrishnan, Assistant Professor of Chemistry at Wellesley College, explains what phase diagrams represent and the differences among solids, liquids, and gases.

Laboratory Demonstration

“A Model of a Gas”

In this demonstration Harvard University Lecture Demonstrator Daniel Rosenberg uses golf balls to represent gas molecules. When the temperature of a gas is raised, the molecules of the gas (in this case the golf balls) move faster, and the volume increases.

Laboratory Demonstration

“Phase Change”

At atmospheric pressure solid carbon dioxide, or dry ice, normally goes directly from a solid to a gas. According to the phase diagram of CO_2 , carbon dioxide will go from a solid to a liquid with an increase in temperature and pressure. This can be seen when dry ice is placed in a tube

with a rubber stopper. As the CO_2 goes from a solid to a gas, the pressure inside the tube increases. As the pressure in the tube increases, the dry ice enters a liquid phase. When the pressure is released, the liquid instantaneously goes back to a solid.

Current Chemistry Research

“Hydrogen Storage”

Hydrogen fuel cells are a promising technology that would limit the burning of fossil fuels. Like fossil fuels, hydrogen fuel cells produce energy, but instead of producing carbon dioxide as a by-product, hydrogen fuel cells produce water. The goal is to manufacture cars that run on hydrogen fuel cells, but storing hydrogen for fuel is inefficient. Cooling hydrogen to a liquid state requires a lot of energy, and hydrogen gas takes up too much space. Dr. Peter Wong, Associate Professor at Tufts University, is seeking a way to use nanofibers to increase the storage capacity of hydrogen gas tanks. Hydrogen molecules adhere to the nanofibers, thus storing more molecules in the tank.

History of Chemistry

“Gas Laws”

Through the use of animations, this segment briefly explains the discovery of the Ideal Gas Law, which combines the work of Robert Boyle, Jacques Charles, Joseph Louis Gay-Lussac, Amedeo Avogadro, and Emile Clapeyron. The Ideal Gas Law summarizes the roles of pressure, volume, and temperature in describing the behavior of a gas.

Lab Demonstration

“Volume = Temperature”

Volume and temperature are directly proportional. For example, a helium balloon shrinks when placed in liquid nitrogen. The balloon's volume decreases due to slower moving helium atoms, which make fewer collisions with the walls of the balloon. After reaching room temperature again, the helium atoms move faster and collide more with the walls of the balloon, so the volume of the balloon increases and it rises to the ceiling.

Real World Application

“Supercritical Fluid”

Cl2 Energy, an independent oil and gas company, is working to store carbon dioxide as a supercritical fluid underground in order to prevent carbon dioxide from polluting the atmosphere and contributing to an increase in greenhouse gases.

Unit Text

Content Overview

Unit 2 begins with a discussion of the characteristics of the states of matter, using the Earth's water cycle as a way to illustrate phase changes. Pressure and temperature are the key var-

UNIT 2 The Behavior of Particles

ables that control states of matter. Phase diagrams are visual representations of the states of a particular substance as a function of pressure and temperature. The text includes analyses of the phase diagrams of water on both Earth and Mars, and the phase diagram of carbon dioxide. The pioneers in this field made huge gains by developing reliable measurement scales and drew important conclusions based on observable phenomena, such as hot air ballooning.

Sidebar Content

Torricelli and the Invention of the Barometer: Barometers allow scientists to quantify pressure. The gas laws are mathematical relationships among pressure, temperature, volume, and number of particles. Without a consistent and reliable measurement of pressure, the gas laws would not have been possible.

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. For centuries, scientists focused on the solid and liquid phases for study. Scientists mentioned in this unit are listed on the timeline.

During the Session

Before Facilitating this Unit

An important aspect of teaching this unit is to uncover the prior knowledge of students. Phase changes are observable phenomena that are very familiar to students. Use the video, text, activities, and demonstrations to strengthen their understanding of the physical changes surrounding them. The video includes a demonstration of creating liquid carbon dioxide at room temperature, which is a great example of how environmental conditions control the state of matter. All matter can exist at the three phases, given the appropriate conditions. The video also provides real-world research applications of phase changes demonstrating that even basic science like phase changes can lead to important technological advances.

Tips and Suggestions

1. **Explain Pressure.** Gas laws are much easier to visualize if pressure is described as a direct result of collisions of particles with the walls of a container.
2. **Use as many real life examples as possible.** There are many examples of the effects of pressure, temperature, and volume on every day scenarios. Have students think about tires expanding when heated from driving, or about cake recipes having alternative

instructions based on altitude, or about basketballs that were left out in the cold overnight and no longer bounce. Many of these gas laws make inherent sense to students given the right context.

3. **More advanced students would benefit from understanding that boiling occurs when the vapor pressure equals the atmospheric pressure.** Cooking temperatures and directions are different high in the mountains compared to sea level due to differences in atmospheric pressure.

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. Fill out a KWL chart about the phases of matter. A KWL chart is a great formative assessment to analyze the prior knowledge of students. The chart has three columns: K for “what you know,” W for “what you want to know” and L for what you learned. The K and W sections are a pre-assessment and should be filled out at the beginning of a unit. The L section is a post-assessment for the end of a unit. Students can fill them in individually or the class can generate a list together.
2. When water boils on the stove, what is in the bubbles?
3. What are the similarities and differences among ice, water, and steam?
4. Why is having a reliable temperature scale useful?
5. What is volume?
6. What is pressure?
7. What is temperature?

Before Watching the Video

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

1. What is temperature?
2. What is the relationship between pressure and volume?
3. What is the relationship between volume and temperature?
4. What is a real life example of Charles’ Law?

Watch the Video

UNIT 2 The Behavior of Particles

After Watching the Video

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

1. What do the terms “directly proportional” and “inversely proportional” mean? Give an example of an equation for each.
2. What does it mean to “boil”? Describe what is happening to the particles.
3. Why do gases make inefficient fuel sources?

Group Learning Activities

Marshmallow Madness

Objective

Observe Boyle’s Law using a small marshmallow and needle-less syringe. In this activity a small marshmallow is inserted into a closed syringe. The marshmallow expands when the volume increases due to a decrease in pressure. The marshmallow shrinks when the volume decreases due to an increase in pressure. The main point of this activity is for students to see the inverse proportionality between pressure and volume.

List of Materials

- Needle-less syringe
- Mini-marshmallows

Procedure

1. Remove the plunger from the syringe and place a mini-marshmallow inside the chamber.
2. Replace the plunger.
3. Plug the tip of the syringe with your finger and move the plunger up and down.
4. Observe the change in the volume of the marshmallow as the pressure changes.

Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. What happens to the marshmallow when the volume decreases? Why?
2. What happens to the marshmallow when the volume increases? Why?
3. What is the relationship between pressure and volume?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Bead Manipulatives for the Phases of Matter

Objective

By modeling solid, liquid, and gas phases with small plastic beads in Petri dishes, students can compare and contrast the freedom of movement and the packing of particles in different phases. The dish for “gas” will only contain a few beads, whereas the dish for “solid” will be completely full. The dish for “liquid” should contain just enough beads to cover the bottom.

List of Materials

- Three Petri dishes for each group of students
- Small plastic beads

Set Up

1. Give each group of students three Petri dishes and beads.
2. Explain that the beads represent particles. Each Petri dish is for one state of matter.

Procedure

1. In a Petri dish, arrange beads to create a model of how particles are arranged in a gas phase.
2. Repeat so that you have one Petri dish that represents a gas phase, one that represents a solid phase, and one that represents a liquid phase.

Discussion

The following questions can guide a discussion and help students reflect on this exercise.

1. How do the Petri dishes that represent solid, liquid, and gas differ from one another?
2. Which phase allows for the greatest freedom of movement?
3. Which phase allows for the least freedom of movement?
4. What do scientists mean by the word “model”?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

In-Class Chemical Demonstrations

The Weight of Air

Objective

This demonstration helps students realize that gases have mass, and that a small amount of gas can occupy a large volume.

List of Materials

- Deflated basketball
- Bike pump
- Balance

Procedure

1. Show the students the deflated basketball and mass it.
2. Ask the students how much the mass will change once it is inflated. The guesses will probably have a wide range.
3. Inflate the basketball and mass it again.

Discussion

Students are always surprised by how little mass air has even with a rather large volume. Another important point to bring up here is what is going on inside the basketball to keep it inflated: The molecules of air are exerting force on the walls of the basketball. Here are some questions that can help guide a discussion.

1. Why do you think the air has such a small mass?
2. What is happening inside the basketball to keep it inflated?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Crushing Soda Cans

Objective

This demonstration shows the relationship between temperature and pressure. A soda can is filled with about a half inch of water and allowed to boil. As the water boils, it turns to steam, pushing air out of the can and filling the can with water vapor. Once the water has boiled, the can is inverted and submerged into a beaker of water. As the can is submerged, the water vapor in the can cools and condenses back to liquid water. This decreases the volume of the water inside the can. As a result, the pressure inside the can decreases, so the air outside of the can has enough pressure to crush the can.

List of Materials

- Two empty soda cans (in order to repeat the demonstration)
- Hot plate
- Tongs
- Large beaker
- Water

Procedure

1. Fill the large beaker with cold water from the tap.
2. Fill the soda can with about half an inch of water.
3. Heat the can until the water has boiled.
4. Using the tongs quickly place the can upside down into the large beaker.
5. Repeat the previous step with the second can.

Discussion

The cans crumple very quickly and loudly upon being placed in the beaker of water. This shows that a quick decrease in temperature of water vapor causes a quick decrease in volume, and thus a decrease in pressure inside the can. Students will want to see it twice, so have two cans on hand. These questions can help guide a discussion.

1. What happens to the air inside the can when the water in the can boils?
2. What happens to the gas inside the can when the can is placed upside down in the cool water?
3. Why do the cans crumple?
4. What would happen if instead of using a soda can, we used a thicker, sturdier soup can?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Hard-boiled Egg Trick

Objective

This demonstration shows the relationship between pressure and temperature.

List of Materials

- A hard-boiled egg without the shell
- Erlenmeyer flask
- Matches

Procedure

1. Light a match and place it in an Erlenmeyer flask.
2. Immediately place a peeled hard-boiled egg into the neck of the flask.

Discussion

The egg is sucked into the flask due to a change in pressure. The match initially heats up the air in the flask, causing an increase in pressure. Once the match goes out, there is a drop in pressure, causing the egg to fall into the flask. These questions can help guide students thinking about this demonstration:

1. What happens to the air inside the flask when the match is lit?
2. Once the match goes out, what happens to the air inside the flask?
3. Why does the egg fall into the flask?
4. Why does the match go out?

Challenge Question

1. How can you get the egg back out of the flask without destroying the egg?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Going Deeper (In-Class Discussion or Reflection)

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

1. Revisit the KWL chart you created before the unit and fill in the “L” section- what you have learned. Did you learn everything you wanted to? Is there anything that you thought you knew but found out was incorrect?
2. Does warm soda or cold soda become flat more quickly? Why?
3. Imagine that you left a basketball outside in the winter. Would you be able to shoot some hoops the next day? Why or why not?
4. Why is it so important for airplanes to have pressurized cabins and oxygen masks?
5. Water boils at lower temperatures at higher elevations. Explain.

Before the Next Unit

Learners should read the Unit 2 text if they haven’t already done so. They may wish to read one or more of the 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 3 video and/or read the Unit 3 text as an assignment before the next session.

References and Additional Resources

Dhar, Deepak. “States of Matter.” *Resonance*, 15.6 (2010): 514-525. Web. 1 July 2013.

NOVA. “Absolute Zero,” NOVA. Accessed February 3, 2013.

<http://www.pbs.org/wgbh/nova/zero/>

Phet Interactive Simulations. “States of Matter,” University of Colorado at Boulder. Accessed February 3, 2013. <http://phet.colorado.edu/en/simulation/states-of-matter>

Lappert, Michael F and John N. Murrell, “John Dalton, the Man and his Legacy: the Bicentenary of his Atomic Theory,” *Dalton Transactions*, 2003, pp. 3811-3820. Accessed October 27, 2014. http://matematicas.udea.edu.co/~carlopez/daton_theory.pdf

For Professional Development

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

Further Reading & Reflection Questions

Barker, Vanessa. "Beyond Appearances: Students' Misconceptions about Basic Chemical Ideas."

A report prepared for the Royal Society of Chemistry. Accessed July 8, 2013.

<http://modeling.asu.edu/modeling/KindVanessaBarkerchem.pdf> : 3-19.

1. How do you think children's views about matter may affect their ability to understand the states of matter as they move into high school chemistry?
2. In your teaching experience, have you encountered students who held any of these "naïve" views? Are there any that are more common than others? How have you addressed those views in the past?
3. How could you incorporate this author's "suggestions for progress" in how to teach the particulate nature of matter and phase changes? Do you foresee any difficulties in implementing any of her suggestions?

Milne, Catherine and Tracey Otieno. Understanding Engagement: Science Demonstrations and Emotional Energy. *Science Education* 2007. Accessed July, 13, 2013. DOI: 10.1002/sce

<https://steinhardt.nyu.edu/scmsAdmin/uploads/000/402/Milne%20%26%20Otieno.pdf>

1. Do you think a student's perception of what chemistry is or who chemists are affects how engaged they are in learning chemistry content? If so, how? If not, why not?
2. Does this paper change the way you think about engagement and how to assess student engagement?
3. What do you think makes a science demonstration successful?
4. Does this study change the way you think about science demonstrations or influence how you will conduct science demonstrations in the future?

Professional Development Assignments

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

UNIT 3

Atoms and Light

Exploring Atomic and Electronic Structure

Unit Overview

This unit covers the progression of the model of the atom from the discovery of the first sub-atomic particle to the development of quantum mechanics. After completing this unit, students will understand the major experiments that changed our understanding of the chemical world over the past few centuries. The major model of the atom was transformed from a simple billiard ball to a complex nuclear model with electrons existing as both particles and waves. Students will gain an understanding of how scientists modify existing theories based on newly available evidence.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Compare and contrast the different models of the atom.
2. Describe the experiments of the major atomic theorists and the progression of the models of the atom over time.
3. Evaluate the evidence that allowed scientists to infer the internal structure of the atom.

Key Concepts and People

1. **J. J. Thomson and Robert Millikan:** Thomson discovered the electron with his cathode ray tube experiment. Later, Millikan determined the charge of the electron based on the results of his oil drop experiment.
2. **Ernest Rutherford:** Rutherford shot positively charged alpha particles at a piece of gold foil. He was able to infer the existence of the nucleus because some of the alpha particles were deflected by areas of dense positive charges within the foil. He then confirmed the existence of the proton, the particle that gives nuclei a positive charge. Later, one of Rutherford's students, James Chadwick conducted experiments in which he discovered the neutron.
3. **The Electromagnetic Spectrum:** The whole range of light, or electromagnetic radiation, is called the electromagnetic spectrum. The spectrum ranges from gamma rays

to radio waves with visible light being in the middle of the spectrum.

4. **Light From Elements:** Scientists burned elements in flames and used spectroscopes to study the emission spectra of various elements. Each element has its own “spectral signature” and they were able to discover new elements using spectroscopes.
5. **Niels Bohr:** Bohr explained the emission spectra for hydrogen. This data had existed for centuries, but no one was able to explain why hydrogen only emits certain wavelengths when excited. Bohr surmised that electrons exist at specific, quantized energy levels. His model is called the planetary model because it shows electrons orbiting the nucleus in set paths.
6. **Max Planck:** Max Planck proposed that light is emitted in discrete packets of energy called “photons.” This idea revolutionized chemistry and physics and launched the field of quantum mechanics.
7. **Quantum Model:** Bohr’s model of the atom only worked for hydrogen, but other elements also have unique spectra. The quantum theorists, including Werner Heisenberg, Louis de Broglie, and Erwin Schrödinger, amended Bohr’s model based on the fact that an electron can exist as both a particle and a wave. We use the quantum model today, which describes the energy and location of electrons as electron clouds, or orbitals.

Video

In the early 20th century, identification of the internal parts of the atom (electrons, protons, and neutrons) led to a modern subatomic theory. Meanwhile, the study of atomic spectra – the light given off by atoms at definite wavelengths – led to the Bohr model of the atom, where electrons exist at distinct energy levels and move between these levels by absorbing and emitting discrete quanta of energy. The measurement of atomic spectra has applications in astrophysics as well as forensic chemistry.

VIDEO CONTENT

Host Science Explanation

“Making Light with Electrons”

Host Dr. Michael C. McCarthy, an astrochemist at the Harvard-Smithsonian Center for Astrophysics, studies the chemical composition of the universe. Here, he introduces the electromagnetic spectrum and explains that we can study matter by seeing how matter absorbs or emits light under different conditions.

Real World Application

Dr. Wayne Stratman is an artist who manipulates the properties of gases to make stunning visual effects. Each element has a unique color, and mixtures of different elements are used to create different color effects.

History of Chemistry**“Atomic Models”**

This segment includes animations and demonstrations of key experiments and findings of atomic theorists. Harvard University Manager of Lecture Demonstration Services Dr. Wolfgang Rueckner demonstrates JJ Thomson’s cathode ray tube experiment, his plum pudding model, and Rutherford’s gold foil experiment. Niels Bohr’s model of the atom, and the quantum mechanical model of the atom are also represented. Dr. Michael C. McCarthy also demonstrates that the nucleus of an atom would be the size of a golf ball if an atom were the size of a stadium.

Laboratory Demonstration**“The Flame Test”**

Harvard University Lecture Demonstrator Daniel Rosenberg demonstrates that each element emits a distinct color of light when burned. The flame excites electrons up to a higher energy level, and photons of light are emitted when the electron falls back down to its original energy level. For example, lithium burns bright red and copper burns green.

Current Chemistry Research**“Observing Sunlight”**

Dr. Kelly Korreck of the Harvard-Smithsonian Center for Astrophysics uses spectroscopy to study the chemical composition of the sun. She studies high-energy emissions from the sun and also builds instruments to take images of the sun from satellites. Her images can be analyzed for individual wavelengths in order to better understand solar events.

Laboratory Demonstration**“Emission Spectra”**

Harvard University Lecture Demonstrator, Daniel Rosenberg uses a diffraction grating to visualize the different colors emitted by various metal salts when the salts are held in a flame. For example, the calcium flame appears orange, but the diffraction grating shows it is also emitting some green.

Real World Application**“Forensic Spectroscopy”**

Dr. Megan Roppolo, an Application Scientist at Olympus NDT, demonstrates the use of a hand-held X-ray Fluorescence Analyzer. She uses it to discover the chemical composition of a historical cannon at the U.S.S. Constitution Museum.

Unit Text

Content Overview

This text provides a roadmap for the changes in the models of the atom over time. Unit 2 ended with John Dalton's theory describing an atom as a billiard ball, and Unit 3 progresses through the subsequent models of the atom and the evidence used in their development. The text is organized by subatomic particle or structure rather than chronologically in order to emphasize the transition between the major models.

Sidebar Content

1. **Robert Millikan's Oil Drop Experiment:** Robert Millikan and his student, Harvey Fletcher, perfected the technique for measuring the mass of an electron by using tiny drops of oil in an electric field.
2. **Discovery of the Neutron Made the Atomic Bomb Possible:** This sidebar briefly discusses the role of the neutron in the development of the atomic bomb. An atomic bomb would not have been possible if scientists were not curious about the inner structure of the atom.
3. **Microwaves:** Every student has seen or used a microwave oven, and the sidebar explains how they work. The scientists covered in Unit 3 used different types of waves to elucidate the internal structure of the atom.
4. **Fraunhofer Lines:** Joseph von Fraunhofer used a spectroscope to study the spectrum of light emitted by stars.
5. **Werner Heisenberg and Niels Bohr:** Heisenberg and Bohr had a complex relationship during World War II. Being on opposite sides of the conflict ruined their previously close relationship.

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. Many of the scientists that contributed to the understanding of the atom are contemporaries so students should note how quickly the model of the atom changed in the late 19th and early 20th centuries.

During the Session

Before Facilitating this Unit

The main lessons in this unit include the progression of the model of the atom over time and how each scientist was able to draw conclusions from his data. Students should also understand that a model is based on data, but is not an exact representation. Models are edited or replaced as new technologies and data become available.

Tips and Suggestions

1. **The iconic picture of the atom with orbiting electrons is actually incorrect. Many students assume this representation is the most accurate.** After completing this unit, students should be able to clearly explain the correct internal structure of the atom.
2. **Chemistry is a progressive discipline that can change based on available technology.** Some of chemistry dates back to ancient times, but it remains a dynamic endeavor.
3. **Models are useful, but often are over-simplifications.**

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 an atom?
2. Draw a picture of the atom.
3. Atomic theorists performed experiments and *inferred* the structure of the atom from their data. What does it mean to infer?

Before Watching the Video

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

1. What is a scientific model?
2. Explain a few different models of the atom.
3. What types of evidence were used to show that the atom is divisible?
4. What are some ways scientists use the electromagnetic spectrum in their work?

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. If you wanted to create a firework display of a green star, what element would you use? Why?
2. If an atom was the size of your classroom, about how big do you think the nucleus would be?
3. Why does Dr. Korreck use telescopes on satellites to take pictures of the sun?
4. What is a diffraction grating? What is it used for?

Group Learning Activities

Making an Atomic Theorist Manipulative

Objective

Students will make a foldable study guide and distill the most important information about the theorists and their experiments. This activity provides a teachable moment to help students focus on the most important information. For example, do you want students to memorize years or just focus on the general chronology? What are the major models of the atom and what do they look like?

List of Materials

- Paper
- Scissors
- Rulers
- Writing utensils
- Textbook or lecture notes on theorists

Set Up

Give each student a piece of paper, scissors, and a ruler. They need access to a textbook or notes on the theorists.

Procedure

1. Fold the paper in half lengthwise.
2. Using the ruler, draw lines to divide the folded flaps into strips. Make sure to have enough strips for all of the theorists.

3. Write the name of a theorist on top of the first strip. Write the important information about that theorist underneath the strip. Draw pictures where appropriate.
4. Repeat for each of the required theorists.

Discussion

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

1. Why were the atomic theorists limited to gathering indirect evidence? Does this mean their conclusions are weak? Why or why not?
2. What are the major changes in the model of the atom over time?
3. Why did the model of the atom change over time?
4. Whose discovery do you think was the most important? Why?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Creative Writing Assignment

Objective

Students will research the lives of the atomic theorists and will write a creative assignment from the point of view of a theorist. This lesson provides opportunities for the more creative types in your class. Possible formats include a letter, journal entry, comic strip, poem, page from the scientist's lab notebook, etc. The written product should show research into the theorist's life and discussion of how his model differed from the previous model. When introducing this assignment, you can use the play *Copenhagen* by Michael Frayn as an example. <http://www.pbs.org/hollywoodpresents/copenhagen/index.html> It is a fictionalization of a meeting between Heisenberg and Bohr that occurred during World War II. In 2002, PBS produced a film version of the play, starring Daniel Craig as Heisenberg. The introduction to the play in the PBS video is a great summary of the lives and times of the two scientists.

Models of the Hydrogen Atom Simulation

Objective

Using a PhET simulation, students will investigate the spectral evidence used to determine the current model of the atom. PhET, a project at the University of Colorado at Boulder, uses education research to design online STEM interactives. The programs are free to use.

UNIT 3 Atoms and Light

List of Materials

- Computers with internet access
- Questions sheet (please see discussion section for a list of suggested questions)

Set Up

Hand out the question sheet to each group of students.

Procedure

1. Log on to the computer and go to the following website:
<http://phet.colorado.edu/en/simulation/hydrogen-atom>
2. Use the simulation to answer the provided questions.

Discussion

Suggested questions to include on handout:

1. Why do the photons of white light have different colors?
2. What wavelengths of light are emitted by the atom?
3. After switching to the prediction mode, which models correctly predict the emission spectra of hydrogen? What do they all have in common? What are the differences?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

In-Class Chemical Demonstrations

Rutherford Scattering Simulation

Objective

Students will learn how to infer the shape of an object based on indirect evidence, and understand how Rutherford concluded that an atom has a nucleus. This activity involves rolling marbles at a mystery shape and using the path of the marbles to infer the shape.

List of Materials

- 10-12 marbles
- A piece of cardboard
- A 3-D shape (packing Styrofoam or wooden shapes work well)
- Analysis sheet

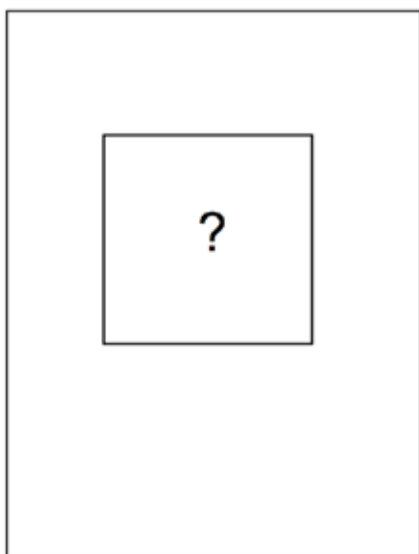
Set Up

1. Without students seeing the shape of the Styrofoam or piece of wood, place that object on the floor with a piece of cardboard on top to hide the shape.
2. Hand out data sheets to each student.

Procedure

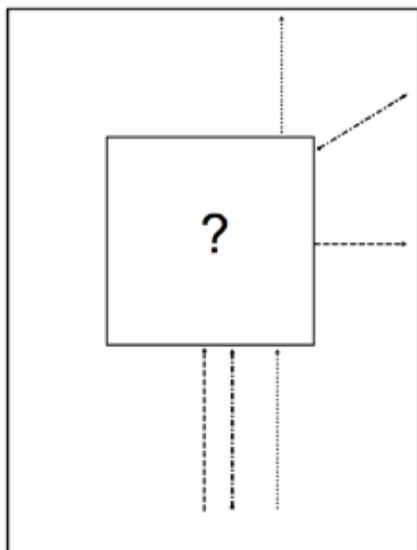
1. Have students take turns rolling marbles toward the shape one at a time and record the approximate path of the marbles on their data sheet.
2. Once there have been enough rolls (usually about 30), have students work in small groups to determine the shape.
3. Reveal the actual shape to the class and discuss the results

The blank data sheet:



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The data sheet with the first few rolls:



Discussion

These questions can help guide a student's thinking during this demonstration:

1. What do the marbles represent in this activity?
2. What does the mystery shape represent?
3. How close was your drawing to the actual mystery shape? Explain.
4. What evidence did you use in your drawing? Explain.
5. How does this activity mimic the experience of Rutherford during his gold foil experiment?
6. Explain the significance of the marbles that went straight through. What did this observation reveal to Rutherford in the actual experiment?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Flame Test

Objective

The flame test is an easy demonstration to perform in order to visualize the differences between elements and the behavior of electrons. There is also a flame test in the video.

Materials

- 25 mL of 1 M sodium chloride solution
- 25 mL of 1 M barium chloride solution
- 25 mL of 1 M calcium chloride solution
- 25 mL of 1 M copper(I) or copper(II) chloride solution
- Wooden coffee stirrers
- Bunsen burner

Procedure

1. Soak the wooden coffee stirrers in the salt solutions for at least 10-15 minutes.
2. Place the soaked stirrers into the flame one at a time, and have students observe the results.

