



CK-12 Chemistry - Intermediate Teacher's Edition



CK-12 Chemistry -Intermediate Teacher's Edition

Wade Baxter, Ph.D.

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AUTHOR Wade Baxter, Ph.D.

EDITORS

Donald Calbreath, Ph.D. Max Helix

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CHAPTER **1** Introduction to Chemistry

Chapter Outline

- 1.1 WHAT IS CHEMISTRY?
- **1.2 THE SCIENTIFIC METHOD**

Contents: CK-12 Chemistry - Intermediate

Unit 1: Overview of Chemistry

- Chapter 1: Introduction to Chemistry
- Chapter 2: Matter and Energy
- Chapter 3: Measurement

Unit 2: Organization of Matter

- Chapter 4: Atomic Structure
- Chapter 5: Electrons in Atoms
- Chapter 6: The Periodic Table
- Chapter 7: Chemical Nomenclature
- Chapter 8: Ionic and Metallic Bonding
- Chapter 9: Covalent Bonding

Unit 3: Quantitative Relationships

- Chapter 10: The Mole
- Chapter 11: Chemical Reactions
- Chapter 12: Stoichiometry

Unit 4: Phases of Matter

- Chapter 13: States of Matter
- Chapter 14: The Behavior of Gases
- Chapter 15: Water
- Chapter 16: Solutions

Unit 5: Chemical Reactions

- Chapter 17: Thermochemistry
- Chapter 18: Kinetics
- Chapter 19: Equilibrium
- Chapter 20: Entropy and Free Energy
- Chapter 21: Acids and Bases
- Chapter 22: Oxidation-Reduction Reactions
- Chapter 23: Electrochemistry

Unit 6: Nuclear, Organic, and Biochemistry

- Chapter 24: Nuclear Chemistry
- Chapter 25: Organic Chemistry
- Chapter 26: Biochemistry

Chemistry Glossary

The Teacher's Edition (TE)

The CK-12 Chemistry - Intermediate Teacher's Edition complements CK-12's Chemistry - Intermediate FlexBook®. The TE comprises six strands: Teaching Strategies & Tips, Differentiated Instruction, Enrichment, Science Inquiry, Common Misconceptions, and Assessments.

Each unit and chapter will have a general overview. Each chapter section will also include an introduction and teaching strategies. The majority of content will be presented by individual lesson.

This Teacher's Edition will focus on eight subtopics for each lesson:

- 1. Key Concept
- 2. Standards
- 3. Lesson Objectives
- 4. Lesson Vocabulary
- 5. Check Your Understanding
- 6. Teaching Strategies
- 7. Reinforce and Review
- 8. Points to Consider

Worksheets and Assessments

- The CK-12 Chemistry Intermediate Workbook complements CK-12's Chemistry Intermediate FlexBook® and contains worksheets for each lesson.
- The CK-12 Chemistry Intermediate Quizzes and Tests complements CK-12's Chemistry Intermediate FlexBook® and contains one quiz per lesson, one chapter test, and one unit test.

Pacing the Lesson

Each chapter has guidelines for the minimum number of class periods needed to teach each lesson. We have strived to keep each chapter under a week of class time, which would cover the complete FlexBook® resource in 25 weeks, providing ample time for flexibility. We realize this is a tremendous amount of material, and many teachers may choose not to utilize the complete FlexBook® resource, providing even more time for flexibility. As the teacher, you can determine if your class needs additional (or less) time on certain lessons/chapters, and adjust the pacing accordingly.

Science Notebook

For a year's study of Chemistry, we recommend a science and/or lab notebook in which students may:

- Answer the Check Your Understanding questions.
- Answer/reflect on the Points to Consider questions.
- Write additional questions about an upcoming lesson, chapter, or unit of study.
- Draw pictures of living organisms and diagrams of life processes.
- Take notes and define academic vocabulary.
- Keep a record of pertinent web sites to access relevant information.
- Write up lab activities.
- Write up ideas for possible long-term projects.
- Keep reflections on what they have learned.

Students should date each entry and refer back to their ideas earlier in the year, reflecting on their deepening understanding.

Teaching Strategies

Throughout the TE, we will provide numerous examples of strategies that can be used to make the content accessible to students. Many strategies and activities have been included as web site links, and we recommend that these be previewed before assigning to the students. Traditional examples of general teaching strategies, differentiated instruction, enrichment, science inquiry, and reinforcement strategies will be provided.

Teaching Strategies: General

- 1. Appreciate what's difficult for students, helping them develop scientific ways of thinking.
- 2. Vary class activities, using a wide variety of resources to aid students in deepening their understanding of scientific issues.
- 3. Give students opportunities to participate in scientific investigations to understand "doing science."

• The Scaffolding Strategy

As in any good teaching, bring up topics with which students are already familiar to give students a context to assimilate new understandings. Give these topics a "twist" to engage student's motivation. Break complex tasks into smaller tasks, show examples of quality outcomes, offer hints or verbal cues, use mnemonic devices, chants and/or songs for activities requiring memorization of facts or procedures. Use graphic organizers such as concept maps; teach key vocabulary before reading the FlexBook® textbook. Continually ask questions to guide and facilitate students in making predication, or to encourage deeper investigations or thinking on a topic. Model activities before students participate. And, ask for student contributions about their past experiences in the field.

• Word Dissection

Chemistry words can be intimidating for students to read, say, and talk about. As teacher, you can make a game of the words, and take a few minutes to do a daily or weekly dissection in class. Make sure students know that it is not a big deal to not know the word when you first come into contact with it, but to learn to break it down into pieces, figure out the meaning of each piece, and then put it all back together again to find the meaning.

• Reading to Learn

Teach your students how to read, comprehend, and summarize scientific text. Each lesson offers an opportunity to use different techniques to guide students to synthesize the core elements of the lesson. Try one or two different techniques each time:

• Using Visuals

Use illustrations in the student edition as a tool for teaching content, exploring ideas, and probing students' misunderstandings.

• Building Science Skills

Have students apply higher-level thinking or other relevant skills as they relate to lesson content (e.g., predicting, forming hypotheses, drawing conclusions, interpreting data, observing, classifying, making inferences, comparing and contrasting, identifying cause and effect, analyzing). This might be achieved through a simple activity, answering questions, class discussion, and/or partner work.

• Discussion

Stimulate class discussion of a topic. This could include scripted questions to ask the class, with expected or sample answers. The discussion tips should be specific and focused. For example, don't write: "Discuss Darwin's theory of evolution." Instead, write: "Guide students in discussing why Darwin's theory was not widely accepted in his own lifetime." Ask: "How did Darwin's theory of evolution conflict with prevailing views of living things?"

• Demonstration

Do (and fully describe) an in-class demonstration to illustrate or explain a process, concept, etc. Keep in mind constraints on classroom time and resources. Include a concluding sentence or scripted question that relates the demonstration to the process or concept.

• Activity

Have students do a simple hands-on activity that will help them better understand a topic or process. Explain fully how the activity is to be done. This could be a pencil-and-paper activity or other activity that does not involve materials, although readily available classroom materials could be used. Again, conclude with a sentence or question that ties the activity with the topic or process being studied.

Teaching Strategies: Differentiated Instruction

These strategies can be used for all three types of student populations that are typically addressed by DI (i.e., ELL, LPR, SN), but a particular population has been specified each time a strategy is used. The strategy can be tailored somewhat to that population, even if it's only by referring to the population type in the strategy (e.g., "Pair English language learners with native speakers of English").

• KWL

Have students make a KWL chart, where K = Know, W = Want to Know, and L = Learned. Students should fill in the K and W columns before reading and the L column after reading a particular passage or lesson.

Cloze Prompts

Give students cloze sentences (basically, fill-in-the-blank sentences) about important lesson concepts. Students are instructed to fill in the missing words as they read the lesson.

• Gallery Walk

Divide the class into groups and have the groups walk around the room to read and discuss posted questions or topics (each on a large sheet of paper). Each group (using a different color pen) answers the questions or writes comments about the topics. They also read and respond to answers/comments written by other groups. This is followed by discussing the answers/comments with the class, reviewing misunderstandings they reveal, or by groups summarizing what they know about one or more questions/topics.

• Think-Pair-Share

Assign questions or topics to individual students to think about. Pair ELL students with native speakers and LPR students with more proficient readers to work together on answering the questions or discussing the topics.

• Frayer Model

Assign this vocabulary strategy, which involves students drawing a large box and dividing it into four parts labeled "Definition," "Drawing," "Example," and "Non-example." Assign students a vocabulary word and tell them to fill in each part of the box for that word.

• Cluster Diagram

Have individual students, pairs, groups, or the class as a whole make a cluster diagram organizing lesson concepts.

• Concept Map

Have individual students, pairs, groups, or the class as a whole make a concept map organizing lesson concepts.

• Venn Diagram

Have individual students, pairs, groups, or the class as a whole make a Venn diagram organizing lesson concepts.

Compare/Contrast Table

Have individual students, pairs, groups, or the class as a whole make a compare/contrast table for specific lesson concepts, processes, etc. (e.g., photosynthesis and cellular respiration; mitosis and meiosis). You may need to provide the column and row headings for the table.

• Cycle Diagram

Have individual students, pairs, groups, or the class as a whole make a cycle diagram to show the steps in a cyclical process (e.g., life cycle of amphibians).

• Flow Chart

Have individual students, pairs, groups, or the class as a whole make a flow chart to show the steps in a process (e.g., photosynthesis).

• Main Ideas/Details Chart

Have students divide a sheet of paper in half, on the left side write the main ideas from a passage or lesson (skipping several lines between the main ideas). On the right side, students are instructed to fill in important details about each main idea as they read.

• Word Wall

Post lesson vocabulary words and their definitions, examples, etc., on a bulletin board or wall. Refer students to the word wall as they study lesson content.

Teaching Strategies: Enrichment

Although online and/or library research is always an option for enrichment, it tends to be overused. Avoid it unless it is really relevant and likely to be helpful for the other students in the class. Whatever students are assigned to do, they should be given a chance to share their work with the class through an informal oral presentation, a written report, etc. In some cases (e.g., making a board game or crossword puzzle), the product can be used by the class to reinforce or review lesson content.

- Research a Topic
- Present a Role-Play
- Teach a Topic
- Create a Video
- Create a Poster
- Debate an Issue
- Interview an Expert
- Create a Model
- Demonstrate a Process
- Take a Survey
- Write an Essay
- Make a Board Game
- Make a Crossword Puzzle
- Create a Web Site
- Make a Diagram
- Make a Diorama
- Make a Display
- Write a Research Proposal
- Make a Video
- Write a Rap (Song)
- Present a PowerPoint Show
- Lead a Discussion

Teaching Strategies: Science Inquiry

These strategies should get students involved in thinking or acting like a scientist. They should help the students learn lesson content by encouraging them to be actively engaged in scientific thinking and/or using scientific methods.

- Ask a Research Question: e.g., based on hypothetical observations.
- Formulate a Hypothesis: e.g., based on a research question. Must be specific and testable; could also ask students to describe data that would support or disprove the hypothesis.
- Develop a Research Plan: e.g., to test a specific hypothesis. Could focus on types of variables, controls, etc.
- Analyze Data: Data could be in a graph or table that is provided in the SE or TE or students could find the data online.

• Solve a Problem: requiring application of lesson concepts, procedures, etc.

Reinforcement Activity

In addition to online quizzes, this could be a quick teacher-directed activity or something students do alone or in pairs to make sure they understand lesson content. It should probably be aimed at the average to below-average students in the class, though reinforcement activities are important for all students. Some suggestions are listed below. The goal is to reveal to the teacher or to the students themselves what they know and what they still don't understand. The activity should include a sentence suggesting a way for students to learn what they don't know (e.g., "Find definitions in the FlexBook® textbook of any vocabulary words you did not know."). This can easily be preceded or followed with a **Lesson Review**. Either you or a student(s) leads a discussion to review the lesson. You can use the Lesson Summary from the student edition. Clarify any issues and answer any questions students may have.

- **Take an Online Quiz**: Have students track their own level of mastery of concepts as measured by quizzes. Additionally, you could have students take the quiz before and after teaching the material and have them track their growth.
- Make Flashcards: This activity could be used for boldface vocabulary words or important concepts; have students use the flashcards to quiz a partner.
- Label a Drawing: The drawing could be art from the SE with the labels deleted.
- Outline the Lesson: This could be done with a partner or as a class using an overhead projector.
- Ask Questions: Each student turns in a question on an index card. Then, the teacher answers or reviews material relevant to those questions that are asked most frequently.
- List and Discuss: Students make a list of something (e.g., reproductive isolating mechanisms), and then partners compare and discuss their lists.
- Use Vocabulary: Students use the lesson vocabulary words in sentences or a brief paragraph.
- Make a Quiz: Students write a few fill-in, matching, or true/false questions and then use them to quiz a partner.
- Make a Drawing: Students create a simple sketch to demonstrate comprehension of a process (e.g., cell division).
- **Complete a Chart**: Students complete missing parts of a diagram or fill in cells of a table that have missing information.

Check Your Understanding

This section includes questions related to previously presented information that the authors consider important for **the student** to have access to the information in the current lesson.

Points to Consider

Questions in this section serve as a segue into the next lesson (or chapter). Ask students to read the Points to Consider at the end of the lesson in their FlexBook® textbook. They can be answered individually or as an opening to lead a class discussion. Use these questions to assess student understanding and misconceptions before beginning the next unit of study.

Making the FlexBook® Textbook Flexible

An important advantage of the FlexBook® textbook is the ability it gives you, the teacher, to choose the chapters and lessons that you think are most important for your own classes. The following information is provided to help you decide whether to include this chapter or certain lessons in this chapter in your students' FlexBook® student edition. You should also consult the standards correlation table when selecting chapters and lessons to include in the FlexBook® resource textbook.

- As the introductory chapter in this FlexBook® resource, CK-12 recommends the inclusion of the material within this chapter in any course on Chemistry.
- Students should read this entire chapter before reading the remaining chapters.
- It is recommended that you include all the lessons of this chapter.



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Chapter Overview

Chemistry is the study of the composition of matter and the changes that matter undergoes. Chemists study both the macroscopic and the microscopic world. The discipline of chemistry can be divided into several branches, such as organic chemistry and analytical chemistry. Chemistry plays a vital role in many aspects of the modern world, from agriculture to medicine to energy production. As do all scientists, chemists employ the scientific method to accumulate information about the particular chemical systems that are being studied.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 1.1: Pacing Lessons for Intro to Chem

Lesson	Class Period(s) (60 min)
1.1 What is Chemistry	2
1.2 The Scientific Method	1

1.1 What is Chemistry?

Key Concept

Chemistry is one of many disciplines of science and contains many sub-disciplines or branches. All chemists study matter at its fundamental level and attempt to understand what is happening when matter undergoes changes. Chemistry impacts all of our lives in many ways.

Standards

Lesson Objectives

- Define chemistry.
- Differentiate between the macroscopic and the microscopic as it relates to chemistry.
- Know the relationship between pure chemistry and applied chemistry.
- Identify and describe the five primary disciplines of chemistry.
- Describe some of the concerns of the modern world in which chemistry has played and will continue to play a role.

Lesson Vocabulary

- analytical chemistry: Branch of chemistry that is the study of the composition of matter, focusing on separating, identifying, and quantifying chemicals in samples of matter.
- applied chemistry: Chemistry that is directed towards a specific practical goal or application.
- biochemistry: Branch of chemistry that is the study of chemical processes that occur in living things.
- chemistry: The study of the composition of matter and the changes that matter undergoes.
- inorganic chemistry: Branch of chemistry that is the study of chemicals that do not, in general, contain carbon.
- macroscopic: Description of substances that refers to objects that can be seen, touched, and measured directly.
- matter: Anything that has mass and takes up space.
- microscopic: Description of substances that refers to the small particles (atoms and molecules) that make up all matter.
- organic chemistry: Branch of chemistry that is the study of carbon-containing chemicals.
- physical chemistry: Branch of chemistry that is the study of the both the macroscopic and the atomic properties and phenomena in chemical systems.
- pure chemistry: Chemistry that is concerned with a greater understanding of the theories behind how matter is changing in chemical reactions.

Teaching Strategies

Introducing the Lesson

Split students up into small groups. Have them brainstorm about typical activities that they engage in on a regular basis. Tell them to decide if each of those activities involves chemistry. On a large sheet of butcher paper, have the groups write down the activities that they think involve chemistry and a brief description of how they know that it does. After 10 minutes, put them up at the front of the room as the leadoff to a class discussion of how chemistry is defined.

Activity

It is important for chemists to know the names and functions of different pieces of lab equipment. Use the link below to create a lab equipment matching activity. Place all of the pieces of lab equipment on the lab benches. In pairs or small groups, students should go around the room and attempt to match each piece of equipment to a written list provided by the teacher. Students should also write a brief description of what they think is the function of each piece of equipment. When all students are finished, lead a class discussion revealing the correct names and functions of each piece. This can be followed up on a later date with a quiz.

http://www.curriki.org/xwiki/bin/view/Coll_rmlucas/LabEquipmentMatching

Common Misconceptions

Students often think of chemicals as only being substances that are in clearly labeled bottles that a scientist may have in the lab. Show a few of these to the class. Then discuss some examples of everyday matter that can also be thought of as a chemical –as long as it is capable of undergoing a change into a different type of matter. Some examples to point out include: foods (changes during digestion or cooking), paper or wood (can be burned), metal objects (can react with air as in the rusting of iron), plastics (can decompose in certain environments). A good demonstration to show is the rapid decomposition of a foam coffee cup in acetone. Place the cup upside-down in a Petri dish, then add a small amount of acetone to the dish. Watch the cup slowly disappear.

Science Inquiry

A non-Newtonian fluid is a substance that behaves in a very peculiar way. One type of non-Newtonian fluid is called oobleck, or sometimes simply ooze. Follow the link below for a recipe for making ooze. Break up the class into small groups and have them each make a sample of ooze and investigate its properties. Use the information on the link to explain what is happening.

https://www.exploratorium.edu/science_explorer/ooze.html

Differentiated Instruction

Make pairs of students with differing abilities. Have each pair produce a concept map of the main topics in the section. "Chemistry" should be in the center and the diagram should include the other words from the vocabulary list.

Enrichment

Break the class into five groups. Have each group create a PowerPoint presentation on one of the five frontiers of chemistry described in the lesson: energy, medicine, materials, the environment, and agriculture. The presentations should focus on the use of chemistry.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 1.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 1.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Chemists, like all scientists, rely on a rigorous and systematic approach to studying the world called the scientific method.

- What are the steps of the scientific method?
- How do chemists and other scientists communicate their findings to the public?

1.2 The Scientific Method

Key Concept

Chemists employ the scientific method in their investigations. The scientific method is a sequence of steps that lead from questions about observable phenomena to the development of scientific theories and scientific laws.

Standards

Lesson Objectives

- Describe how the Renaissance period in history changed the approach to science.
- Identify the steps of the scientific method.
- Differentiate between the independent variable and the dependent variable in an experiment.
- Describe how scientists work in research groups and communicate their results.

Lesson Vocabulary

- control group: A group that experiences the same conditions as the experimental groups with the exception of the variable that is being tested.
- dependent variable: The experimental variable that is observed during an experiment.
- hypothesis: A testable conjecture that proposes an answer to a research question.
- independent variable: The experimental variable that is changed during an experiment.
- scientific law: A statement that summarizes the results of many observations and experiments and to which there are no known exceptions.
- scientific method: A systematic and logical approach to the acquisition of knowledge.
- theory: An explanation that has been repeatedly tested and confirmed by multiple researchers and has gained wide acceptance.

Teaching Strategies

Introducing the Lesson

The majority of students in a high school chemistry course will be familiar with the scientific method. Begin the topic by asking students for key terms that pertain to the scientific method. Write these on the board and follow with the activity below.

1.2. The Scientific Method

Activity

A hands-on activity is a great way to teach the steps of the scientific method. The link below is to a lab that uses bubble gum as the experimental material in an investigation that requires students to use the scientific method.

http://serc.carleton.edu/sp/mnstep/activities/27600.html

Common Misconceptions

One misconception is that the scientific method is a rigid and linear sequence. In reality, many scientists may be involved in tackling a particular research problem and each one may be following the method in different ways. Not every experiment that a scientist does results in a clear "answer" to a hypothesis. In practice, experimental results often open up more questions that must be addressed.

Another misconception involves the timeframe that scientific research follows. Students become conditioned to think that science experiments can be done in one or two hours (class periods), when real research may extend over months or even years. To be certain of one's conclusions, scientists must usually repeat experiments and modify experimental conditions over and over again.

A useful site to visit in order to see the complexity of scientific study and its relationship to the scientific method is listed below.

http://undsci.berkeley.edu/article/0_0_0/howscienceworks_02

Science Inquiry

To give students a fun opportunity to put their skills with the scientific method to the test, have the class perform the activity at the link below. The activity involves students engaging in "thumb wars" against other students. Experimental variables include thumb length, wrist length, and thumb circumference. Students are required to make hypotheses and graph results. Make sure to download the lab and make handouts for each student.

http://www.lessonplansinc.com/biology/detail/scientific_method_activity

Differentiated Instruction

Have English learners and less proficient readers work with a partner to create an artistic rendition of the scientific method. Provide a large sheet of butcher paper along with markers and/or colored pencils for the students to use.

Enrichment

The web site below is a terrific expansion on the scientific method and how it has been used in the past for some of the most important scientific discoveries. Have students read the article and write a one or two page summary of what they've learned about how the scientific method was used in the development of modern cell theory. Note: The cell theory begins on page 5 of the article.

http://science.howstuffworks.com/innovation/scientific-experiments/scientific-method.htm

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 1.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 1.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Chemistry is the study of matter and the changes that it undergoes. Before beginning to examine those changes, it is important to be able to classify matter according to certain common characteristics and be able to recognize when a chemical change is occurring.

- What properties of matter allow it to be classified?
- What clues allow a chemist to recognize a chemical reaction?



Matter and Change

Chapter Outline

- 2.1 **PROPERTIES OF MATTER**
- 2.2 CLASSIFICATION OF MATTER
- 2.3 CHANGES IN MATTER



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Chapter Overview

Chemistry is the study of matter and so it is important to be able to understand the properties of matter and to classify matter based on type. All samples of matter have both physical and chemical properties, and thus can undergo physical and chemical changes. Classifications of matter include elements, compounds, and mixtures. Careful observation of changes in matter must be made in order to determine if each change is physical or chemical in nature.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 2.1	1:	Pacing Lessons
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Lesson	Class Period(s) (60 min)
2.1 Properties of Matter	1
2.2 Classification of Matter	1
2.3 Changes in Matter	1

2.1 Properties of Matter

Key Concept

Any sample of matter can be described in part by its physical properties, those that can be observed or measured without altering the substance. Physical properties can be extensive or intensive, depending on whether the property is dependent on the amount of matter in the sample. States of matter are an important physical property. A change to a sample of matter that does not alter its identity is a physical change.

Standards

Lesson Objectives

- Describe a substance of matter according to its physical properties.
- Distinguish extensive and intensive properties.
- Describe the three states of matter.
- Identify physical changes to matter.

Lesson Vocabulary

- extensive property: Property that depends on the amount of matter in a sample.
- gas: A form of matter that has neither a definite shape nor a definite volume.
- intensive property: A property of matter that depends only on the type of matter in a sample and not on the amount.
- liquid: A form of matter that has a definite volume, but an indefinite shape.
- mass: Measure of the amount of matter that an object contains.
- physical change: A change to a sample of matter in which some properties of the material change, but the identity of the matter does not.
- physical property: The characteristic of a substance that can be observed or measured without changing the identity of the substance.
- pure substance: Matter that has a uniform and definite composition.
- solid: A form of matter that has a definite shape and volume
- vapor: The gaseous state of a substance that is solid or liquid at room temperature.
- volume: Measure of the space occupied by a sample of matter.

Teaching Strategies

Introducing the Lesson

Have students brainstorm all of the physical properties of matter that they can think of. Write the results on the board. They may need some nudging to think of things such as solubility and electrical conductivity. Define extensive and intensive properties for them. Allow the students a few minutes to split the list on the board correctly into the two categories. Reveal the answer and see how they did.

Common Misconceptions

Students often have trouble understanding that dissolving a substance in water is a physical change and not a chemical change. To emphasize this point, you can completely dissolve some salt into a beaker of water. Pour the salt water solution into an evaporating dish and heat with a Bunsen burner. Show the students that the salt is left behind when the water is boiled away.

Some students misunderstand that the identity of a substance is not altered by a change of state. In particular, emphasize that the boiling of water (or any substance) simply spreads the molecules apart from one another. It does not break chemical bonds. Molecular models may be useful to make this point.

Science Inquiry

Density is one of the physical properties that students learn about in this lesson. A fun way to see density in action is to prepare a density column. The link below gives instructions on how to prepare one possible column. Either prepare one as a demonstration or allow students to prepare their own in small lab groups. Discuss how density is an intensive property because the density of a given substance is the same no matter how much of it is present in the sample.

http://www.stevespanglerscience.com/lab/experiments/seven-layer-density-column

Differentiated Instruction

Students of lower ability may benefit from making particle models of the three states of matter. Small foam balls or miniature marshmallows can be used to represent the particles. A solid should have the particles packed together in an orderly arrangement. A liquid, should have the particles be close together, but more random. The gas should have the particles spread far apart.

Enrichment

Have more advanced students research two states of matter other than solid, liquid, and gas. The states are plasmas and Bose-Einstein condensates. Students can either write up a summary of their findings or prepare a brief PowerPoint and an oral report.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 2.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 2.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

A mixture is different from a pure substance because it does not have a uniform and definite composition. Pure substances include both elements and compounds.

- Can mixtures be further classified into more than one type?
- How are compounds different than elements?

2.2 Classification of Matter

Key Concept

All matter can be characterized as either a pure substance or a mixture. Mixtures may be either homogeneous or heterogeneous. Parts of a mixture can be separated from one another by physical means such as filtration or distillation. A pure substance has a fixed composition and can either be a compound or an element. Compounds can only be broken down into their elements by a chemical change. Chemical symbols and chemical formulas are used by chemists to represent elements and compounds.

Standards

Lesson Objectives

- Define a mixture and understand why mixtures are different than pure substances.
- Classify mixtures as homogeneous or heterogeneous.
- Describe several ways to separate mixtures.
- Distinguish between elements and compounds.

Lesson Vocabulary

- chemical change: A change that produces matter with a different composition.
- chemical formula: An expression that shows the elements in a compound and the relative proportions of those elements.
- chemical symbol: One- or two-letter designation of an element.
- compound: A substance that contains two or more elements chemically combined in a fixed proportion.
- distillation: Method of separating a mixture based on differences in boiling points.
- element: The simplest form of matter that has a unique set of properties
- filtration: A technique that separates the solid from the liquid in a heterogeneous mixture.
- heterogeneous mixture: A mixture in which the composition is not uniform throughout the mixture.
- homogeneous mixture: A mixture in which the composition is uniform throughout the mixture.
- mixture: A physical blend of two or more components, each of which retains its own identity and properties in the mixture.
- phase: Any part of a sample that has a uniform composition and properties.

Teaching Strategies

Introducing the Lesson

The primary point of emphasis in this lesson is for students to understand the difference between a pure substance and a mixture. Produce several examples to show to the class and ask them to correctly identify each as a pure substance or mixture. Examples should include at least one element, one compound, one homogeneous mixture, and one heterogeneous mixture.

Common Misconceptions

Some students have difficulty understanding the difference between a compound and a mixture. Ask them what two gases make up most of our atmosphere (nitrogen and oxygen). Write some of the formulas for various nitrogen oxides on the board (NO, NO₂, N₂O) and discuss how it is very fortunate for living things that the nitrogen and oxygen in air do not react and are instead simply a mixture. Ask them if they can think of a natural event that causes nitrogen and oxygen to chemically react (lightning strikes).

Science Inquiry

Separating parts of a mixture is a great way for students to gain a hands-on understanding of homogeneous mixtures. Many online sites have example activities of separating the inks in a black marker using paper chromatography.

A basic approach that can be done rather quickly: https://www.exploratorium.edu/science_explorer/black_magic.h tml

A more advance treatment involved calculation of retention factors: http://blogs.oregonstate.edu/hydroville/files/20 14/06/paper_chrom1.pdf

Distillation is a separation technique that can also be performed in the laboratory. This can easily be done as a demonstration since a proper distillation requires a great deal of equipment. The link below describes the purification of salt water by distillation.

http://www.education.com/science-fair/article/salt-water-distillation/

Differentiated Instruction

Have lower level students and English language learners make a concept map of all the vocabulary terms in this lesson. Have students compare their maps with those of others.

Enrichment

Another technique commonly used by chemists to enact a separation of the parts of a mixture is called chromatography. Have advanced students prepare a poster showing different types of chromatography along with descriptions of how chromatography causes separation to occur. Assign different students or groups of students a different type of chromatography (column, thin layer, paper, gas, etc.)

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 2.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 2.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

In a chemical reaction, one or more substances are changed into different substances.

- How are chemical changes different from physical changes?
- What are the visual clues that a chemical reaction is taking place?

2.3 Changes in Matter

Key Concept

Chemical properties of a substance describe the ability of that substance to undergo chemical reactions with other substances. When a chemical reaction occurs, reactants are converted into products. Evidence that a chemical reaction is occurring includes a color change, the production of a gas, the production of a precipitate, or a transfer of energy.

Standards

Lesson Objectives

- Identify the chemical properties of a substance.
- Describe chemical changes and differentiate them from physical changes.
- Use various visual clues to identify whether a chemical reaction is taking place.

Lesson Vocabulary

- chemical property: Describes the ability of a substance to undergo a specific chemical change.
- chemical reaction: Process that occurs when one or more substances is changed into one or more new substances.
- precipitate: Solid product that forms from a reaction and settles out of a liquid mixture.
- product: Substance that is present at the end of a chemical reaction.
- reactant: Substance that is present at the start of a chemical reaction.

Teaching Strategies

Introducing the Lesson

Discuss with the students how chemical reactions occur all around them in everyday life and not just in a laboratory. Give them a few minutes to think of as many examples of these everyday chemical reactions. Ask for the class to share what they think and then ask them how it is possible to tell that a chemical reaction is occurring.

Common Misconceptions

Students sometimes have difficulty understanding that when two pure substances combine, they can do so either physically or chemically and that the end result is very different in both cases. Show them a sample of zinc (as powder or small pellets) and some sulfur powder. Mix them together at room temperature and discuss how no chemical reaction has taken place –so the result is a heterogeneous mixture. If you can, heat the mixture in a crucible over a Bunsen burner for several minutes. Discuss how the heat provokes a chemical reaction, forming zinc sulfide. Note: This reaction is described in the student text.

Science Inquiry

Students will benefit from observing chemical and physical changes in a laboratory setting. The two links below are activities in which students perform quick experiments at a number of different lab stations. They make observations and decide whether a chemical or physical change has occurred in each case. Either could be modified based on available chemicals or performed as teacher demonstrations.

http://www.slcschools.org/departments/curriculum/science/Grade-7-to-8/Grade-8/documents/Physical-and-Chemica 1-Changes.pdf

http://www1.umn.edu/ships/modules/scimath/mendlev3.htm

Differentiated Instruction

Assign students a different element and ask them to research the physical and chemical properties of their element. Using a large sheet of butcher paper, have them make two columns, titled "Physical" and "Chemical." List the properties for their element. Emphasize that the chemical properties will be descriptions of some of the chemical reactions of their element. Display the results around the room.

Enrichment

Advanced students can begin to investigate the types of chemical reactions: synthesis, decomposition, single replacement, double replacement, and combustion. Have them find an example of each reaction and relate it to the current lesson. In other words, how can an observer tell that a chemical reaction is occurring in each case?

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 2.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 2.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

A chemist must do more than simply make observations about chemical changes that are occurring in his or her experiments. He or she must also take measurements in order to fully understand what is taking place during chemical reactions.

- What types of measurements are routinely made by chemists?
- How does a chemist indicate the degree of confidence in a measurement?
- How are different numerical measurements manipulated when performing calculations?



Measurements

Chapter Outline

- 3.1 THE INTERNATIONAL SYSTEM OF UNITS
- 3.2 UNIT CONVERSIONS
- 3.3 UNCERTAINTY IN MEASUREMENTS



Tweenk. www.commons.wikimedia.org/wiki/File:Lab_glassware.jpg. CC BY 3.0

Chapter Overview

Chemistry is the study of matter and so it is important to be able to understand the properties of matter and to classify matter based on type. All samples of matter have both physical and chemical properties, and thus can undergo physical and chemical changes. Classifications of matter include elements, compounds, and mixtures. Careful observation of changes in matter must be made in order to determine if each change is physical or chemical in nature.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

Lesson	Class Period(s) (60 min)
3.1 The International System of Units	1
3.2 Unit Conversions	2
3.3 Uncertainty in Measurements	2

TABLE 3.1: Pacing Lessons

3.1 The International System of Units

Key Concept

The International System of Units (SI) consists of seven base units from which all other units are derived. The SI is based on the metric system, which uses prefixes to represent various powers of ten. Scientific notation is used in the case of very large or very small numbers. The SI unit of length is the meter and the SI unit of volume is the cubic meter. The SI unit of mass is the kilogram. Weight is a measure of force that gravity exerts on an object and thus depends on location. The temperature of an object is related to its kinetic energy. While Celsius is the metric unit of temperature, the SI unit is the Kelvin.

Standards

Lesson Objectives

- Identify the seven base units of the International System of Units.
- Know the commonly used metric prefixes.
- Convert between the Celsius and Kelvin temperature scales.
- Understand volume and energy as combinations of SI units.
- Distinguish between mass and weight.