Discussion

Many students are exposed to the flame test during middle school, but have not discussed energy levels of electrons before chemistry class. While the splints are burning, explain to the students that they are actually seeing a combination of different wavelengths, and these wavelengths are the products of electrons changing energy levels within the atom.

Hazards

Barium chloride is toxic when ingested. Be careful with the flames.

Disposal

Check local regulations for proper disposal procedures for any leftover barium or copper salts.

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. What does it mean to “infer”? Why did atomic scientists have to use indirect data to draw conclusions about the structure of the atom?
2. How does the flame test relate to pigments used in clothing or art?
3. In science, what does “model” mean? Can you think of common models used in science?

Before the Next Unit

Learners should read the Unit 3 text if they haven't already done so. They may wish to read one or more of the 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 4 video and/or read the Unit 4 text as an assignment before the next session.

References and Additional Resources

Frayn, Michael. *Copenhagen*. London: Methuen, 1998.

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Godbole, Rohini. "An Experiment that Shaped the Physics of the Century." *Resonance*. 16:11 (2011): 1019-1028.

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Phet Interactive Simulations. "Rutherford Scattering," University of Colorado at Boulder. Accessed March 29, 2013. <http://phet.colorado.edu/en/simulation/rutherford-scattering>

Phet Interactive Simulations. "Models of the Hydrogen Atom," University of Colorado at Boulder. Accessed March 29, 2013. <http://phet.colorado.edu/en/simulation/hydrogen-atom>

Rhodes, Richard. *The Making of the Atomic Bomb*. New York: Simon and Schuster, 1986.

For Professional Development

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

Further Reading & Reflection Questions

Viana, Helio Elael Bonini and Paulo Alves Porto. "The Development of Dalton's Atomic Theory as a Case Study in the History of Science: Reflections for Educators in Chemistry." *Science & Education*, 19 (2010): 75–90. Accessed July 7, 2013.

<http://www.iq.usp.br/palporto/VianaPortoSci%26Educ2010.pdf>

1. Page 76 discusses the reasons for the shift to include the history of science in science education in Denmark. Do you agree with any of these reasons? If so, which ones and

why? If not, why not?

2. Did you find the case study of the development of Dalton's theory helpful for your own understanding of chemistry, and science in general? If so, how? If not, why not?
3. Do you include the history of science in your curriculum? Why or why not?
4. Does this paper influence how you will approach teaching chemistry? If so, how? If not, why not? Has this paper convinced you that understanding the history of science is important to understanding scientific principles themselves? Why or why not?

Gilbert, John K., "The role of visual representations in the learning and teaching of science: An introduction." *AsiaPacific Forum on Science Learning and Teaching*, Volume 11, Issue 1, Forward, p. 1. (Jun., 2010). Accessed October 24, 2014. http://www.ied.edu.hk/apfslt/download/v11_issue1_files/foreword.pdf

1. How do you currently use models to teach certain chemistry concepts? What are some of the challenges you face in using models to teach chemistry concepts? What are some of the challenges your students have had in learning about models?
2. Did this paper help clarify for you how models can be used to teach scientific concepts? If so, how? If not, why not? Was there anything in this paper that stood out for you? Explain.
3. Does this paper influence how you will use models in teaching chemistry? 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 in your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?

UNIT 4

Organizing Atoms and Electrons

The Periodic Table and the Formulation of Compounds

Unit Overview

Unit 4 covers the most important and useful document in chemistry, the periodic table. Students will learn the history of the periodic table and how to decode the large amount of information provided by the table, including atomic mass, relative atomic size, and electron configuration. The unit also delves into the formation of ions and ionic compounds.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Classify elements as metals, non-metals, or metalloids.
2. Synthesize the electron configuration and predict the reactivity of any element based on its position on the periodic table.
3. Identify trends in effective nuclear charge, atomic size, ionization energy, electron affinity, ionic size, and electronegativity.
4. Provide the name and formula of any ionic compound.

Key Concepts and People

1. **Dmitry Mendeleev and Henry Moseley:** Mendeleev is the father of the modern periodic table. He used relative masses and reactivity trends to make his version. Moseley ranked elements based on atomic number.
2. **A Tour of the Periodic Table:** The periodic table conveys important information about the elements, including atomic number, mass, and metallic character. The text provides a quick guide for interpreting this information.
3. **Periodic Trends:** Certain properties of elements, such as atomic radius, electronegativity, and ionization energy, change in predictable patterns from left to right and top to bottom on the periodic table. These trends allow us to rank elements with respect to these properties based on their relative position on the periodic table.

4. **Energy Levels, Orbitals, and Electron Configuration:** An atom's electrons are located outside its nucleus in energy levels and orbitals. There are rules that dictate in which energy level and orbital electrons of a given atom can be found.
5. **Forming Compounds:** Atoms will gain, lose, or share electrons to obtain a more stable electron configuration. To achieve this goal, atoms will form two major types of bonds: ionic and covalent. There are particular naming conventions that govern how ionic and covalent compounds are named.

Video

For centuries, chemists tried different methods to organize elements around patterns of chemical and physical trends, or regularities, eventually leading to the modern periodic table. Electron configuration is a powerful predictive tool, a simple extension from the periodic table. Physical characteristics, including atomic radius and reactivity, all depend on electron configuration and can be teased from a careful understanding of the periodic table. A living document, the periodic table is continually updated as new manmade heavy elements are discovered in research laboratories.

VIDEO CONTENT

Host Science Explanation “Arranging the Elements”

Dr. Mala Radhakrishnan, Assistant Professor of Chemistry at Wellesley College, discusses the organization of the periodic table and the information that can be learned from it.

Laboratory Demonstration “Periodic Trends in Reactivity”

Reactivity is the tendency of an element to undergo a chemical reaction. It increases as atomic number goes up in each group or family of elements. Harvard University Lecture Demonstrator Daniel Rosenberg drops lithium, sodium, and potassium into water to show that reactivity increases as you move down the alkali metal group. The demonstration ends with historic footage of 20,000 pounds of metallic sodium being dumped into a lake.

Host Science Explanation “Electron Configuration”

Dr. Mala Radhakrishnan explains that the reactivity of an element depends on its atomic structure, and in particular, its electron configuration. Only the outermost, or “valence,” electrons are involved in reactions. The formation of compounds is based on the fact that elements are more stable with a full outer shell of electrons. Elements in the same group in the periodic table have the same number of valence electrons, and thus will react in similar ways to obtain a full outer shell.

Laboratory Demonstration

“Alkaline Earth Metals”

Daniel Rosenberg demonstrates that the alkaline earth metals are less reactive than the alkali metals. He also shows that magnesium does not react readily with room temperature water, unlike lithium, sodium and potassium. Calcium, which is below magnesium on the periodic table, is more reactive in water than magnesium, but not as reactive as lithium, sodium, or potassium.

History of Chemistry

“Seaborg’s Elements”

The periodic table is arranged in orbital blocks. The *s*, *p*, and *d* blocks demonstrate predictable patterns, but discovery of *f* block elements led to a rearrangement of the periodic table.

Dr. Darleane C. Hoffman recounts the discovery of elements synthesized in the Lawrence Livermore National Laboratory in California during World War II. Dr. Glenn T. Seaborg, her collaborator, added the actinide series at the bottom of the periodic table because plutonium acted more like a lanthanide than a transition metal. Seaborg’s team discovered many of the elements in the actinide series and beyond. Element 106 is named seaborgium in honor of Dr. Seaborg’s many contributions to chemistry.

Current Chemistry Research

“Synthesizing New Elements”

Scientists at Lawrence Livermore National Laboratory are working on synthesizing superheavy elements, which are defined as elements above atomic number 107. They use particle accelerators to combine two smaller nuclei. Research teams analyze data about new, superheavy elements to determine physical and chemical properties.

Unit Text

Content Overview

The unit begins with a look at the scientists responsible for the development of the periodic table: Dmitri Mendeleev and Henry Moseley. The text then provides a tour of the periodic table, explaining the information contained in the individual boxes in the table as well as how to identify elements as metals, non-metals, and metalloids based on their location within the table. The text defines atomic mass, atomic number, and isotopes. The unit then discusses electron configuration and describes how a full valence shell is related to the stability of a given atom. The subsequent three sections focus on periodic trends, including effective nuclear charge, atomic radii, ionization energy, and electron affinity. The last two sections describe forming and naming ionic compounds.

Sidebar Content

- I. **The Last Natural Element:** Rhenium represents the last element that was discovered in nature. The periodic table has continued to expand due to the discovery of

UNIT 4 Organizing Atoms and Electrons

man-made elements in the laboratory.

2. **The Periodic Table as Art:** The periodic table can have many different arrangements, including artistic renderings. *Divining Nature: An Elemental Garden* is a 3-D art installation by Washington, D.C. artist Rebecca Kamen.
3. **Athletes, Artificial Steroids, and Carbon-14:** Different isotopes of carbon have different properties. By testing for the presence of carbon-14, carbon-12, and carbon-13 in the blood stream, sporting organizations can detect if an athlete has been using synthetic steroids.
4. **Mass Spectrometry:** Mass spectrometry identifies compounds based on their mass-to-charge ratio. Compounds are vaporized, ionized, and sent through magnetic plates toward a detector.
5. **Absorption of Strontium in the Aftermath of Chernobyl:** This sidebar explains the health risks of strontium-90 absorption in the area surrounding Chernobyl following the nuclear disaster in 1986. Strontium-90 can replace calcium in bones and cause many health problems.

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. It might be interesting to point out that Mendeleev developed his version of the periodic table well before the atomic theorists knew the internal structure of the atom.

During the Session

Before Facilitating this Unit

Some of the topics in this unit can be more methodical and less relatable to everyday life, such as electron configuration and nomenclature. It is important to give students lots of practice problems, but also to provide students with opportunities to move around the classroom and use their hands. These topics will often elicit questions such as, “When am I ever going to use this?” Use the video to show real-world applications.

Tips and Suggestions

1. **It is not necessary to memorize the periodic table.** It is more important to understand what information is contained within the periodic table and how to use the

periodic table to find it.

2. **When analyzing the data on trends, many students will notice that there are quite a few exceptions.** The point of this unit is to see the general trends and not become bogged down in the outliers. Assess students based on their knowledge of the trends, not on their ability to memorize exceptions.
3. Emphasize the fact that **Roman numerals convey the charge on a transition metal**, not the number of the atom in a molecule.
4. **Students will need lots of practice problems for nomenclature.** It can be very frustrating, so be creative when giving formative assessments or activities. See additional references for online tools to make quizzing students more interesting.

Starting the Session: Checking Prior Thinking

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

1. What does it mean to be “periodic”? List two examples of phenomena that occur periodically.
2. What is an element?
3. What are some properties of metals and non-metals?

Before Watching the Video

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

1. What is the trend in reactivity for the alkali metals?
2. List two factors that led to the development of the current periodic table.
3. How are new elements made and named?

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 are some of the advantages to the common form of the periodic table?
2. Explain what the following statement means: The periodic table is a living document.
3. What is the valence shell? Why is it important?
4. What is needed for magnesium to react with water? Why isn't that required for calcium?

5. What is a homolog? Why is it necessary to use a homolog in the study of super heavy elements?

Group Learning Activities

Make Your Own Periodic Table Project

Objective

Students will gain the experience of organizing related objects of their choosing. This activity emphasizes that the arrangement of similar objects can take many forms, but some are more useful than others. Many students are mystified by the periodic table, and this activity helps students to realize that it is just one way to organize elements.

List of Materials

- Student handouts outlining the project
- Computers
- Blank paper
- Markers
- Rulers

Setup

Make a handout detailing the instructions for the project as outlined in the procedure below.

Procedure

1. Divide the class into small groups and have each group select a topic to research. Make sure to obtain the teacher's permission before continuing. Example topics include shoes, fruit, vegetables, flowers, musical instruments, etc.
2. Create your own periodic table on the provided blank paper, using the rulers and markers.
3. Arrange examples of these objects into a periodic table that has 8 groups and 4 periods. The arrangement must demonstrate trends going from top to bottom within the groups and from left to right within the periods.
4. Projects will be graded on creativity, thoroughness, and a clear demonstration of trends.

Discussion

The following questions can help students think about what they have learned by doing the activity. Students may discuss the questions in-class or reflect on them in a journal.

1. What have you learned about organizing objects?
2. When your group was trying to determine its organizational scheme, was there only one correct answer? How did you decide which organizational scheme to use?

3. What patterns are evident in your periodic table? Try to think of another example of your chosen object and determine where it fits in your table.

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Color-Coding the Periodic Table

Objective

Students will analyze the trends in atomic radius, ionization energy, and electronegativity by color-coding the periodic table. Students will then be able to visualize where on the table the higher values tend to be, and where the lower values tend to be.

List of Materials

- Packet including values of atomic radius, ionization energy, and electronegativity for each element and three copies of the periodic table
- Colored pencils
- Calculators

Setup

Create packets including values of atomic radius, ionization energy, and electronegativity for each main group element and three copies of the periodic table. The periodic tables can be abbreviated to only include the eight main group elements, such as the one shown below. Each copy of the periodic table will be color coded for one of the trends.

H							He
Li	Be	B	C	N	O	F	Ne
Na	Mg	Al	Si	P	S	Cl	Ar
K	Ca	Ga	Ge	As	Se	Br	Kr
Rb	Sr	In	Sn	Sb	Te	I	Xe
Cs	Ba	Tl	Pb	Bi	Po	At	Rn

Procedure

1. Develop a color scheme to represent large, medium, and small values for electronegativities. Electronegativities range from 0 to 4.0. Calculators may be needed to determine

UNIT 4 Organizing Atoms and Electrons

reasonable ranges for each trend. An example of a potential range and color scheme could be:

0-1.2 = light green

1.2-2.6 = medium green

2.7 to 4.0 = dark green

2. Color each element on the periodic table according to the coloring scheme. On the first copy of the periodic table, fluorine would be shaded dark green because its electronegativity value is 4.0, and aluminum would be shaded a medium green because its value is 1.6.
3. Repeat by color-coding a new periodic table based on values of atomic radii, and then another for ionization energy. Make sure to include a key for the coloring scheme for each trend.

Discussion

Through this activity students will be able to see regions on the periodic table that tend to have larger or smaller values for each trend. The following questions can guide a discussion and help students reflect on this exercise.

1. What advantages does color-coding these trends provide?
2. What trends do you notice for electronegativity, atomic radius and ionization energy? Are there any similarities?
3. Which family of elements has the smallest atomic radii?
4. Based on the trend in atomic radius, rank the following elements from smallest to largest: Si, K, F, Al.
5. If an atom formed a positive ion by losing an electron, would you expect the atom to become larger or smaller? Why?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

The Average Atomic Mass of Banium (or Candium)

Objective

Students will practice calculating the average atomic mass of an element based on relative abundance. For this activity, a sample of different dried beans or small candy represents an element with multiple isotopes. Students will be given a random sample of a mixture of isotopes of the element and will calculate the average atomic mass of the element. At the end of this lesson,

students should have a strong understanding of what an isotope is and how to calculate an average atomic mass.

List of Materials

- Samples of the element: either dried beans (kidney beans, navy beans, and chickpeas) or candy (Plain M&Ms, Peanut M&Ms, and Skittles)
- Balance
- Calculators
- Student handout

Setup

1. Divide the class into small groups.
2. Explain to students that the different beans or types of candy represent different isotopes of the element *beanium* or *candium*.
3. Provide each group of students with a sample of the element (about 50-100 pieces).

Procedure

1. Count the number of total particles and the number of each isotope.
2. Find an average mass of each isotope by weighing the total sample of the isotope and dividing by the number of particles.
3. Calculate an average atomic mass of the element based on the average mass and relative abundance of each isotope.

Discussion

The following questions can guide a discussion to help students reflect on this exercise.

1. How can atoms of carbon-13 exist if the mass doesn't match the one listed on the periodic table.
2. What is an isotope?
3. How is an average atomic mass determined?

Hazards

Be aware of any students with allergies before allowing students to eat any leftover candy. Do not eat candy if activity was performed on a lab bench.

Disposal

Beans can be saved for use in subsequent years.

Bond with a Classmate

Objective

Students require lots of practice problems to master ionic nomenclature. Bonding with a classmate allows students to move around the room while practicing these key skills.

List of Materials

- 25-35 ion cards. Make sure there is an even distribution of positive and negative charges. For smaller classes, it is helpful to have extra cards on hand, and to allow students to switch their cards. Make sure to include transition metal ions and polyatomic ions in addition to the usual monatomic ions. For example, include Li^+ , Mg^{+2} , Al^{+3} , N^{-3} , O^{-2} , F^- , Zn^{+2} , Fe^{+2} , Fe^{+3} , Cu^+ , Cu^{+2} , SO_4^{-2} , SO_3^{-2} , CH_3COO^- , NO_3^- , NO_2^- , etc.
- Student handout: a chart with the following columns: Cation, Anion, Compound Formula, and Compound Name.

Setup

1. Randomly distribute one ion card to each student.
2. Provide each student with a copy of the handout to record his or her compounds.

Procedure

1. Find a classmate with an ion of the opposite charge.
2. Record the identity of the individual ions and the name and formula of the compound they form.

Discussion

Ask students to share with class and write on the blackboard the name and formula of the compound that they formed. Check to ensure that students named compounds with transition metals appropriately and used parentheses when needed with polyatomic ions.

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Polyatomic Ion Bingo!

Objective

Bingo is used to reinforce the names and identities of the polyatomic ions. Many students have trouble remembering the difference between similar ions, for example nitrate, nitrite, and nitride, and this activity can help students become more familiar with the ions before using them to make compounds.

List of Materials

- 25-35 Bingo cards with the names of ions on them. Alternately, the card can be filled with formulas of ionic compounds, and you can call out names of ions like nitrate and iron(II)
- Bingo chips (optional)

Teacher Setup

1. Randomly distribute one Bingo card to each student.
2. Make sure to establish rules such as how many rounds each student can win, whether you will continue a round until there are two, three, or four winners, and what the prize will be for winning.
3. Call the formula of the ion. For the first few rounds, it is useful to write the formula on the board and allow students to use a chart with a complete list of the ions. Later rounds should be much faster.

Procedure

1. Use the chips to cover your card as the teacher calls the formula of the ion.
2. Yell out “Bingo” when you have five in a row.

Discussion

Teachers can use the following for a discussion or students can reflect on these

1. What is the difference between nitrate, nitrite, and nitride?
2. What do you think the suffixes –ate and –ite mean?
3. The ion arsenide was not used in the Bingo game. What do you think its formula would be? What about arsenate and arsenite?

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

In-Class Chemical Demonstrations

Reactivity of Alkali Metals

Objective

This demonstration shows that as a group of metals descends on the periodic table, the reactivity increases.

WARNING: Sodium and potassium are highly reactive. Safety precautions need to be taken. Consult the MSDS or SDS for each chemical.

List of materials

- Large beaker (600 or 1000mL)
- Small knife or spatula
- Watch glass
- Small, pre-cut pieces of sodium and potassium are available for purchase as part of a periodic reactivity kit from some chemical suppliers, such as Flinn Scientific. These will allow you to do similar demonstrations as in the video, but on a smaller scale. The pre-cut pieces can be further divided into two or three pieces to obtain a visible but controlled reaction.
- Phenolphthalein to observe the formation of a base in water. (*optional*)

Procedure

1. Fill a large beaker (600 or 1000mL) two-thirds full of water. Add phenolphthalein if desired.
2. The sodium and potassium pieces are stored in oil. Remove one piece of sodium from the container and place on a watch glass. Cut it in half with a small knife or spatula.
3. Drop the small piece of sodium in the water and observe the reaction. The water will turn pink if phenolphthalein was used.
4. Place the second piece of sodium in the water and again observe the reaction.
5. Repeat the process with potassium.
6. Return the sodium and potassium to a secure storage area.

Discussion

The following questions can help guide students to think about what they learned by observing this demonstration.

1. What is the trend in reactivity for the alkali metals?
2. Why is potassium more reactive than sodium?

3. Describe what happens at the atomic level in the reaction for both sodium and potassium. What are the products?

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS or SDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Sodium and potassium are both very dangerous. The reaction should be done behind a blast shield and under a hood. Everyone in the room should be wearing safety goggles.

Disposal

Keep the remaining sodium and potassium in a secure location. The reacted water will be slightly basic. Check the pH and neutralize if necessary before disposing of it down the drain. Check local regulations for disposal criteria.

Going Deeper (In-Class Discussion or Reflection)

Teachers can use the following for a discussion or students can reflect on these questions in their journals.

1. What characteristics make the periodic table so useful for scientists and chemistry students alike?
2. Why are electron configurations of atoms important?
3. Why do atoms form ions and why do ions form compounds?
4. If you were to make a new element, what would you name it?
5. List 5 ways that chemistry students or scientists use the periodic table.
6. Would you rather win the Nobel Prize in chemistry or have an element named after you?

Before the Next Unit

Learners should read the Unit 4 text if they haven't already done so. They may wish to read one or more of the 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 5 video and/or read the Unit 5 text as an assignment before the next session.

References and Additional Resources

Chemistry Interactive Review Activities. "Periodic Behavior and Ionic Bonding." ScienceGeek.net. Accessed July 23, 2013. <http://www.sciencegeek.net/Chemistry/taters/directory.shtml>

"Chemical Compounds Practice Quiz." Mr. Carman's Blog. Accessed July 23, 2013. <http://www.kentschools.net/ccarman/cp-chemistry/practice-quizzes/compound-naming/>

"Create Your Own Sporcle Games!" Sporcle.com Accessed July 23, 2013. <http://www.sporcle.com/create/>

Kean, Sam. *The Disappearing Spoon*. New York: Little Brown, 2010.

Kulis, Mark. *The Alien Periodic Table Challenge*. Accessed March 7, 2013. http://www.nclark.net/alienperiodictable___kulis.pdf

Logerwell, Mollianne G. and Donna R. Sterling. "Fun with Ionic Compounds." *The Science Teacher*, December 2007: 27-33.

Radioactive Wolves. Dir. Klaus Feichtenberger. By Klaus Feichtenberger. Perf. Harry Smith. PBS, 2011. DVD.

For Professional Development

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

Further Reading & Reflection Questions

Salame, Issa., Samema Sarowar, Sazea Begum, and David Krauss. "Students' Alternative Conceptions about Atomic Properties and the Periodic Table." *The Chemical Educator* 16 (2011): 190-194. Accessed July 22, 2013. <http://faculty.bmcc.cuny.edu/faculty/upload/Article-atomic-properties.pdf>

1. Do you find the results of this study surprising? Why or why not?
2. What measures could you take to help students gain a conceptual understanding of periodic trends?
3. Do the results of this study change the way you think about assessing knowledge and understanding of periodic trends?

Talanquer, Vicente. "Recreating a Periodic Table: A Tool for Developing Pedagogical Content Knowledge." *The Chemical Educator*. 10(2005): 95-99. Accessed July 23, 2013. <http://www.cbc.arizona.edu/tpp/TCEPeriodic.pdf>

NOTE: Before reading the "Discussion and Reflection" section of this paper, think about, brainstorm, and write down ways that you would approach this activity. What strategies do you think you'd use to determine how to organize these elements?

1. In your brainstorming did you encounter any of the challenges encountered by the teachers in the study?
2. What does this activity offer that a simple matching of "alien" elements to the current periodic table does not?
3. What does this study reveal about the difficulty that high school students may have in understanding the periodic table?
4. How would you adapt this activity for high school classroom use?

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 in your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?

UNIT 5

Making Molecules

Lewis Structures and Molecular Geometries

Unit Overview

This unit covers the basics of bonding and takes us from a two-dimensional model of atoms and molecules into three dimensions. It begins by examining valence electrons and why atoms lose, gain, or share electrons to fill their outer shell. Understanding how individual atoms combine and the three-dimensional shapes of molecules enables students to predict many physical and chemical properties of substances.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Predict the Lewis dot structure for covalent or ionic molecules.
2. Identify the geometry of molecules.
3. Describe a compound's physical properties, such as boiling point and solubility, based on intermolecular forces.

Key Concepts and People

1. **Valence Electrons:** Atoms act to obtain a full outer shell of electrons, which confers stability.
2. **Lewis Dot Structures:** The number of valence electrons controls the number of electrons each atom tends to lose, gain, or share in a chemical reaction. Lewis dot structures are a simple, useful method for visualizing the number of valence electrons.
3. **Ionic Bonding:** Ionic bonds occur when atoms have very different electronegativities, and one atom gives up electrons to the other to form ions. Oppositely charged ions attract each other and pair up due to Coulombic forces.
4. **Covalent Bonds and the Octet Rule:** Covalent bonds form when atoms share electrons. The octet rule states that elements will work to get a total of eight valence electrons. Elements can share electrons through covalent bonds to achieve a complete octet.
5. **Polarity and Basic Lewis Structures:** In covalent bonds, the more electronegative

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atom pulls the electron clouds closer than the less electronegative atom, which leads to polar covalent bonds. In a polar covalent bond, the more electronegative atom will have a slightly negative charge, while the other atom will have a slightly positive charge.

6. **Advanced Lewis Structures:** Sometimes, when molecules form where the normal patterns of bonding are not followed, or if the molecules are actually charged molecular ions, the atoms themselves can have a formal charge on them.
7. **VSEPR and Molecular Geometries:** Valence shell electron pair repulsion (VSEPR) theory predicts the 3-D shape of molecules based upon the extent of electron-pair repulsion.
8. **Hybrid Orbitals:** When an electron is shared between two different atoms in a covalent bond, the two electron clouds that correspond to each of the atomic orbitals merge to become a larger cloud. This combination of atomic orbitals is called a “molecular orbital.”
9. **Intermolecular Forces:** There are forces of attraction that exist between molecules themselves. These intermolecular forces, while much weaker than ionic and covalent bonds, have important effects on the way molecules, once formed, interact with one another. Molecular forces include hydrogen bonding, dipole-dipole forces, and London dispersion forces.
10. **Physical Properties of Molecules:** Molecular structure determines the types of intermolecular forces that hold molecules together. The stronger these forces, the more energy it takes to move the molecules apart from each other. This affects physical properties of molecules such as melting point, boiling point, viscosity, and solubility.

Video

Molecules can form when atoms bond together by sharing electrons and can be represented by a useful shorthand called Lewis structures. These visual representations provide information to predict the three-dimensional shapes of molecules using valence shell electron pair repulsion (“VSEPR”) theory. Understanding how atoms bond within molecules provides insight into cell replication. Building on this knowledge, the shapes of molecules reveal the effectiveness of important antibiotics such as penicillin, and scientists can manipulate shapes of molecules to help design new cancer-treating drugs.

VIDEO CONTENT

Host Science Explanation

“Lewis Structures and Radicals”

Making molecules of a particular shape and function can have important implications for fighting

disease. Dr. Beth Taylor, Chemistry Instructor at the Massachusetts Institute of Technology, explains Lewis structures using water and a hydroxyl radical as examples. Radicals have a lone, unpaired electron and thus are highly reactive. Hydroxyl radicals can damage DNA, but are also part of the mechanism by which white blood cells destroy bacteria.