Lesson Vocabulary

- energy: The capacity to do work or to produce heat.
- International System of Units (SI): A system of measurement based on the metric system and adopted by international agreement in 1960.
- joule: SI unit of energy; in terms of base units, a joule is equal to a kg \cdot m²/s².
- kinetic energy: Energy that an object has due to its motion.
- liter: The volume of a cube that measures 10 cm (1 dm) on each side.
- measurement: A quantity that has both a number and a unit.
- scientific notation: A way to express numbers as the product of two numbers: a coefficient and the number 10 raised to a power.
- temperature: Measure of the average kinetic energy of the particles in matter.
- weight: Measure of force that is equal to the gravitational pull on an object.

Teaching Strategies

Introducing the Lesson

Begin a lesson by asking the class if they can name all of the seven units of the SI. They are likely to name some that are derived units like volume or speed. They might need hints for mole and ampere and will likely have trouble coming up with candela. After the central unit (with no prefix) is established, encourage the students to memorize the most commonly used prefixes and their relationship to the central unit (mega-, kilo-, centi-, milli-, micro-, and nano-). Perform the science inquiry activity described below.

Common Misconceptions

Some students think of mass and weight as being the same thing. If you have access to a hanging spring scale, demonstrate its use compared to a balance. Discuss how weight is a force that is measured in newtons, which are equal to mass multiplied by the acceleration of gravity.

Heat and temperature are also sometimes misunderstood. Emphasize that temperature is something that an object "has", while heat is the energy that follows between two objects that are at a different temperature.

Science Inquiry

Give students practice making measurements by having them perform a simple activity. Prepare a list of measurements to make –the list below provides some ideas, but you can come up with more.

- height of the lab bench
- width of their desk
- mass of a pencil
- temperature of a beaker of water
- volume of a plastic cup
- volume of a regular object such as a block of wood

Provide the students with the necessary variety of measuring tools: ruler, meter stick, thermometer, electronic balance, graduated cylinders of different sizes. The students should report the measurements along with the correct units.

Differentiated Instruction

Students that may need more practice with measurements can benefit from the web site shown below. It has a variety of practice exercises, games, and quizzes that can be assigned as needed.

http://www.softschools.com/measurement/

Enrichment

Have students research the standards of measurement for time, length, and mass and how each of those standards has changed over the years. For example, the meter was originally defined (by the French) as one ten-millionth of the distance from the equator to the North Pole along the meridian that passes through Paris. That standard is no longer used, but has changed to being based on the distance that light travels in a second.
Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 3.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 3.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Conversions between units of the metric system are made easy because they are related by powers of ten and because the prefixes are consistent across various types of measurement (length, volume, mass, etc.)

- What is the mass in grams of a 2.50 kg book?
- What is the length in cm of a field that is 0.65 km?

3.2 Unit Conversions

Key Concept

A conversion factor is a ratio of two equivalent quantities, each with a different unit. Conversion factors are used to solve problems with the mathematical technique of dimensional analysis. Unit conversions within the metric system are made simple because all units are related by factors of 10. Derived units such as area and volume can also be converted using dimensional analysis. The physical property of density can be measured or expressed in different units and dimensional analysis can be used to convert from one density unit to another. Dimensional analysis can also be used to calculate an unknown mass or volume of a substance with a known density.

Standards

Lesson Objectives

- Identify and use conversion factors.
- Use the method of dimensional analysis to convert between units.
- Understand density as a physical property of matter.
- Perform calculations with derived units, including density.

Lesson Vocabulary

- conversion factor: Ratio of equivalent measurements used to convert a quantity from one unit to another.
- density: The ratio of the mass of an object to its volume.
- derived unit: The unit that results from a mathematical combination of SI base units.
- dimensional analysis: A technique that uses the units (dimensions) of the measurement in order to correctly solve numerical problems.

Teaching Strategies

Introducing the Lesson

Start with a conversion factor that is familiar such as 12 donuts = 1 dozen. Do an example of converting a number of donuts to dozens as well as a number of dozens to number of donuts. Stress the canceling of the units every time you do a problem on the board. Give the students some practice problems with everyday conversions like this so that they get used to the method before proceeding to the metric system, which intimidates some students. Proceed from simple to more complex types of problems: (1) one-step metric conversions, (2) two-step metric conversions,

(3) area and volume conversions, (4) derived units such as speed and density. Make sure the students master each type before you move on.

Common Misconceptions

Some students have trouble with the notion that a fraction can be "equal" to 1 when the numerical portion of the fraction is not. Stress that the numerator and the denominator simply have to represent equal quantities.

Converting area and volume units are often very troublesome because the conversion unit has to be squared or cubed. Write out the conversion in two steps as shown below for the conversion of cm^3 to m^3 .

Science Inquiry

The link below is for a density-based inquiry lab activity. Students must collect a set of mass and volume data for an unknown liquid, graph the results, and use their graph to make predictions. This is a good activity to use for students to practice their graphing skills. Ideally, have them use a program such as Microsoft Excel for their graphs rather than drawing them by hand. It's more accurate.

https://www.flinnsci.com/media/621577/91655.pdf

Differentiated Instruction

Train students to keep lists of what they know and what they don't know for each problem. For example, suppose they need to convert 200 m to km. Under "known", they would write "200 m". Under "unknown", they would write "? km". Then write a connecting equation to go from the known to the unknown. In this example, "1000 m = 1 km". Starting problems like this will help students from getting confused.

Enrichment

A number of online tutorials and quizzes can provide students who need extra practice with instead feedback on how they are doing with unit conversions. The first site listed below has an excellent tutorial on single unit, double unit, and squared and cubed unit conversions. It also has a worksheet that can be assigned to the students so that they must show their work on paper as they go along.

http://joneslhs.weebly.com

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 3.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 3.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Measurements all must have a certain amount of uncertainty in them, since no measuring tool is 100% accurate. The uncertainty in a measurement must be considered both when reporting measured values and when doing calculations.

- How is the uncertainty in a given measurement indicated in the reported value?
- When a quantity such as density is calculated from two measurements (mass and volume), is it important to measure both accurately or is just one sufficient?
- What are the meanings of the terms precision and accuracy when dealing with measurements?

3.3 Uncertainty in Measurements

Key Concept

Chemists strive for both accuracy and precision in all measurements. Percent error is a quantity that can be calculated and gives an indication of how close a measured value is to the true value. All measuring tools have a degree of uncertainty to them. The significant figures of a measured quantity include all of the certain digits plus one estimated digit. When performing calculations involving two or more measured quantities, specific rules must be followed in order to determine how to properly round off the answer.

Standards

Lesson Objectives

- Distinguish between accuracy and precision in measurements.
- Calculate the percent error of a measured quantity.
- Report measured values to the correct number of significant figures based on the measuring tool.
- Perform calculations with measured quantities, rounding the answers to the correct number of significant figures.

Lesson Vocabulary

- accepted value: The true or correct value of a measurement based on general agreement with a reliable reference.
- accuracy: Measure of how close a measurement is to the correct or accepted value of the quantity being measured.
- error: The difference between the experimental value and accepted value of a measurement.
- experimental value: The quantitative value of a measurement obtained during an experiment.
- percent error: The absolute value of the error divided by the accepted value and multiplied by 100%.
- precision: The measure of how close a series of measurements are to one another.
- significant figures: All the certain digits in that measurement plus one uncertain or estimated digit.

Teaching Strategies

Introducing the Lesson

On the board, draw the two rulers from Figure 3.9 in the text. Ask the class which ruler is a better measuring tool and why. Follow that with a discussion of how the ruler with the millimeter markings yields a measurement that can be estimated to the hundredth of a centimeter. Stress that significant figures always include exactly one uncertain or estimated digit. This is true for any measuring instrument.

Common Misconceptions

Students often believe that an electronic balance (or pH meter) gives an exact measurement since they don't have to read it visually like a graduated cylinder. Emphasize that the last digit given by the balance is still uncertain and that the significant figure rules apply.

Students may believe that both accuracy and precision can be judged based on a single measurement. Stress that precision can only be determined from multiple measurements.

Science Inquiry

The following activity requires students to make multiple mass and volume measurements and to calculate densities. Then, they calculate average values for their data and percent error. This allows them to judge the accuracy and precision of their work.

http://serc.carleton.edu/sp/mnstep/activities/36239.html

The activity found at the link below is a good introduction to accuracy and precision and gives students a chance to practice percent error calculations. Rather than collecting their data, they use data provided for them, so it will take less time than the previous activity. It could also be assigned as homework.

http://www.sciencelearn.org.nz/Science-Stories/Measurement/Precision-and-accuracy

Differentiated Instruction

Have less proficient readers write the rules for significant figures in their own words. Have them provide examples demonstrating each of the rules. The students can compare their results in small groups.

Enrichment

The percent error calculation is how the accuracy of a measurement is determined and is an important calculation that all students should be able to do. Have advanced students research the calculation called average deviation. Give them a series of measurements as an example and have them calculate the average deviation and be able to explain the purpose of the calculation.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 3.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 3.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Measurements will be a constant consideration throughout your study of chemistry. Next you will begin a study of the atom, its component parts, and the evolution of the atomic model.

• Atoms are extremely small and extremely light. How do you think that the mass and size of an atom can be measured? Do you think the accuracy of these measurements has improved over time?

Atomic Structure

Chapter Outline

CHAPTER

- **4.1 А**томѕ
- 4.2 THE NUCLEAR MODEL OF THE ATOM
- 4.3 **ISOTOPES AND ATOMIC MASS**



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Chapter Overview

The atom as a concept dates back to the philosophers of ancient Greece. Beginning in the later 1700s, scientific experimentation began to pave the way for a succession of models of the atom. John Dalton's model viewed the atom as a tiny, indestructible sphere. The discovery of subatomic particles necessitated several alterations of the model. Thomson's plum pudding model had electrons embedded within a sphere of positive charge. Rutherford's gold foil experiment resulted in the nuclear model of the atom, where most of the mass and all of the positive charge of an atom is found in a tiny nucleus. All atoms can be described by their atomic number and their mass number. Most elements have two or more isotopes, which differ in mass because of variation in the number of neutrons in the nucleus.

Online Resources

See the following Web sites for appropriate activities:

For more information about the origins of atomic theory, from Democritus to Dalton, the following website is an excellent source: http://www.visionlearning.com/en/library/Chemistry/1/Early-Ideas-about-Matter/49

Pacing the Lessons

TABLE 4.1	1:	Pacing Lessons
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Lesson	Class Period(s) (60 min)
4.1 Atoms	1-2
4.2 The Nuclear Model of the Atom	2
4.3 Isotopes and Atomic Mass	2

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 4.2:	Chemistry	/ Concepts
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FlexBook® Lesson	Chemistry Concepts
4.1 Atoms	4.1 -4.6
4.2 The Nuclear Model of the Atom	4.7 -4.14
4.3 Isotopes and Atomic Mass	4.15 - 4.19

4.1 Atoms

Key Concept

Democritus, an ancient Greek philosopher, suggested the concept of an atom as the smallest unit of matter. John Dalton, an English schoolteacher, formulated an atomic model based on scientific experimentation. Those experiments led to the law of conservation of mass, the law of definite proportions, and the law of multiple proportions. Dalton's model viewed the atoms as very small, indivisible particles which participated in chemical reactions by combining, separating, and rearranging.

Standards

Lesson Objectives

- Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
- Describe John Dalton's atomic theory.

Lesson Vocabulary

- atom: The smallest particle of an element that retains the properties of that element.
- law of conservation of mass: A law that states that during a chemical reaction, the total mass of the products must be equal to the total mass of the reactants.
- law of definite proportions: A law that states that a given chemical compound always contains the same elements in the exact same proportions by mass.
- law of multiple proportions: A law that states that whenever the same two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.

Teaching Strategies

Introducing the Lesson

Begin by asking the students about their current knowledge of atoms. They will most likely name some of the parts of the atom and give some indication of its extremely small size. Tell them that the focus of the first two lessons of this chapter is to examine how the understanding of the atom has changed over time. Discuss why it is important for chemists to use models: because atoms and molecules are so small that it is impossible to directly observe their behavior.

Models

Even though a full treatment of chemical formulas and bonding does not come until later chapters, use some simple molecular models and chemical formulas to help explain the laws of definite proportions and multiple proportions. Show the students how certain elements are capable of combining together chemically to make different compounds. Examples include CO and CO₂ (discussed in the text), H_2O (water) and H_2O_2 (hydrogen peroxide), and multiple oxides of nitrogen (NO, N₂O, NO₂, N₂O₅).

Common Misconceptions

Some students may get the mistaken impression that the law of multiple proportions contradicts the law of definite proportions. Stress that they mean different things. The law of definite proportions applies to just one particular compound at a time and says that the mass ratio of the elements is always the same within that one compound. Because of our current understanding of atoms and molecules, that may seem obvious. Remind them that the laws from this lesson came about before even Dalton's model of the atom. Mass could be measured and understood so that is why all three laws are based on mass and that the concept of atoms is unnecessary for the laws.

Science Inquiry

The law of conservation of mass can easily be studied with a lab activity. The link below is to one that involves dissolving an Alka-Seltzer tablet in water in a closed container. The follow-up review portion of the lab involves balancing a chemical equation for the reaction. As equation balancing and formula writing are covered in a later chapter, it could easily be omitted here.

http://serc.carleton.edu/sp/mnstep/activities/35943.html

Differentiated Instruction

The mathematics behind the law of multiple proportions can be difficult for students of lower ability. Pair up students with stronger math skills with lower ability students and have them share their problem solving techniques. The following videos may also be viewed by students who need additional help with multiple proportions problems.

https://www.youtube.com/watch?v=D6HbmG8nIrU

https://www.youtube.com/watch?v=5zwk4PZi5x0

Enrichment

Have students research the life and scientific pursuits of John Dalton. In addition to his atomic theory, he also made advancements in the fields of physics, meteorology, and even color blindness. Groups of students could prepare a slideshow and oral report of their findings.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 4.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 4.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

All atoms are known to consist of several subatomic particles. Elements are distinguished from one another by the numbers of those particles that are present.

- Which three subatomic particles are present in all atoms?
- What is the general arrangement of these particles within atoms?

4.2 The Nuclear Model of the Atom

Key Concept

John Dalton's model was incorrect because experiments showed that all atoms could be divided into smaller particles. J.J. Thomson discovered the electron with a cathode ray tube. Robert Millikan's oil-drop experiment allowed the determination of the mass and charge of the electron. Further experiments proved the existence of the proton and the neutron. Ernest Rutherford performed the gold foil experiment, proving that most of the atom was empty space and that most of the mass and the positive charge of every atom were concentrated in a very small and dense nucleus.

Standards

Lesson Objectives

- Distinguish between the three main subatomic particles.
- Understand the contributions of J.J. Thomson, Robert Millikan, and Ernest Rutherford to atomic theory.
- Describe the structure of the nuclear atom.

Lesson Vocabulary

- atomic model: A way of accurately representing the experimental evidence about atoms in the simplest way possible
- cathode ray: Glowing beam that travels from the cathode to the anode in a cathode ray tube
- cathode ray tube: Device in which an electric current is passed through gases at low pressure
- electron: Negatively charged subatomic particle present in all atoms.
- neutron: A subatomic particle with no electrical charge and a mass that is approximately the same as the mass of a proton.
- nucleus: The tiny, dense, central core of the atom that is composed of protons and neutrons.
- proton: A positively charged subatomic particle present in all atoms.

Teaching Strategies

Introducing the Lesson

Students will most likely come into the course with some knowledge of the primary three subatomic particles (electron, proton, neutron). Begin the lesson by having students write down their recollections of these particles in terms of their charge, mass, and location within the atom. Summarize by making a table on the board with the

correct information. Then you can discuss how the scientific experiments led to the discovery of each of these particles.

Demonstration

If possible, demonstrate the use of a cathode ray tube to the class. Discuss the parts of the tube and show how the ray can be deflected with a magnet.

Common Misconceptions

Students sometimes fail to realize the purpose of learning about models that were later proven to be incorrect, such as Dalton's model and the plum pudding model. Emphasize that science is not just a collection of facts, but is an ongoing process. Within all science disciplines, new discoveries are constantly being made and that constantly changes our understanding. So even though some of the atomic models were "wrong", it is still important to learn the process of science.

Science Inquiry

Rutherford's experiment can be simulated in the lab with a fun activity that the students will enjoy. Students randomly roll a marble (alpha particle) at two or three target marbles (gold nuclei) and count the number of times the targets are hit. Using a statistical analysis, the diameter of the target marbles can be calculated. The link below describes the lab.

http://quarknet.fnal.gov/dnload/RollingWithRutherford.pdf

Differentiated Instruction

Have students create a timeline of major atomic discoveries described in the first two lessons, from Democritus to Rutherford. In the next lesson, the atomic model will be furthered by additions of discoveries up through the middle of the twentieth century. So the students should leave enough space to add to their timeline. Have them include brief descriptions of the contributions of each scientist on to the timeline.