Current Chemistry Research

“Free Radicals in Cell Replication”

By understanding how radicals work, scientists can come up with solutions to shut down the ones that harm our health. Christina Zimanyi, a graduate student in the laboratory of Dr. Catherine Drennan at the Massachusetts Institute of Technology, studies free radicals. She uses X-ray crystallography to image an enzyme called ribonucleotide reductase (RNR). RNR has an oxygen radical that helps start the reactions that make DNA, which is essential for cell division. By understanding the pathway of this radical, researchers could investigate ways to inhibit its function in cancer cells, thus inhibiting the cancer cell’s ability to divide.

Laboratory Demonstration:

“The Geometry of Molecules”

Electron pairs repel each other, and thus will automatically arrange themselves in a molecule as far away from each other as possible. Valence Shell Electron Pair Repulsion (VSEPR) Theory allows us to predict the geometries of molecules based on this repulsion. Harvard University Lecture Demonstrator Daniel Rosenberg uses balloons to represent pairs of electrons to visualize how they mutually repel. He uses the arrangement of balloons to represent the major shapes in VSEPR: octahedral, trigonal bipyramidal, tetrahedral, trigonal planar, and linear.

History of Science

“Penicillin and VSEPR”

Alexander Fleming discovered the antibiotic penicillin in 1928. It is the geometry of penicillin that makes it an effective antibiotic. The molecule contains a four-membered ring called a “beta-lactam.” The geometry of this ring causes the penicillin molecules to form covalent bonds with an enzyme that is responsible for building the cell walls of bacteria cells. This reaction kills the activity of the enzyme and causes the bacterial cells to perish.

Real World Application

“Designing New Cancer Drugs”

Molecular geometry is an important factor in the development of drugs to fight diseases such as cancer. Dr. Alexander Taylor and the other scientists at Constellation Pharmaceuticals wish to specifically target essential proteins in cancer cells. They adjust the geometry of molecules to specifically fit into the binding pockets of these proteins. Polarity, the distribution of electrons in a molecule, is also a consideration in drug development because polar molecules will bind more tightly within a polar binding pocket.

Unit Text

Content Overview

Unit 5 expands on the discussion of electron configuration in Unit 4 by focusing on the importance of valence electrons and their role in bonding. The text explores using Lewis dot symbolism to represent valence electrons on single atoms and bonding in molecules, and includes many examples of Lewis dot structures for molecules involving multiple bonds, expanded octets, and resonance structures. Valence Shell Electron Pair Repulsion Theory, or VSEPR, allows students to predict the shape of molecules in three dimensions, which relates directly to the geometry of hybrid or molecular orbitals. The arrangement of orbitals also has an impact on how well molecules stick to each other. These intermolecular forces control physical properties, such as boiling point and solubility.

Sidebar Content

1. **Linus Pauling:** Linus Pauling made many important contributions to chemistry research in the 20th century, including the concept of hybrid orbitals, which is critical to understanding chemical bonding.
2. **DNA and Mg^{2+} :** DNA is a polymer that contains negatively charged phosphate groups that should repel each other. Magnesium cations help to neutralize the environment in the nucleus of a cell and allow the DNA strand to remain intact.
3. **Resonance:** When it is possible to draw more than one acceptable Lewis structure for a given molecule, where all the connections are the same in both structures but the *location of the electrons* is different, we call these structures “resonance structures.”
4. **Radicals:** Radicals are molecules or atoms that have an unpaired electron.
5. **Expanded Octet Geometries:** Some compounds can have more than 8 electrons in their valence shell, like PF_5 or SF_6 . This creates different kinds of geometries such as trigonal bipyramidal and octahedral.
6. **Oil Spills and Tar Balls:** Crude oil spills contain a mixture of different sized hydrocarbons. Hydrocarbons only have London forces between them, so the rate at which they evaporate is dependent on the size of the hydrocarbon. When an oil spill happens, the smaller hydrocarbons evaporate first, but the larger ones can form together creating big tar balls, which can persist for a long time.

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. Key figures mentioned in this unit are listed on the timeline.

During the Session

Before Facilitating this Unit

Molecular geometry can be a difficult concept for students to grasp. The goal of this unit is to give students a way to visualize the shapes of molecules. The course guide outlines several options for students to see and touch models, depending on the technology and equipment available.

Tips and Suggestions

1. **A symmetrical 2-D structure does not mean there is a symmetrical 3-D structure.** For example, the Lewis dot structure for water can be drawn linearly, but water actually has a bent geometry.
2. **Every chemistry student should understand why bent geometries correspond to polar molecules, even though the Lewis dot structure appears symmetrical.** Bent geometries are especially important for students to understand hydrogen bonding.
3. **Advanced level classes should discuss examples of expanded octets and more complicated geometries, such as square planar and T-shaped.** High school classes do not necessarily need to delve into as much detail, but should understand the differences among tetrahedral, trigonal pyramidal, and bent.
4. **Many students confuse intermolecular forces and bonding.** Emphasize the fact that breaking intermolecular forces does not mean breaking covalent bonds. Boiling water creates steam, not oxygen and hydrogen gas.

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. List five objects whose shape determines its function.
2. What would happen if you tried to push two electrons together? Explain.
3. Why don't oil and water mix?

Before Watching the Video

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

1. What is an octet? Why do atoms want a full octet?

2. Why would scientists care about the 3-D structures of molecules?
3. What is polarity?

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. Why are radicals so reactive?
2. Why would stopping the activities of the enzyme RNR help to fight cancer?
3. What does Christina Zimanyi use to visualize the shape of RNR?
4. How do the bond angles in penicillin cause it to function as an antibiotic medicine?

Group Learning Activities

Two Molecular Shapes Activities

Objective

The following two activities—**PhET Simulation: Molecular Shapes** and **Building Molecules**—allow students to visualize the 3-D shapes of molecules. PhET, a project at the University of Colorado at Boulder, uses education research to design online STEM interactives. The programs are free to use. **Building Molecules** allows students to physically hold and make molecules, which may help some students who have difficulty visualizing 3-D shapes in 2-D spaces.

PhET Simulation: Molecular Shapes

List of Materials

- **Computer or tablet with internet access** (one per group). If computers or tablets are unavailable, this could be done as a demonstration. The simulations' website is <http://phet.colorado.edu/en/simulation/molecule-shapes>
- **Chart with the following columns:** Chemical Formula, Lewis Dot Structure, 3-D Sketch, # of Electron Densities, # of Lone Pairs, and Molecular Geometry. Some suggested molecules include CO_2 , CH_2O , CH_4 , NH_3 , H_2O , BF_3 , SF_6 , PCl_5 , and XeF_4 . The entire chart will fit on a single, double-sided sheet of paper if the document is in a landscape orientation.

Set Up

Create handout as described in the list of materials.

Procedure

1. Log on to <http://phet.colorado.edu/en/simulation/molecule-shapes>
2. While working in small groups, fill out the chart.

Discussion

After performing this activity, ask students if they noticed any patterns. Emphasize the fact that patterns exist in geometries, so once you know the geometry of CH_4 , you know the geometry of CF_4 , CBr_4 , Cl_4 , NH_4^+ , etc. For each of the molecules listed, have students brainstorm other molecules that would have the exact same geometry. Discuss the list as a class. Have students explain why a given molecule has a particular shape and why it's important to know the shape of a molecule.

Hazards

There is no increased risk of harm to do this activity.

Disposal

There are no special disposal considerations.

Building Molecules**List of Materials**

- **A molecular modeling kit or toothpicks and gum drops.**
- **Chart with the following columns:** Chemical Formula, Lewis Dot Structure, 3-D Sketch, # of Electron Densities, # of Lone Pairs, and Molecular Geometry. Some suggested molecules include CO_2 , CH_2O , CH_4 , NH_3 , H_2O , BF_3 , SF_6 , PCl_5 , and XeF_4 . The entire chart will fit on a single sheet of paper double-sided if the document is in a landscape format.

Set Up

Create handout described in the list of materials.

Procedure

While working in small groups, use the materials to make the 3-D molecules listed on the sheet and fill out the chart.

Discussion

After performing this activity ask students if they noticed any patterns. Emphasize the fact that patterns exist in geometries, so once you know the geometry of CH_4 , you know the geometry of CF_4 , CBr_4 , Cl_4 , NH_4^+ , etc.

Hazards

There is no increased risk of harm to do this activity.

Disposal

Throw all disposable materials into the garbage and wipe down tables or desks following the activity.

IMF and Surface Tension Activity

Objective

In this activity, students count the number of drops of three different types of liquids (water, methanol, and acetone) that can fit on a penny without spilling over the sides. The number of drops that fit on the surface of the penny relates to the intermolecular forces holding the molecules together. Water and methanol both exhibit hydrogen bonding, whereas acetone has only dipole-dipole interactions. Water's hydrogen bonds are stronger than methanol's intermolecular forces, so it will have the largest number of drops. Stronger IMFs correspond to more drops being able to fit on the surface of the penny.

List of Materials

- One penny per group
- Small dropper bottle of water (one per group, or groups can share)
- Small dropper bottle of methanol (one per group, or groups can share)
- Small dropper bottle of acetone (one per group, or groups can share)
- Rinse bottle of water
- An organic waste container (fume hood).

Procedure

1. One drop at a time, count the number of drops of water that will fit onto the surface of a penny before the water pours off. Record results. Repeat for a total of three trials for consistency.
2. Rinse the penny and work area with water and dry with a paper towel.
3. Repeat for methanol and acetone.

Discussion

The following questions can help students reflect on this exercise:

1. The penny held the greatest number of drops of which liquid?
2. The penny held the least number of drops of which liquid?
3. Which liquid has the strongest intermolecular forces? Explain your reasoning.
4. What intermolecular forces are present in each of the three liquids? It may help to draw

a diagram to explain your answer.

5. Intermolecular forces in molecules affect their surface tension. Surface tension is defined by Merriam-Webster as: “the attractive force exerted upon the surface molecules of a liquid by the molecules beneath that tends to draw the surface molecules into the bulk of the liquid and makes the liquid assume the shape having the least surface area.” How does this relate to how much of each liquid fits on a penny?

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 safety glasses/ goggles and gloves while performing this experiment. The room should be well ventilated.

Disposal

Use caution disposing of acetone and methanol. They should be placed in the appropriate hazardous waste container.

In-Class Chemical Demonstrations

Balloon Demonstration of Molecular Geometries

Objective

Electrons automatically rearrange themselves to spread out as much as possible. We can easily simulate this using balloons to represent areas of electron density around a central atom.

List of materials

- 6 balloons

Procedure

1. Inflate 6 balloons.
2. Tie them together in the middle so that they make an octahedral shape.
3. Pop one balloon and shake the balloons until they resemble a trigonal bipyramidal geometry.
4. Continue until all of the balloons are popped.

Discussion

The following questions can help guide students to think about what they learned by observing this demonstration.

1. What do the balloons represent?

2. Why do any given number of balloons always take the same shape?

Bending Water

Objective

Water is a polar molecule, and can be attracted by static electricity. Students sometimes have a difficult time remembering that water is polar, and this demonstration will give them visual evidence that a water molecule acts like a tiny magnet.

List of materials

- A balloon, nylon comb, or any item that can become charged by static electricity
- A water faucet or buret filled with water

Procedure

1. Inflate the balloon, if necessary.
2. Adjust the faucet or buret to produce a small, even stream of water.
3. Charge the balloon, comb, or other item by rubbing it against your clothing or through your hair.
4. Hold the charged item about an inch away from the stream of water.
5. Repeat the process while varying the size of the stream of water, the amount of charge on the item, and the distance between the charged item and the stream of water.

Discussion

The following questions can help guide students to think about what they learned by observing this demonstration.

1. Why does the water bend toward the charged item?
2. What was the effect of having a larger or smaller stream of water?
3. What was the effect of having more charge or holding the item closer to the stream of water?
4. What do you think would happen if we had a stream of oil or hexanes, a non-polar molecule?

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 their journals.

1. Which geometries are always polar? Why?
2. Which is stronger a covalent bond or an IMF? Why?
3. Why do chemists focus on valence electrons?
4. Draw a diagram on the molecular level of water boiling.

Before the Next Unit

Learners should read the Unit 5 text if they haven't already done so. They may wish to read one or more of the 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 6 video and/or read the Unit 6 text as an assignment before the next session.

References and Additional Resources

Gillespie, Ronald James and Hargittai, Istvan. *The VSEPR Model of Molecular Geometry*. Boston, MA: Dover, 1991, 2012.

PhET Interactive Simulations. "Molecule Shapes." University of Colorado at Boulder. Accessed February 3, 2013. <http://phet.colorado.edu/en/simulation/molecule-shapes>

Schmidt, Hans-Jurgen, et al. "Students' understanding of boiling points and intermolecular forces." *Chemistry Education Research and Practices*. 10 (2009): 265-272. Accessed June 9, 2013. <http://pubs.rsc.org/en/content/articlelanding/2009/rp/b920829c>

Tan, Kim-Chwee Daniel and David F.Treagust, "Evaluating Students' Understanding of Chemical Bonding," *School Science Review* 81 (294), September 1999. Accessed October 24, 2014. <https://www.ase.org.uk/journals/school-science-review/1999/09/294/1186/SSR294Sept1999p75.pdf>.

For Professional Development

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

Further Reading & Reflection Questions

Nahum, Tami Levy, Rachel Mamlok-Naaman, and Avi Hofsein. "A New 'Bottom-Up' Framework for Teaching Chemical Bonding." *Journal of Chemical Education* 85:12 (2008): 1680-1685. Accessed July 23, 2013. http://stwww.weizmann.ac.il/menu/staff/Rachel_Mamlok/Bonding%20%20A%20new%20framework.pdf

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Nahum et al, “Developing a New Teaching Approach for the Chemical Bonding Concept Aligned With Current Scientific and Pedagogical Knowledge,” *Science Education* (2007). Accessed October 24, 2014. http://deepblue.lib.umich.edu/bitstream/handle/2027.42/56057/20201_ftp.pdf

1. Do you agree with any of the difficulties that the authors present about the traditional approach to teaching chemical bonding? If so, which ones and why? If not, why not?
2. What advantages do you see in using the “bottom-up” framework presented in this paper to teach chemistry? Do you see any disadvantages? What may be challenging in using this bottom-up approach?
3. Does this paper influence how you will approach teaching chemical bonding in the future? Why or why not? How, if at all, would you incorporate the bottom-up framework?

Cooper, Melanie M., Nathaniel Grove, Sonia M. Underwood, and Michael W. Klymkowsky. “Lost in Lewis Structures: An Investigation of Student Difficulties in Developing Representational Competence.” *Journal of Chemical Education*. 87:8 (2010): Accessed July 25, 2013. <http://educationgroup.mit.edu/HHMIEducationGroup/wp-content/uploads/2011/04/13-Cooper-et-al-2010.pdf>

1. Were you surprised by the results of this study? Why or why not?
2. Do you agree with the authors’ suggestions on how to better teach Lewis structures and their significance? Why or why not?
3. How would you use the constructivist model to teach Lewis structures? Does this paper influence how you will teach or assess these topics?

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 in your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?

UNIT 6

Quantifying Chemical Reactions

Unit Overview

Chemists use numbers and symbols in the form of chemical formulas to represent chemical reactions. Atoms, elements, and molecules combine together in whole number ratios relative to each other. These whole number ratios are determined using physically measurable quantities, and these measurable quantities contain a huge number of atoms. This unit focuses on quantifying chemical reactions by exploring moles, atomic and molar masses, balanced chemical equations, and reaction yields. Being able to quantify chemical reactions is important in many practical applications of chemistry.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Differentiate between empirical and molecular formulas and understand how they are determined experimentally.
2. Balance equations for basic chemical reactions.
3. Calculate chemical quantities from measurable quantities using mole ratios.
4. Understand how limiting reagents and incomplete reactions influences reaction yields.

Key Concepts and People

1. **Empirical Formula:** The simplest whole number ratio of elements in a compound (the empirical formula) can be determined by analyzing the products of combustion of the compound.
2. **Molecular Formula:** For many compounds, the empirical formula alone is not enough to determine the identity or structure of the compound. The molecular formula contains the actual number of each type of atom in a compound.
3. **Avogadro's Number and Atomic Mass:** The mole is the basic unit in chemical calculations. The mole represents physically measurable quantities of atoms and molecules, and thus is a huge number. The officially defined quantity of a mole is the number of carbon-12 atoms in exactly 12 grams of carbon-12.
4. **Molecular Masses and Mass Percentages:** The molecular mass of a given molecule is the sum of the atomic masses of its components. Mass percentages reveal the

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fraction of mass that each type of atom contributes to the molecular mass of the overall molecule.

5. **Chemical Reactions and Equations:** Chemical equations may be written either in formulas or in chemical structures. Equations must be balanced to account for every atom on each side. A balanced equation is the basis of all reaction analysis.
6. **Stoichiometry:** The amounts of reactants and products used in a reaction can be compared in mole ratios using the coefficients of the balanced equation, and converted to measurable quantities as needed.
7. **Limiting Reagents:** Most reactions are not performed with the exact amounts of reactants necessary to consume each other completely; one or more usually will be left over. The maximum possible yield of a reaction is based on the limiting reagent.
8. **Percent Yield:** In a given chemical reaction, the percent yield reveals how much of the reactants were converted to products compared to what would be expected if 100% of the reactants were converted to 100% of the products.

Video

Stoichiometry gives us the quantitative tools to figure out the relative amounts of reactants and products in chemical reactions. Balancing the number of atoms on each side of the equation, calculating the amount of each reactant, and figuring out which reactant will run out first are all fundamental principles when designing any chemical reaction. These principles are applied when splitting water into hydrogen and oxygen for energy, manufacturing sodium iodide for radiation detectors, and producing common chemicals from renewable resources.

VIDEO CONTENT

Host Introduction

“Quantifying Chemical Reactions”

The Nocera Laboratory at Harvard University researches how to create cheap, clean, and renewable energy. Host David Song, a graduate student in the Nocera Laboratory, introduces the idea that a gallon of water could provide enough energy to power a house for an entire day.

Current Chemistry Research

“A Balanced Equation for Fuel”

The Nocera Laboratory at Harvard University has developed the “artificial leaf,” a new technology that mimics photosynthesis and generates hydrogen gas from water. This system is designed to generate hydrogen for energy use by using sunlight to “split” water. The amount of hydrogen produced in the process can be calculated based on the balanced equation for the chemical reaction.

Laboratory Demonstration**“Making Water 2:1”**

The combination of hydrogen gas and oxygen gas to form water requires twice as much hydrogen as oxygen, but what happens if you mix equal parts of the two gases? Dr. Wolfgang Rueckner, Manager of Lecture Science Demonstration Services at Harvard University, shows what happens.

Host Science Explanation**“Calculating Molar Mass”**

In addition to a balanced equation, chemists need a way to deal with extremely large quantities of atoms and molecules in order to calculate chemical quantities. David Song introduces the concept of the mole and uses the reaction that forms sodium iodide, a compound used in many radiation detectors, to show how to calculate molar mass.

Host Introduction to Science Explanation**“Limiting Reagent”**

Usually, one reagent in a reaction is consumed completely before the remaining reagents are exhausted. This reagent is referred to as the limiting reagent. Dave Song demonstrates this concept by baking cookies in which the chocolate chips serve as the limiting reagent.

Laboratory Demonstration**“Finding the Limiting Reagent”**

Methanol vapor will ignite on a wire of catalytic metal when allowed to react in the presence of oxygen. Harvard University Lecture Demonstrator Daniel Rosenberg shows that when the methanol vapor is reacted in an open container, the reaction proceeds continuously. When the container is closed, however, the reaction ultimately stops due to a lack of oxygen before all the methanol is consumed.

Real-World Application**“Sustainable Chemistry”**

Myriant Corporation, a global renewable chemicals company located in the Boston area, uses renewable resources to make chemical products that would ordinarily be made from the non-renewable resource, petroleum. Researchers at Myriant show that in order to produce environmentally friendly chemical products, it is important to understand how to balance chemical equations, how to calculate molecular masses, and how to work to your advantage with limiting reagents.

Unit Text

Content Overview

This unit introduces the basics of stoichiometry: how chemists are able to determine molecular quantities of reactants and products in chemical reactions. First, the text explains that atoms combine in whole number ratios to form molecules, which are represented by empirical and molecular formulas. In order to deal with the large quantities of atoms involved in a given

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chemical reaction, chemists use a unit of measurement called the “mole” to represent amounts of a chemical substance. The text then explains Avogadro’s number, atomic masses, and molar masses. This information leads to a discussion of chemical reactions and how they are represented by balanced chemical equations, followed by an explanation of stoichiometry. The unit ends by discussing limiting reagents and percent yields of reactions.

Sidebar Content

1. **Stinking Up the House:** Oliver Sacks recounts a time when he produced a large amount of hydrogen sulfide in his house.
2. **Berzelius and the Birth of Formulas:** Jons Jacob Berzelius is the man behind our current system of writing chemical formulas.
3. **Isomers and Lewis Structures:** Butanol and diethyl ether are isomers, which means they have the same molecular formula, but have different structures, and thus very different properties. Lewis structures help represent the differences in the isomers’ structures.
4. **How Many Items Are in a Mole?** This sidebar shows examples of how big a number one mole really is.
5. **Avogadro’s Law Example:** Avogadro’s 1811 theory predicted that a given volume of any gas has the same number of particles.
6. **Converting Moles to Atomic Mass:** There are two major ways to use atomic masses to determine how much of an element to use in a given chemical reaction.
7. **Equilibrium Reactions:** Even a small amount of carbon dioxide dissolved in water (even a huge body of water!) can make the solution acidic by the formation of carbonic acid.

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

Equilibrium, particularly the equilibrium of hemoglobin and oxyhemoglobin, heavily influences the physiology of exercise. Although many biochemical reactions have very simple stoichiometry, the actual amounts of the various reactants and products can heavily influence the body’s ability to perform strenuous activity.

Control a Haber-Bosch Ammonia Plant Interactive

The Haber process is based on an equilibrium process. Reaction yields are dependent not only upon the reaction conditions overall, but also fundamentally upon the stoichiometry of the reaction. For example, for the production of two molecules of ammonia, one molecule of nitrogen and three molecules of hydrogen are required. Please note that a lesson plan and student worksheet are available online with this interactive.

During the Session

Before Facilitating this Unit

Few topics are as critical to chemistry as the mole concept, but few topics are more intimidating to students. Often, so much emphasis is placed on the sheer size of Avogadro's number that students don't realize how rarely they will actually need to use it directly. They invoke it far too frequently in their calculations. The focus in this unit is on chemical proportions and on expressing chemical quantities in terms that we can easily describe.

Tips and Suggestions

1. Although Avogadro's number is intimidatingly large, **it is actually not needed as frequently in calculations as students often think.** Most calculations in chemistry use mole proportions with respect to measurable quantities such as masses. While the scale of the number is important, and the actual value is used under some circumstances, the focus of sample calculations should be slanted more toward mole ratios in compounds and equations, and the relationship of those quantities to masses, volumes, and other readily measured parameters. Encourage students to first set up multi-step calculations in terms of the kinds of quantities they are attempting to convert, and then to include the numerical values once they have developed a strategy.
2. **The empirical formula for a compound is not sufficient to determine the identity of a molecular compound,** but even when the molecular formula is known, many molecular formulas can be drawn with different isomeric structures. A discussion of how much information can be obtained from a given type of experiment is useful not only for introduction of the different types of formulas but also for introduction to the link between experiment design, execution, and conclusions.
3. **The determination of a limiting reagent may be done in various ways,** but the simplest way is to calculate how much of the desired product can be formed in a reaction using each of the known quantities of reactant. Even if one reagent is obviously limiting (e.g., a reaction with oxygen is performed in open air), it is useful to emphasize the limiting reagent so that students acquire the habit of looking for it before beginning calculations.

UNIT 6 Quantifying Chemical Reactions

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. How many molecules of water do you think are in a given glass of water?
2. What are some ways chemists measure how much of a given compound they have?
3. What does a chemical equation represent?
4. What does this chemical equation show: $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$? Draw a visual representation of this.
5. Is it important for chemical equations to be balanced? Why or why not?
6. Which is heavier: one atom of iron or one atom of gold? If you have 10g each of iron and gold, which one will contain more atoms?
7. Suppose you had 50 atoms of iron and 50 molecules of water. Which sample would weigh more? Which sample would contain more atoms total?
8. If you have 20 atoms of hydrogen and 10 atoms of oxygen, how many molecules of water can you make? What if you have 20 atoms of hydrogen and 20 atoms of oxygen? What if you have 40 atoms of hydrogen and 10 atoms of oxygen?

Before Watching the Video

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

1. What is the balanced chemical equation for splitting water?
2. What happens when two parts hydrogen and two parts oxygen combine?
3. What does a mole represent?
4. How do you use the periodic table to determine the molar mass of sodium iodide (NaI)?
5. What is a limiting reagent?
6. In the reaction of methanol vapor on the catalytic wire what reaction is happening? How do you know methanol is the limiting reagent?

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. Why is having a balanced reaction important in chemical calculations?

2. What is the ideal way to minimize waste when performing chemical reactions? If it is not possible to eliminate waste entirely, what factors might influence the strategy to minimize it?
3. Are masses like grams and kilograms always the easiest way to measure chemical quantities? How else might you need to measure amounts of reactants and products in order to perform chemical calculations?
4. Many reactions, especially in industry, are monitored for their progress over time. Is it necessary (or sensible) to do this by measuring the amount of reactant still present? Why or why not?

Group Learning Activities

Building Isomers

Objective

For this activity, students will use models to explore different isomeric forms of chemical formulas. For example, C_2H_6O may be the formula for either ethanol or dimethyl ether, and the models of the two structures may be switched simply by the interchange of the $-H$ and the $-CH_3$.

List of Materials

- Chemical modeling kits
- If chemical modeling kits aren't available, marshmallows and toothpicks can be used to model molecules. For example, large marshmallows can represent carbon atoms, small marshmallows can be hydrogen atoms, and different colored marshmallows or other soft candies can be other atoms. Although marshmallows do not hold bond angles particularly firmly, the models should give students a reference from which to draw the Lewis structures of different isomers, or vice versa.

Set Up

List molecular formulas on the board. Suitable molecular formulas include, but are not limited to: C_2H_6O , C_6H_{14} , $C_4H_{11}N$, $C_6H_{12}O$, and $C_6H_{13}Br$.

Procedure

1. If using molecular modeling kits, connect atoms to build as many different configurations of molecules represented by the molecular formulas on the board, keeping in mind the limit of how many bonds each atom can make.
2. If using marshmallows and toothpicks, gather enough large marshmallows (for carbon atoms), small marshmallows (for hydrogen atoms), and colored marshmallows or other candies (for atoms other than C and H) to account for every atom in one of the

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molecular formulas listed on the board. Using toothpicks to connect the marshmallows, build as many different configurations as possible of the different atoms, keeping in mind the limit of how many bonds each atom can make.