Enrichment

The discovery of the neutron came much later than the discovery of the electron and of the nucleus. The reason is that its lack of charge made it difficult to detect. Have a group students research the experiments by James Chadwick that led to the discovery of the neutron. The students can summarize their findings in a report to the class.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 4.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 4.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The number of protons in a given atom must be equal to the number of electrons in that atom so that the atom is electrically neutral. The composition of the nucleus of an atom yield important information about the atom.

- What is the meaning of the atomic number of an atom?
- What are isotopes?

4.3 Isotopes and Atomic Mass

Key Concept

The atomic number of an element is the number of protons found in the nucleus of the atoms of that element. The mass number is equal to the sum of the protons and neutrons of an atom. Isotopes are different forms of a given element whose mass numbers are different because of a different number of neutrons. The average atomic mass for any element can be calculated from the mass of each individual isotope combined with the percent natural abundances of each of the isotopes.

Standards

Lesson Objectives

- Define atomic number.
- Define mass number.
- Understand how isotopes differ from one another and be able to designate them by various methods.
- Be able to calculate the average atomic mass of an element.

Lesson Vocabulary

- atomic mass: The weighted average of the atomic masses of the naturally occurring isotopes of an element
- atomic mass unit: A mass equal to one twelfth the mass of an atom of carbon-12.
- atomic number (Z): The number of protons in the nucleus of an atom.
- isotopes: Atoms that have the same atomic number but different atomic masses, due to a change in the number of neutrons.
- mass number: The total number of protons and neutrons in an atom.
- nuclide: The nucleus of a given isotope of an element; and atom that is defined by the number of protons and neutrons in its nucleus.

Teaching Strategies

Introducing the Lesson

Hold up same-size pieces of two metals that are very similar in appearance, such as zinc and aluminum. Ask the students what makes them different. They may say things that are macroscopic properties like strength or conductivity. Lead them into thinking about the microscopic. What is different about the atoms of zinc compared to

the atoms of aluminum? This leads to the definition of atomic number, the lead-off concept of the lesson. Follow up with the realization that the piece of zinc is heavier than the piece of aluminum, indicating a greater density. This is because of the greater number of protons and neutrons in each atom of zinc, which leads to the concept of mass number.

Simulation

The computer simulation below allows students to investigate isotopes and atomic mass. If used with a projector, the teacher could lead a class activity by demonstrating the simulation. Alternatively, if a computer lab is available, the class could be taken there for each student to use the simulation independently or with a partner. The site has many activity handouts (under Teaching Ideas) that can be used with the simulation. Choose whichever one works for your class.

http://phet.colorado.edu/en/simulation/isotopes-and-atomic-mass

Common Misconceptions

Students often confuse atomic mass with mass number. Stress that mass number is by definition an integer because is equal to the number of protons + neutrons. Atomic mass is not an integer because it is a weighted average of multiple isotopes. Even for elements that only have one naturally occurring isotope, the atomic mass on the Periodic Table is still not a whole number. For example, the only natural isotope of arsenic is arsenic-75, yet its atomic mass from the Periodic Table is 74.92 amu. Why? Atomic masses of isotopes vary slightly and only carbon-12 has an atomic mass that is an integer since the amu is defined relative to the mass of carbon-12.

Science Inquiry

A great activity for the class is to calculate average atomic mass of an element called "vegium", which consists of three isotopes –beans, peas, and corn. Each isotope has a different mass. The handout for the lab can be found at the link below.

http://www.dsisd.txed.net/documentcenter/view/15736

An alternative to this activity uses candy instead of vegetables, so the element is called "cadmium". You can use plain m & m's, skittles, and Reese's pieces as the three isotopes. The procedure of the lab is the same as for the vegium lab above.

Differentiated Instruction

Some students will grasp the parts of the atom better by making models. This can be done very simply or with much more flair. A simple way is to use paper and three different colored stickers purchased from an office supply store. The paper needs to have a circle in the middle to represent the nucleus. Assign students a particular isotope (the most common one is best), and instruct them to use the stickers to represent protons, neutrons, and electrons.

The link below is for a much more elaborate model-building activity using fishing line and marshmallows. Adapt it to other materials if you like.

http://serc.carleton.edu/sp/mnstep/activities/26414.html

Enrichment

Use the internet to search out elements that have more than one naturally occurring isotope. Assign different students or pairs of students different elements. Have them prepare a large summary of all the naturally occurring isotopes

of their element on butcher paper. Under the listing of each isotope, they should note what distinct use (if any) that isotope has. For example, iodine-131 is used extensively in the medical field for diagnosis and treatment of thyroid disorders. Finally, have them use their listing to calculate the average atomic mass of their element and compare to a periodic table. Web Elements is a great internet tool for this type of activity.

http://www.webelements.com

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 4.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 4.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The electrons in an atom are particularly important because it is their arrangement and behavior that dictates the types of compounds that element is likely to form. The atomic model has changed over time in terms of the understanding of electrons.

- How is an understanding of the behavior of light related to electrons?
- What is the Bohr model of the atom?



Electrons in Atoms

Chapter Outline

- 5.1 LIGHT
- 5.2 THE QUANTUM MECHANICAL MODEL
- 5.3 ELECTRON ARRANGEMENT IN ATOMS



Chip Griffin. www.flickr.com/photos/chipgriffin/2210297177/. CC BY 2.0.

Chapter Overview

Rutherford's gold foil experiment established the nucleus of the atom as the source of the positive charge and most of the mass. The nature of the electrons was still unknown. Experiments by Planck and Einstein led to the notion that light exhibits a wave-particle duality. The particle nature of light means that the energies that can be absorbed or released by atoms are quantized. The Bohr model of the atom explained the atomic emission spectrum of the hydrogen atom. Electrons also behave like both waves and particles. The quantum mechanical model of the atom is the modern accepted version of the atom, where electrons occupy orbitals of various shapes and energies, but their exact location is never known. An electron configuration is the specific arrangement of all the electrons in any atom.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 5.	1:	Pacing Lessons
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Lesson	Class Period(s) (60 min)
5.1 Light	2
5.2 The Quantum Mechanical Model	2
5.3 Electron Arrangement in Atoms	2-3

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 5.2 :	Chemistry	Concepts
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FlexBook® Lesson	Chemistry Concepts
5.1 Light	5.1 –5.7
5.2 The Quantum Mechanical Model	5.8 -5.13
5.3 Electron Arrangement in Atoms	5.14 -5.19

5.1 Light

Key Concept

In order to understand the nature of electron behavior in atoms, it is necessary to understand that light possesses both a wavelike nature and a particle nature. Visible light is one small part of the entire electromagnetic spectrum. Quantum theory states that the energy that is lost or gained by atoms occurs in discrete units called quanta. The photoelectric effect demonstrates the particle nature of light. The Bohr model of the atom is one in which electrons travel around the nucleus in well-defined circular orbits. Transitions between energy levels are possible and produce atomic emission spectra.

Standards

Lesson Objectives

- Describe the relationships between speed, wavelength, and frequency of light.
- Understand the photoelectric effect and how it is related to the wave-particle duality of light.
- Describe how changes in electron energies lead to atomic emission spectra.
- Describe the Bohr model of the atom.

Lesson Vocabulary

- atomic emission spectrum: A pattern of lines formed when the light emitted from an atom passes through a prism to separate it into the different frequencies of light it contains.
- electromagnetic radiation: A form of energy that exhibits wavelike behavior as it moves through space.
- electromagnetic spectrum: All forms of electromagnetic radiation.
- excited state: A state of an atom where its potential energy is higher than the ground state.
- frequency: The number of waves that pass a certain point in a specified amount of time.
- ground state: The lowest energy state of the atom.
- photoelectric effect: A phenomenon that occurs when light shined onto a metal surface causes the ejection of electrons from that metal.
- photon: A particle of electromagnetic radiation that has zero mass and carries a quantum of energy.
- quantum: The minimum quantity of energy that can either be gained or lost by an atom.
- wavelength: The distance between any two corresponding points on adjacent waves.

Teaching Strategies

Introducing the Lesson

Students probably have some familiarity with waves, but you will want to make sure that they understand the variables involved with light waves: wavelength, frequency, and amplitude. Draw several different waves with differences in those quantities. Discuss the electromagnetic spectrum, stressing the very large variation in wavelength and frequency from one end to the other

Demonstrations

Several demonstrations will help your students understand continuous spectra and atomic emission spectra.

Pass a source of white light though a small hole and then through a prism. When the light is then displayed onto a white background, the continuous spectrum or rainbow will be visible.

If possible, obtain a gas discharge lamp and a set of gas discharge tubes filled with various gases. When discussing the excitation of electrons and the Bohr model, demonstrate with the use of the tubes. Students can view the colored gas tubes with a prism or inexpensive spectrograph in order to see the atomic emission spectrum for the particular element. The following video shows how the demonstration works.

https://www.youtube.com/watch?v=ioSiSA9spfQ

Common Misconceptions

The units for frequency sometimes cause problems for students when doing calculations. Stress the dimensional analysis of problems involving frequency and that the unit for frequency (hertz) is actually a "per second" (s⁻¹). For example, speed in m/s divided by frequency in s⁻¹ yield the wavelength in m because the per second unit cancels.

 $\frac{m/s}{s^{-1}} = m$

When Planck's constant, with a unit of joule-seconds (J•s) is multiplied by frequency in s^{-1} , the resulting unit is energy in joules (J).

 $J \cdot s \times s^{-1} = J$

Science Inquiry

Flame tests are an indispensable activity for your students to perform in the lab. Salts or salt solutions of various metals are introduced into a Bunsen burner flame, then identified by the various colors that are produced. The colors are generated by the heat from the flame exciting electrons, which fall back to the ground state and emit light. There are several ways to do flame test labs, so choose whichever fits the materials in your lab.

These two labs use nichrome wires which are dipped into aqueous solutions: http://kaffee.50webs.com/Science/lab s/Chem/Lab-Flame.Tests.html

http://www.creative-chemistry.org.uk/activities/flametests.htm

This lab instead uses wood splints that have been soaked in distilled water. The splints are then dipped directly into the solid metal salts, so no solutions need to be made:

http://serc.carleton.edu/sp/mnstep/activities/34745.html

Differentiated Instruction

The calculations in this section could prove difficult for some students, especially due to the use of exponential notation. Pair students of lower math ability with more advance students when practicing the calculations using $c = \lambda v$ and E = hv. Make sure to help students with what keys to use on their calculators in order to correctly do the calculations with scientific notation. They may not know about the "EE" or "EXP" buttons and how to use them.

Enrichment

As stated in the text, Bohr was able to exactly predict the wavelengths of light in the atomic emission spectrum for the hydrogen. However, for atoms with more than one electron, the mathematics did not work. The Rydberg equation allows students to calculate the wavelengths of light that are produced as a result of any electron transitions for the hydrogen atom, based only on the initial and final principal quantum number, n_i and n_f respectively.

$$\Delta E = R_H (\frac{1}{n_i^2} - \frac{1}{n_j^2}), \text{ where } R_H = 2.178 \times 10^{-18} \text{ J (the Rydberg constant)}$$

Solving the above equation yields the difference in energy between the two levels, which can be followed by calculation of the frequency and wavelength of light that result from any electron transition. Note that for electromagnetic radiation to be emitted, n_i must be larger than n_f . This results in a negative value for ΔE , meaning that energy is being lost by the falling electron.

Have advanced students use the Rydberg equation to calculate the energies and wavelengths for the four electron transitions from the Balmer series that yield the four visible lines of hydrogen's atomic emission spectrum. Compare to Figure 5.8. There may be a very small difference in the results due to the speed of light being slightly less in air than in a vacuum.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 5.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 5.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

As with light, the electron has been shown to exhibit both wavelike behavior and particle behavior. The quantum mechanical model of the atom fundamentally changed the way in which electrons were viewed, as compared to the Bohr model.

- How can the wavelength of an electron be determined?
- What is the modern interpretation of where electrons exist within the atom?

5.2 The Quantum Mechanical Model

Key Concept

Just as with light, electrons also behave as both a wave and as a particle. The wavelength of an electron can be determined from the de Broglie wave equation. Quantum mechanics is used to describe the motion of subatomic particles. As a result of the Heisenberg uncertainty principle, the specific location of the electrons in an atom cannot be known, only the probability that an electron may be found in a certain region. The energies of all of the electrons in any atom can be described by the use of four quantum numbers.

Standards

Lesson Objectives

- Understand the de Broglie wave equation and how it illustrates the wave nature of the electron.
- Explain the difference between quantum mechanics and classical mechanics.
- Understand how the Heisenberg uncertainty principle and Schrödinger's wave equation led to the idea of atomic orbitals.
- Know the four quantum numbers and how they are related to the arrangement of electrons in an atom.
- Describe the interrelationships between principal energy level, sublevel, orbital and electron spin and how they relate to the number of electrons of an atom.

Lesson Vocabulary

- angular momentum quantum number: Indicates the shape of the orbital, symbolized by the letter l.
- Heisenberg uncertainty principle: A principle that stated that it is impossible to determine simultaneously both the position and the velocity of a particle.
- magnetic quantum number: Indicates the orientation of the orbital around the nucleus, symbolized by the letter ml.
- orbital: The three-dimensional region of space that indicates where there is a high probability of finding an electron.
- principal quantum number: Indicates the principal or main energy level occupied by the electron, symbolized by the letter n.
- quantum mechanical model: An atomic model that comes from the solution to Schrödinger's equation.
- quantum mechanics: The study of the motion of objects that are atomic or subatomic in size and thus demonstrate wave-particle duality.
- quantum numbers: The numbers that specify the properties of the atomic orbitals and the electrons in those orbitals.
- spin quantum number: Indicates the direction of electron spin, symbolized by the letter m_s.

Teaching Strategies

Introducing the Lesson

Set up the lesson with a brief discussion of the difference between classical mechanics and quantum mechanics. Ask the students about the concepts of classical mechanics that they may learn about in a physics class –velocity, acceleration, etc. With large everyday objects (cars, thrown balls, etc.) that travel much slower than the speed of light, the laws of physics are pretty straightforward. However, with submicroscopic objects traveling very fast, the "laws" are changed and these objects simply don't behave in the same ways that large objects do. There is a process called tunneling, which is where a particle can move from one place to another without ever being anywhere in between. The students may think of it as teleportation. Electrons and other small particles frequently move by tunneling.

Common Misconceptions

Some students may misunderstand the meaning of the electron cloud, as in the lab activity (under Science Inquiry here) or Figure 5.14 in the text. Stress that the dots they see in an electron cloud do not represent different electrons, but rather, places where an electron could be at any given moment. Heisenberg's uncertainty principle means that it's not possible to know the exact location. Hydrogen, with one electron, has an electron cloud.

Some students may get hung up on all of the various shapes of the orbitals. Inform them that those shapes come about from solutions to the wave function for many electron atoms, math that is far beyond the scope of an introductory course. If they pursue chemistry in college, they will learn more about orbital shapes in an upper level course in quantum mechanics.

Science Inquiry

A good way for students to visualize the quantum mechanical model of the atom is to perform the following activity. The students generate a facsimile of an electron cloud by dropping a marker on to a target and graphing the results. The dots simulate electron density, which should decrease as distance from the nucleus increases.

download the resource from: http://mrsj.exofire.net/chem/docs/eprob_lab.doc

Differentiated Instruction

Some students may benefit from making models in order to help visualize the shapes of the atomic orbitals. Assign pairs or small groups different orbitals. Keep in mind that d and f orbitals are more complicated and will take more time and care to build accurately. Students can use materials of their own choosing.

Enrichment

Instruct interested students to research the life and work of Erwin Schrödinger. His work on quantum mechanics is most responsible for the modern model of the atom. The students can prepare an oral report with a PowerPoint presentation.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 5.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 5.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Chemistry and chemical reactions are very much concerned with electrons. Electrons are the particles that are responsible for chemical bonds between atoms. During some reactions, certain elements lose electrons, while other elements gain electrons.

- How are electrons arranged around the nuclei of atoms?
- Can chemical and physical properties of elements be explained by their electron arrangement?

5.3 Electron Arrangement in Atoms

Key Concept

Electron configurations refer to the specific arrangement of electrons in every atom. Electron configurations can be determined by following the Aufbau principle, the Pauli exclusion principle, and Hund's rule. Each element has a unique electron configuration. Valence electrons, the electrons in the outermost principal energy level, are responsible for the chemical behavior of elements.

Standards

Lesson Objectives

- Understand how to apply the Aufbau principle, the Pauli exclusion principle, and Hund's rule to determine ground state electron configurations.
- Be able to write correct orbital filling diagrams and electron configurations for all elements.
- Know how to use the noble gas notation shorthand method.
- Be able to determine the number of valence electrons and the number of unpaired electrons in any atom.
- Understand that some electron configurations are exceptions to the normal Aufbau process.

Lesson Vocabulary

- Aufbau principle: A principle which states that an electron occupies orbitals in order from lowest energy to highest.
- electron configuration: The arrangement of all of the electrons of that atom.
- Hund's rule: A rule that states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and that each of the single electrons must have the same spin.
- noble gas notation: Consists of the elemental symbol of the last noble gas prior to that atom, followed by the configuration of the remaining electrons.
- Pauli exclusion principle: A principle that states that no two electrons in an atom can have the same set of four quantum numbers.
- valence electrons: Electrons in the highest occupied principal energy level of an atom.

Teaching Strategies

Introducing the Lesson

Review the four quantum numbers and have copies of Table 5.2 available for reference or recreated on the board. Progress by increasing atomic number, starting with hydrogen and ending with carbon. Draw the orbital filling diagrams for each on the board. Emphasize the difference between principal energy level and sublevel. Discuss Hund's rule, which becomes relevant starting with carbon. Have the students practice with a few larger elements, using Figure 5.19 as a guide. Once students become comfortable with drawing orbital filling diagrams, proceed on to writing the electron configurations.

Common Misconceptions

Students may be confused about the sublevel order of fill after argon (Z = 18). At this point, the 4s fills before the 3d because it is of slightly lower energy. Beyond this point, the order of fill becomes irregular. Encourage students to use the diagram in Figure 5.19 to help with the electron configurations of larger elements. Either provide this diagram for student use during assessments or teach them how to recreate it.

Make sure to emphasize that valence electrons are only those in the outermost principal energy level. This can be confusing for elements whose d or f sublevel is partially full. Since those sublevels are never in the outermost principal energy level, they are never valence electrons. The maximum number of valence electrons is eight, corresponding to ns^2np^6 .

Science Inquiry

The website below is a comprehensive examination of electron configurations, meant to be performed by the class as a group activity. Students write out several orbital filling diagrams, electron configurations, and noble gas configurations. They answer questions based on their work.

http://kaffee.50webs.com/Science/activities/Chem/Activity.Electron.Coonfiguration.html

Differentiated Instruction

The link below is to another electron configuration activity. It uses the analogy of an atom being like a boarding house with the electrons as the boarders. It may be helpful for students having difficult with orbital filling diagrams or as extra practice for the entire class.

http://gradnation.org/sites/default/files/Electron%20Configurations.pdf

Enrichment

Advanced students can learn to assign the four quantum numbers to any electron in any atom. The website below is a good place to visit to learn how to do this.

http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch6/quantum.html

For example, the last electron of an iron atom is the 6th electron of the 3d sublevel. In an orbital filling diagram, it would be in the first d orbital (of five), which can be assigned the m_l value of -2. Recall that the possible ml values for a set of five d orbitals is -2, -1, 0, +1, +2. The arrow in the orbital would point downward, corresponding to an m_s vale of $-\frac{1}{2}$. The four quantum numbers for this electron can then be written in the following format of (n, l, m_l, m_s) .

26th electron of Fe = $(3, 2, -2, -\frac{1}{2})$

Assign students or pairs of students a collection of different elements and have them write the four quantum numbers for the last electron in each of their configurations.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 5.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 5.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The periodic table is a diagram showing each and every known element in a very specific arrangement. That arrangement is related to electron configurations and chemical reactivity.

- How was the periodic table developed and by whom?
- What tools are used to identify the placement of elements in the modern periodic table?

The Periodic Table

Chapter Outline

CHAPTER

- 6.1 HISTORY OF THE PERIODIC TABLE
- 6.2 ELECTRON CONFIGURATION AND THE PERIODIC TABLE
- 6.3 PERIODIC TRENDS



User:HazelBlk/Wikimedia Commons. commons.wikimedia.org/wiki/File:PknMitre10MEGAinterior.JPG. CC BY 3.0.

Chapter Overview

The origin of the periodic table came about as Mendeleev and other scientists attempted to organize elements that displayed similarities in chemical reactivity. The modern periodic table is arranged by increasing atomic number and is organized into periods and groups. Electron configurations help to explain why the table has the exact form that it does. Periodic trends are systematic variations in properties of elements and include, atomic radius, ionization energy, ionic radius, and electronegativity.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

 TABLE 6.1: Pacing Lessons

Lesson	Class Period(s) (60 min)
6.1 History of the Periodic Table	1
6.2 Electron Configuration and the Periodic Table	2
6.3 Periodic Trends	2-3

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 6.2: Chemistry Concepts

FlexBook® Lesson	Chemistry Concepts
6.1 History of the Periodic Table	6.1 –6.7
6.2 Electron Configuration and the Periodic Table	6.8 -6.14
6.3 Periodic Trends	6.15 -6.21

6.1 History of the Periodic Table

Key Concept

Scientists in the 1800s began to recognize that some elements seemed to belong in the same category because of their similarity in chemical reactivity. Dmitri Mendeleev proposed a periodic table arranged in order of increasing atomic mass of the elements, resulting in columns of elements with similar chemical behavior. Improvements were made to the table, ordering the elements by atomic number rather than atomic mass. Elements in the modern periodic table belong to specific periods and groups. Elements can also be categorized as metals, nonmetals, or metalloids.

Standards

Lesson Objectives

- Describe some of the early attempts to organize the chemical elements.
- Understand how Mendeleev organized his periodic table.
- Know the improvements that Moseley made to Mendeleev's table.
- Know the periodic table.
- Describe various components of the modern periodic table, including periods, groups, metals, nonmetals, and metalloids.

Lesson Vocabulary

- group: A vertical column of the periodic table.
- metal: An element that is a good conductor of heat and electricity.
- metalloid: An element that has properties that are intermediate between those of metals and nonmetals.
- nonmetal: Element that is, generally, a poor conductor of heat and electricity.
- period: A horizontal row of the periodic table.
- periodic law: A law that states that when elements are arranged in order of increasing atomic number, there is a periodic repetition of their chemical and physical properties.
- periodic table: The arrangement of the elements in order of their atomic numbers so that elements with similar properties appear in the same vertical column or group.

Teaching Strategies

Introducing the Lesson

Make sure that you have a large periodic table on the wall in the classroom. It is also useful to have a laser pointer, but a long pointer such as a meter stick will also work. Begin with open-ended questions to the class about why the periodic table has such a strange shape. See if you can lead them into figuring out why the first row of the table has only two elements (H and He), while the second row has eight elements. Point out that the two rows that are pulled out at the very bottom actually fit in to the rest of the table when atomic numbers are taken into account. Count the number of elements in each block (s, p, d, f) and see if anyone can figure out the significance of the numbers 2, 6, 10, and 14. Finally, open a discussion of the history of the periodic table by pointing out that early efforts to organize the table were attempted in the early 1800s, well before the discovery of subatomic particles. Ask the class how they think that could have been done? (Answer: observations of chemical reactions)

Demonstration

Obtain small samples of sodium and potassium metal. Carefully show the samples to the class, but do not allow anyone to touch them. Discuss how Mendeleev and others before him arranged the known elements into groups based on chemical reactivity. Place the metal samples into separate beakers of water. This should be done under a fume hood. The metals will each react very rapidly with the water, giving off hydrogen gas. Ask if other well-known metals react with water in this way. Metals with which they will be familiar, such as iron and copper, do not react with water.

Common Misconceptions

Since Mendeleev has garnered the vast majority of the credit for being the "father of the periodic table", students may get the mistaken notion that he originated the idea. Stress to them that there were others who had attempted to produce a table before Mendeleev (Dobereiner, Newlands). Mendeleev's idea for a periodic table was accepted while those of the others was not. Part of the reason was simply the passage of time, but it was also due to the fact that more elements were being discovered and included in the table and because of the gaps that Mendeleev correctly put into his table.

Science Inquiry

Perform the following activity as a lead-in to looking at physical properties of elements and trends, much as the early chemists did.

- 1. Have the students make a five-column table. In the first column, list the elements lithium, sodium, potassium, rubidium, and cesium. The other four columns should be titled: atomic number, density, melting point, and boiling point.
- 2. Repeat step 1 for these four elements: fluorine, chlorine, bromine, and iodine.
- 3. Use the internet to look up the physical property data to fill in the table. The Web Elements site will provide the answers.

http://www.webelements.com

With the class, discuss the trends that they see for density, melting point, and boiling point. In other words, do the values for these properties increase or decrease as the atomic number within the group increases? Ask the class why

6.1. History of the Periodic Table

they think the densities for the second table show a much wider variation than the densities for the first table. (The physical states are different.)

Differentiated Instruction

The following website contains some games based on the periodic table that students may find entertaining and informative. Games include hangman, word searches, crossword puzzles, and others.

http://education.jlab.org/itselemental/

Enrichment

Have interested students prepare a biographical summary of the life of Dmitri Mendeleev. He made many contributions to life and science besides the periodic table. Students can make a poster or write a report on his accomplishments.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 6.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 6.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The periodic table is more than simply a collection of elements arranged by similarities in chemical reactivity. The information you know about quantum numbers and electron configurations can be integrated into a more complete understanding of the table.

- What is the order of fill of electron sublevels?
- What are valence electrons and how can the number of valence electrons in an atom be determined?

6.2 Electron Configuration and the Periodic Table

Key Concept

Each successive period on the periodic table is the beginning of a new principal energy level of electrons. The number of elements in each period corresponds to the number of electrons that are required to fill the sublevels. The periodic table can be broken up into four blocks, indicating what type of sublevel is in the process of being filled. Representative elements include the s-block and the p-block. Transition elements occupy the d-block. The f-block contains the lanthanides and actinides, which are pulled out below the main body of the periodic table.

Standards

Lesson Objectives

- Understand the relationship between the number of orbitals in various energy sublevels and the length of the periods in the periodic table.
- Identify each block of the periodic table and be able to determine which block each element belongs to based on its electron configuration.
- Describe the relationship between outer electron configuration and group number. Be able to determine the number of valence electrons for any element.
- Locate the following groups on the periodic table: alkali metals, alkaline earth metals, halogens, and noble gases.

Lesson Vocabulary

- actinide: A series of 14 elements from thorium (atomic number 90) to lawrencium (atomic number 103).
- alkali metal: An element in Group 1 (lithium, sodium, potassium, rubidium, cesium, and francium).
- alkaline earth metal: An element of Group 2 (beryllium, magnesium, calcium, strontium, barium, and radium).
- halogen: An element of Group 17 (fluorine, chlorine, bromine, iodine, and astatine).
- inner transition metal: An element in the lanthanide or actinide series.
- lanthanide: A series of 14 elements from cerium (atomic number 58) to lutetium (atomic number 71).
- noble gas: An element of Group 18 (helium, neon, argon, krypton, xenon, and radon).
- representative (main-group) elements: All of the elements that comprise the *s*-block and the *p*-block.
- transition element: an element that is found in Groups 3-12 on the periodic table.

Teaching Strategies

Introducing the Lesson

Divide the class into small groups. Assign each group three elements. Make sure that the elements are from the same group, such as boron, gallium, and indium. Ask them to write out the electron configurations for each element. When everyone is finished, see if they can detect similarities in the electron configurations. Make sure that you avoid any of the electron configuration exceptions like chromium or copper.

Common Misconceptions

Students are prone to making mistakes determining information about an element solely from its electron configuration, but this is an important skill. For example, suppose that a student is examining the configuration of the element technetium. In some textbooks or other sources, the configuration may be written as $[Kr]5s^24d^5$, where the principal energy levels are in the order of fill rather than in ascending order. Students may see this configuration and mistakenly think that it belongs to period 4. They may have difficulty with correctly determining the group number of p-block elements from the configuration because they forget to add ten to the sum of the superscripts. For example, the element with a configuration of $[Ne]3s^23p^5$ belongs to group 17 and not group 7.

Science Inquiry

A number of online activities are available to initiate a study of the periodic table and its organizational principles. Two are listed below. The first is a more basic activity so that students can learn the different parts of the table. The second activity is more advanced and links elemental electron configurations to periodic table placement.

http://geraghty.weebly.com/uploads/1/3/5/6/13564507/chemistry-periodic_table_activity.pdf

Download the resource: http://storerscience.weebly.com/uploads/2/1/3/4/21347444/lab_electron_configuration_activity.doc.

Differentiated Instruction

Allow small groups of less proficient students to use Figure 6.8 to practice writing full electron configurations for elements. Once they have written the configuration, have them identify the period, group, and block in which the elements is found (using Figure 6.9 as another aid). Have them explain to each other how the configuration matches up with the other information.

Enrichment

Once students have become comfortable with the periodic table and the electron configurations, give them a blank periodic table –only empty boxes. Project a copy on to the wall. Pick various boxes at random and challenge students to be able to write the electron configuration of the element that belongs in that box. Do not choose configuration exceptions.

Have advanced students use a periodic table with electron configurations provided. Assign individuals one or two elements that have unusual electron configurations that do not strictly follow the Aufbau principle. Have them research the reason why their configurations are different and how that changes their chemical properties.
Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 6.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 6.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

So far you have learned that the periodic table is based on both electronic structure and similarities in chemical reactivity. Next we will consider how some important physical properties of elements show distinct trends within periods and groups.

- What determines the size of an atom?
- If a valence electron were removed from an atom, what charge would the resulting particle have?

6.3 Periodic Trends

Key Concept

The structure of the periodic table is such that certain trends exist in the physical properties of the elements. Atomic radii decrease left to right across periods and increase top to bottom down groups. Ionization energy, the energy required to remove an electron from an atom, increases across periods and decreases down groups. Removing multiple electrons always takes more energy, especially once the noble gas core is reached. Electron affinity is the energy change that occurs when an atom gains an electron. Cations are always smaller than the parent atom, while anions are always larger. Electronegativity, the ability to attract shared electrons, increases across periods and decreases down groups. The most reactive metals are in the lower left portion of the table, while the most reactive nonmetals are in the upper right portion.

Standards

Lesson Objectives

- Learn the periodic trends for atomic radius.
- Know the relationship between group number and valence electrons.
- Describe how ions are formed.
- Learn the periodic trends for ionization energy.
- Explain how multiple ionization energies are related to noble gas electron configurations.
- Describe electron affinity.
- Predict the effect that ion formation has on the size of an atom.
- Learn the periodic trends for electronegativity.

Lesson Vocabulary

- anion: A negatively charged ion.
- atomic radius: One-half the distance between the nuclei of identical atoms that are bonded together.
- cation: A positively charged ion.
- electron affinity: The energy change that occurs when a neutral atom gains an electron.
- electronegativity: A measure of the ability of an atom to attract the electrons when the atom is part of a compound.
- ion: An atom or group of bonded atoms that has a positive or negative charge.
- ionization energy: The energy required to remove an electron from an atom.

Teaching Strategies

Introducing the Lesson

Discuss the meaning of trend and emphasize that trends can have exceptions. The key to understanding the trends is to know the meaning of nuclear charge and to remember that successive principal energy levels consist of larger orbitals. Use visuals on the board to emphasize atomic radii differences relating to nuclear charge and number of energy levels. For this purpose, atoms can be drawn in a Bohr-model fashion.

Make sure that students understand ion formation before discussing ionization energy and ionic sizes. Stress that ions will be very important in future study of chemical compounds.

Demonstration

The link below is for a lab activity that shows the reactivity of the halogen group. The goal is to demonstrate that chlorine is more reactive than bromine, which is more reactive than iodine. The specifics of the lab involve being able to write equations for single replacement reactions, which is covered in a later chapter. However, the activity can be adapted to a demonstration here. If so, perform the displacement reactions in test tubes rather than well plates. The demonstration reinforces the information from the end of the lesson that the more reactive nonmetals are located at the top of a given group.

http://www.nuffieldfoundation.org/practical-chemistry/reactions-aqueous-solutions-halogens

Common Misconceptions

Students tend to have difficulty with the concept of electron shielding and its effect on the properties discussed in the lesson. Pick a student from the back of the room and call him/her the valence electron. Call yourself the nucleus. There is an attraction between the two of you. However, all the other students are the inner electrons and are getting in the way of that attraction. Therefore, the "valence electron" student will find it much easier to leave (be ionized).

Taking it Further

The concept of effective nuclear charge (Z_{eff}) can be used in concert with electron shielding. In simplified form the effective nuclear charge is equal to the atomic number (Z) minus the number of inner or non-valence electrons. For example, lithium has 3 protons, 2 inner electrons, and 1 valence electron. For lithium, the $Z_{eff} = 3 - 2 = 1$.

The group trend for effective nuclear charge is that it is constant within a group. The element cesium is in the same group as lithium. Its effective nuclear charge is $Z_{eff} = 55 - 54 = 1$. This is the same value as lithium. Despite the huge increase in nuclear charge, the inner electrons "cancel out" most of that attractive force and leave the valence electron of cesium very susceptible to removal: even more so than for lithium. The low Z_{eff} of the alkali metals is why they have the lowest ionization energies of any group.

The period trend for effective nuclear charge is that it increases by one across a period. For example, the element nitrogen is in the same period as lithium. It has 7 protons, 2 inner electrons, and 5 valence electrons. The $Z_{eff} = 7 - 2 = 5$. The higher effective nuclear charge means that nitrogen holds on to its valence electrons more tightly and thus has a higher ionization energy. Ionization energy increases from left to right across a period because the effective nuclear charge increases.