3. From each model, draw a Lewis structure of the compound.

Discussion

The following questions can guide a discussion and help students reflect on this exercise.

1. Which molecular formula gave the most possible isomers? Which one gave the fewest?
2. Did you have to dismantle the models completely in order to build new isomers of the same compound? Explain.
3. Why do you think different isomers have different chemical properties?

Hazards

There are no hazards associated with this exercise; however, if the exercise is performed on a laboratory bench, the candies should not be eaten afterwards.

Disposal

Marshmallows and toothpicks may be disposed in the trash.

Building Reagents and Products with Toys (May also be done as a demo)

Objective

The variations of this exercise all address different principles of chemical proportions: specific numbers of atoms are necessary to form specific compounds, but in some cases the same numbers and types of atoms may be able to combine in different ways. In all cases, simple models (not necessarily molecular models) are built from commercially available construction toys.

List of Materials

- LEGO® Bricks, blocks, TINKERTOYS® or other construction toys; each group should be given enough pieces to construct a few model structures with some extra pieces left over.

Set Up for Variation I: Limiting Reagent Part I

Build simple models/structures from the building materials for the students to copy. For example, you could build a structure with one red brick and two small yellow bricks attached to one side, etc.

Procedure for Variation I: Limiting Reagent Part I

1. Using the pieces available in your set, build as many replicas as possible of the model displayed at the instructor's desk.

2. Once you have built as many of the models as your available pieces will allow, make a note of which piece(s) remain and which were completely consumed in the building of the target models.
3. Record what the limiting reagents were for each model.

Set Up For Variations II and III

As indicated for variation I, build models/structures for the students to copy, but instead of providing individual building pieces to the students to build their models, provide smaller, simple models/structures as starting materials.

Variation II: Limiting Reagents Part 2

1. Students will be shown simple models of “reactants” and “products” at the instructor’s desk, and given several of each model of the different “reactant” constructions.
2. Using only the pieces available in your set, build as many sets of the models at the instructor’s desk as possible. Dismantle models from your set only if you will be able to construct “product” models and not leave behind portions of the original “reactant” models.
3. Once you have built as many of the “product models” as your available “reactant” models will allow, make a note of which original model(s) remain and which were completely consumed. Which were the limiting reagents?

Variation III: Reactants to Products

1. You will be given two or more pre-constructed simple models.
2. Rearrange the pieces from your given models into as many different combinations of new models as possible, with no individual pieces left over.
3. Make a note of each combination you come up with before dismantling the models to generate new combinations.

Discussion

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

1. In Variation I: Limiting Reagent Part I, what represents the limiting reagent? How do you know?
2. The building of models from individual blocks is analogous to the formation of compounds from individual atoms, while the building of models from components of other models is analogous to the formation of compounds from existing compounds. Which is more common for actual chemical reactions? Explain.
3. In variation III, you could have created “products” from the original models you were

given by dismantling all the models completely into individual blocks. How is this similar to decomposition of molecules into their component elements? How is it different?

Hazards

There are no hazards associated with this activity.

Disposal

The materials for this exercise are reusable.

In-Class Chemical Demonstrations

When Oxygen Is Limiting

Objective

In this demonstration, the burn times for a candle under a flask of air and under a flask of pure oxygen are compared and used to estimate the percentage of oxygen in the atmosphere.

Materials Needed

- Two short taper candles, at least one of a diameter and height to be able to fit under an inverted 250mL Erlenmeyer flask
- Two 250mL Erlenmeyer flasks
- Oxygen tank or lecture bottle, with rubber tubing to bubble oxygen into underwater flask
- Basin to fill with water
- Stopper for one of the flasks
- Stopwatch
- Matches or lighter

Procedure

1. Fill the basin with water and submerge one of the two flasks to fill it.
2. Open the regulator on the oxygen tank and bubble oxygen into the submerged flask through the hose.
3. Close the flask with the stopper, and then remove the flask from the basin.
4. Dry the outside of the oxygen-filled flask.
5. Secure the taper candles to stand up (either use a candle tray or soften the wax at the base to stick to the table). Light both candles.
6. Cover one of the candles with the NON-stoppered Erlenmeyer flask (i.e., the one not filled with pure oxygen). Have students measure and record the time from the covering of the candle to the extinguishing of the flame.

7. Once the flame is extinguished and data are recorded, light the candle again, but this time cover the candle with the flask containing pure oxygen. Again, have students measure and record the time required to extinguish the flame under the flask.
8. Point out to the students that the second candle (not put under the flasks) has been burning the entire time.

Discussion

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

1. Why did the candle on the desk stay lit while the candles put under flasks went out?
2. Calculate the ratio between extinguishing times for the candle under the flask filled with air to the candle under the flask filled with pure oxygen. How close does this ratio come to 21% (the approximate percentage of the atmosphere occupied by oxygen)?
3. If the flasks used to cover the candles had been twice as large, what do you think the lengths of time to extinguish the candles would be? Explain.

Hazards

Gas cylinders must be secured appropriately before use. Keep the gas cylinder away from the matches/lighters and lighted candles. Be careful not to burn yourself when lighting candles. Oxygen is highly flammable.

Disposal

Candles are reusable. Wet matches before disposing in the trash.

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. What happens to the mass of material inside an open container if a reaction releases gas-phase products? What about in a closed container? What happens to the pressure inside a closed container if a reaction releases gas-phase products? What about in an open container?
2. A compound containing only C and H is burned for combustion analysis. If the reaction produces equal moles of carbon dioxide and of water, what is the empirical formula of the compound? If the compound's formula weight is 84 g/mol, what is the molecular formula of the compound?
3. You are working in a lab, and need to synthesize 10 g of a compound. The procedure says that the reaction you are using gives only a 70% yield. How much of the compound do you need to plan to make in order to have an actual yield of 10 g?

UNIT 6 Quantifying Chemical Reactions

4. You are doing a multi-step procedure in a lab. Step 1 has a 70% yield, step 2 has a 20% yield, and step 3 has a 95% yield. What will the total percent yield for the process be?
5. Suppose you worked in a lab and were studying the synthesis described above. If you were trying to develop a way to improve the overall yield of the synthesis, which step would you first attempt to improve, and why?
6. Suppose you are setting up a reaction and need to determine which reagent should be limiting. What factors should you consider in choosing the limiting reagent?
7. In many chemical reactions, more than one product is produced. If only one of the products is desired, what are some of the ways that the product might be isolated?

Before the Next Unit

Learners should read the Unit 6 text if they haven't already done so. They may wish to read one or more of the 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 7 video and/or read the Unit 7 text as an assignment before the next session.

References and Additional Resources

Thompson, Stephen. *Chemtrek: Small-Scale Experiments for General Chemistry*. Vol 2. Boston: Allyn and Bacon, 1990.

Cacciatore, Kristen L. and Hannah Sevan. "Teaching Lab Report Writing through Inquiry: A Green Chemistry Stoichiometry Experiment for General Chemistry." *Journal of Chemical Education*. 83: 7 (2006). 1039-1041. Accessed August 9, 2013. <http://pubs.rsc.org/en/content/articlelanding/2007/rp/b6rp90017h/unauth#!divAbstract>

Fach, Martin, Tanja de Boer and Ilka Parchmann. "Results of an Interview Study as Basis for the Development of Stepped Supporting Tools for Stoichiometric Problems." *Chemistry Education Research and Practice*. 8:1 (2007) 13-31. Accessed August 9, 2013. http://www.rsc.org/images/Fach%20paper%20final_tcm18-76278.pdf

For Professional Development

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

Further Reading & Reflection Questions

DeMeo, Stephen. "Revisiting Molar Mass, Atomic Mass, and Mass Number: Organizing, Integrating, and Sequencing Fundamental Chemical Concepts." *Journal of Chemical Education*. 83:4 (2006). 617-621. Accessed August 9, 2013. <http://www.bruderchemistry.com/core/content/files/Revisiting%20Mass%20article%201.pdf>

1. Did this paper convince you that it makes the most sense to teach molar mass first, then atomic mass, then mass number? Why or why not? What do you think are some benefits and challenges to teaching the content in this order?
2. Do you think the questions presented in "The Creation of New Problems" could assess a student's conceptual understanding of molar mass, atomic mass, and mass number? Why or Why not?
3. Did this paper help clarify any confusion you may have had with these terms? If so, how?

Arasasingham, Ramesh D., Mare Taagepera, Frank Potter, Stacy Lonjers. "Using Knowledge Space Theory to Assess Student Understanding of Stoichiometry." *Journal of Chemical Education*. 81: 10 (2004). Accessed August 9, 2013. <http://pubs.rsc.org/en/Content/ArticleLanding/2011/RP/clrp90039k#!divAbstract> (article available via free login)

1. Were you surprised by the results of this study? Why or why not? What was most surprising and why? What was least surprising and why?
2. Have you encountered students that held misconceptions similar to the ones revealed in the study? What are some ways you could address these types of misconceptions?
3. Do you agree with the authors' suggestions for how to improve the teaching of these concepts? How might you apply some of their suggestions to how you teach these materials?

Sanger, Michael J. "Evaluating Students' Conceptual Understanding of Balanced Equations and Stoichiometric Ratios Using a Particulate Drawing." *Journal of Chemical Education*, v82 n1 p131 Jan 2005. Accessed October 24, 2014. <http://pubs.acs.org/doi/abs/10.1021/ed082p131> (available via ACS) and <http://eric.ed.gov/?id=EJ726143>

1. Were you surprised by the results of this study? Why or why not? What was most surprising to you and why?
2. How have you assessed students' knowledge of stoichiometric principles in the past? In teaching stoichiometry, have you encountered students who are able to solve stoichiometric calculations, but who lacked conceptual understanding of stoichiometry? How have you addressed this?
3. Does this study influence how you will teach or assess stoichiometric principles? If so, how? If not, why not?

Professional Development Assignments

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

UNIT 7

The Energy in Chemical Reactions

Thermochemistry and Reaction Energies

Unit Overview

Unit 7 introduces students to thermochemistry, the study of energy in chemical reactions. After completing this unit, students should be able to understand the various types of energy along with basic thermodynamic terms: system, surroundings, heat, and work. The unit also explains the first law of thermodynamics and provides an introduction to enthalpy and calorimetry.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Describe the different kinds of energy: kinetic, potential, chemical, and thermal.
2. Describe how the internal energy of a system can increase or decrease through heat and work.
3. Explain the First Law of Thermodynamics (the conservation of energy) and give examples of one form of energy converting into another.
4. Define the terms endothermic and exothermic.
5. Describe Hess's Law and what it is used for.
6. Explain how bond enthalpies can be used to find the change in enthalpy (ΔH) of a reaction.

Key Concepts and People

1. **Kinetic and Potential Energy:** Objects can possess both kinetic and potential energy. When a substance goes through a phase change, its particles gain or lose kinetic energy.
2. **Energy Changes in Chemistry:** Chemical energy is potential energy that is stored in chemical bonds. This energy can be released by exchanging high energy bonds for low energy bonds. Thermal energy is the total energy of all of the molecules in a substance.
3. **Heat, Work, and Internal Energy:** Heat and work are not energy, but rather describe ways energy enters and leaves a system.

UNIT 7 The Energy in Chemical Reactions

4. **Energy Conversions and Conservation:** While energy frequently changes from one form to another, the total amount of energy in an isolated system does not change. The First Law of Thermodynamics explains that energy cannot be created nor destroyed.
5. **Enthalpy:** The change in enthalpy (ΔH) expresses the amount of heat produced or absorbed by a chemical reaction at constant pressure. Chemical reactions either release heat (exothermic) or absorb heat (endothermic).
6. **Calorimetry:** A calorimeter is a container that can measure the amount of heat released or absorbed by any chemical process.
7. **Hess's Law:** Hess's Law explains that the change in enthalpy of a chemical reaction depends only on the enthalpy of the reactants and the products, and not on anything that happens in between.
8. **Bond Enthalpies:** Breaking chemical bonds requires enthalpy (endothermic), and forming chemical bonds releases enthalpy (exothermic). Determining if a chemical reaction as a whole is endothermic or exothermic depends on whether or not breaking the old chemical bonds requires more enthalpy than the enthalpy released from forming the new chemical bonds. The amount of enthalpy associated with a bond is called "bond enthalpy."
9. **Standard Enthalpies of Formation:** The standard enthalpy of formation refers to the enthalpy change of the formation of one mole of a compound from its constituent elements in their most stable state at 1 atm of pressure and 25°C.

Video

The phrase "chemical reaction" conjures up images of explosions, bubbling gases, flames, and smoke. So many chemical reactions have visible results because energy is being transferred from one form to another—the realm of thermodynamics. Thermodynamics provides rules for predicting the progress of a reaction and for harnessing the energy released. It is key to solving pressing engineering problems, such as making the next generation of cleaner, and more efficient, automobile engines.

VIDEO CONTENT

Host Introduction "Exothermic Reactions"

Dr. Nicole Labbe, a chemical engineer at the Argonne National Laboratory at the University of Chicago, introduces the first law of thermodynamics and enthalpy by describing what happens when a log burns.

Laboratory Demonstration

“Flame Tornado”

Daniel Rosenberg, Lecture Demonstrator at Harvard University, creates a flame tornado to show a visually stunning example of an exothermic reaction.

Laboratory Demonstration

“Heat Absorber”

To show an endothermic reaction, Daniel Rosenberg reacts barium hydroxide and ammonium chloride. To prove the reaction is endothermic, he places water between a block of wood and the beaker in which the reaction is taking place. Because the reaction is endothermic, it takes enthalpy from the surroundings, which means it takes some enthalpy from the water between the beaker and the block. This causes the water to freeze.

Host Science Explanation

“Measuring Heat”

Dr. Nicole Labbe demonstrates how a coffee cup calorimeter is used to measure the heat released when sodium hydroxide dissolves in water.

Real World Application

“Mississippi Biofuels”

By using bomb calorimetry, Dr. Gretchen Sassenrath is leading a study for the U.S. Department of Agriculture to determine which crops in Mississippi could best serve as a potential alternative fuel source.

Host Science Explanation

“Bond Enthalpy”

Dr. Nicole Labbe explains how to calculate values of enthalpies of a given reaction by looking up the bond enthalpies of the products and reactants of that reaction. Chemists have been doing experiments for hundreds of years to determine bond enthalpies for various chemicals, which are now compiled in tables. Bond enthalpies can help predict the mechanism of a given reaction.

History of Chemistry

In thermodynamics, work is any transfer of energy in a system that is not a heat transfer. Work can create heat, and heat can create work. During the days of steam power, scientists wondered if an engine could perform a heat-to-work conversion with 100% efficiency. Using a theoretical piston engine, Nicolas Leonard Sadi Carnot (1798-1832), a French military engineer and physicist, was the first to realize that it is impossible. Carnot's discovery led to the second law of thermodynamics.

Current Chemistry Research

“Fighting the 2nd Law”

Kevin Cedrone, a graduate student at MIT's Sloan Automotive Lab, studies how to get as much work out of an engine as possible. He strives to get engines to run as cleanly and efficiently as possible.

Unit Text

Content Overview

Energy is the ability to do work, and it comes in many forms. The unit begins with a discussion of kinetic energy, potential energy, thermal energy, latent heats of fusion/vaporization, and chemical energy. These last three forms of energy make up the internal energy of a system. A system can exchange energy with the surroundings in the forms of heat and/or work.

Next, the unit discusses the work of Thomson and Joule, which demonstrated that energy cannot be created nor destroyed (the First Law of Thermodynamics), but it can be converted from one form to another.

The First Law of Thermodynamics is followed by a discussion of enthalpy. The enthalpy change (ΔH) of a system is the change in internal energy plus the work done by or to the system. ΔH can be positive (endothermic) or negative (exothermic). Calorimeters are used to measure the enthalpy change of a reaction. The final sections of the unit cover bond enthalpy and the energy values of food.

Sidebar Content

1. **State Functions:** This sidebar explains the difference between state and path-dependent functions.
2. **Food Energy:** Gram for gram, fat stores more energy than protein and carbohydrates.
3. **Nitrogen's Triple Bond:** The Haber-Bosch process is an industrial process that converts nitrogen gas (N_2), into fertilizer, which contains nitrogen that plants can use to grow. The Haber-Bosch process has also made the production of explosives more economical.

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.

Control a Haber-Bosch Ammonia Plant Interactive

This interactive allows students to control an ammonia plant to maximize profit. Students will be able to observe the effect of changing the temperature on the production of ammonia. The background information also discusses the role of bond enthalpy in the productions of artificial fertilizer. Please note that a lesson plan and student worksheet are available online with this interactive.

Chemistry of Running Interactive

Students will try to maximize the performance of their avatar in a marathon using information from this unit. The body can extract the most energy from food using oxygen in aerobic respiration (glycolysis). When oxygen delivery is maxed out, the body also performs anaerobic respiration, which causes the buildup of lactic acid.

During the Session

Before Facilitating this Unit

There are many concepts in this unit that are traditionally difficult for students to understand. For instance, there are a lot of common misconceptions that make learning this material challenging. Pay close attention to the tips and suggestions section and use the activities and demonstrations to help reinforce these concepts. Be sure to leave plenty of time for class discussion. The video highlights the main ideas in this unit and can provide students with real world applications to these challenging concepts.

Tips and Suggestions

1. **Students often think heat and temperature are the same.** Be sure to clearly and explicitly explain the difference between heat (thermal energy measured in Joules) and temperature. This guide provides a demonstration that helps show the difference. (See In-Class Demonstrations.)
2. **A common misconception is that breaking bonds releases energy, and making bonds takes energy.** In fact, breaking bonds takes energy and making bonds releases energy. To drive this point home, use the analogy of breaking a stick; breaking the stick requires effort (an input of energy).
3. **Some students find it hard to remember that exothermic reactions warm up and endothermic reactions cool down.** Because we describe exothermic reactions as releasing energy, students conclude that if something loses energy, it should cool down. Emphasize that the chemical bonds lose energy in an exothermic reaction, and the surrounding solution then heats up.
4. **Some students are confused by the definition of “moles of reaction.”** The units of enthalpy (ΔH) are kilojoules per mole, and students sometimes ask, “per mole of what?” The moles in this case are moles of the reaction as written. In other words, the ΔH for the reaction $2A \rightarrow 2B$ will be double the ΔH for the reaction $A \rightarrow B$.

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.

UNIT 7 The Energy in Chemical Reactions

1. What is energy? What different kinds of energy can you name?
2. What happens to the energy of gasoline after it burns inside a car engine?
3. What chemical reactions can you think of that release energy? What chemical reactions can you think of that absorb energy (or require energy to happen)?
4. Why do you think some reactions release energy, and some absorb energy?
5. Why does fatty food contain a lot of calories? And why does your body store extra calories as fat?

Before Watching the Video

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

1. What is the First Law of Thermodynamics?
2. What is “enthalpy”?
3. What is an exothermic reaction? Is the change in enthalpy positive or negative?
4. What is an endothermic reaction? Is the change in enthalpy positive or negative?
5. What are calorimeters used for?
6. What is the difference between food calories and chemistry calories?
7. What is bond enthalpy, and how are the values used?
8. What is work?
9. What is the Second Law of Thermodynamics?
10. What is a PV diagram used for?
11. If an engine produces a lot of heat, what does that tell you about its mechanical efficiency?

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 practical uses can you think of for exothermic chemical reactions? For endothermic chemical reactions?
2. Why is it technically incorrect to say that a power plant or engine “produces” energy?

3. Electrical energy is generated in different ways, such as burning coal, hydroelectric dams, nuclear power plants, wind turbines, and solar panels. What forms of energy do these use as their sources?
4. What other uses could you think of for calorimetry?
5. Why might a chemist use bond enthalpies to calculate a ΔH value instead of calorimetry?
6. As you saw in the video segment about the car engine, some energy is always lost in the form of heat. How could a car engine or a power plant make use of this energy instead of letting it go to waste?

Group Learning Activities

Finding the Enthalpy of Reaction (ΔH)

Objective

In this activity students will determine the ΔH values for two chemical reactions.

List of Materials

Each group requires:

- 1 Styrofoam coffee cup
- 50 mL of 1 M HCl
- 50 mL of 1 M NaOH
- 15 g sodium acetate
- Thermometer
- Scale
- Weigh boats

Procedure (Part I: The reaction of NaOH(aq) and HCl(aq))

1. Write the balanced equation for the following reaction:
$$\text{NaOH(aq)} + \text{HCl(aq)} \rightarrow$$
2. Weigh an empty coffee cup. Record the mass.
3. With a graduated cylinder, obtain 50 mL of 1 M HCl and pour it into the coffee cup.
4. Rinse the graduated cylinder and obtain 50 mL of 1 M NaOH.
5. With a thermometer, measure the temperature of the HCl.
6. Pour the NaOH into the coffee cup and stir gently with the thermometer.

UNIT 7 The Energy in Chemical Reactions

Record the highest temperature that the thermometer reaches.

7. Weigh the cup and the liquid.

Discussion

The following data table can help students when recording information for this experiment.

Mass of empty cup = _____	Starting temperature = _____
Mass of cup with liquid = _____	Ending temperature = _____
Mass of liquid = _____	Change in temperature = _____

The following questions can help guide students thinking about this activity:

1. Was the reaction exothermic or endothermic? How do you know?
2. Describe the flow of energy in the system.
3. Calculate the amount of heat the reaction gave off using $q = mc\Delta T$ ($c = 4.18$)
4. Calculate the moles of HCl and NaOH that reacted.
5. Calculate the ΔH of this reaction using $q = n\Delta H$
6. If the solutions you used were more concentrated, would the value of q that you calculated in question 2 be higher or lower?
7. If the solutions you used were more concentrated, would the value of ΔH that you calculated in question 4 be higher or lower?

Procedure (Part 2: The Dissolving of Sodium Acetate ($\text{NaC}_2\text{H}_3\text{O}_2$))

1. Write the balanced equation for this reaction.
$$\text{NaC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \rightarrow$$
2. Obtain 15 g of $\text{NaC}_2\text{H}_3\text{O}_2$ (sodium acetate) on a weighing boat.
3. Break any large chunks of the $\text{NaC}_2\text{H}_3\text{O}_2$ into pea-size pieces. (It will not dissolve fast enough if the pieces are too large.)
4. With a graduated cylinder, obtain 100 mL of tap water. Pour it into the coffee cup and measure the temperature.
5. Place the $\text{NaC}_2\text{H}_3\text{O}_2$ in the water, stir gently with the thermometer, and record the lowest temperature reached.

6. Weigh the cup and the liquid.

Discussion

The following questions can help guide students thinking about this activity:

1. Was the reaction exothermic or endothermic? How do you know?
2. Calculate the ΔH of the reaction of sodium acetate with water.
3. What practical use might this reaction have?

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, lab coat, and gloves. This activity should be performed in a fume hood. Hydrochloric acid: Toxic by ingestion or inhalation; severely corrosive to skin and eyes. Sodium hydroxide: Corrosive liquid; skin burns are possible; very dangerous to eyes; wear gloves. Sodium acetate: Skin, eye and respiratory irritant. LD50 3530 mg/kg.

Disposal

Check local regulations for proper disposal of the chemicals.

In-Class Chemical Demonstrations

Difference Between Heat (Thermal Energy) and Temperature

Objective

In this demonstration the same amount of thermal energy is added to different volumes of water. The temperature of each volume of water will be different even though the same amount of heat is added to each container of water. The smaller volume will have a higher temperature than the larger volume because you are adding more heat per molecule into the system.

List of Materials

- 1 large heating surface (large enough to hold two large beakers)
- 2 large glass beakers of water
- 2 thermometers
- Heat resistant gloves to handle hot beakers

Procedure

1. Fill beakers with water so that one has twice as much water as the other.
2. Place each beaker on the heating surface and turn it on.
3. Indicate that the same amount of thermal energy is being added to each beaker of water.

UNIT 7 The Energy in Chemical Reactions

Ask students what they think will happen to the temperature of each beaker of water. Will they be the same temperature? Will one heat up faster than the other?

4. Before the water boils, take the temperature of both beakers of water. The smaller beaker should be at a lower temperature than the larger beaker.

Discussion

The following questions can help students understand what they observe.

1. If the thermal energy being added to this system is the same, why does one have a higher temperature than the other?
2. Describe the flow of energy in this system.
3. What is the difference between heat and temperature?
4. How are heat and temperature related?
5. Can these two beakers of water ever reach the same temperature? If so, under what conditions?
6. What would happen to the temperature of each beaker of water if you added the same amount of ice to each beaker? Would the ice melt faster in one beaker than in the other? (You may wish to demonstrate this as well)
7. What would happen to the temperature of the water if you had the same volume of water, but one container was made out of copper and the other was made of glass? Describe the energy flow in this system. (You may wish to perform this demonstration as well.)

Hazards

Be careful when using the burners. Wear thermal gloves to handle hot beakers.

Disposal

There are no special disposal considerations.

The Conversion of Mechanical Work to Thermal Energy

Objective

In this demonstration running a blender causes the temperature of the water inside to increase, showing an example of mechanical work converting to thermal energy.

List of Materials

- Blender
- Water
- Thermometer

Procedure

1. Fill a blender half way with room temperature water.
2. Take the temperature of the water.
3. Run the blender on the highest speed for a few minutes.
4. Take the temperature of the water.

Discussion

The following questions can help students understand what they observe.

1. In this example, what is the mechanical work?
2. Why does the temperature of the water increase when the blender is turned on?
3. Describe the flow of energy in the system.

Hazards

Be careful when operating the blender. Make sure the lid is on before starting the blender. Do not operate the blender when the thermometer is in it.

Disposal

There are no special disposal considerations.

Endothermic Reaction

Objective

This demonstration provides a safe example of an endothermic reaction of citric acid solution and sodium bicarbonate:



List of Materials

- 8 g sodium bicarbonate
- 25 mL of 1 M citric acid
- 400 mL beaker
- Thermometer

Procedure

1. Add 8 g of sodium bicarbonate to 25 mL of 1 M citric acid in a 400-mL beaker.
2. Add the solid slowly to avoid bubbling over.
3. Monitor the temperature with a thermometer, and/or have students feel the bottom of the beaker to see that it is cold.

UNIT 7 The Energy in Chemical Reactions

Discussion

1. Is this an example of an endothermic or exothermic reaction?
2. Describe the flow of energy in this system.
3. What happens in an endothermic reaction?
4. What happens in an exothermic reaction?

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, lab coat and gloves. Citric acid: Severe eye irritant. Sodium bicarbonate: Slightly toxic by ingestion. LD50 4220 mg/kg.

Disposal

Check local regulations for proper disposal of the chemicals.