Effective nuclear charge is more complicated than what has been discussed here, and takes into account sublevel differences as well. However, the use of the concept may help students understand electron shielding.

Science Inquiry

Students can gain insight into the trends by making graphs of the various properties as a function of atomic number. Since these graphs are shown and discussed in the FlexBook® student edition, it is best to do this activity as an introduction to the chapter. The link below contains directions for making four graphs. If time is limited, do only atomic radius and ionization energy.

Download the resource: http://mcs.monet.k12.ca.us/schools/TeacherWebsite/7-12/Durham.K/Worksheets%20preAP/C hapter%204/Periodic%20Trends%20graphing%20lab.doc

Differentiated Instruction

Have students practice their knowledge of trends by giving them trios of elements, either all in the same group or all in the same period. For example, Na, S, and Al or P, Sb, and N. Have them order the elements by increasing atomic radius, increasing ionization energy, increasing electronegativity. When they have mastered that, challenge them with more elements that they can still figure out just by knowing the trends.

Enrichment

The Alien Periodic Table is an entertaining activity for all students. It allows them to use their knowledge of the periodic table to construct one from clues given to them about so-called alien elements.

http://www.nclark.net/alienperiodictable___kulis.pdf

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 6.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 6.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Compounds result from the chemical combinations of elements. The nature of chemical compounds depends on the types of elements that are combining.

- What type of compounds results when a metal reacts with a nonmetal?
- What type of compound results when two nonmetals react with each other?



Chemical Nomenclature

Chapter Outline

- 7.1 IONS
- 7.2 IONIC COMPOUNDS
- 7.3 ACIDS AND BASES



Jacques-Louis David. commons.wikimedia.org/wiki/File:Portrait_of_Antoine-Laurent_Lavoisier_and_his_wife.jpg. Public Domain.

Chapter Overview

Nomenclature is the systematic technique of naming compounds based on well-established rules. The nomenclature of inorganic compounds depends on whether the substance is an ionic compound or a molecular compound. The formula of an ionic compound must involve a ratio of cations to anions that makes the compound electrically neutral. Ionic compounds may be binary or ternary, in which one or more polyatomic ions is present. Molecular compound nomenclature uses prefixes to indicate the number of each atom present in the compound. Acids and bases are an important class of compounds that have specific nomenclature rules as well.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

Pacing the Lessons

TABLE 7	1: Pacing	Lessons
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Lesson	Class Period(s) (60 min)
7.1 Ionic Compounds	3-4
7.2 Molecular Compounds	1
7.3 Acids and Bases	1

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 7.2: Chemistry Concepts

FlexBook® Lesson	Chemistry Concepts
7.1 Ionic Compounds	7.1-7.10
7.2 Molecular Compounds	7.11
7.3 Acids and Bases	7.12-7.13

7.1 lons

Key Concept

An electron diagram is a representation of an atom and its valence electrons. When metals react with nonmetals, one or more electrons are transferred from the metal atoms to the nonmetal atoms. The number of electrons transferred is dictated by the octet rule. The most common ions formed by elements is constant within a group on the periodic table. Atoms and ions that have the same electron configuration are isoelectronic. Transition metals form positive ions by losing electrons, but often cannot become isoelectronic with a noble gas.

Standards

Lesson Objectives

- Be able to determine the number of valence electrons for any element and draw an electron dot diagram for any atom.
- Use the octet rule to predict the charges of the most common ions formed by the representative elements.
- Write electron configurations for ions.
- Identify other atoms or ions that are isoelectronic with a particular ion.
- Know that transition metal ions with either half-filled or completely filled d sublevels are particularly stable.

Lesson Vocabulary

- electron dot diagram: A diagram in which the valence electrons of an atom are shown as dots distributed around the element's symbol.
- isoelectronic: An atom and an ion of a different atom, or two different ions, that have the same electron configuration.
- octet rule: A rule that states that atoms tend to form compounds in ways that give them eight valence electrons and thus the electron configuration of a noble gas.

Teaching Strategies

Introducing the Lesson

Students learned about ions in *Chemical Nomenclature* in the context of writing correct formulas for ionic compounds. Start the lesson by naming these six representative elements (Na, F, Mg, O, Al, N) and having students identify their most common ion charge. Ask why there is only one charge possible for these elements (noble gas electron configuration). Which noble gas? Write the electron configuration and make the point that ions for all of these elements are isoelectronic with neon. Ask which other set of six ions would be isoelectronic, this time with argon.

Common Misconceptions

Students generally grasp the concept of common ions and noble gas electron configuration easily. However, as noted in the text, make sure to stress that two ions/atoms that are isoelectronic are not identical. The charge difference is very important and so the physical and chemical behavior of isoelectronic substances is different as well.

Students will often make errors when writing electron configurations of transition metal ions by removing electrons from the (n-1)d sublevel before removing the valence electrons form the ns sublevel. They will need to practice these configurations.

Science Inquiry

Students can begin this lesson by practicing the common ion skills that they achieved during *Chemical Nomenclature* when writing formulas for ionic compounds. The links below are worksheets that students should be able to complete effectively.

https://colemanchemistry.wikispaces.com/file/view/IonWorksheet.pdf

Download the resource: http://www.gcisd-k12.org/cms/lib/TX01000829/Centricity/Domain/745/writing%20binary%2 0ionic%20formulas.doc.

Differentiated Instruction

For students that may have difficulty envisioning the relationship of ion charge to the relative numbers of protons and electrons in an atom, the simulation below may prove useful. It allows students to "build" atoms and ions by adding in subatomic particles.

http://phet.colorado.edu/en/simulation/build-an-atom

There are many possible assignments already made that accompany the simulation. These can be found on the main page above. The link below takes you to one assignment that required students to build specific atoms and ions and asks questions concerning the atomic structures that result. The assignment may work well here, or could be broken up into sections and assigned earlier with *Electrons in Atoms* and *The Periodic Table*.

http://phet.colorado.edu/en/contributions/view/3613

Enrichment

Advanced students can investigate the transition metals and the ions that they form in more detail. Assign students a set of several transition metals. Students can use the internet to find out what are all the possible ion charges for those metals. Write the electron configurations that would result. Identify ionic compounds in which the transition metals have adopted each of those charges.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 8.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 8.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Ionic compounds adopt the structure of an extended, three-dimensional lattice of alternating positive and negative ions held together by electrostatic attractive forces.

- How strong is an ionic crystal?
- Is an ionic crystal malleable or brittle? Why?
- Will ionic compounds conduct an electric current?

7.2 Ionic Compounds

Key Concept

Chemical compounds are either organic if they contain carbon, or inorganic if they do not contain carbon. Inorganic compounds may either be molecular or ionic. Ionic compounds, which consist of oppositely charged ions, attain the structure of an extended three-dimensional crystal lattice. The overall charge of any compound is neutral, so the ratio of ions in the crystal must reflect that neutrality. Binary ionic compounds consist of a monatomic metal cation and a monatomic nonmetal anion. Many transition metal ions may adopt multiple charges, so the Stock system is used for naming compounds with transition metal ions. A polyatomic ion is one that contains two or more atoms bonded together. Ionic compounds with polyatomic ions are called ternary ionic compounds. The crisscross method can be used to write correct formulas for all ionic compounds.

Standards

Lesson Objectives

- Distinguish between inorganic and organic chemistry.
- Interpret a molecular formula.
- Explain why an ionic compound is represented by an empirical formula.
- Be able to determine the charges of monatomic ions formed by the representative elements from the position of each element on the periodic table.
- Use the Stock system to identify the charge of transition metal ions.
- Name an ionic compound given its formula.
- Write the correct formula for an ionic compound given its name.

Lesson Vocabulary

- binary ionic compound: An ionic compound composed of a monatomic metal cation and a monatomic nonmetal anion.
- empirical formula: A formula that shows the elements in a compound in their lowest whole-number ratio.
- inorganic chemistry: A branch of chemistry that is the study of chemicals that do not, in general, contain carbon.
- molecular formula: Formula of a molecular compound that shows the kinds and number of atoms present in a molecule of the compound.
- monatomic ion: Ion that consists of a single atom with either a positive or negative charge.
- organic chemistry: A branch of chemistry that is the study of carbon-containing chemicals.
- polyatomic ion: An ion composed of more than one atom.
- ternary ionic compound: An ionic compound composed of three or more elements.

Teaching Strategies

Introducing the Lesson

Establish that this chapter is all about the nomenclature of inorganic compounds and that organic chemistry is covered in a later chapter. Have several models on hand: (1) molecular compounds like water and ammonia, and (2) an ionic compound like sodium chloride. Ask the class what is different about the models. Write the chemical formulas on the board. Ask the class what is different about the elements that make up each compound. Lead them to seeing that ionic compounds consist of a metal and a nonmetal. Write several more formulas on the board, some of molecular compounds and some of ionic. Have the students classify them correctly. This leads into the teaching of ions, how they form, and how they are named.

Common Misconceptions

Some students may have trouble with the formulas of ternary ionic compounds. If using the crisscross method, emphasize that the subscripts of polyatomic ions are a permanent part of that ion and cannot be changed in any way. As an example, consider the method as applied to iron(III) sulfate:

 Fe^{3+} $\mathrm{SO_4}^{2-}$

Some students may make an error with the crisscross and come up with Fe_2SO_3 , by replacing the 4 subscript on the O with the 3 from the charge of the iron. Stress that the 4 has to stay with the oxygen and the 3 goes outside parentheses: $Fe_2(SO_4)_3$.

Most students do well with each category of nomenclature in this chapter when taught on its own. The challenge becomes when all types of names and formulas are brought together. This is an important skill. When finishing this lesson, make sure to combine all types of ionic compounds together and require students to write correct names and formulas.

Science Inquiry

Chemical nomenclature can be taught as an inquiry activity. Students should have an introduction to ions, how they are formed, and how they are named. The focus of the activity is for students to learn how to put ions together to make neutral compounds and to distinguish between the different types of inorganic compounds: binary ionic, binary molecular, ternary ionic, acids and bases. The activity thus fits with all of *Chemical Nomenclature* of the FlexBook® student edition, not just lesson 7.1.

http://serc.carleton.edu/sp/pkal/mnscu/activities/38478.html

Differentiated Instruction

Have students make flashcards with the names and formulas of ions. Pair students and have them quiz each other with the flashcards. Memorizing the polyatomic ions is a challenging task. You may decide to limit the number of polyatomic ions that students are required to memorize.

Enrichment

Advanced students can research the concept of lattice energy. Lattice energy is defined as the energy that is required to break up an ionic crystal lattice into its gaseous ions.

$$NaCl_{(s)} + energy \rightarrow Na^+_{(g)} + Cl^-_{(g)}$$

Lattice energy is what leads to the stability of a crystal lattice. The higher the lattice energy is for a compound, the more stable the compound. This stability means that the melting point for the compound is higher. Have students prepare a brief report on lattice energy and compare the values for a variety of ionic compounds.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 7.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 7.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The large class of inorganic chemical compounds can be sub-classified into ionic compounds and molecular compounds.

- How is the structure of a molecular compound different from the structure of an ionic compound?
- Relate molecular compounds to the law of multiple proportions.
- What is the nomenclature for molecular compounds?

7.3 Acids and Bases

Key Concept

Acids are molecular compounds that ionize in aqueous solution to a hydrogen ion and a corresponding anion. Bases are ionic compounds that dissociate in aqueous solution to a cation and the hydroxide ion. Acid nomenclature is based on the suffix of the name of the anion, with three rules to follow. Base nomenclature is like other ionic compounds: the name of the cation followed by "hydroxide".

Standards

Lesson Objectives

- Define acids and bases in terms of the ions that are produced when each type of compound is dissolved in water.
- Be able to name acids when given their chemical formulas.
- Write the correct formula for an acid when given its name.
- Name and write formulas for bases.

Lesson Vocabulary

- acid: A molecular compound that contains one or more hydrogen atoms and produces hydrogen ions when dissolved in water.
- base: An ionic compound that produces hydroxide ions when dissolved in water.
- binary acid: An acid that consists of hydrogen and one other element.
- oxoacid: Acid that consists of hydrogen, oxygen, and a third element.

Teaching Strategies

Introducing the Lesson

Have an example of an acid and a base on hand when introducing the topic. Hydrochloric acid and sodium hydroxide will work. Most students will be familiar with the concept of pH. Ask if they know what is neutral on the pH scale, what is acidic, and what is basic. Ask if they know of any everyday substances that are acids and bases. This lesson only concerns the nomenclature of these important classes of compounds. You will return to acids and bases in detail in a later chapter.

Demonstration

Show that acids and bases are somewhat "opposite" in their behavior. Use litmus paper and pH paper to illustrate the idea of indicators. Show that equal amounts of acid and base will combine to make a neutral solution (pH = 7). Make sure that the concentrations of your acid and base solutions are the same.

Common Misconceptions

The nomenclature of acids is far more challenging than the nomenclature of bases. For some reason, many students seem to get attached to the "hydro-" prefix and want to use it for naming all acids. Stress the three rules and that for acids derived from anions that have "-ate" or "-ite" suffixes, there is no "hydro-" prefix.

Because acids are described as breaking apart into ions, students assume that acids are ionic compounds. Actually, they are molecular compounds that break into ions when dissolved in water. The nature of the dissolving process is covered in a later chapter. In fact, the acid naming rules apply only to the substance when dissolved in water. Consider the example below:

HBr(g) = hydrogen bromide (pure molecular compound)

HBr(aq) = hydrobromic acid (acid)

To become an acid, hydrogen bromide gas must be dissolved into water. At that point, the molecules break apart into hydrogen ions (H+) and bromide ions (Br-).

Science Inquiry

Continue with the activity described in Lesson 7.1: Ionic Compounds. This inquiry activity describes the nomenclature rules for acids and gives students practice naming and writing their formulas.

http://serc.carleton.edu/sp/pkal/mnscu/activities/38478.html

Differentiated Instruction

Less advanced readers should use the tables of anions from lesson 7.1 (Tables 7.1 and 7.5) to assist them in writing acid names and formulas. Have them make three columns on a sheet of paper, titled "ide", "ate", and "ite". Put all of the anions from the tables into the correct column. Then, with a partner, use the acid naming rules to turn each anion into the correct formula for the acid and the correct name. Share the results with other students to check for accuracy.

Enrichment

Provide advanced students with a selection of acids that are commonly found in foods. Have them research the structures of the acids and draw them on a butcher paper poster. They should include explanations of what the acids are used for and in what foods they are commonly found. Most of these will be organic acids. Some examples include citric acid, tartaric acid, acetic acid, lactic acid, malic acid, and benzoic acid.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 7.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 7.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The extended three-dimensional structure of ionic compounds is commonly called a crystal lattice. The interactions of the ions within a crystal lattice lead directly to many important physical and chemical properties of ionic compounds.

- How is a crystal lattice formed?
- Why is a crystal lattice the most stable structural form for an ionic compound?
- What physical properties do ionic compounds have in common as a result of their structure?

8 Ionic and Metallic Bonding

Chapter Outline

CHAPTER

- 8.1 IONS
- 8.2 IONIC BONDS AND IONIC COMPOUNDS
- 8.3 METALLIC BONDS



John W. Schulze (Flickr: Gruenemann). www.flickr.com/photos/gruenemann/5559115094/. CC BY 2.0.

Chapter Overview

The valence electrons of an atom can be shown with the use of an electron dot diagram. Ions form in order for atoms to obey the octet rule. Metal atoms lose electrons, while nonmetal atoms gain electrons. Transition metal atoms often cannot lose enough electrons to obtain a noble gas electron configuration. The transfer of one or more electrons from a metal atom to a nonmetal atom results in an ionic bond between the two oppositely charged ions. Ionic compounds have specific properties such as hardness and brittleness that result from the ionic bonding and crystalline structure. Pure metals also adopt a crystalline structure where valence electrons are delocalized and can move around the stationary metal cations.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 8.1: Pacing Lessons

Lesson	Class Period(s) (60 min)
8.1 Ions	1-2
8.2 Ionic Bonds and Ionic Compounds	1-2
8.3 Metallic Bonds	1

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 8.2: Chemistry Concepts

FlexBook® Lesson	Chemistry Concepts
8.1 Ions	8.1 -8.5
8.2 Ionic Bonds and Ionic Compounds	8.6 -8.9
8.3 Metallic Bonds	8.10 - 8.12

8.1 lons

Key Concept

An electron diagram is a representation of an atom and its valence electrons. When metals react with nonmetals, one or more electrons are transferred from the metal atoms to the nonmetal atoms. The number of electrons transferred is dictated by the octet rule. The most common ions formed by elements is constant within a group on the periodic table. Atoms and ions that have the same electron configuration are isoelectronic. Transition metals form positive ions by losing electrons, but often cannot become isoelectronic with a noble gas.

Standards

Lesson Objectives

- Be able to determine the number of valence electrons for any element and draw an electron dot diagram for any atom.
- Use the octet rule to predict the charges of the most common ions formed by the representative elements.
- Write electron configurations for ions.
- Identify other atoms or ions that are isoelectronic with a particular ion.
- Know that transition metal ions with either half-filled or completely filled d sublevels are particularly stable.

Lesson Vocabulary

- electron dot diagram: A diagram in which the valence electrons of an atom are shown as dots distributed around the element's symbol.
- isoelectronic: An atom and an ion of a different atom, or two different ions, that have the same electron configuration.
- octet rule: A rule that states that atoms tend to form compounds in ways that give them eight valence electrons and thus the electron configuration of a noble gas.

Teaching Strategies

Introducing the Lesson

Students learned about ions in *Chemical Nomenclature* in the context of writing correct formulas for ionic compounds. Start the lesson by naming these six representative elements (Na, F, Mg, O, Al, N) and having students identify their most common ion charge. Ask why there is only one charge possible for these elements (noble gas electron configuration). Which noble gas? Write the electron configuration and make the point that ions for all of these elements are isoelectronic with neon. Ask which other set of six ions would be isoelectronic, this time with argon.

Common Misconceptions

Students generally grasp the concept of common ions and noble gas electron configuration easily. However, as noted in the text, make sure to stress that two ions/atoms that are isoelectronic are not identical. The charge difference is very important and so the physical and chemical behavior of isoelectronic substances is different as well.

Students will often make errors when writing electron configurations of transition metal ions by removing electrons from the (n-1)d sublevel before removing the valence electrons form the ns sublevel. They will need to practice these configurations.

Science Inquiry

Students can begin this lesson by practicing the common ion skills that they achieved during *Chemical Nomenclature* when writing formulas for ionic compounds. The links below are worksheets that students should be able to complete effectively.

https://colemanchemistry.wikispaces.com/file/view/IonWorksheet.pdf

Differentiated Instruction

For students that may have difficulty envisioning the relationship of ion charge to the relative numbers of protons and electrons in an atom, the simulation below may prove useful. It allows students to "build" atoms and ions by adding in subatomic particles.

http://phet.colorado.edu/en/simulation/build-an-atom

There are many possible assignments already made that accompany the simulation. These can be found on the main page above. The link below takes you to one assignment that required students to build specific atoms and ions and asks questions concerning the atomic structures that result. The assignment may work well here, or could be broken up into sections and assigned earlier with *Electrons in Atoms* and *The Periodic Table*.

http://phet.colorado.edu/en/contributions/view/3613

Enrichment

Advanced students can investigate the transition metals and the ions that they form in more detail. Assign students a set of several transition metals. Students can use the internet to find out what are all the possible ion charges for those metals. Write the electron configurations that would result. Identify ionic compounds in which the transition metals have adopted each of those charges.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 8.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 8.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Ionic compounds adopt the structure of an extended, three-dimensional lattice of alternating positive and negative ions held together by electrostatic attractive forces.

- How strong is an ionic crystal?
- Is an ionic crystal malleable or brittle? Why?
- Will ionic compounds conduct an electric current?

8.2 Ionic Bonds and Ionic Compounds

Key Concept

An ionic bond is an electrostatic attraction between oppositely charged ions. These ions form as a result of a transfer of one or more electrons from a metal atom to a nonmetal atom. The lowest whole-number ratio of ions is called a formula unit. Ionic compounds adopt the structure of a crystal lattice and the coordination number of each ion in the crystal is the number of nearest neighbor ions of the opposite charge. Ionic crystals are hard and brittle. Ionic compounds will conduct electricity, but only when melted or dissolved in water so that the ions are capable of movement between electrodes.

Standards

Lesson Objectives

- Explain how an ionic bond is formed when electrons are transferred from one atom to another in terms of the resulting electrostatic attraction. Draw diagrams showing this process.
- Describe the structural arrangements of ions in a crystal, including coordination number and its relationship to a given compound's formula unit.
- Explain how various physical properties result from the ionic crystal lattice, such as strength, hardness, high melting points, brittleness, and electrical conductivity.

Lesson Vocabulary

- coordination number: The number of ions (or atoms) that immediately surround an ion of the opposite charge (or other atoms) within a crystal lattice.
- formula unit: The lowest whole number ratio of ions represented in an ionic compound.
- ionic bond: The electrostatic force that holds ions together in an ionic compound.
- ionic compound: An electrically neutral compound consisting of positive and negative ions.

Teaching Strategies

Introducing the Lesson

Ions have been introduced in *Chemical Nomenclature* in the context of writing correct formulas for ionic bonds. What this lesson does is introduce the concept of ion formation as a transfer of electrons from metal atom to nonmetal atom. This is an important distinction to make, as electrons cannot simply be "lost" or "gained" out of thin air. Show

the students several examples of ionic bond formation by the transfer of electrons, using different ion ratios. Provide practice problems for the students. They should be able to show the transfer of electrons so that each ion obeys the octet rule and write the ion formulas that result.

Demonstrations

Have a molecular model of a sodium chloride crystal and point out the alternating positive and negative ions. Show how the coordination numbers of each ion can be determined from the model. Demonstrate the brittleness of ionic compounds by striking a large crystal of rock salt with a hammer. Show the students how all of the fragments have smooth surfaces and compare this to the sodium chloride model.

Obtain an electrical conductivity tester as shown in Figure 8.9. The tester consists of two electrodes and a light bulb. Show the students that neither distilled water nor solid salt will conduct electricity. Mix the salt into the water and show how the conductivity of the solution increases as the salt dissolves.

Common Misconceptions

The electron dot diagrams which show one or more electrons being transferred from one atom to another is a good way to explain ionic bond formation. Have the students practice these with a variety of cation and anion charges. The danger is that students may not connect that process to the formation of a crystal lattice. Stress that when two ions attract one another, the total energy of the system decreases, which is desirable. In order to maximize the decrease in energy, many such ionic bonds form simultaneously and the crystal lattice results. Have a model of an ionic crystal to show to the students. The tutorial under "Science Inquiry" also makes this point. Above all, they must not make the error of thinking that a "molecule" of sodium chloride (or any other ionic compound) exists.

Science Inquiry

The computer-based tutorial below is a terrific visualization of how ionic bonds form and can be used as an introduction to this lesson. It shows how metal atoms lose electrons to nonmetal atoms and the resulting ionic bonds lead to the formation of a crystal.

http://www.pbslearningmedia.org/resource/lsps07.sci.phys.matter.ionicbonding/ionic-bonding/

Differentiated Instruction

Some students may benefit by making drawings or models of the ionic crystal lattice for various compounds. Models can be made relatively inexpensively by using Styrofoam balls and toothpicks. The foam balls can be painted with different colors to represent the cations and anions. Have students make models of sodium chloride or of other ionic compounds that may have different structures. Another possible example to try is calcium fluoride (CaF_2), which adopts a crystal structure called the fluorite structure.

https://en.wikipedia.org/wiki/Calcium_fluoride

Enrichment

Have students research natural minerals and their ionic structures (see Figure 8.7). They can make a PowerPoint presentation showing the beautiful forms that the crystals take and their chemical formulas. You may choose to have different students focus on particular elements and instruct them to find the naturally occurring minerals that contain these elements. Their presentations should also include the physical properties of the minerals.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 8.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 8.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Metals are also crystalline materials. For a pure metal, each lattice point is an atom of the metal, rather than cations and anions as in an ionic crystal lattice. The familiar properties of metals result from its crystalline structure.

- What are the physical properties of metals?
- What is an alloy?

8.3 Metallic Bonds

Key Concept

Metallic bonding is a type of bonding that occurs in pure metals or in alloys. A metal consists of mobile valence electrons surrounding stationary metal cations. Since the valence electrons are not attached to one given atom, they are able to move freely throughout the metal. This leads to common metal properties such as malleability and conductivity. Most metal atoms pack in one of three arrangements: body-centered cubic, face-centered cubic, or hexagonal. Alloys are mixtures of metals and can have advantageous properties compared to pure metals.

Standards

Lesson Objectives

- Describe the electron-sea model of metallic bonding.
- Explain how metallic bonding is responsible for the conductivity and luster of metals.
- Explain why metals are malleable and ductile, while crystalline ionic compounds are not.
- Describe how metal atoms are arranged, including the three most common packing systems.
- Identify some common alloys and explain their importance.

Lesson Vocabulary

- alloy: A mixture composed of two or more elements, at least one of which is a metal.
- closest packing: The most efficient arrangements for packing atoms, ions, or molecules in a crystal.
- metallic bond: The attraction of the stationary metal cations to the surrounding mobile electrons.

Teaching Strategies

Introducing the Lesson

Show the students samples of several different metals such as copper, iron, zinc, magnesium, lead, and tin. Ask them to create a list of the physical properties that all metals have in common. Demonstrate the malleability of metals. Ask why metals all have these properties, whereas nonmetals and ionic compounds do not. Stress that physical properties arise from the bonding and structure of the atoms and/or molecules in a substance.

Taking It Further

The interactive tutorial below explains the basics of metallic structure along with animations of how metal atoms behave when subjected to stress. The stresses include a physical force that may bend the metal as well as heating the metal.

http://rmpbs.pbslearningmedia.org/resource/phy03.sci.phys.matter.metal/the-structure-of-metal/

Common Misconceptions

Students can become confused when metal atoms are described as being stationary cations surrounded by mobile electrons. This is because the idea of a cation may lead them to think of the loss of electrons and ionic bonds. Rather than use the word cations, you can refer to the stationary metal nuclei and inner electrons.

Seeing the coordination numbers of the different metals from the unit cell diagrams (Figure 8.14) can be difficult. If models are available, they could help. Unit cells and the structure of crystalline solids will be discussed in more detail in *States of Matter*.

Science Inquiry

The link below is to a computer-based activity that describes metallic bonding and compares it to other types of bonding. Several links are embedded within the worksheet, allowing students to see animations such as metals conducting electricity and what happens when different types of compounds are heated to their melting points. Students are required to answer questions based on what they see.

http://hrsbstaff.ednet.ns.ca/butlerb1/Chem11/metallic_bonding_activity%20version%202010.htm

Differentiated Instruction

The concept of closest packing can be applied to many more situations than just the packing of metal atoms. Students can brainstorm other situations in which they have encountered the close packing of same-shaped objects. A good example is the arrangement of the hexagons in the honeycomb of a beehive. Bring in a set of billiard balls or marbles for students to use to make different packing arrangements and compare their densities.

Enrichment

Have interested students do research on alloys. Many alloys, such as bronze, have been in use for thousands of years. Assign different alloys to different students or groups of students. Their research should focus on the history of the alloy, its composition, how it is made, and its uses. Format could be a written report, poster, or PowerPoint presentation.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 8.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 8.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Molecular compounds are a class of substances that take the form of individual molecules.

- Which types of elements make up molecular compounds?
- How is the chemical bonding within a molecular compound different from the bonding that occurs in ionic compounds or metals?
- How are the physical properties of molecular compounds different from those of ionic compounds or metals?



Covalent Bonding

Chapter Outline

- 9.1 LEWIS ELECTRON DOT STRUCTURES
- 9.2 MOLECULAR GEOMETRY
- 9.3 POLARITY AND INTERMOLECULAR FORCES
- 9.4 HYBRIDIZATION OF ATOMIC ORBITALS
- 9.5 REFERENCES



Ben Mills (User:Benjah-bmm27/Wikimedia Commons). commons.wikimedia.org/wiki/File:Water-elpot-transparent-3D-balls.png. Public Domain.

Chapter Overview

Covalent bonding occurs between two atoms that share one or more pairs of electrons in order to satisfy the octet rule. Bond formation releases energy, while the breaking of a covalent bond absorbs energy. Lewis electron dot structures are drawn in order to visualize the bonding patterns and shapes of molecules. Valence-shell electron pair repulsion theory describes the electron domain geometries and resultant molecular geometries of molecules. Covalent bonds may be either polar or nonpolar, depending on the equality of sharing of valence electrons. Molecular polarity is dependent on bond polarity and on molecular geometry. Intermolecular forces are those that exist between molecules and are responsible for physical state, melting points, and boiling points. Valence bond theory describes covalent bonds occurring as a result of overlap of atomic orbitals. Some atomic orbitals must undergo hybridization in order to generate the necessary types of orbitals formed in certain molecules.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE	9.1:	Pacing	Lessons
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Lesson	Class Period(s) (60 min)
9.1 Lewis Electron Dot Structures	3-4
9.2 Molecular Geometry	1-2
9.3 Polarity and Intermolecular Forces	2-3
9.4 Hybridization of Atomic Orbitals	1-2

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE	9.2:	Chemistry	Concepts
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FlexBook® Lesson	Chemistry Concepts
9.1 Lewis Electron Dot Structures	9.1 –9.9
9.2 Molecular Geometry	9.10 -9.12
9.3 Polarity and Intermolecular Forces	9.13 -9.16
9.4 Hybridization of Atomic Orbitals	9.17 -9.20

9.1 Lewis Electron Dot Structures

Key Concept

Chemical bonds between atoms form so that the potential energy of the system decreases to a minimum. A covalent bond forms when two atoms share one or more pairs of electrons. The bonding scheme of a molecule or a polyatomic ion can be determined by drawing Lewis electron dot structures and following the octet rule. Covalent bonds may be single, double or triple. A bond is coordinate covalent when one of the atoms contributes both of the shared electrons. A molecule displays resonance when two or more valid Lewis structures can be drawn. Exceptions to the octet rule include incomplete octets, odd-electron molecules, and expanded octets. Bond energy is the energy required to break a covalent bond.

Standards

Lesson Objectives

- Describe how a covalent bond forms, including the energy change involved in the process.
- Use the octet rule to draw Lewis electron dot structures for simple molecules. Know how and when to incorporate double and triple bonds into the structures.
- Understand how a coordinate covalent bond differs from other covalent bonds.
- Be able to draw Lewis structures for polyatomic ions.
- Understand the concept of resonance.
- Know some common exceptions to the octet rule.
- Relate bond energy to the stability and reactivity of molecules.

Lesson Vocabulary

- bond energy: The energy required to break a covalent bond between two atoms.
- coordinate covalent bond: A covalent bond in which one of the atoms contributes both of the electrons in the shard pair.
- covalent bond: A bond in which two atoms share one or more pairs of electrons.
- diatomic molecule: A molecule containing two atoms.
- double covalent bond: A covalent bond formed by atoms that share two pairs of electrons.
- Lewis electron dot structure: Notation that depicts valence electrons as dots around the atomic symbol of that element; used to depict bonding patterns in molecules.
- lone pair: A pair of electrons in a Lewis electron-dot structure that is not shared between atoms.
- resonance: The use of two or more Lewis structures to represent the covalent bonding in a molecule.
- single covalent bond: The bond that forms when two atoms are joined by the sharing of one pair of electrons.

- structural formula: A formula that shows the arrangement of atoms in a molecule and represents covalent bonds between atoms by dashes.
- triple covalent bond: A covalent bond formed by atoms that share three pairs of electrons.

Teaching Strategies

Introducing the Lesson

Remind the students that everything that happens in a chemical system does so in order to lower the energy of the system –from electron transitions to bond formation. Use Figure 9.1 to justify why covalent bonds occur at all and that the only elements that can exist naturally as individual atoms are the noble gases. Metals have the extended three-dimensional system described by metallic bonds. Nonmetals and metalloids all have a different sort of bonding –the covalent bond. Explain how two nonmetals cannot both satisfy the octet rule by transferring electrons as in an ionic bond. The numbers won't work. Progress through examples of Lewis structures of simple molecules, starting with single bonds only. Always emphasize that the total number of electrons in the Lewis diagram must be equal to the sum of all of the valence electrons of the atoms in the molecule.

Taking It Further

Occasionally, more than one valid Lewis structure can be drawn for a molecule, but the structures are not equivalent. In other words, the structures cannot properly be called resonant. The concept of formal charge allows one to determine which structure is preferred in this case. The formal charge is a number that is applied to each individual atom in the molecule. The definition is shown below:

Formal charge = Number of valence electrons in free atom –(number of nonbonding electrons + $\frac{1}{2}$ of the number of bonding electrons)

The sum of all the formal charges for a molecule is equal to zero. For a polyatomic ion, the sum of the formal charges is equal to the charge of the ion.

For more information about formal charges, see the links below:

http://www.mhhe.com/physsci/chemistry/carey/student/olc/ch01lewis.html

https://www.youtube.com/watch?v=lbrhuMqcCwI

The molecule N_2O can be drawn three ways. In each case, one of the nitrogen atoms is the central atom because it is less electronegative than oxygen. The figure below shows the three structures, along with the formal charges for each atom.



The most preferred structure is the one which minimizes the nonzero formal charges. This eliminates the first structure because of the -2 formal charge on the singly bonded nitrogen. Of the other two structures, each has one +1 and one -1 formal charge. The structure with the -1 formal charge on the oxygen is preferred because oxygen has a higher electronegativity than nitrogen. Structural studies of N₂O show that this is the bonding pattern –a triple bond between the two nitrogen atoms and a single bond to the oxygen atom.