Going Deeper (In-Class Discussion or Reflection)

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

1. How many different forms of energy can you find in the classroom right now?
2. The lights in the classroom release energy as light. Can you trace all of the energy conversions from the local power plant to the light? Can you trace all of the energy conversions from the power plant back to the sun?
3. If you charge an electric car using electricity from a coal-powered plant, is that better for the environment than a regular gas-powered car? What information would you need to answer this question?
4. How many real-life examples of exothermic and endothermic reactions can you think of?
5. If a calorimeter is not insulated well, what effect would that have on your calculated value of ΔH ? (Assume the water in the calorimeter starts at room temperature.)
6. Do you think that drinking ice water would be an effective way to lose weight because your body would expend energy to warm it up inside your body? Would it be possible to counteract the calories in a scoop of ice cream if the ice cream were made cold enough? What information would you need to answer these questions?
7. Suppose you put two pots of water on the stove. One is a large pot for pasta; the other

is a small pot for a cup of coffee or tea. Which one boils faster (assume that the burners are on the same setting)? What is the temperature of the boiling pasta water? What is the temperature of the boiling coffee/tea water?

8. Write the chemical equation for the combustion of octane (the principal hydrocarbon in gasoline). What will happen to the pressure if this reaction is performed in a closed container? What if the container is closed, but can change its volume?
9. Investigate how instant hot- and cold-packs (such as those available in first-aid kits) work.

Before the Next Unit

Students should read the Unit 7 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 8 video and/or read the Unit 8 text as an assignment before the next session.

References and Additional Resources

Lightman, Alan P. "The Conservation of Energy." *Great ideas in physics: the conservation of energy, the second law of thermodynamics, the theory of relativity, and quantum mechanics*. 3rd ed. New York: McGraw-Hill, 2000.

For Professional Development

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

Further Reading and Reflection Questions

Barker, Vanessa. "Students' Ideas about Thermodynamics." *Beyond Appearances: Students' Misconceptions about Basic Chemical Ideas*. A report prepared for the Royal Society of Chemistry: 58-63. Accessed July 8, 2013. <http://modeling.asu.edu/modeling/KindVanessaBarkerchem.pdf>.

1. Were you surprised by these common misconceptions about thermodynamics? Have you previously held any of these views yourself or encountered students that held these view? How have you addressed them in the past?

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2. Do you agree with the author's suggestions for progress? Do you foresee any difficulties in implementing any of her suggestions? Can you think of any other ways to improve the teaching of thermodynamic principles?

Le Merechal, Jean-Francois and Rania El Bilani. "Teaching and Learning Chemical Thermodynamics in School." *International Journal of Thermodynamics* 11:(2) 2008. 91-99. Accessed July 31, 2013. <http://ijoticat.com/article/view/1034000216>

1. Have you encountered students who had good problem solving skills but little conceptual understanding? What are some ways you could assess students to determine if they have both good problem solving skills and conceptual understanding?
2. How could you incorporate or adapt the "energy chain" model in your teaching in a way that you think could increase conceptual understanding of energy and thermodynamics?
3. Has this study influenced how you think about teaching energy and thermodynamics? Are there any other ways you could increase conceptual understanding of these topics? Explain.

Professional Development Assignments

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

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**“CO₂ 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?

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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 , $\text{Zn}(\text{NO}_3)_2$, KNO_3 , $\text{Sn}(\text{NO}_3)_2$, $\text{Ca}(\text{NO}_3)_2$, NaOH , Na_2SO_4 , Na_2CO_3 , and NaCl
- 10 mL each of 1.0 M solutions of AgNO_3 , $\text{Zn}(\text{NO}_3)_2$, KNO_3 , $\text{Sn}(\text{NO}_3)_2$, and $\text{Ca}(\text{NO}_3)_2$, 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 , $\text{Zn}(\text{NO}_3)_2$, KNO_3 , $\text{Sn}(\text{NO}_3)_2$, $\text{Ca}(\text{NO}_3)_2$, 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.

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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 bromocresol 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

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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: $[\text{Ag}^+, \text{Ba}^{2+}, \text{Cu}^{2+}, \text{Ca}^{2+}]$. 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?

UNIT 9

Equilibrium and Advanced Thermodynamics

The Delicate Balance of Chemical Reactions

Unit Overview

Unit 9 introduces the concept of entropy and its role as the driving force for chemical reactions. Students will learn the Second Law of Thermodynamics in terms of entropy rather than energy and work as it had been previously presented in Unit 7. The middle sections of the unit explain spontaneity and Gibb's free energy, explaining why some chemical reactions occur spontaneously while others do not. Even nonspontaneous reactions can still occur when coupled to spontaneous reactions. The final sections introduce chemical equilibrium, the equilibrium constant, and equilibrium shifts (Le Chatelier's Principle).

Learning Objectives and Applicable Standards

Participants should be able to:

1. Define microstate, macrostate, and entropy.
2. Explain how spontaneous processes increase the entropy of the universe.
3. Perform calculations with the Gibbs free energy equation to determine if a reaction is spontaneous.
4. Explain how nonspontaneous reactions can still happen when they are coupled with a spontaneous reaction.
5. Define the equilibrium state as the one where the Gibbs free energy is at a minimum.
6. Write the equilibrium constant expression for a chemical reaction.
7. Predict the direction a reaction will shift when: reactants/products are added/removed, the reaction vessel is compressed, or the temperature changes.

Key Concepts and People

1. **Microstates, Macrostates, and Entropy:** Ludwig Boltzmann formulated an equation for entropy based on microstates and macrostates. The more microstates in a given macrostate, the more likely that macrostate is to occur.

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2. **Entropy of Energy Quanta:** Quanta of energy tend to distribute themselves throughout matter evenly.
3. **Entropy and States of Matter:** Different states of matter have different amounts of entropy associated with them; in general, solids have the least entropy and gases have the most.
4. **Spontaneity and Gibbs's Free Energy:** The Gibbs free energy equation allows us to determine the spontaneity of a chemical reaction.
5. **Coupling Reactions:** A nonspontaneous reaction can still occur if it is coupled to a spontaneous reaction.
6. **Chemical Equilibrium:** When a reaction has proceeded to the point where the Gibbs free energy has reached a minimum, the reaction has reached equilibrium.
7. **The Equilibrium Constant:** The equilibrium constant expresses the ratio of products to reactants at equilibrium. Reactions with high equilibrium constants favor the product side at equilibrium, while reactions with low equilibrium constants favor the reactant side.
8. **Le Chatelier's Principle:** A chemical reaction at equilibrium will react to resist any changes made to the system and return to equilibrium.
9. **Temperature and Equilibrium:** A change in temperature affects the equilibrium constant. Whether the equilibrium constant increases or decreases is a function of whether the reaction is endothermic or exothermic.

Video

Video Content

Light a match and chemical change happens in a one-way process: Reactants are transformed into products. But there are many chemical reactions called "equilibrium reactions" that operate in both directions: with reactants and products always present. The Unit 9 video will show how chemical equilibrium works, the essential role it plays in the function of the human body, and how it is exploited in chemical processes such as ammonia synthesis, a process that provides food for up to half the world's population.

Host Introduction

Dr. Wilton Virgo, assistant professor at Wellesley College, introduces spontaneous and nonspontaneous chemical reactions by discussing the rusting of iron and the smelting of iron ore.

Host Science Explanation

“Dynamic Equilibrium”

Dr. Wilton Virgo explains chemical equilibrium as a state in which a chemical reaction runs forward and backward simultaneously. At equilibrium, both reactions continue to occur, but because they occur at the same rate, the concentrations of reactants and products remain constant. While the reaction rates are equal at equilibrium, the concentrations of reactants and products need not be. The equilibrium constant K is the ratio of product concentration to reactant concentration. When the value of K is high, products are favored at equilibrium. When the value of K is low, reactants are favored.

Laboratory Demonstration

“Le Chatelier’s Principle”

Dr. Wilton Virgo demonstrates Le Chatelier’s Principle using the reaction that forms the iron(III) thiocyanate complex: $\text{Fe}^{3+} + \text{SCN}^- \rightleftharpoons \text{FeSCN}^{2+}$. Adding the reactant Fe^{3+} shifts the reaction to the right, and precipitating the Fe^{3+} shifts the reaction to the left.

Real World Application

“Counteracting Poison”

Dr. Daniel Deschler at the Massachusetts Eye and Ear Infirmary shows how manipulating two important equilibrium reactions in the blood can treat carbon monoxide poisoning.

Laboratory Demonstration

“Pressure and Le Chatelier’s Principle”

Daniel Rosenberg, Lecture Demonstrator at Harvard University, demonstrates how changing pressure can shift the gaseous equilibrium reaction $2\text{NO}_2 \rightleftharpoons \text{N}_2\text{O}_4$ according to Le Chatelier’s Principle.

Lab Demonstration

“Temperature and Equilibrium”

Daniel Rosenberg demonstrates the effect of temperature on the equilibrium constant by heating and cooling tubes containing the $2\text{NO}_2 \rightleftharpoons \text{N}_2\text{O}_4$ reaction.

History of Chemistry

“The Haber-Bosch Process”

Authors Thomas Hager and Sam Kean, and Nobel prize-winning MIT professor Dr. Richard Schrock, describe the race to produce artificial fertilizer in the early 1900s. Natural sources of fixed nitrogen were drying up, and without the ability to harness atmospheric nitrogen to fertilize crops, the world faced potential mass starvation. Fritz Haber found a way to react nitrogen and hydrogen gas at high temperature and pressure to create ammonia.

Unit Text

Content Overview

The text begins with an illustration of Ludwig Boltzmann's work in statistical mechanics using a simple example of gas particles distributing themselves throughout a container. This leads to definitions of the terms "microstate" and "macrostate." "Entropy" represents a quantity that reflects the number of microstates in a given macrostate. The entropy of energy quanta and states of matter is discussed. The text then introduces Gibb's free energy and shows how the spontaneity depends on the enthalpy change and entropy change of the system. Spontaneous reactions can drive non-spontaneous reactions; entropy can decrease in one part of a system as long as the overall entropy of the system increases.

The text then explains equilibrium and the equilibrium constant. It discusses Le Chatelier's Principle, which states that a chemical reaction at equilibrium will shift to counteract any changes (concentration, pressure, temperature) made to the system. The text concludes by examining how temperature affects equilibrium.

Sidebar Content

1. **Osmosis and Entropy:** This sidebar explores the role of entropy in osmosis. It further explains the role of osmosis in food preservation and in the diseases cholera and cystic fibrosis.
2. **Q, the Reaction Quotient:** This sidebar introduces the reaction quotient and shows how to calculate it to predict the direction of a reaction.

Interactives

Historical Timeline of Chemistry

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

Control a Haber-Bosch Ammonia Plant Interactive

In this interactive, students attempt to maximize the profit of an ammonia plant. Ammonia production stops when the reaction reaches equilibrium. Using the principles in this unit, students will be able to manipulate the equilibrium to keep the ammonia flowing. Students must find the best combination of pressure and temperature to maximize production, as well as apply Le Chatelier's Principle. Please note that a lesson plan and student worksheet are available online with this interactive.

During the Session

Before Facilitating this Unit

The video provides students with a qualitative introduction to chemical equilibrium, the equilibrium constant, and Le Chatelier's Principle. The text approaches equilibrium from the point of view of thermodynamics. Students will learn that equilibrium is a state in which entropy is maximized; the Gibbs free energy has reached a minimum. After watching the history of the Haber-Bosch process in the video, students can see it in action with the Control a Haber-Bosch Ammonia Plant interactive.

Tips and Suggestions

1. **The concept of entropy is difficult for many students to grasp.** Many teachers equate entropy with disorder, which can lead some students astray. To avoid this, emphasize the statistical origins of the concept of entropy.
2. **Entropy works on the level of particles.** While analogies like shuffling cards can be useful to get the concept of entropy across, some students can come away with the mistaken impression that entropy applies to macroscopic objects. The statistics behind entropy require that particles and quanta of energy move around randomly, and large everyday objects do not.
3. **This text approaches equilibrium from a thermodynamic point of view.** Many explanations of equilibrium start with kinetics; the kinetics approach is more accessible for many students and may be appropriate for introductory classes.

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 does the word "equilibrium" mean to you? What does it mean in everyday life? What do you think equilibrium means in chemistry?
2. What is a spontaneous reaction?
3. Why do some chemical reactions occur and others do not? Why is it impossible to "unburn" something?
4. Our bodies burn glucose for energy in a spontaneous reaction. Plants "unburn" glucose when they photosynthesize. How can both these reactions happen, seemingly spontaneously?
5. If ten people flipped ten coins simultaneously, what would be the percentages of heads and tails? Would you expect it to be exactly 50/50? Why or why not? What if 100 people flipped the coins? A million?

Before Watching the Video

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

1. Why doesn't pure iron metal exist in nature?
2. Does a reaction at equilibrium stop?
3. At equilibrium, are the amounts of products and reactants equal?
4. What is the equilibrium constant (K)?
5. What is the role of hemoglobin in the blood?
6. How does the reaction shift when hemoglobin is in the lungs? In the tissues?
7. How does a change in temperature affect an equilibrium reaction?
8. Why can't plants use atmospheric nitrogen?
9. What was Haber's major contribution to chemistry? Why is it so important to humanity?

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 an equilibrium reaction?
2. What is Le Chatelier's Principle?
3. How does hyperbaric oxygen treat carbon monoxide poisoning?
4. How does a change in pressure affect a gas-phase equilibrium reaction?
5. Hyperbaric oxygen is effective in treating carbon monoxide poisoning because pressure can have a large effect on gas-phase equilibria. Why does pressure NOT have the same effect on aqueous reactions?
6. While Fritz Haber's process benefitted humanity greatly by harnessing nitrogen for fertilizer, what else was the process used for? What is the other side of Haber's legacy?

Group Learning Activities

Graphing Equilibrium

Objective

This activity simulates a chemical system coming to equilibrium. Students will observe that the ratio of products to reactants at equilibrium is the same regardless of starting amounts.

List of Materials

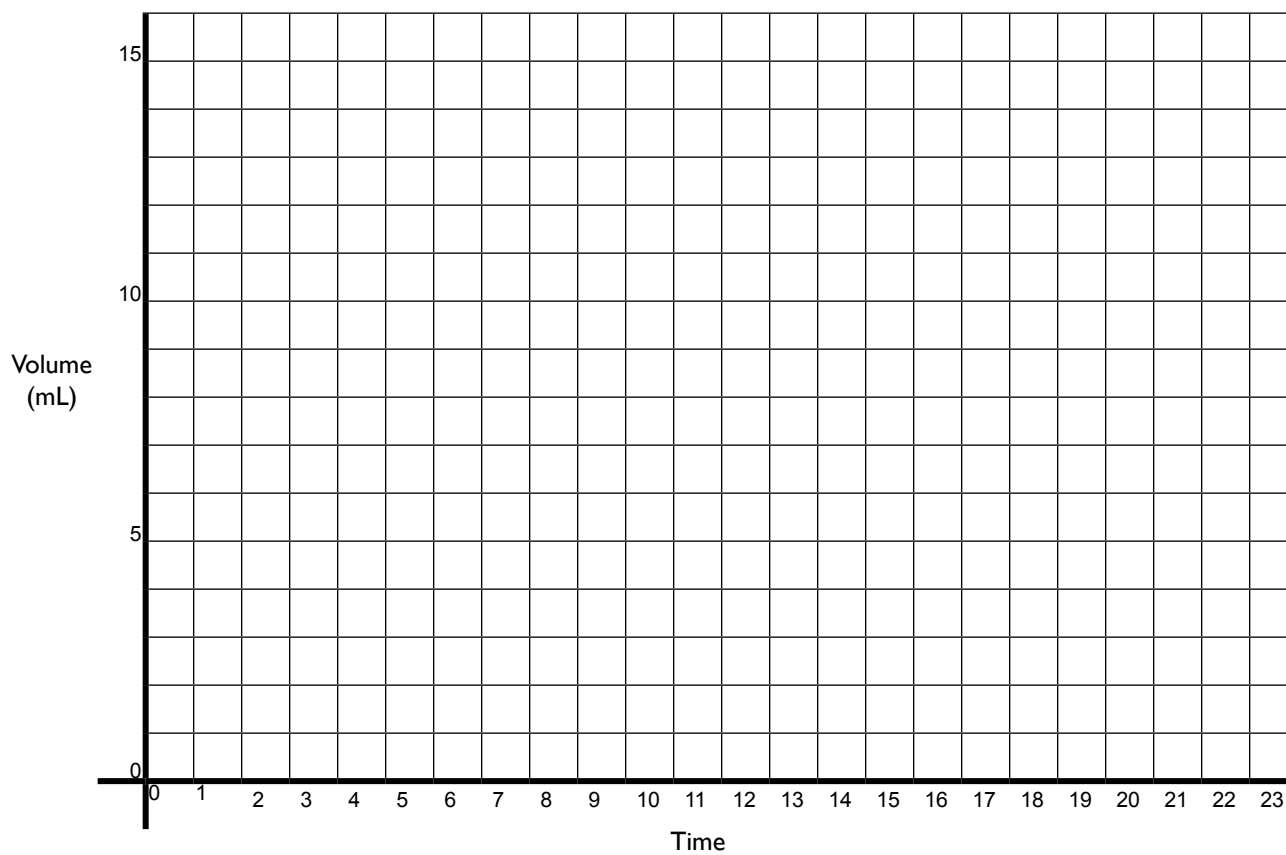
- Two 25-mL graduated cylinders
- Two different sized pipettes (one about twice the size of the other) or two straws or two tubes (one about twice the diameter of the other)
- Water

Procedure I

1. Place the two graduated cylinders on the bench; the one on the left will be designated “reactant” and the one on the right will be designated “product.”
2. Add 15 mL of tap water to the reactant cylinder. Mark the volume that each cylinder contains on the graph below at time = 0.
3. SIMULTANEOUSLY insert the large pipette or tube into the reactant cylinder all the way to the bottom and the small pipette/ tube all the way into the product cylinder. Fill the pipettes/ tube with the water. To fill the tubes, squeeze the tubes with your thumb and middle finger and cover the tops of each tube with your index finger, then lift out the tubes. Transfer the water to the OPPOSITE cylinder.
4. Record the volume of water in both tubes and record on the graph for time = 1. Use different colors on the graph to distinguish products and reactants.

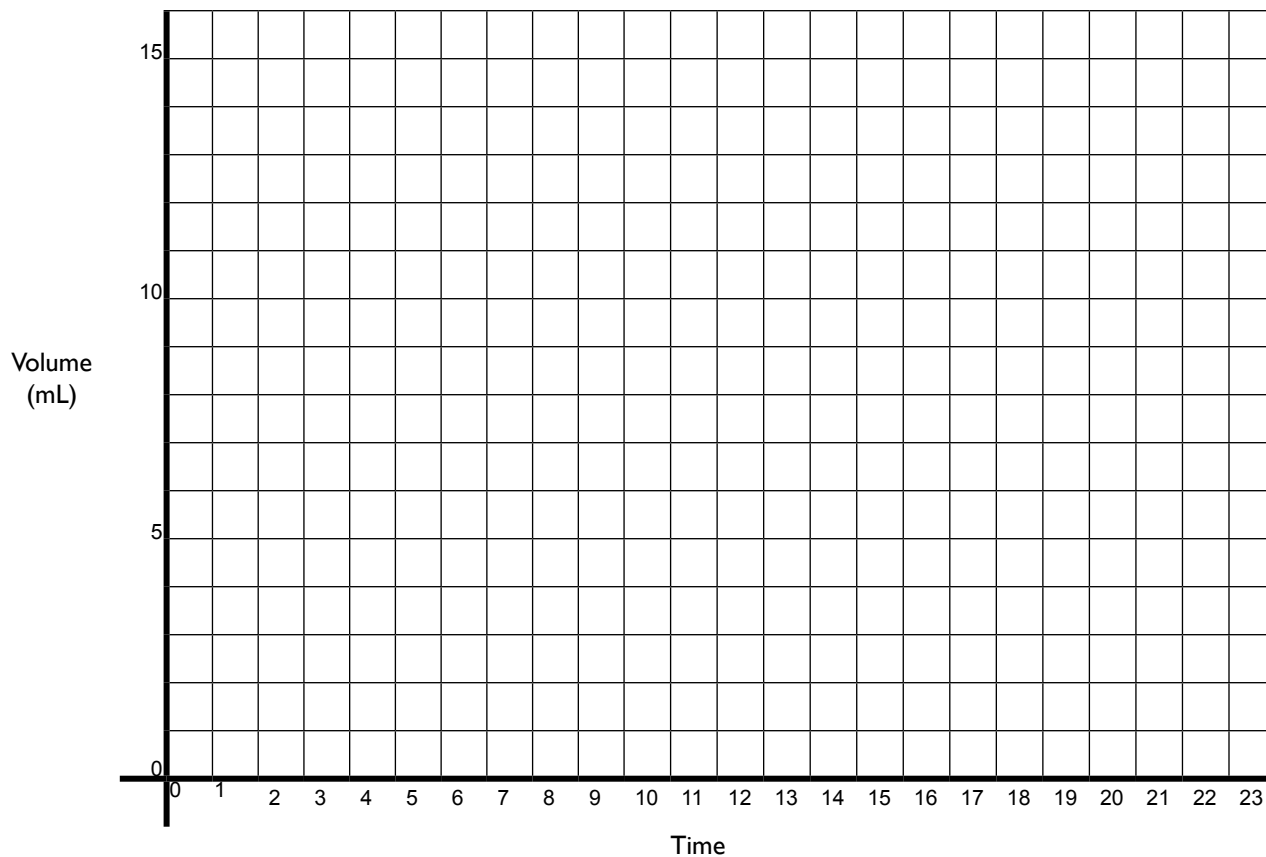
UNIT 9 Equilibrium and Advanced Thermodynamics

- Repeat steps 3 and 4 over and over until the amounts of products and reactants remain constant.



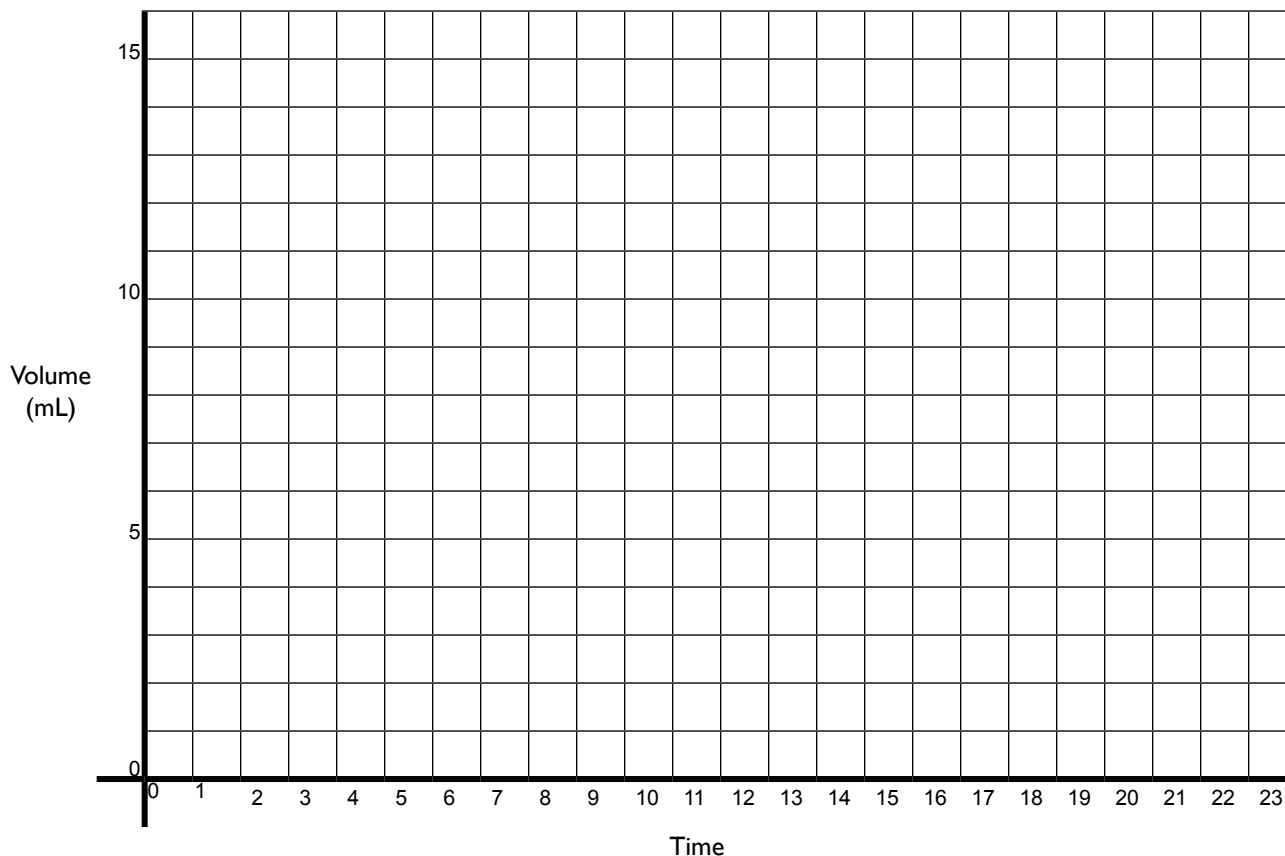
Procedure 2

Repeat the same procedure, this time starting with the **reactant** cylinder **empty** and the **product** cylinder with 15 mL. Graph your data below.



Procedure 3

Repeat the same procedure. This time, start out both cylinders with 7.5 mL of water. Graph your data below.



Discussion

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

1. How do you know when the reaction reaches equilibrium?
2. Once the reaction reaches equilibrium, do the reactions stop?
3. At equilibrium, are the rates of the forward and reverse reactions the same or different?
4. At equilibrium, do the amounts of reactants and products change?
5. Calculate the ratio of products to reactants for each reaction at equilibrium. Does this ratio depend on the starting amounts of products and reactants?

6. If the reaction has reached equilibrium, what would happen if more water were added to the reactant cylinder? The product cylinder?
7. What could you change about the experiment to make the final equilibrium ratio different from what you found?

Hazards

Wear goggles to prevent water splashing into eyes.

Disposal

There are no special disposal considerations.

Gibb's Free Energy and Le Chatelier

Objective

The contraction of a stretched rubber band is a spontaneous process. This demonstration allows students to figure out the signs of ΔH and ΔS for the following process:

stretched \rightarrow contracted

The demo works best with thick, heavy rubber bands.

List of Materials

- Thick heavy rubber bands
- Ring stand
- Hair dryer

Procedure

1. Working in pairs, one student should stretch the rubber band and hold it in a stretched position for 20 seconds.
2. The other student should touch the middle of the rubber band, while the one holding the rubber band allows it to quickly contract. Be careful not to snap the rubber band on the other student's hand.
3. Repeat so the other student can feel the temperature of the rubber band as it expands and contracts.

Discussion

The contracting band should feel cool, indicating that contraction is an endothermic process (ΔH is positive). If contraction is spontaneous but endothermic, the spontaneity must be due to an increase in entropy (ΔS is positive). The polymer strands have greater entropy when contracted, and less entropy when stretched; stretching the band puts the strands in a more parallel, ordered arrangement. The following questions can help guide students' thinking during this activity:

UNIT 9 Equilibrium and Advanced Thermodynamics

1. How does the temperature change when the rubber band contracts?
2. Does the temperature change indicate that the contraction is an exothermic or endothermic process? Why?
3. Is the rubber band contracting a spontaneous or nonspontaneous process?
4. If the contraction of the rubber band is spontaneous and it's an endothermic process, what must happen to the entropy? Why?

To introduce Le Chatelier's Principle you can also hang a rubber band from a ring stand and suspend a weight from the bottom so the band is partially stretched. Ask the students to predict whether the band will become longer or shorter when heated with a hair dryer. (The band will shorten according to Le Chatelier's Principle.)

Hazards

Students should not snap rubber bands at other students. Wear safety glasses for added protection in the case of flyaway rubber bands.

Disposal

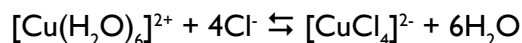
There are no special disposal considerations.

In-Class Chemical Demonstrations

Le Chatelier's Principle with Copper(II) Chloride Complexes

Objective

Students will see the following reaction shift according to Le Chatelier's Principle:



Adding chloride ions will shift the reaction to the right; diluting it will shift it left. This reaction also demonstrates how changes in temperature can change the equilibrium constant (K).

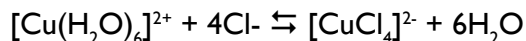
List of Materials

- 0.5 M Solution of CuCl_2
- HCl
- Stirring rod
- Beaker

Procedure

1. In a 0.5 M solution of CuCl_2 the copper ions form a complex with water which gives the solution a light blue color $[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$.
2. While stirring constantly, slowly pour in HCl.

3. The following equilibrium will shift to the right, and the solution will turn light green due to the presence of $[\text{CuCl}_4]^{2-}$ ions.



4. The reaction will shift the other way when diluted with distilled water.
5. The above reaction is endothermic; you can demonstrate this by heating the solution to make it greener and chilling it to make it bluer.

Discussion

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

1. When the chloride ions were added by adding HCl, which way did the equilibrium shift, and why?
2. When the solution was diluted, which way did the equilibrium shift, and why?
3. Silver ions combine with chloride ions to form the insoluble precipitate AgCl. If silver ions were added to the copper(II) chloride solution, which way would the equilibrium shift? What color change would you expect?
4. Compare the heated solution with the cooled solution. At which temperature were there more products?
5. Based on your answer to question 4, does the equilibrium constant for this reaction increase or decrease as temperature increases?
6. Based on your answer to question 4, is the reaction endothermic or exothermic?

Hazards

It is good lab practice to review a chemical's Material Safety Data Sheet (MSDS) before working with any chemical. Follow instructions on the MSDS and encourage students to review them. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, lab coat, and gloves. Copper(II) chloride: Highly toxic by ingestion and inhalation. LD50 140 mg/kg. Hydrochloric acid: Toxic by ingestion or inhalation; severely corrosive to skin and eyes.

Disposal

Check local regulations for proper disposal of chemicals.

Going Deeper (In-Class Discussion or Reflection)

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

1. Someone throws four coins in the air and they all end up showing heads. Does this

UNIT 9 Equilibrium and Advanced Thermodynamics

contradict the Second Law of Thermodynamics? Why or why not? Explain in terms of microstates and macrostates.

2. A ten-year-old boy tells his mother that there's no point to cleaning up his room because the Second Law of Thermodynamics says it will inevitably get messed up again. Is this argument valid?
3. What real-life example of equilibrium can you think of? If this equilibrium were disturbed, would it resist the change according to Le Chatelier's Principle?
4. Find examples of spontaneous and nonspontaneous chemical reactions or processes. Can you think of a process that can be spontaneous in the forward direction under certain conditions, and spontaneous in the reverse direction under different conditions?

Before the Next Unit

Students should read the Unit 9 text if they haven't already done so. They may be assigned one or more reading assignments or activities from the list below, or if you choose to have them use the course materials outside of class, they can watch the Unit 10 video and/or read the Unit 10 text as an assignment before the next session.

References and Additional Resources

Lightman, Alan P. "The Conservation of Energy." *Great ideas in physics: the conservation of energy, the second law of thermodynamics, the theory of relativity, and quantum mechanics*. 3rd ed. New York: McGraw-Hill, 2000.

Lower, Stephen. "Thermodynamics of chemical equilibrium." *Steve Lower Stuff*. Accessed August 7, 2012. <http://www.chem1.com/acad/webtext/thermeq/index.html>

Phet Interactive Simulations. "Energy Forms and Changes." University of Colorado at Boulder. Accessed August 12, 2013. <http://phet.colorado.edu/en/simulation/energy-forms-and-changes>

Phet Interactive Simulations. "Reversible Reactions." University of Colorado at Boulder. Accessed August 12, 2013. <http://phet.colorado.edu/en/simulation/reversible-reactions>

The Royal Society of Chemistry. "The quantum casino." *Presenting Science*. Accessed August 7, 2012. <http://www.presentingscience.com/quantumcasino/index.html>

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

Lambert, Frank L. "Disorder- A Cracked Crutch for Supporting Entropy Discussions." *Journal of Chemical Education*. 79: 2 (2002): 187-192. Accessed August 5, 2013. http://www.researchgate.net/publication/231265843_Disorder_-_A_Cracked_Crutch_for_Supporting_Entropy_Discussion

1. Do you agree with the author that teaching entropy as disorder is misleading and damaging? Why or why not? Use specific examples in this paper or from your teaching experience to support your position.
2. Does this paper influence how you will approach teaching entropy? If so, how? If not, why not?

Tyson, Louise and David F.Treagust. "The Complexity of Teaching and Learning Chemical Equilibrium." *Journal of Chemical Education*. 76: 4 (1999): 554-558. Accessed on August 5, 2013. http://www3.fsa.br/fabricio/images/Chemical_equilibrium_JCE_Tyson_et_al.pdf

1. Compare and contrast the benefits and challenges associated with using Le Chatelier's Principle, Equilibrium Law, and collision theory in teaching about equilibrium? Do you agree with the author's recommendations for how to use these explanations?
2. Have you encountered students who were confused by equilibrium terminology such as equilibrium, closed system, or equilibrium shift/position? What are some ways you might be able to address these discrepancies in language use?
3. Were you surprised that students are confused by the difference between physical and chemical equilibrium? Why or why not? How could you avoid introducing this kind of confusion to your students?

Professional Development Assignments

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

UNIT 10

Acids and Bases

The Voyage of the Proton

Unit Overview

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

Learning Objectives and Applicable Standards

Participants should be able to:

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

Key Concepts and People

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

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

- The Dissociation of Water:** Water molecules exist in equilibrium with hydronium ions and hydroxide ions. Arrhenius acids and bases form water when they neutralize each other.
- The pH Scale:** The acidity of a solution is measured in pH. A solution's pH indicates the concentration of hydronium ions. The pH scale is a logarithmic scale running from zero to 14. Anything with a pH below seven is considered an acid, and anything with a pH above seven is considered a base. Solutions with a pH of seven are neutral.
- Strengths of Acids and Bases:** Acids and bases come in different strengths. The stronger the acid or base, the more it dissociates in aqueous solutions. The dissociation constant (K_a) reflects the extent of dissociation.
- Measuring pH:** pH meters use electrochemistry to measure the concentration of hydronium ions (H_3O^+) and thus can provide the pH of solutions. Indicators change color at certain pH levels, so chemists use them to estimate the pH of solutions.
- Neutralization and Acid Base Titrations:** In general, when acids and bases react together, they neutralize each other, and the pH of the solution gets closer to seven. A titration curve charts the pH of a neutralization reaction as the acid and base slowly combine. Titration curves are used to determine the concentration of an unknown and to find pK_a values.
- Conjugates of Acids and Bases:** A Brønsted-Lowry acid is anything that can donate a proton (H^+), and a Brønsted-Lowry base is anything that can accept a proton. When the proton is dissociated from an acid, the remaining molecule is then called the conjugate base of the acid. Similarly, once a proton is picked up from a base, the molecule is now called a conjugate acid.
- Buffers:** Indicators and buffers are examples of shifting equilibria. Buffers stabilize pH levels at a certain value.

Video

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

ment, acid rain plagues industrial portions of the world; the chemical nature of acid rain reactions and the environmental response and impact are part of acid-base chemistry.

VIDEO CONTENT

Host Introduction

“What is an Acid, What is a Base?”

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

Laboratory Demonstration

“Corrosive Acids and Bases”

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

Host Science Explanation

“A Cheesy Weak Solution”

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

Real World Application

“Baked Reactions”

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

Host Science Explanation

“Acidic Pond”

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

Laboratory Demonstration

“Buffered Lemonade”

Harvard University Lecture Demonstrator Daniel Rosenberg makes two solutions. One is pure water and the other is water with equal amounts of weak acid and weak base. When lemonade is introduced into the pure water solution, it instantly turns acidic from the citric acid in the lemonade. When the lemonade is added to the buffered solution with a weak acid and weak

base, the pH doesn't change substantially. The buffered solution is able to take on the added acid without becoming acidic.

Real World Application

“Removing Acids from the Body”

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

Unit Text

Content Overview

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

Sidebar Content

1. **Svante Arrhenius:** This Swedish chemist created a definition of acids and bases still used today.
2. **Calculating pH:** This sidebar describes the origins of the pH scale and how the pH of a solution is calculated.
3. **Cooking Fish With Citrus Juice – Ceviche:** The acids in citric juice denature the proteins in fish and make the fish appear cooked.
4. **The Dissociation Constant of Water:** The dissociation of water is an equilibrium reaction, with an equilibrium constant called K_w .
5. **Color Changing Flowers:** Many plants contain various forms of anthocyanins and respond to changing pH differently, so they can be used as indicators in a broad range of solutions of different acidities. For example, hydrangea flowers actually change color depending on the acidity of the soil they're growing in. In acidic soils they are blue and in basic soils they appear pink.

6. **Acid Rain:** The release of sulfur dioxide by burning coal leads to acid rain and widespread environmental damage.
7. **Amphoteric Compounds:** Amphoteric compounds are molecules that can act as an acid or a base depending on the situation.
8. **Resonance and Acidity:** Carboxylic acids are common in organic and biochemistry and they can form resonance structures and give up a proton easily.
9. **Acidosis and the Body:** Changes in the pH of blood can lead to problems such as rickets and the bowing of legs.

Interactives

Historical Timeline of Chemistry

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

Chemistry of Running Interactive

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

During the Session

Before Facilitating this Unit

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

Tips and Suggestions

1. **Chemically speaking, acid strength (the K_a value) is different from concentration.** A highly concentrated solution of a weak acid may have a lower pH than a very dilute solution of a strong acid.
2. **Indicators and Buffers are applications of Le Chatelier's principle.** At this point in most chemistry courses, students have learned about chemical equilibrium.

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

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

Starting the Session: Checking Prior Thinking

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

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

Before/While Watching the Video

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

Watch the Video

After Watching the Video

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

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

Group Learning Activities

Buffer Activity

Objective

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

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

List of Materials

Each group requires:

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

Procedure (Part I)

1. In a small Erlenmeyer flask, pour approximately 10 mL of distilled water. Add four drops of universal indicator. The solution should be yellow to green (about pH 5 to 7).
2. Drop by drop, add 0.1 M HCl to the flask. Record how many drops it takes to decrease the pH by one unit.
3. Rinse out the flask and repeat the procedure with 0.1 M NaOH. Record how many drops it takes to increase the pH by one unit.

Procedure (Part II)

1. Using a pipette bulb and pipette, put 5 mL of 0.1 M NaH_2PO_4 solution into a small Erlenmeyer flask. To the same flask, add 5 mL of 0.1 M Na_2HPO_4 . Add four drops of universal indicator. The solution should be green (about pH 7).
2. Drop by drop, add 0.1 M HCl to the flask. Record how many drops it takes to decrease the pH by one unit.
3. Rinse out the flask and repeat the procedure using 0.1 M NaOH instead of HCl. Record

UNIT 10 Acids and Bases

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

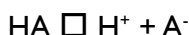
Discussion

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

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

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

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



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

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

Hazards

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

Universal indicator: Alcohol-based solution; flammable liquid.

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

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

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

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

Disposal

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

Indicators and pH**Objective**

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

List of Materials

- 0.02 M HCl
- 0.02 M $\text{HC}_2\text{H}_3\text{O}_2$
- Distilled water
- 0.02 M NH_3
- 0.02 M NaOH
- Turmeric extract (ground turmeric dissolved in ethanol)
- Methyl red
- Thymolphthalein
- Well plate

Procedure

- I. Pre-lab Calculations: Calculate the pH of the following solutions. The K_a of $\text{HC}_2\text{H}_3\text{O}_2$ is 1.8×10^{-5} ; the K_b of NH_3 is 1.8×10^{-5} .

0.02 M HCl	0.02 M $\text{HC}_2\text{H}_3\text{O}_2$	H_2O	0.02 M NH_3	0.02 M NaOH

UNIT 10 Acids and Bases

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

	HCl	HC ₂ H ₃ O ₂	H ₂ O	NH ₃	NaOH
turmeric extract					
methyl red					
thymolphthalein					

Discussion

The following questions will help students think about this activity.

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

Hazards

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

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

Acetic acid: Corrosive liquid.

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

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

Ethanol: Dangerous fire risk; flammable; addition of denaturant makes the product poisonous-it cannot be made nonpoisonous. TLV 1880 mg/m³.

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

Thymolphthalein: Alcohol solution; flammable liquid.

Disposal

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

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

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

List of Materials

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

Set Up

1. Before the session, chop the cabbage, cover with boiling water, and allow to come to room temperature. Strain the liquid.
2. If desired, soak strips of filter paper in the liquid and allow to dry for pH paper.

Procedure: Making a set of pH standards:

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

2. Transfer 1.0 mL of this diluted solution into another test tube, and add 9mL of distilled water. Mix well.
3. Repeat this procedure five more times, then repeat entire procedure for the other compound. Each tenfold dilution will produce a solution with a pH approximately one unit different from the previous test tube. 1.0 M HCl (pH 0) can be diluted down to pH ~7, as can 1.0 M NaOH (pH 14).
4. Label each test tube with the approximate pH, and place a few drops of the indicator in each tube to show the color for that pH.

Procedure: Testing the pH of household liquids or beverages:

1. If testing baking soda or antacids, grind antacid tablets or clumps of baking soda into a fine powder and dissolve in a few milliliters of distilled water.
2. Pour a few milliliters of each liquid into a clean test tube or beaker.
3. Add a few drops of indicator to each liquid, and record the color and the approximate pH (or whether the liquid is acidic or basic).
4. If using strips of pH paper, a drop of each liquid may be placed on a strip of the paper instead of filling a test tube and adding indicator.

Procedure: “Titrating” an antacid tablet

1. Grind an antacid tablet into a fine powder and dissolve in water.
2. Add a few drops of indicator to the beaker or test tube.
3. Using a graduated disposable pipet, add vinegar a few drops at a time and stir.
4. Add vinegar a drop at a time and record the volume added until the indicator turns pink/red (showing the presence of excess acid).

Discussion

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

1. Why did you make a pH standard scale? What information does that scale provide? Calculate the concentration of acid and base for each dilution. How does this reflect a logarithmic scale?
2. Which household liquids were acidic? Which ones were basic? Which ones appeared close to a pH of 7? How did your results compare with your expectations?
3. What information does the volume of vinegar needed to titrate the antacid tell you?
4. Vinegar is diluted acetic acid (CH_3COOH , one ionizable proton). If the main component

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

- Vinegar is typically 5% w/v, meaning that every 100 mL of solution contains 5 g of acetic acid. Look at the mass of CaCO_3 per tablet as listed on the antacid bottle, and use your equation above to determine how many milliliters of vinegar should be required to react completely with one tablet. How well does this calculation agree with your experimental results?
- All substances that behave as indicators are weak acids. How do indicators work? Why do they change color at different pH levels? (Think about Le Chatelier's principle.)

Hazards

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

Disposal

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

In-Class Chemical Demonstrations

Acid and Base Safety

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

Hazards

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

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

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

Disposal

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

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

Strong vs. Weak Acids

Using litmus paper, indicators, or a pH meter, measure the pH of four solutions: 0.02 M HCl, 0.02 M HC₂H₃O₂, 0.02 M NH₃, and 0.02 M NaOH. Emphasize that the concentrations of all the solutions are the same, and ask the students to explain why the weak acid and weak base do not produce such large changes in pH.

Hazards

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

Disposal

Check local regulations for proper disposal of chemicals.

Going Deeper (In-Class Discussion or Reflection)

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

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

Before the Next Unit

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

References and Additional Resources

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

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

For Professional Development

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

Further Reading and Reflection Questions

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

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

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

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

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

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

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

1. Were you surprised by the results of this study? Why or why not? What was particularly surprising to you?
2. Does this study influence how you will approach teaching acid-base chemistry or chemistry in general? If so, how? If not, why not?

Professional Development Assignments

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

UNIT II

The Metallic World

Unit Overview

This unit provides an overview of both electrochemistry and basic transition metal chemistry. Oxidation-reduction reactions, also known as redox reactions, drive electrochemistry. In redox reactions, one compound gains electrons (reduction) while another one loses electrons (oxidation). The spontaneous directions of redox reactions can generate electrical current. The relative reactivities of substances toward oxidation or reduction can promote or prevent processes from happening. In addition, redox reactions can be forced to run in their non-spontaneous direction in order to purify a sample, re-set a system, such as in a rechargeable battery, or deposit a coating of another substance on a surface.

The unit also explores transition metal chemistry, both in comparison to principles of main-group chemistry (e.g., the octet rule) and through various examples of inorganic and bioinorganic compounds.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Describe the difference between spontaneous and nonspontaneous redox processes in terms of both cell EMF (electromotive force) and DG.
2. Recognize everyday applications of spontaneous redox processes as well as applications that depend on the forcing of a process to run in the non-spontaneous direction.
3. Understand the basics of physiological redox processes, and recognize some of the enzymes that facilitate electron transfer.
4. Compare basic transition-metal chemistry to main-group chemistry in terms of how ions are formed and the different types of bonding in metal complexes.

Key Concepts and People

1. **Redox Reactions:** Redox reactions can be analyzed systematically for how many electrons are transferred, whether or not the reaction happens spontaneously, and how much energy can be transferred or is required in the process.

- The Activity Series:** The activity series is a list of metals ranked by how easily a metal can be oxidized (lose electrons); the higher the metal is on the list, the more easily it is oxidized. This can be used to predict if a simple redox reaction will happen spontaneously.
- Half-reactions:** Half-reactions can be written to illustrate more clearly the reduction and oxidation processes in a redox reaction. In a half-reaction, either the oxidation or the reduction reaction is described and the electrons are explicitly included.
- Standard Reduction Potentials:** Standard reduction potentials are used to determine the potential difference between two half-reactions. This is the basis not only for determining whether the reaction will run in the first place but also for determining the work that can be done by the reaction.
- Batteries and Galvanic Cells:** Batteries are arguably the most recognizable applications of redox reactions. Non-standard conditions (such as concentrations different from 1.0 M for solution-phase species) influence the observed potential difference between half-cells in a galvanic cell. Since most applications of redox chemistry actually take place under non-standard conditions, the influence of the conditions is an important factor in designing applications that use redox reactions.
- Electrolytic Reactions:** Processes that represent the nonspontaneous direction of a redox reaction are referred to as “electrolytic.” Electrolytic processes will not run without the application of a voltage larger than that predicted for the reaction, but as long as that condition is met, the process is governed by stoichiometry.
- Transition Metals:** Transition metals share many patterns of reactivity with main-group elements, but also can react in unique ways due to the *d* electrons in their valence shell.
- Colorful Chemistry:** Many transition metals are intensely colored. When ligands bond to transition metals, the metals’ *d* orbitals split into at least two different energy levels. What makes transition metal complexes colored is that these gaps between the sets of *d* orbitals have energies that fall in the visible region of light.
- Metal Biochemistry:** Although metals make up a very small percentage of elements in physiological systems, they play some very specific and critical roles in a variety of processes. In addition to hemoglobin, a number of enzymes contain various metals; many of these enzymes facilitate steps in redox processes.

Video

Metals allow the transfer of electrons through a process called oxidation-reduction, or “redox,” when one species gains electrons while another loses them. Chemists take advantage of this

process by using electron transfers to power the batteries in our flashlights, phones, or cars. In biochemistry, trace metals, such as cobalt in Vitamin B12, often drive chemical reactions that are essential for human health. Redox reactions also occur without metals, as is the case when lightening hair color.

VIDEO SEGMENT DESCRIPTIONS

Host Science Explanation

“The Metallic World and Electrochemistry”

Dr. Catherine Drennan of MIT introduces metals and the various things they can do, and defines redox.

Real World Application

“Powered by Redox”

A research group at Worcester Polytechnic Institute is programming high-functioning robots to assist people in basic activities. Because these robots need to be very mobile, they are powered by batteries—and the performance of these batteries is based on oxidation-reduction chemistry.

Laboratory Demonstration

“Oxidation-Reduction Reaction”

The reaction of copper ions with zinc metal is a single-displacement reaction, one of the simplest forms of a redox reaction. It can be performed in a single beaker, which produces a copper film on the surface of the zinc metal.

Laboratory Demonstration

“Redox: Creating an Electric Current”

Redox reactions can be set up into two “half-reactions” in separate beakers. Connection of the beakers by a salt bridge and a wire generates current through the wire. This is the basic setup of the Daniell cell, the classic spontaneous redox reaction and the basis of an early battery.

Current Chemistry Research

“Vitamin B12”

Metals are found in the body only in trace amounts, but these trace amounts play crucial physiological roles. Dr. Catherine Drennan’s research group studies the chemistry of vitamin B12, specifically the reactivity of the cobalt ion in the center of the protein. The determination of which oxidation state, Co^{+1} , Co^{+2} , or Co^{+3} , is present can be as simple as a test that takes advantage of a common transition-metal phenomenon: the different oxidation states have different colors.

Real World Application

“Folic Acid in Fetal Development”

The chemistry of vitamin B12, as studied by the Drennan group, involves a reaction with vitamin B9, or folic acid. Folic acid plays a critical role in the prevention of neural tube defects in

developing fetuses. Because the development of the neural tube occurs so early in pregnancy and the folic acid must already be in the mother's system by the time she is pregnant, many commercially available foods are fortified with folic acid. This has helped to dramatically decrease the percentage of babies born in the US with neural tube defects.

Laboratory Demonstration

“Non-spontaneous Redox Reaction”

Non-spontaneous redox reactions can be forced to occur by the application of an electric current. In this demonstration, an iron screw is plated with zinc metal by using the screw and a piece of zinc as electrodes in a solution containing zinc ions. This process, called “galvanizing,” is often used to make metals (especially iron) more resistant to corrosion.

Real World Application

“Hair, Highlights, and Redox”

Many redox reactions involve metals, but some do not. One common cosmetic redox process is the bleaching of hair using hydrogen peroxide. The hydrogen peroxide oxidizes melanin molecules in hair, which then absorbs higher-energy wavelengths of light, causing the molecules (and the hair) to appear lighter.

Unit Text

Content Overview

Redox reactions are among the simplest cases in which students can explain observed reactivity and also predict reactivity based on established patterns. Redox reactions are also the basis for a wide array of recognizable applications, and the convenient relationships between electron transfer and energy transfer allow for straightforward calculations to support how the various applications work. The text presents both the activity series for metals and the tables of standard reduction potentials in the context of predictions and quantification of reactivity, with several examples of applied redox processes.

Although the chemistry of transition metals is sufficiently rich and unique to warrant a course of its own, the behavior of transition metals is analogous in some ways to that of main group metals. Some of the classic characteristics of transition metals (their multiple stable oxidation states, their ability to form brightly-colored compounds, their ability to participate in multiple types of bonding) are introduced in comparison and contrast to the analogous main-group behaviors.

Sidebar Content

1. **The Permanganate Ion: An Excellent Oxidizing Agent:** The permanganate ion is an example of a colorful oxidizing agent. This sidebar uses the permanganate ion to illustrate how to balance redox reactions using half reactions.
2. **pH and Redox Reactions:** This sidebar describes how pH meters use redox reactions to measure the concentration of H^+ in a given solution.

3. **Galvanization:** Zinc can be electroplated onto the surface of iron, which will help prevent the iron from oxidizing (rusting). This is common practice to prevent nails and screws from rusting.
4. **Nickel Hydrazine Perchlorate:** The explosive compound nickel hydrazine perchlorate contains both an oxidizing agent and a reducing agent. If one were to simply strike it with a mortar, a violent explosion would ensue, which is exactly what happened when one graduate student was working with it. Luckily the student survived, but he ended up blowing up a corner of the lab and he lost three fingers and suffered eye damage in the process.
5. **Titanium White and Zinc Oxide:** This side bar explains why titanium white and zinc oxide appear white.
6. **Polydentate Ligands:** When one molecule donates more than one lone pair of electrons to a metal, the molecule is referred to as a “polydentate ligand,” literally meaning “having many teeth” in Latin. EDTA is a common polydentate ligand and is often used to render metals unreactive, such as when someone is suffering from heavy metal poisoning.

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 the text are represented on the timeline.

Control a Haber-Bosch Ammonia Plant Interactive

The conversion of hydrogen and nitrogen into ammonia is a nonmetallic redox reaction. While this reaction itself does not involve a metal, an iron-based catalyst helps speed up the reaction. Please note that a lesson plan and student worksheet are available online with this interactive.

During the Session

Before Facilitating this Unit

Generally, we recommend watching the videos first before reading the written units. Since there are two main components for this unit—redox reactions and transition metals—it might be helpful to teach the unit in segments. The video focuses heavily on redox processes, but includes a bit of transition metal chemistry in the segment on physiological redox. The demonstrations and activities presented in this guide are also primarily redox-focused (including an example of a redox reaction that does not require a metal). The course guide includes one demonstration of the effect of ligands on color in transition metal complexes.

Tips and Suggestions

1. **Redox processes require both oxidation and reduction.** Remind students that the half-reaction method used to balance redox equations is bookkeeping only, and that the half-reactions happen in tandem, even if the reagents are separated into half-cells as for a galvanic cell.
2. **Any reaction can be written; not all will proceed as written.** The activity series and the series of standard reduction potentials both allow for predictions of which direction a reaction will proceed spontaneously. Many students struggle with predictions of reactivity for anything more complicated than a single displacement; encourage them to write the half-reactions, find them on the table, and identify whether the reduction half-reaction as written in the equation also has the larger of the two reduction potentials on the table.
3. **d-electrons make transition metals unique.** The *d*-electrons in the valence shell of transition metals contribute heavily to the metals' patterns of reactivity, as well as to their tendency to form highly colored compounds. Emphasize that transition metals follow patterns of reactivity analogous to patterns the students have already learned, even though the specifics of transition metal reactivity are beyond the scope of this course.

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. Some metals can dissolve in acid. What do you think happens to the metal during this process?
2. Hydrogen burns readily in air. The primary product of this reaction is water, even though air contains more nitrogen than oxygen. Why do you think this happens?
3. Sodium metal reacts with water to form sodium hydroxide and hydrogen gas. How many protons and electrons does sodium have before the reaction? What about hydrogen? How many protons and electrons do the sodium and hydrogen have after the reaction?
4. Suppose you attempted to react gold, rather than sodium, with water. Would this reaction occur as vigorously?

Before Watching the Video

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

1. What makes a good battery?

2. How do biological redox reactions differ from inorganic redox reactions?
3. Can reactions be forced to run backwards?
4. Do redox reactions always have to involve metals? Why or why not?

Watch the Video

After Watching the Video

The redox reactions discussed in the video include very simple (zinc with copper ions) and extremely complex redox processes (the cycle between Co(I) and Co(III) in reactions with folic acid and homocysteine). Encourage students, especially in more complex systems, to focus on the movement of electrons more than on the exact pattern of bond breakage and formation. Similarly, encourage students to view systems such as vitamin B12 in the context of the atoms immediately surrounding the metal center, rather than trying to envision the entire molecule. Use these additional questions as follow-up, either as a group discussion or as short writing assignments.

1. The Daniell cell (the copper/zinc reaction in the demonstrations) was an early battery system. What are some differences between the Daniell cell and more modern batteries, like the ones used to power the robots in the video?
2. Transition metals appear colored because of the number of *d*-electrons in the metal center and because of the ions or molecules immediately surrounding the metal. Which of these two factors is responsible for the different colors used to analyze samples in the Drennan lab? Explain why this is a useful characterization test.
3. In a beaker, a redox reaction can be forced to run in the non-spontaneous direction by applying an electric current to the components of the reaction. This is, of course, not possible in the human body. What has to happen in the body in order to force a reaction to go in the non-spontaneous direction?

Group Learning Activities

Dissolving a Penny from the Inside Out

Objective

This activity will illustrate the relative redox activity of the metals used in the Daniell Cell, which is highlighted in section 5 of the Unit 11 text.

Pennies minted after 1982 are not pure copper, but rather a zinc core with a thin copper coating. Because zinc is much more reactive than copper, if a penny is scored along the edges to reveal some of the zinc, the zinc can be dissolved in acid to leave the copper behind. The reactions of the zinc core will be measurable after a few minutes, but can also be left out long enough to consume the entire core and leave the copper shell behind. Although this experiment

technically defaces US currency, it is not done for the purpose of defrauding anyone, and therefore is a mainstay demonstration and laboratory experience in general chemistry courses.

List of Materials

- Pennies, some scored on the edge to reveal zinc core, one scored and one unscored per group
- Copper pieces
- Zinc pieces
- 6 M HCl
- Water
- Beakers, 3 per group
- Balance
- Tweezers
- Watch glasses, 3 per group

Procedure

1. Measure and record the masses of pennies, both scored and non-scored.
2. Place both a scored and a non-scored penny into a beaker of acid. Observe and record any changes that occur.
3. While the pennies react, measure the mass of a piece of copper and the mass of a piece of zinc.
4. Place the piece of copper into a beaker of acid and the piece of zinc into a separate beaker of acid. Observe and record any changes that occur.
5. Use tweezers to remove the pennies, copper piece, and zinc piece from the acid solutions. Carefully rinse the pennies and metal pieces with distilled water and dry them on paper towels.
6. Measure the masses of the pennies and metal pieces after the reactions and compare.

Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. Which pieces of metal reacted with acid? Why do you think they reacted in that way? What does this reveal about their relative reactivities? Explain the reactions that were happening.
2. Did both pennies react when placed in acid? If not, which one did? Explain how you determined your answer.
3. Copper is a relatively unreactive metal. Why is it used only for the outside shell of pennies instead of the entire penny?

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. 6 M HCl is highly corrosive, and the reaction with zinc may be very vigorous. Wear eye protection, gloves, and lab coats while performing this experiment. If possible, while metals are dissolving, cover beakers with a watch glass in order to prevent splashes. Use tweezers to remove any pieces of metal from beakers.

Disposal

Check local regulations for proper disposal of HCl.

Solids may be discarded in the trash, though unreacted copper and scored pennies are reusable for multiple lab sessions.

“Lemon Cell Battery”

Adapted from K. R. Muske, et al. in *Journal of Chemistry Education*, April 2007

Objective

A simple galvanic cell can be assembled using strips of different metals and a strong electrolyte solution. These cells can provide power for small electrical components, and may be connected in series to provide even more power. The lemon juice/salt solution serves as both a reagent and a salt bridge.

List of Materials

- Lemon juice (~3 mL per cuvette)
- Measuring spoons or graduated cylinders, one per group
- Beakers, one to three per group
- Table salt
- Stirring rods, one per group
- Funnels, one per group
- Mg ribbon, ~5 cm length per cuvette
- Cu sheet and wire, ~5 cm sheet per cuvette
- UV/vis cuvettes, one per group (more if groups are connecting cells in series)
- Double-ended alligator clip wire leads, two per group
- Single alligator clips, as many as needed to connect cells if groups are making multiple cells
- Small motor, light bulb or other electrically-powered component, or voltmeter; one per group

Procedure

1. Measure 3 mL of lemon juice into a beaker.
2. Stir 0.5-1.5 tsps. table salt into the lemon juice and allow to dissolve. Record the

amount used to make the solution.

3. Fold a magnesium strip over the top of one side and into the cuvette, and a copper strip over the top of the opposite side into the cuvette
4. Attach one end of an alligator clip lead to each of the strips, and attach the other end of each lead to one of the connections to a motor, LED, or other electronic component. [NOTE: if connecting multiple cells, use alligator clips to connect the magnesium side of one cuvette to the copper side of the next in the series, and connect the wire leads at the ends of the entire series.]
5. Position the motor slightly below the cuvette(s), and as far away as the wires will allow in order to prevent sparks and ignition of hydrogen from the salt solution.
6. Pour the lemon juice/salt solution into the cuvette leaving about a centimeter of the top free from the solution.
7. If using a voltmeter, record the voltage for the cell, then prepare solutions of different concentrations, replace the solutions in the cuvettes, and determine the optimum salt concentration for high cell EMF.
8. If using a motor, use the motor to power a model vehicle!

Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. If hydrogen gas is generated at the copper electrode, which electrode was the anode and which was the cathode? Write the half-reactions occurring at each electrode.
2. Based on your answers to question 1, look up the standard reduction potentials for the half-reactions and calculate the standard EMF of the cell. Compare your results with the observed EMFs if you used a voltmeter.
3. Draw a diagram of this setup and explain what is happening.

Hazards

Hydrogen gas is generated at the copper electrode. Make sure area is well ventilated or perform under a fume hood to avoid inhalation of the gas. Fill the cuvettes only when the electrical connections have been made, and keep motors and other electrical components below the cuvettes to avoid contact with the gas. Wear eye protection, gloves, and a lab coat.

Disposal

Lemon juice/salt solutions may be disposed in the sink. Copper and unreacted magnesium may be reused.

In-Class Chemical Demonstrations

Iodine-Starch-Vitamin C Demo

Adapted from “Dr. Campbell’s Favorite Demos.”

Objective

This is a very quick demonstration of a redox reaction that does not involve metals. Starch forms a blue complex when combined with elemental iodine. When iodine reacts with the ascorbic acid in a vitamin C tablet, the iodine is oxidized to iodide ions (and the ascorbic acid is reduced to dehydroascorbic acid). Iodide ions do not form a complex with starch, and so the color is discharged as the iodine is oxidized. The “ingredients” in this demo can be used on a variety of surfaces (including paper!) and the vitamin C tablet can be used to discharge the color either across the entire surface or selectively.

List of Materials

- Tincture of iodine
- Spray starch
- vitamin C tablets from drugstore

Procedure

1. Swab tincture of iodine solution on paper or a watch glass.
2. Before the iodine solution dries fully, spray starch over the solution to generate a dark blue color.
3. Rub a vitamin C tablet in the iodine/starch complex to discharge the color.

Discussion

The following questions can help guide a discussion.

1. Starch forms a dark blue complex in the presence of iodine, but is colorless in the presence of iodide. Identify the species being oxidized and reduced in this demonstration.
2. Why is vitamin C sometimes referred to as an antioxidant?

Hazards

None of the components in the demo are hazardous, but tincture of iodine solution can stain just about anything. The iodine/starch complex is very goopy, so if demonstrating the reaction on your hand, wash your hands immediately afterward.

Disposal

All the reagents in this demo may be disposed in the trash or down the drain.

Cobalt Painting

Adapted from *Chemical Magic* by Leonard A. Ford.

Objective

This demonstration illustrates the change in color of cobalt (II) ions in the presence of different ligands. Both the chloride and the acetate can be dissolved in water to give a pinkish solution from the complex, but will be colored differently when the water is driven off. Solutions of the two cobalt complexes are used to paint on regular paper. When the picture is warmed near a Bunsen burner flame, the dehydration of the compounds turns blue-green and violet where the original painting was pink. If the paper is subsequently sprayed with water, the compounds rehydrate and return to pink.

List of Materials

- Aqueous CoCl_2 (~0.5 M, 10 mL; solids available from Aldrich or other chemical supply company)
- Aqueous $\text{Co}(\text{C}_2\text{H}_3\text{O}_2)_2$ (~0.5 M, 10 mL)
- Stirring rods or paintbrushes, one per solution
- Paper
- Bunsen burner
- Spray bottle of water

Procedure

1. Use the two different solutions to paint a picture on a piece of regular paper. The picture should appear pink.
2. Hold the paper near a Bunsen burner flame and allow the water to be driven off. The two solutes will appear different colors (blue-green for the chloride, violet for the acetate).
3. If desired, spray the paper with water to demonstrate the return of the pink color upon re-hydration of the cobalt ions.

Discussion

The following questions can help guide a discussion.

1. Why do the two different salts produce different colors when heated but the same pink color when wet?
2. Based on your answer to question 1, would other metal salts be usable for this demo? What would be necessary in order for the demo to work?
3. Could you reheat and then re-moisten the same painting used in this demo? Explain.

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 safety glasses and gloves while preparing solutions and while performing the demonstration. Wash hands thoroughly after performing the demo. Use caution when holding paper near the Bunsen burner flame so as not to burn the paper.

Disposal

Check local regulations for proper disposal of chemicals. Cobalt compounds may be recyclable. Check with local hazardous waste removal companies for availability. Heavy metal waste should, whenever possible, be separated from other waste material.

Halogen Redox Chemistry and Solubility**Objective**

This demonstration takes advantage of the different halogens' colors in their respective elemental forms, and their solubility in nonpolar solvents. Single displacement reactions between organic-soluble halogens and water-soluble halide salts are visible by the color changes of the organic layer. The halogens will displace halides according to their relative reactivities: chlorine will displace both bromide and iodide, leading to brown and purple hexane layers, respectively. Bromine will displace iodide but not chloride; iodine will not displace either.

List of Materials

- ~20 mL each of 1.0M aqueous solutions of sodium chloride, sodium bromide and sodium iodide
- ~20 mL hexane
- ~10 mL elemental bromine (or bromine in water)
- ~10 mL chlorine bleach (or chlorine in water)
- ~1 g elemental iodine
- 6 Test tubes or disposable bulb pipets

Procedure

1. Have students make observations of all six solutions before mixing.
2. Mix water and hexane to show that it forms two layers instead of mixing.
3. Dissolve each of the halogens in hexane in two test tubes apiece, and have students note the colors of the hexane when each halogen is dissolved in it.
4. To each of the test tubes, add the solution of one of the sodium salts (containing the anion of one of the two other halogens; e.g., add NaBr to one tube with Cl_2 and NaI to the other).

5. Shake each tube, allow the layers to re-form and have students make note of any color changes.

Discussion

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

1. Is hexane polar or non-polar? How can you tell?
2. For each test tube, explain what is happening.
3. If this demo had included a solution of sodium fluoride, do you think you would have needed to worry about generating fluorine gas? Why or why not?
4. In Unit 4 you learned about electronegativity. Compare the trend in electronegativity with the trend in halogen displacement reactions.

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. Halogens are toxic, albeit minimally so in the quantities described here. Hexane is flammable. Do not perform this experiment near an open flame. Wear proper protective equipment when performing this demonstration.

Disposal

Check local regulations for proper disposal. Bromine, iodine, and hexane require hazardous waste disposal according to applicable federal, state, or local regulations. Hexane should be separated from aqueous reagents as much as possible.

Going Deeper (In-Class Discussion or Reflection)

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

1. Why is it not enough merely to set up two half-reactions with significantly different reduction potentials if you're trying to design a battery?
2. Based on the form of the Nernst equation, why does a reaction coming to equilibrium mean a "dead battery?"
3. Why do you suppose the 18-electron rule is analogous to the octet rule?
4. Look up the structure of the heme group of hemoglobin. Draw the Lewis structures of O_2 , CN^- , and CO , and show how they might be able to bond to the iron in the heme group. How many electrons does this give to the metal?

Before the Next Unit

Learners should read the Unit 11 text if they haven't already done so. They may wish to read one or more of the 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 12 video and/or read the Unit 12 text as an assignment before the next session.

References and Additional Resources

Levi, Primo. *The Periodic Table*. 1st American ed. New York: Schocken Books, 1984.

Dr. Campbell's Favorite Demos. "Vitamin C Cleans Up Iodine Stains on Hand," Accessed August 28, 2013. <http://campbelldemo.blogspot.com/search?q=Vitamin+C+Cleans+Up+Iodine+Stains+on+Hand>

Ford, Leonard Augustine *Chemical Magic*. Revised by E. Winston Grundmeier. 2nd ed. New York: Dover Publications, 1993.

Muske, K. R., C.W. Nigh and R. D. Weinstein, "A Lemon Cell Battery for High-Power Applications." *Journal of Chemical Education* 84: 4 (2007): 635-638.

For Professional Development

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

Further Reading & Reflection Questions

Acar, Burcin and Leman Tarhan. "Effect of Cooperative Learning Strategies on Students' Understanding of Concepts in Electrochemistry." *International Journal of Science and Mathematics Education*. 5 (2007), 349-373. Accessed August 13, 2013.

<http://link.springer.com/content/pdf/10.1007%2Fs10763-006-9046-7.pdf>

Chemical Heritage Foundation, "Electrochemistry: Electrolysis of Solutions – Teacher Activity." Accessed October 24, 2014. <http://www.chemheritage.org/Downloads/Classroom-Activities/Teacher-Guide--Electrolysis.pdf>

1. Were you surprised by the results of this study? Why or why not?
2. Do you agree with the author that if cooperative learning is properly applied, as in this study, the formation of students' misconceptions can be prevented? Why or why not?
3. Does this paper influence how you will approach teaching redox chemistry? If so, how? If not, why not?

Barke, Hans-Diester. "Two Ideas of the Redox Reaction: Misconceptions and Their Challenge in Chemistry Education." *African Journal of Chemical Education*. 2:2 (2012), 32-50. Accessed August 26, 2013. <http://www.ajol.info/index.php/ajce/article/view/82445/72600>

1. Have you encountered students holding any of the same misconceptions as those discussed in this paper? Are there some that are more prevalent than others? Were you surprised by any of the misconceptions? If so, which ones and why? If not, why not?
2. Do you agree with the author's suggestions for teaching redox chemistry? If so, which ones, and why? If not, why not?
3. Does this paper influence how you will approach teaching redox chemistry? 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 in your classroom. Where would it fit into your curriculum or standards? Would you change the demonstration or activity in any way? How would you assess student learning?

UNIT 12

Kinetics and Nuclear Chemistry

Rates of Reaction

Unit Overview

In this unit, students will learn about chemical kinetics and nuclear chemistry; both these topics share the concept of half-life. The text and video introduce the basics of collision theory, and show how increasing the number of productive collisions speeds up reactions. Understanding chemical kinetics is vital to maximizing the efficiency of industrial chemical production and to understanding the regulation of biochemical reactions.

In the latter half of this unit, students will learn about the early groundbreaking research in radiation and nuclear chemistry. They should understand the basic types of radiation and be able to explain the nuclear reactions that produce them. After learning about fission and fusion, students should be able to explain how nuclear weapons and nuclear power plants operate. The material in this unit will give them a deeper awareness of important issues such as nuclear waste disposal and nuclear weapons proliferation.

Learning Objectives and Applicable Standards

Participants should be able to:

1. Describe the different factors that influence the rate of a chemical reaction, referencing the basics of collision theory.
2. Explain what a rate law is and what kind of data one needs to determine what the rate law for a reaction is.
3. Read and understand a potential energy diagram, define activation energy, and explain how higher temperatures and catalysts speed up a reaction.
4. Explain reaction mechanisms and rate-limiting steps.
5. Trace the history of the discovery of radioactivity.
6. Describe the different types of radioactive decay and write balanced nuclear equations.
7. Explain the concept of half-life and describe the process of radiometric dating.
8. Relate the history of the discovery of nuclear fission and fusion, and explain how these reactions produce nuclear explosions and nuclear power.

Key Concepts and People

1. **Collision Theory:** Molecules must collide in order to react. Collision theory underlies the study of chemical kinetics. To react, molecules must collide in certain ways. Reactions proceed faster when productive collisions occur more often.
2. **Potential Energy/Reaction Coordinate Diagrams:** These diagrams show whether a reaction is endothermic or exothermic and the activation energy of the reaction.
3. **Reaction Mechanisms:** Reactions usually happen in a series of simpler reactions. This series of steps is called the “reaction mechanism,” and the slowest step determines the rate of the overall reaction.
4. **Catalysts:** Catalysts speed up chemical reactions without themselves being changed by the chemical reaction. They increase the reaction rate by providing an alternate mechanism with lower activation energy.
5. **The Discovery of Radioactivity:** Wilhelm Röntgen, Henri Becquerel, and Marie and Pierre Curie pioneered research in radioactivity. Together these scientists discovered radioactivity and described its properties. Different modes of decay produce particles with different energies and properties.
6. **Radiation:** Radioactive material decays in a predictable way. Radioactive substances have half-lives. Because the decay is predictable, scientists can use radiometric dating to determine the age of ancient material.
7. **Nuclear Fission:** Lisa Meitner and Otto Hahn discovered nuclear fission. The discovery of fission leads to the development of nuclear chain reactions, nuclear bombs, and nuclear power plants.

Video

From an instantaneous explosion to the slow rusting of iron, the rates at which different chemical reactions proceed can vary tremendously depending on several factors, including temperature and concentration. Sometimes, like with the rotting of food, chemists want to slow down reaction rates. But often, the goal is to speed them up—and one way to do this is to use a catalyst. In this video, we will learn about catalysts and how using them can lead to cheaper, more effective, and more sustainable drug production processes. We will also discover how the rates of some reactions, like nuclear decay, are unchangeable, and how scientists take advantage of this, using PET scans to reveal the presence of disease.

VIDEO CONTENT

Host Introduction “Chemical Reactions”

Dr. Wilton Virgo, assistant professor at Wellesley College, demonstrates the basics of chemical reactions. He rides bumper cars at an amusement park to explain that chemical reactions happen when molecules collide, and that increasing the concentrations will increase the chances of a reaction happening. He then rides a rollercoaster to discuss activation energy—the energy needed to make a given reaction “go.”

Laboratory Demonstration “Activation Energy”

Harvard University Lecture Demonstrator Daniel Rosenberg demonstrates activation energy by mixing hydrogen and chlorine gases in a test tube stopped by a cork. He keeps a dark cloth over the tube because light can provide the activation energy needed to make the reaction happen. By shining lights with different energy levels on the tube, he discovers what it will take to make the reaction happen and to shoot the cork across the room.

Host Science Explanation “Catalysts”

Dr. Wilton Virgo is on a train in the amusement park and talks about how the train is like a catalyst because it is providing an alternative pathway that will take him from one side of the park to another in a shorter amount of time. Dr. Virgo expands on the analogy by explaining that at each stop, the train empties and is available to take more passengers back to the other side of the park. A catalyst acts in the same manner – it changes during a reaction, but once the reaction is over, the catalyst returns to its original state and can be used in subsequent reactions.

Laboratory Demonstration “Elephant’s Toothpaste”

Daniel Rosenberg demonstrates a catalyst in action using potassium iodide to catalyze the decomposition of hydrogen peroxide. He then does the same reaction but mixes in dish soap to create some visually interesting results.

Current Chemistry Research “Molecular Architects”

At Boston College, Professor Amir Hoveyda and his team have developed a set of catalysts that may provide a new blueprint for how we approach catalyst development for years to come. This segment provides an interesting definition of what a catalyst is by discussing what they are looking for and what they have discovered.

History of Chemistry “Nuclear Decay”

Wilton Virgo explains that the rates of some chemical reactions are unchangeable: the most common example of reactions that cannot be slowed down or speeded up are nuclear or

radioactive decay. Scientist Marie Curie coined the term “radioactivity” to describe the steady emission of rays she observed in uranium and thorium. Along with her husband Pierre, Curie accurately proposed what was a revolutionary idea for its time—that this behavior was the result of something happening inside the atom.

Real World Application “PET SCANS”

At Massachusetts General Hospital, radioactivity and half-lives play a crucial role in performing PET (Positron Emission Tomography) scans. PET scans are a crucial tool in detecting diseases at an early stage and can also be used by doctors to target their treatment in more specific and effective ways.

Unit Text

Content Overview

The first half of this unit introduces chemical kinetics. The basics of collision theory and rate laws are discussed first, followed by the effects of concentration, temperature, and catalysts on reaction rate. The second half of the unit covers nuclear chemistry and begins with a brief history of the discovery of radioactivity. The next sections describe basic types of radiation and nuclear equations, followed by nuclear stability, the strong force, and half-lives. The final sections discuss fission, fusion, nuclear weapons, and nuclear power plants.

Sidebar Content

1. **Rate Laws:** This sidebar shows how an equation (a rate law) represents the rate of a chemical reaction in terms of reactant concentrations.
2. **The Arrhenius Equation:** Collision frequency, temperature, and activation energy all appear in this equation for the rate constant.
3. **Nuclear Stability:** There are four fundamental forces of nature: gravity, electromagnetism, the strong force, and the weak force. The strong force holds protons together in a nucleus, while the repelling forces of the positive charges are constantly trying to force protons apart from one another. However, the strong force holds the protons together and binds neutrons to protons and other neutrons.
4. **Decay Chains and Radon:** This explains nuclear decay chains and how uranium-238 leads to radon gas in houses.

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.

Control a Haber-Bosch Ammonia Plant Interactive

In this simulation, students try to maximize the efficiency of an ammonia plant. Students will be able to watch the interaction of H_2 and N_2 molecules and see the effect of concentration and temperature on reaction rate. They will also be able to choose from three catalysts to increase the plant's productivity. Please note that a lesson plan and student worksheet are available online with this interactive.

During the Session

Before Facilitating this Unit

The kinetics activity provided below gives students a good introduction to some of the factors that influence reaction rates. The cornstarch demo reinforces the basics of collision theory in a dramatic way, and the hydrogen peroxide demo goes hand in hand with the text's discussion of catalysis.

The second half of the unit text covers nuclear chemistry, and because hands-on activities with radioactive material are not safe, the teacher will need to rely on simulations such as the half-life activity below. Many simulations for radioactive decay, radiometric dating, and fission chain reactions are also available online. (See References and Additional Resources.)

Tips and Suggestions

1. **The rate law is not derived from the net chemical equation.** Some students assume that all reactants in the equation can affect the rate, or that the coefficients in the equation determine the rate law. Rate laws can only be determined experimentally.
2. **Individual nuclei decay at random.** Many students at this level have had some exposure to the concept of half-life, and mistakenly believe that small numbers of nuclei will behave like large numbers. Emphasize that the decay of each individual nucleus is completely unpredictable, and only large numbers of nuclei behave predictably.

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.

UNIT 12 Kinetics and Nuclear Chemistry

1. What chemical reactions happen very slowly? Very quickly? Why do you think the rates are different?
2. Why might someone want to speed up a chemical reaction? Why might someone want to slow down a chemical reaction?
3. How do you think changing the rate of a reaction could be changed? What factors might influence reaction rates?
4. What do you think of when you hear the word “radioactivity”?
5. What are some uses of radioactive materials and radiation?
6. How is it possible that radiation can be both a cause of and a treatment for cancer?

Before Watching the Video

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

1. What is activation energy?
2. What role does a catalyst play in the “elephant’s toothpaste” reaction?
3. What is radioactivity?
4. What does the term half-life mean?

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 catalyst? How do catalysts work?
2. Provide some examples of how catalysts can be useful.
3. What role does radioactivity play in PET scans at Massachusetts General Hospital?

Group Learning Activities

Kinetics

Objective

Students will control the rate of the reaction between Alka-Seltzer tablets and water to make the top of a film canister pop off after exactly 30 seconds.

Materials

Each group of students will require:

- A film canister
- Some Alka-Seltzer tablets
- Water (cold and hot)

Procedure

When Alka-Seltzer dissolves in water, it produces CO_2 gas. You will perform this reaction inside a closed film canister. Your task is to control the rate of this reaction so that the top of the canister pops off after exactly 30 seconds. Try adjusting the amount of Alka-Seltzer, the size of the Alka-Seltzer pieces, and the temperature of the water.

Discussion Questions

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. What effect did the amount of Alka-Seltzer have on the reaction rate? Why?
2. What effect did the size of the Alka-Seltzer pieces have on the reaction rate? Why?
3. What effect did the temperature of the water have on the reaction rate? Why?

Hazards

Goggles should be worn at all times to protect students from lids and bubbling Alka-Seltzer.

Disposal

Water and Alka-Seltzer may be poured down drain. Unused tablets may be stored for later use or disposed in trash.

Half-Life**Objective**

Students will simulate nuclear decay and graph the data in order to gain an understanding of half-life.

Materials

- One coin per student

Procedure

Start with the entire class standing up; each standing student represents a radioactive nucleus. Have all students flip their coins once. If a student gets heads, he/she remains standing (the nucleus doesn't decay); if a student gets tails, the student sits (the nucleus decays). Record the number of standing students in a table on the board. Have the remaining students flip their coins one more time. Those that get tails must sit. Record the number of standing students. Repeat this cycle until all students are seated. Each round of coin flipping represents one half-life. Graph

UNIT 12 Kinetics and Nuclear Chemistry

the data from the table. To get a smoother graph, you will probably need to repeat the procedure a few times and take an average of the data.

Discussion Questions

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. The data from this experiment probably did not create a perfectly smooth graph, where exactly half the sample decays after each half-life. If you started with 10 grams of a real radioactive element, however, the graph would look almost perfectly smooth. Why is there a difference?
2. What determined the half-life of the element in this simulation?
3. What determines the half-life of actual radioactive elements?
4. After each half-life, the amount of a radioactive element decreases by half. Will the element ever decay completely, if the amount just keeps getting cut in half?

Hazards

There is no increased risk of harm in this activity.

Disposal

There are no special disposal considerations.

In-Class Chemical Demonstrations

Reaction Rate and Surface Area

Objective

This demonstration shows that surface area can affect the rate of a reaction.

List of Materials

- Cornstarch
- Bunsen burner
- Drinking straw

Procedure

1. Show that a small pile of cornstarch will not catch fire when the surface is burned with a Bunsen burner.
2. Place some cornstarch in a drinking straw. From the side and from below, blow the cornstarch into the Bunsen burner flame; it will burn rapidly in a fireball.

Discussion

These questions can help guide a discussion.

1. What is the difference between the pile of cornstarch and when the cornstarch is blown?
2. Why does the blown cornstarch catch on fire, but the pile of cornstarch does not?

Hazards

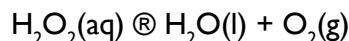
Make sure you blow the cornstarch away from people and flammable materials.

Disposal

There are no special disposal considerations.

Catalysis**Objective**

This demonstration shows that yeast serves as a catalyst to the decomposition of hydrogen peroxide. The hydrogen peroxide decomposes slowly in the following reaction:

**List of Materials**

- 3% hydrogen peroxide solution
- Large beaker
- Dry yeast

Procedure

1. Pour about 50 mL of 3% hydrogen peroxide solution into a large beaker.
2. Sprinkle some dry yeast into the beaker; the catalase enzyme in the yeast speeds up the decomposition, and the mixture will rapidly foam up.
3. Light a wooden splint and blow it out, leaving the end smoldering. Put the burning end into the gas inside the beaker to show that the splint burns faster in the oxygen produced.

Discussion

These questions can help guide a discussion.

1. Before the yeast is added, is any reaction occurring?
2. Why does the yeast cause the reaction to speed up?
3. Hydrogen peroxide is used in first aid to clean cuts and scrapes. When it comes in contact with damaged skin, it foams up just like the demonstration. 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. Wear proper protective gear at all times: chemical splash goggles, chemical-resistant apron, and gloves.

Hydrogen peroxide: Many substances will cause hydrogen peroxide to decompose into water and oxygen gas. It deserves the science teacher's special handling and storage attention. Substance is severely corrosive to skin, eyes, and respiratory tract; very strong oxidant. Dangerous fire and explosion risk. Do not heat this substance. While a 3% solution of hydrogen peroxide is very weak, it is still an oxidizer and a skin and eye irritant.

Disposal

Check local regulations for proper disposal of hydrogen peroxide.

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. Find real-life examples of reaction rates being influenced by concentration, surface area, temperature, and catalysts.
2. Radiocarbon dating is not accurate when applied to objects that are very young or very old. Explain.
3. A fission chain reaction is fastest when the fissile material is in a spherical shape. If the material is pounded into a flat sheet, the chain reaction will slow down and possibly stop. Explain.
4. After each half-life, the amount of the radioactive substance decreases by half. Will the substance ever completely disappear?
5. Some of the scientists working on the Manhattan Project felt a lot of guilt for having worked on creating a weapon of mass destruction. Others felt proud of their role in bringing World War II to an end. Do you think that working on the Manhattan Project was morally right?
6. Why does refrigeration of food prevent spoilage?
7. Thermodynamics actually predicts that diamonds are less stable than graphite, and therefore should turn into graphite readily. Does this happen as predicted? Why or why not?
8. Suppose you have two gas phase molecules, A and B. They react to form product AB by colliding into one another. Based on your understanding of gas behavior, think about a

way in which you could improve the probability of molecule A colliding with molecule B in a closed reaction container.

9. Suppose you have two molecules, C and D. Like molecules A and B in example 3, they can form a product, CD, by colliding into one another. Molecule C, however, is a large, complicated molecule; the product CD forms only when molecule D hits one specific spot on molecule C. How can you improve the probability of this reaction occurring? How is this the same or different from your answers to the previous question?

Before the Next Unit

Students should read the Unit 12 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 13 video and/or read the Unit 13 text as an assignment before the next session.

References and Additional Resources

"Marie Curie and The Science of Radioactivity." The American Institute of Physics -- Physics Publications and Resources. <http://www.aip.org/history/curie/> (accessed August 10, 2012).

The Day After Trinity. DVD. Directed by Jon Else. Chatsworth, CA: Image Entertainment, 2002.

"NOVA | Island of Stability." PBS: Public Broadcasting Service. <http://www.pbs.org/wgbh/nova/physics/stability-elements.html> (accessed August 10, 2012).

Phet Interactive Simulations. "Alpha Decay." University of Colorado at Boulder. Accessed August 29, 2013. <http://phet.colorado.edu/en/simulation/alpha-decay>

Phet Interactive Simulations. "Beta Decay." University of Colorado at Boulder. Accessed August 29, 2013. <http://phet.colorado.edu/en/simulation/beta-decay>

Phet Interactive Simulations. "Nuclear Fission." University of Colorado at Boulder. Accessed August 29, 2013. <http://phet.colorado.edu/en/simulation/nuclear-fission>

Phet Interactive Simulations. "Radioactive Dating Game." University of Colorado at Boulder. Accessed August 29, 2013. <http://phet.colorado.edu/en/simulation/radioactive-dating-game>

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

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following papers and answer the corresponding reflection questions. Participants should then complete the accompanying professional development assignments.

Further Reading and Reflection Questions

Kolomuc, Ali and Sehar Tekin. "Chemistry Teachers' Misconceptions Concerning Concept of Chemical Reaction Rate." *Eurasian Journal of Physics and Chemistry Education*. 3:2 (2011): 84-101. Accessed August 29, 2013. <http://eurasianjournals.com/index.php/ejpce/article/viewFile/686/394>

1. Were you surprised by the results of this study? Why or why not?
2. Have you held any of the same misconceptions as the teachers in this study? If so, which ones? How did you overcome these misconceptions? Have any of your students held the same misconceptions uncovered by this study? Which ones? How have you addressed those in the past?
3. What are some strategies you could employ to help address the misconceptions uncovered by this study?

Viola, V. E. "Teaching Nuclear Science: A Cosmological Approach." *Journal of Chemical Education*. 71: 10 (1994): 840-844. Accessed August 29, 2013. http://depa.fquim.unam.mx/amyd/archivero/nucleosin2_20646.pdf

1. What are some of the benefits and challenges of presenting nuclear chemistry in the context of cosmology for high school audiences?
2. How could you use the ideas presented in this paper and adapt them to teach nuclear chemistry at a high school level?

Professional Development Assignments

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

UNIT 13

Modern Materials and the Solid State

Unit Overview

This unit explores advances in modern materials in the context of their underlying chemistry. Many common materials represent significant technological advances but are still based on basic chemical principles. Therefore, an understanding of chemistry is an excellent entry point not only to understand how a material behaves or a device works, but also to understand the ways in which cutting-edge advances are made.

Learning Objectives and Applicable Standards

Participants will be able to:

1. Differentiate between types of solid structures in terms of atomic interactions and order (or lack thereof) in the structure.
2. Recognize the conditions that influence phase changes in materials.
3. Identify different types of alloys in terms of their constituents and behavior.
4. Recognize and characterize the different allotropes of carbon.
5. Identify various naturally occurring and synthetically created polymers.

Key Concepts and People

1. **Modern Materials:** The chemistry of matter in the solid state and the applications of solid materials are areas of some of the most cutting-edge research in chemistry and engineering.
2. **What Is a Solid?:** Both the identity of the elements and the arrangement of atoms in a solid affect the properties of the solid. At the most basic level, solids are classified as either amorphous or crystalline, depending on whether or not they exhibit long-range order at the atomic level.
3. **Ionic and Covalent-Network Solids:** Coulombic forces hold together ionic solids while covalent bonds hold together covalent solids.

4. **Molecular Solids:** Molecular solids, such as ice, are held together by intermolecular interactions of the molecules themselves. These interactions are weaker than the ionic or covalent interactions in ionic and covalent solids, and the physical characteristics of the solids reflect this.
5. **Metallic Solids:** Metals have readily accessible valence electrons, and in metallic solids the electrons are shared readily between atoms. Metals can be used in their pure form but can also combine to form alloys, whose properties are influenced by the identity and proportion of the components.
6. **Phase Changes in Solids:** Temperature and pressure influence the state of matter in which a material will be found. Phase diagrams summarize the behavior of materials under different temperature and pressure conditions.
7. **Natural and Synthetic Polymers:** Carbon is not only the basis of life but also the basis of polymer chemistry. Both naturally-occurring and synthetically-created polymers are tremendously important materials in modern life.
8. **Alloys:** Metals form alloys by combining in various proportions, and occasionally with trace amounts of non-metals.

Video

While chemical reactions in gases and liquids are essential to the understanding of chemistry, the chemistry of solid-state materials characterizes most of the interactions we have with matter on a daily basis. Chemists take advantage of the complexity of solids to engineer new materials, including nanoparticles, polymers, and advanced metal alloys. These new materials have many potential applications in sensors, advanced drug delivery systems, and space exploration. Today, modern materials are following a heritage—one that can be traced back to earlier civilizations—in which the properties of solids are manipulated to advance human needs.

VIDEO CONTENT

Host Science Introduction

Materials scientist and science writer Dr. Ainissa Ramírez, introduces the unique properties of solid-state materials through the example of a “memory metal,” an alloy that can be deformed at room temperature, but then returns to its original shape when heated. This segment emphasizes a central idea of the unit: The observable properties of bulk solid materials are the result of the interactions of their molecular components at the microscopic level.

Real World Example

“Atomic Arrangements”

Dr. Raquel Alonso Pérez, curator of the Mineralogical and Geological Museum at Harvard University, uses examples from the minerals in her collection to point out the different types

of bonding that hold solids together. Comparing the different forms of carbon - diamond and graphite - which differ greatly in their properties because of differences in their internal bonding arrangements, she reinforces that the external properties of solids result from these atomic and molecular structures.

Current Chemistry Research

“Carbon Nanotube Sensors”

Joseph Azzarelli’s research in the Swager group at MIT focuses on developing new sensors for agricultural, food, industrial, and environmental applications. He has found ways to use carbon nanotubes, yet another molecular form of carbon, as the heart of a sensor that can measure ethylene gas, a plant hormone that controls the ripening of fruit.

History of Chemistry

“Early Plastic”

In the 1800s, Leominster, Massachusetts became known as the “Comb City,” because factories there specialized in the manufacture of combs and other items made from natural materials, such as bone and horn. Paul Benoit, a curator at the Leominster Historical Society, describes how Leominster switched to man-made materials shortly after the turn of the 20th century. One of the first plastics was celluloid, a polymer that could be molded and cooled to make objects of many different shapes. Leominster made important contributions to the rise of today’s multi-billion dollar plastics industry.

Laboratory Demonstration

“Making Polymers”

One of the best-known plastics is nylon, which, like all polymers, is formed from linking together smaller molecules, or monomers, into large molecules. Harvard University Lecture Demonstrator Daniel Rosenberg shows how to react two monomers, diacid chloride and hexamethylenediamine, and come away with a long strand of nylon.

Current Chemistry Research

“Polymer Research in the Pharmaceutical Industry”

Dr. Trevor Castor, Chief Scientist at Aphios Corporation, is enveloping insulin inside biodegradable polymers. The polymer protects the insulin, a protein, from the harsh chemical environment in the digestive system. There, the encapsulated nanoparticles are so small that they are absorbed directly into the bloodstream. Once inside the blood, the polymer degrades, releasing the insulin. This technology has the potential to deliver insulin without injections, thereby improving quality of life for millions of people with diabetes.

Current Chemistry Research

“High Temperature Alloys”

Engineers at the Smithsonian Astrophysical Observatory are building components for a new spacecraft, one that will travel close enough to the Sun to take actual samples of the Sun’s corona. During the mission, the spacecraft will experience solar radiation that is 500 times more intense than it is on Earth, so the materials used must be able to withstand extremely high

temperatures. Smithsonian Astrophysicist Dr. Anthony Case and technology and product development manager Stephen McCrossan of Plansee USA, describe how alloys of molybdenum with tungsten and zirconium have the necessary characteristics.

Unit Text

Content Overview

The unit begins with a brief review of chemical bonding to introduce chemistry as the foundation for the performance of materials in the solid state. The development of solid materials, and the behavior of naturally occurring solid materials, depends not only on the chemistry but also the structure of the solid itself. Different types of solids exhibit different types of atomic interactions and different microscale geometry, and these both influence the behavior of the solids. At the very basic level, solids may be either amorphous or crystalline. Within those broad definitions, however, particularly for crystalline solids, a vast array of geometric arrangements and chemical interactions are possible.

Alloys such as bronze and steel are among the earliest examples of the engineering of solids, but the development of new alloys continues, and remains a significant contributor to industry. Carbon is not only the centerpiece of organic chemistry but also polymer chemistry, and is therefore another centerpiece of industrial and commercial applications. Both naturally occurring and synthetically created polymers are relevant to such applications. The production and the recycling of these products are major contributors to industry.

Sidebar Content

1. **The Chemistry of Steelmaking:** The chemical composition of steel dramatically affects its properties. Beginning in the mid-19th century with the Bessemer converter, production methods for steel have become significantly more controlled allowing for more variations on the composition of the alloy.
2. **Recycling Plastics:** Plastics can be recycled, but because the polymers that make them come in a wide range of chemical structures, different types must be processed differently.

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. It is interesting to note that even before the era of modern chemistry, advances in materials science helped shape civilization (e.g., Bronze Age, Iron Age).

Control a Haber-Bosch Ammonia Plant Interactive

Although the reaction in this interactive is not the manufacture of any of the materials discussed in the unit, the principles behind the successful operation of a chemical plant are similar to those for the successful manufacture of various modern materials. Please note that a lesson plan and student worksheet are available online with this interactive.

During the Session

Before Facilitating this Unit

Generally, we recommend watching the videos first before reading the written units. The sections of the video for this unit correlate very closely with the order of the topics in the text, and so the text may be brought in one segment at a time to address each topic in more detail. Encourage students to take the examples given in the text and video and think about as many other real-world examples as they can of the same sorts of materials. This unit can help bridge the gap between laboratory chemistry and real world applications, and provides opportunities for fun hands-on learning.

Tips and Suggestions

1. Materials science may be one of the easiest topics in which to emphasize the interdisciplinary nature of scientific research. Many of the principles described in this unit, however, are based on science beyond the scope of this course.
2. Understanding the chemistry of the components of a solid is a critical first step in understanding how the solid itself will behave.
3. Designing new materials is a challenge of chemistry (to know the properties of the raw materials), engineering (to determine the optimal combination of properties necessary to solve a particular problem), and occasionally physics (to understand on the microscopic level how the structure of a material will influence its function).

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. Compare the geometry around a carbon atom with four single bonds; one double and two single bonds; one triple and one single bond; two double bonds. How do these shapes influence the structure of the molecule? How do you think they influence the behavior and appearance of a sample of the compound?
2. Suppose you have a sample of iron, a sample of rock salt, a sample of ice, and a sample of rubber. Which will melt the most easily? Which will break the most easily? Why?

Before Watching the Video

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

1. How are pure metals different from ionic compounds? Do they look and behave the same?
2. Think about different types of plastic. Are they all equally rigid? Do they all qualify as being solid?
3. What is steel? Is it made up of one element or several?

Watch the Video

After Watching the Video

The various solids described in the video encompass all of the major types of bonding as well as many of the major classes of modern solid-state materials. Encourage students to draw connections between the current-research examples (plastics, alloys) with simpler, already-familiar examples (e.g., polyester, steel). Use these additional questions as follow-up, either as a group discussion or as short writing assignments.

1. In the video, you saw that different types of bonding can hold solids together in their most favorable structures, and that this type of bonding in turn influences the properties of the solid. Can you think of examples, besides the ones given, of different types of solids whose behavior reflects the kind of forces holding the solid together?
2. What sorts of forces hold samples of polymers together? What about carbon nanotubes? What about alloys? How are these different?
3. How are unit cells similar to molecular formulas? How are they different?

Group Learning Activities

“[Marshm]allotropes”

Objective

In this exercise, students use marshmallows and toothpicks to build models of the two main allotropes of carbon—diamond and graphite—in order to compare their structures. Students may also want to build a buckyball, a spherical allotrope of carbon with a molecular formula of C_{60} . You may wish to have each group responsible for the model of a single allotrope. If multiple groups build the same model, they may combine their models once they are built.

List of materials

- Marshmallows or other soft candies, 50 per group (at least 60 if building a buckyball)
- Toothpicks, 50 per group

- Pictures of the solid-state structures of graphite, diamond, and buckyball for reference (may be at the front of the room)

Procedure

1. Examine the structures of diamond, graphite, and a buckyball.
2. Using marshmallows as carbon atoms and toothpicks to bond between them, connect the marshmallows to mimic the structure of whichever allotrope you are building. For diamond, you may need to have one person hold the bond angles in place while others build up around the initial tetrahedra. For buckyballs, you may wish to build several hexagons and then connect them. For graphite, make several sheets so that the sheets can be laid on top of one another.
3. If multiple groups have made models of the same allotrope, combine the models once each group has exhausted their supply of materials.
4. Compare the models of the different allotropes for how easily they deform under pressure and under side-to-side shear.

Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. Compare the macroscopic appearances of graphite and diamond. How do you think their atomic-level structures affect their observed macroscopic structure and properties?
2. Graphite is an excellent electrical conductor, while diamond is an insulator. What do you think is responsible for this? (Hint: what does it mean for something to conduct electricity?)

Hazards

There is no increased risk of harm to do this activity.

Disposal

Marshmallows and toothpicks may be disposed of in the trash.

Advances in Packing Peanuts

Adapted from Penn State's Materials Research Institute's "Packing Peanuts" activity.

Objective

Water-soluble packing peanuts made from cornstarch are now available as an alternative to the traditional Styrofoam. In this activity, students will compare the peanuts made from a natural polymer to those made from the synthetic polymers.

List of materials

- Styrofoam packing peanuts, 3-5 cups per group
- Cornstarch packing peanuts, 3-5 cups per group
- Large (~1 L) beakers or clear-sided boxes
- Eggs, 2 per group
- Meter stick, one per group
- Weight plates (or weights plus a watch glass to place directly onto the peanuts)

Procedure

1. Fill one beaker or box with Styrofoam packing peanuts, the other with cornstarch peanuts. Make sure that the two containers are filled as close to the same level as possible.
2. Place weight plates onto the top of the pile of Styrofoam peanuts until the peanuts are compressed by approximately 25% from their original height in the beaker. Make a note of how much weight was required to do this.
3. Repeat step 2 for the cornstarch peanuts, recording the weight required.
4. Remove the weights, and fill additional peanuts into the beakers (to equivalent heights) if they appear permanently compressed.
5. Drop an egg from 25 cm above the surface of the Styrofoam peanuts. Does it break?
6. Drop an egg from progressively higher above the Styrofoam peanuts. Note when the egg finally breaks (or when you run out of measurable height on the meter stick!).
7. Repeat steps 5 and 6 with an egg over the cornstarch peanuts.
8. Pour water into each beaker. Allow the beakers to stand for several minutes and note the appearance of the packing peanuts.

Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. How are the two types of packing peanuts similar? How are they different?
2. Which of the two materials provided the best resistance to compression under loading of weight?
3. Which of the two materials provided the best protection for dropping an egg?
4. What are some of the advantages and disadvantages of these types of packing peanuts? When might you want to use one over the other?

Hazards

There is no increased risk of harm to do this activity.

Disposal

Cornstarch packing peanuts may be rinsed down the drain. Styrofoam peanuts may be washed and reused, or disposed of in the trash.

In-Class Chemical Demonstrations

Close-Packed Candy

Adapted from Penn State's Materials Science Research Institute's "Amorphous Solids" activity.

Objective

This demonstration will illustrate the two types of close-packing, hexagonal and cubic. The two structures have the same packing efficiencies, but differ in the pattern of the spheres after the first two layers. While the packing efficiency can be proven mathematically, comparing the numbers of spheres used to build four layers of each type of structure in the same size container, and then comparing the amount of sand, sugar, or water necessary just to fill the spaces between those spheres can also illustrate packing efficiency.

List of materials

- Spherical candies or marbles, 50 each of three different colors (or 50 each of one color and 150 each of two colors if demonstrating both hexagonal and cubic close-packing simultaneously)
- Clear box (two if demonstrating two packing structures simultaneously)
- Sand or sugar, enough to cover 4 layers of candy/marbles (depending on size of box and candies). If using marbles for crystal structures, use water to illustrate packing efficiency.
- Measuring cup

Procedure

1. Using one color of candy, fill a single layer on the bottom of the box. Gently shake the box to ensure that the layer is fully packed.
2. Using a second color of candy, build a second layer on top of the first in the interstitial sites. Ensure that the layer is fully packed.
3. For hexagonal close-packing, return to the first color of candy and build a third layer like the first two, with the spheres positioned in the interstitial sites directly above the spheres in the first layer. Build a fourth layer in the second color, with the spheres directly above the spheres in the second layer.
4. For cubic close-packing, build a third layer with the third color of candy, with the spheres

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in the interstitial sites offset from the first two layers. Build a fourth layer in the first color, with the spheres directly above the spheres in the first layer.

5. Have the students examine the differences between the two structures from both the side and the top, making note of the space between the spheres.
6. Pour sand, sugar, or water into the boxes and carefully spread around just enough to cover the layers and fill the gaps (do not shake the boxes). Once the space is completely filled, remove the spheres and measure the amount of sugar/sand/water used. This should be the same amount, or close to it, for the two structures.
7. Count the total number of spheres used to build each of the two models. These numbers should also be similar.

Discussion

The following questions can help students reflect on this exercise. If possible, have students record their responses to the questions in a journal and spend some time discussing as a class.

1. How do the two structures differ when viewed from the side? What do they have in common? Sketch what you see.
2. How do the two structures differ when viewed from the top? What do they have in common? Sketch what you see.
3. Sand and sugar were used to illustrate packing efficiency because they are significantly smaller particles than the candies/marbles used to build the structures. Could you have placed larger particles in between the spheres? Would this have given an accurate view of the amount of space between the spheres? Explain.
4. What role does packing efficiency play in the property of a given material?

Hazards

There is no increased risk of harm to do this activity.

Disposal

Marbles, sand, and any other inedible material may be reused. Candy and sugar may be disposed of in the trash.

Instant Polymers: Making Nylon, Making Rubber

Adapted from *Chemical Magic* by Leonard Augustine Ford.

Objective

This demonstration combines two classic demonstrations of rapid polymerization reactions, with two very different results. The reaction to generate nylon produces a filament that can be drawn out, while the coagulation to generate rubber produces a blob.

List of materials

- 250 mL beakers, 5
- Stirring rods, 2 (3 if spooling the nylon)
- Tweezers
- 10% acetic acid solution in water, ~50 mL
- Synthetic latex, ~50 mL
- 6 g sebacoyl chloride in 70 mL heptane (available individually from Aldrich or other chemical suppliers; prepare solution in advance of demo)
- 3 g 1,6-diaminohexane in 70 mL water (diamine available from Aldrich or other chemical suppliers; prepare solution in advance of demo)
- Methanol, ~200 mL

Procedure

1. Have the students observe the appearances of beakers filled with latex, acetic acid, sebacoyl chloride, and diaminohexane solutions.
2. Pour the latex solution into the acetic acid and stir vigorously with a stirring rod. The latex should form a ball.
3. Remove the ball from the beaker, squeeze out the liquid with your hands, and bounce the ball. Allow students to pass the ball around and inspect it.
4. Tilt the beaker containing the diaminohexane solution and carefully pour the solution of sebacoyl chloride down the lower wall to form two layers.
5. Dip a glass rod into the combined solutions at the interface of the layers, and carefully pull the film that forms at the interface. Pick this film up into the tweezers and pull the strand out from the beaker.
6. If desired, spool the resulting strand of nylon around two glass rods.
7. Wash the nylon with methanol to remove the acid generated in the reaction.

Discussion

These questions can help guide a discussion.

1. If the two solutions used to form the nylon had not been mixed together carefully, what do you think would have happened? (Hint: what happened with the other demo?)
2. The formation of the rubber is due to coagulation of the latex, while the formation of the nylon is due to a reaction between the two chemicals. Do you think that the rubber could have been formed into a filament the way the nylon was if the latex and acid solutions had been combined more carefully in the same the way the two solutions were for the nylon?

3. Look up the structures of sebacoyl chloride and 1,6-diaminohexane. If the reaction between these to produce nylon also produces HCl, how do you suppose the reaction proceeds? Draw a structure for the repeating unit of this polymer.

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 safety goggles, a lab coat, and gloves for both portions of this experiment. The experiments may be performed in a fume hood or on a desktop, but should be performed in a well-ventilated area if not in a hood. Methanol is toxic and flammable. Caution anyone who might have a latex allergy against handling the rubber ball. Latex will stain clothing permanently.

Disposal

Check local regulations for the safe disposal of methanol and any un-reacted organic acids and bases. An excess of acetic acid may be used to coagulate any residual latex. Any remaining nylon reagents may be stirred together to allow coagulating, and then rinsed with methanol and disposed of. The latex and nylon may be disposed in the trash.

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. Discuss how different types of polymer structures and processing techniques (amorphous vs. crystalline, crosslinked vs. not, etc.) influence the macroscopically observed properties of the polymers (pliability, conductivity, permeability, etc.).
2. Tin exists in two different crystal structures. Old organ pipes made of tin were discovered many years later to be off-key. Why did this happen?

References and Additional Resources

Ford, Leonard Augustine, and E. Winston Grundmeier. *Chemical Magic*. 2nd ed. New York: Dover Publications, 1993.

Penn State Materials Research Institute. Center for Nanoscale Science. "Packing Peanuts Activity." Accessed August 15, 2013, http://www.mrsec.psu.edu/education/nano-activities/polymers/packing_peanuts/packing_peanuts.pdf

Penn State Materials Research Institute. Center for Nanoscale Science. "Amorphous Solids." Accessed August 15, 2013, http://www.mrsec.psu.edu/education/nano-activities/liquid_metal/amorphous_solids/amorphous_solids.pdf

Penn State Materials Research Institute. Center for Nanoscale Science “Nano-Activities for Kids.” Accessed August 28, 2013. <http://www.mrsec.psu.edu/education/nano-activities/>

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. We recommend that participants answer these questions in a journal.

Further Reading and Reflection Questions

Hodgson, Steven C., Sephen W. Bigger, and Norman C. Billingham. “Studying Synthetic Polymers in the Undergraduate Chemistry Curriculum.” *Journal of Chemical Education*. 78:4 (2001). Accessed August 28, 2013. <http://ncsu.edu/project/chemistrydemos/Organic/PolymerLitReview.pdf>

1. Do you think it is important to expose students to polymers at the high school level? Why or why not?
2. The authors suggest that in introductory college level chemistry courses it is not always practical to devote a whole unit to polymers. Do you think this is true for the high school level? Why or why not?
3. To incorporate polymer chemistry into the already established curriculum, the authors suggest using polymer-related examples to teach general chemistry concepts. How might you incorporate polymers into your curriculum? What are some examples you could use?

Bodner, George M. “Why Lecture Demonstrations Are ‘Exocharmic’ for Both Students and Their Instructors.” *University Chemistry Education- Royal Society of Chemistry*. 5 (2001): 31-35. http://chemed.chem.purdue.edu/chemed/bodnergrouop/PDF_2008/71%20Demos%20UCEd.pdf

1. Do you agree with the author that demonstrations should contain “an element of the unexpected”? Why or why not?
2. Do you agree with the authors’ conclusions? Which ones in particular resonate with you? Which ones do you disagree with? Explain your reasoning.
3. What are some strategies you might use when conducting demonstrations to help bring about conceptual change for the learner?

Professional Development Assignments

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