Common Misconceptions

Students may have trouble with the octet rule as it applies especially to molecules with double or triple bonds. Make sure to explain that when two atoms are double bonded, both of the atoms get to "count" all four electrons toward their total. When it's a triple bond, both atoms "count" six electrons.

Emphasize that drawing two or more resonance structures, with a double-sided arrow between the structures, does not mean that the molecule "flip-flops" back and forth between them. It is just that this is the easiest way to show the bonding and that the true structure is an average of the resonant forms.

Science Inquiry

The link below is to a comprehensive, inquiry-based web quest on covalent bonding. Students view tutorials, practice drawing Lewis structures, and take quizzes on what they have learned. Students will need access to a computer for this activity.

http://www.millcreekhs.com/attachments/article/673/Covalent-Bonding-Lewis-Structure-Webquest.pdf

Differentiated Instruction

The tutorial at the link below may help students having difficulty understanding why covalent bonds form. Have students view the tutorial and summarize briefly what they have learned.

http://www.pbslearningmedia.org/asset/lsps07_int_covalentbond/

The following link is to another tutorial which walks students through how to construct Lewis electron dot diagrams. At the end, students practice writing proper structures, guided by questions from the tutorial.

https://www.wisc-online.com/learn/natural-science/chemistry/gch6404/lewis-dot-structures-of-covalent-compounds

Enrichment

As discussed in the "Taking It Further" section above, formal charge is a concept which can deepen students' understanding of Lewis electron dot structures. Have advanced students assign formal charges to Lewis dot structures that they produce. Polyatomic ions provide more good practice for assigning formal charges. Often, a given structure will be the preferred structure according to formal charge, even though it violates the octet rule. The chlorate ion (ClO_3^-) is an example. It can be drawn with all single bonds, or one or two Cl-O double bonds. The best structure is one with two double bonds because it minimizes the formal charge.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 9.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 9.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Sample answers coming soon.

Points to Consider

The physical and chemical properties of various substances are dependent upon their structures and bonds. Some of these properties are related to the arrangement of the atoms within a molecule and the resulting molecular geometry.

- How do the covalent bonds in a molecule allow one to predict molecular shapes?
- How many different basic shapes are possible?

9.2 Molecular Geometry

Key Concept

The geometry of molecules can be predicted by using valence shell electron pair repulsion theory. The electron domain geometry of a molecule refers to the number of groups of electrons around the central atom, whether the groups are single bonds, multiple bounds (double, triple), or lone pairs. There can be anywhere from two to six groups of electrons around the central atom, corresponding to electron domain geometries of linear, trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral respectively. Lone pairs of electrons change the molecular geometries and slightly alter bond angles. Possible molecular geometries of molecules with lone pairs include bent, trigonal pyramidal, seesaw, T-shaped, linear, square pyramidal, and square planar.

Standards

Lesson Objectives

- Explain the basis of VSEPR theory.
- Predict the shapes of molecules and polyatomic ions using VSEPR theory.
- Account for variations in bond angles based on the relative repulsive forces exerted by lone pairs and bonding pairs of electrons.
- Describe the relationship between molecular geometry and electron domain geometry.

Lesson Vocabulary

- electron domain geometry: Geometry based only on the total number of electron pairs around the central atom, both bonding pairs and lone pairs.
- molecular geometry: The three-dimensional arrangement of atoms in a molecule.
- valence shell: The outermost electron-occupied shell of an atom.
- valence shell electron pair repulsion: A theory for predicting molecular geometries; it states that electron pairs will repel each other such that the valence electron-pairs stay as far apart from each other as possible.

Teaching Strategies

Introducing the Lesson

Ask the class what will happen if two electrons are brought near to one another. (They repel each other.) Describe how this simple concept is the basis for an entire theory about the shapes of molecules. Groups of electrons around

the central atom of a molecule repel each other and so will arrange themselves in such a way as to be as far apart from one another as possible. Make sure to have molecular models on hand when discussing all of the various shapes described in the lesson.

Common Misconceptions

One troublesome misconception that many students have concerns the shape of the water molecule. This arises because students can draw the Lewis structure in a way that makes it appear to be linear.

FIGURE 9.1

To avoid this problem, start with the molecular model of a methane molecule (tetrahedral). Demonstrate that when an atom is removed, but the electron pair is "left behind" the shape is trigonal pyramidal. Now show that the removal of a second atom, no matter which one, results in a bent structure. Make the point that when the Lewis structure is drawn as shown above, the oxygen atom is actually coming out of the plane of the paper (or board). Hold up a bent water molecule to the board to emphasize this. From that point on, always draw the Lewis structure of the AB_2E_2 molecules to make them look bent instead of linear.



Science Inquiry

It is extremely beneficial for the students to practice building models of simple molecules and determining the geometries by using VSEPR theory. There are many examples of lab activities available on the internet. Some require molecular model kits, while others are written in a way that toothpicks and different colored gumdrops can be used instead. Decide what is best for your time and materials.

This activity includes extensive notes and explanation on VSEPR and how geometries are determined. It then includes a modeling activity at the end. It includes information on molecular polarity and hybridization.

http://swc2.hccs.edu/pahlavan/intro_labs/Exp_17_Lewis_Dot_Structures_and_Geometry.pdf

This activity also requires students to know the hybridizations of the central atoms, which is covered in Lesson 9.4.

http://bcpshelpdeskhighschoolscience.weebly.com/uploads/6/3/4/6/6346142/lab_-_vsepr_modeling_ltf.pdf

This activity does not mention polarity or hybridization. However, it does not included expanded octet geometries.

https://staff.rockwood.k12.mo.us/grayted/generalchemistry/General%20Chemistry%20Documents/U4%20Covalent% 20Bonding%20and%20Compounds/LAB%20VSEPR%20Model%20Building.pdf

Differentiated Instruction

The link below is for a simulation that can be used as an introduction for all students or for a follow-up exercise for students that may be having difficulty envisioning the shapes. It allows students to rotate molecules on the screen and see how bonded pairs and lone pairs affect directly affect the shapes and the bond angles. There are many choices for downloadable lesson plans to suit your own class.

http://phet.colorado.edu/en/simulation/molecule-shapes

Enrichment

There are several larger molecular geometries that are not discussed in this lesson. The acronyms are AB5E2, AB6E, AB7, AB8, and AB9. Instruct advanced students to research these geometries and give examples of molecules or ions that adopt each of them.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 9.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 9.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The electronegativity of an atom describes its ability to attract shared electrons. Nonmetal atoms generally have high electronegativities, while metals generally have low electronegativities.

- How can electronegativity be used to predict whether a bond between two atoms will be ionic or covalent.
- Is the sharing of electrons in a covalent bond always equal, or can the sharing be dominated by one of the atoms? How would this affect the properties of the molecule?

9.3 Polarity and Intermolecular Forces

Key Concept

The electronegativity difference between the two atoms in a bond determines its character. Bonds with a small difference represent equal sharing of electrons and are called nonpolar covalent. A larger difference represents unequal sharing and a polar covalent bond, while a very large difference gives rise to an ionic bond. The geometry of a molecule must also be considered when determining the molecular polarity. Intermolecular forces are those that exist between molecules. London dispersion forces are relatively weak forces, while dipole-dipole forces are slightly stronger. The strongest of the forces is the hydrogen bond, where an electron deficient hydrogen atom is attracted to a lone pair of electrons on a neighboring molecule. They type and strength of the intermolecular forces affects the melting and boiling points of a substance.

Standards

Lesson Objectives

- Describe how the electronegativity difference between two atoms in a covalent bond results in the formation of a nonpolar covalent, polar covalent, or ionic bond.
- Describe how molecular geometry plays a role in determining whether a molecule is polar or nonpolar.
- Distinguish between the following three types of intermolecular forces: dipole-dipole forces, London dispersion forces, and hydrogen bonds.
- Describe how chemical bonding and intermolecular forces influence the properties of various compounds.

Lesson Vocabulary

- diatomic element: An element whose natural form is of a diatomic molecule.
- dipole: A molecule with two poles, a positive pole and a negative pole; also called a polar molecule
- dipole-dipole forces: The attractive forces that occur between polar molecules; the positive end of one dipole is attracted to the negative end of a neighboring dipole.
- hydrogen bond: An intermolecular attractive force in which a hydrogen atom that is covalently bonded to a small, highly electronegative atom is attracted to a lone pair of electrons on an atom in a neighboring molecule.
- intermolecular forces: Attractive forces that occur between molecules.
- London dispersion forces: The intermolecular forces that occur between all atoms and molecules due to the random motion of electrons.
- nonpolar covalent bond: A covalent bond in which the bonding electrons are shared equally between the two atoms.
- polar covalent bond: A covalent bond in which the atoms have an unequal attraction for electrons, and so the sharing is unequal.
- polar molecule: A molecule in which one end of the molecule is slightly positive, while the other end is slightly negative.
- van der Waals forces: The weakest intermolecular forces; they consist of dipole-dipole forces and dispersion forces.

Teaching Strategies

Introducing the Lesson

Make sure that students have a copy of a chart of electronegativities of the elements. Start the discussion with simple diatomic molecules, focusing on the electronegativity values, the unequal sharing of the electrons, and the resultant charge separation. Make sure to use a δ + and a δ - in drawings, emphasizing that the size of the partial charge is dependent on the magnitude of the electronegativity difference. Give students practice, including nonpolar covalent, polar covalent, and ionic examples. The simulation below would fit in well after this introduction.

Simulation

The simulation at the link below allows students to study the bond polarity and molecular polarity of many different molecules. Simple two-atom and three-atom molecules can be constructed where the atoms have a range of electronegativities. The different molecules can be put into a simulated electric field to see how they would behave. Electron density distributions can be superimposed on to the molecular structures. A wide variety of downloadable activities accompanies the simulation on the same page. The simulation would work best if students have been given an introduction to electronegativity, bond polarity, covalent and ionic character, and molecular polarity.

http://phet.colorado.edu/en/simulation/molecule-polarity

Demonstration

An easy demonstration of the high degree of polarity of water is the following: Run a thin stream of water either from a tap or from a buret. Rub a balloon with some fur or on a student's hair to give it a charge. Hold the balloon very close to the stream of water, but don't allow it to touch. The stream will be "pulled" towards the balloon because of the charge separation in the water molecule. Discuss how this will not work with a nonpolar liquid. If possible, try it with mineral oil or another nonpolar liquid.

Common Misconceptions

Students may struggle with the idea that bond character is a continuum from covalent to ionic and that many bonds are not 100% one or the other. For example, a bond with an electronegativity difference of 1.7 is about 50% ionic and 50% covalent. This concept is difficult for students to grasp. Make sure to differentiate between bond character and the overall physical structure of a compound. A compound adopts either a crystal lattice structure or that of discrete individual molecules. Bond character is a major determining factor, but other factors, such as particle size, also contribute. Bond character would predict that HF is ionic, but the compound hydrogen fluoride is actually molecular. This is because of the small size of the hydrogen ion and its inability to attain a crystalline structure. Copper(II) chloride, CuCl₂, is a crystalline compound even though the difference in electronegativity between copper and chlorine is only 1.3, suggestive of a covalent Cu-Cl bond.

London dispersion forces are present in all molecules. Students may get the mistaken idea that dispersion forces only exist in nonpolar molecules and that they are always weaker than dipole-dipole forces. In reality, a nonpolar molecule with many electrons may have stronger intermolecular forces than a smaller polar molecule. Cl₂ is nonpolar, but has

a higher boiling point than HCl, which is polar. So the dispersion forces of Cl_2 are stronger than the van der Waals forces (dispersion and dipole-dipole) of HCl. This is because of the much larger electron cloud in Cl_2 .

Science Inquiry

The following activity is designed as an introductory inquiry to molecular polarity. Two liquids, hexane and water, are examined by the students. The goal is to for the students to describe their physical properties and account for the differences.

http://www.iacad.org/istj/32/3/polar.pdf

The activity below serves as a good, quick introduction to some of the unique properties of water that are a result of its hydrogen bonding capability. Students observe cohesion, adhesion, capillary action, and surface tension.

http://serc.carleton.edu/NAGTWorkshops/oceanography/activities/hydrogen_bond.html

Differentiated Instruction

Have students create a flowchart showing the relationships of bond polarity to molecular polarity to the various types of intermolecular forces. The students should write yes/no questions leading through the flowchart.

Enrichment

Paper chromatography is a method of separating components of a liquid mixture based on differences in polarity. There are many on-line activities centered around separating black marker ink into different colors. Have students perform the activity, then research how the technique works.

http://www.teachengineering.org/view_activity.php?url=collection/wst_/activities/wst_environmental/wst_environmental_lesson02_activity1.xml

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 9.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 9.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Lewis electron-dot structures and VSEPR theory explain the bonding patterns and molecular geometries of molecules, but not the specific way in which electrons in atomic orbitals interact with one another as a covalent bond forms.

- How can atomic orbitals be modified to explain how electrons are arranged in a molecule?
- How are single bonds different from double or triple bonds?

9.4 Hybridization of Atomic Orbitals

Key Concept

Valence bond theory describes a covalent bond as an overlap of atomic orbitals of the atoms participating in the bond. Some covalent bonds can only be properly understood if the atomic orbitals on the central atom mix together to form hybrid orbitals. Hybridization of orbitals can only involve mixing of nonequivalent orbitals. The type of hybrid orbitals that are formed by a given molecule can be determined by knowing the electron domain geometry. Linear, trigonal planar, and tetrahedral domain geometries give rise to sp, sp^2 , and sp^3 hybridization respectively. Trigonal bipyramidal and octahedral domain geometries use sp^3d and sp^3d^2 hybrids. Molecules with double or triple bonds display two types of covalent bonding. A sigma bond is an end-to-end overlap of orbitals, while a pi bond is a side-to-side overlap of unhybridized *p* orbitals.

Standards

Lesson Objectives

- Describe valence bond theory as it pertains to the formation of a covalent bond between atoms.
- Describe the process of electron promotion and hybridization during the formation of hybrid orbitals.
- Explain the relationship between electron domain geometry and the various types of hybrid orbitals.
- Distinguish between sigma and pi bonding.

Lesson Vocabulary

- hybridization: The mixing of the atomic orbitals in an atom to produce a set of hybrid orbitals.
- hybrid orbitals: The atomic orbitals obtained when two or more nonequivalent orbitals from the same atom combine in preparation for bond formation.
- pi bond (π): A bond formed by the overlap of orbitals in a side-by-side fashion with the electron density concentrated above and below the plane of the nuclei of the bonding atoms.
- sigma bond (σ): A bond formed by the overlap of orbitals in an end-to-end fashion, with the electron density concentrated between the nuclei of the bonding atoms.
- valence bond theory: A theory which states that the electrons in a molecule are assumed to occupy atomic orbitals of the individual atoms and that a covalent bond results from overlap of those orbitals.

Teaching Strategies

Introducing the Lesson

It is essential to establish the concept of a covalent bond being any overlap of atomic orbitals, accompanied by a decrease in the energy of the system. Discuss the covalent bonds that occur in simple diatomic molecules like H_2 and F_2 before moving on to hybridization, which is the more difficult aspect of the lesson. Use models to show a variety of carbon-containing (organic) compounds. Ask the students how many bonds carbon seems to always form. Ask them why carbon makes four bonds (because of its four valence electrons). Draw the valence shell orbital filling diagram for carbon and see if a student can figure out how carbon can form four bonds when it has only two unpaired electrons. Describe the electron promotion and why the geometry is incorrect if the electrons remain in s and p orbitals. Describe the formation of the sp^3 , sp^2 , and sp hybrids. You will need to have visuals so that students can see the shapes. The link below is a quick animation that can be shown to the class.

http://www.learnerstv.com/animation/animation.php?ani=52

Common Misconceptions

Once students learn about hybrid orbitals, they may make the mistaken assumption that all covalent bonds require hybridization. Remind them of the earliest examples (H_2, F_2) where overlap was able to occur of unhybridized atomic orbitals.

Emphasize that pi bonds are a different type of bond than sigma bonds. Use models to show how there is no free rotation of a double bond. Students sometimes get confused about the two different types of bonds. A key point for the students to know is that single bonds are sigma bonds, while pi bonds are the second bond of a double bond or the second and third bonds of a triple bond.

Science Inquiry

The link below describes an activity where balloons can be used as an aid to help students visualize the shapes of hybridized atomic orbitals.

https://www.flinnsci.com/media/620832/91330.pdf

Differentiated Instruction

Table 9.7 is essential for the students to know –that the type of hybrid orbitals is determined by the electron domain geometry as predicted by VSEPR theory. Have students recreate the table in large formal using visuals. Ask them to draw the hybrid orbitals and draw the models of the example molecules.

Enrichment

Molecular orbital (MO) theory is an alternative to valence bond theory in describing the nature of covalent bonding. In MO theory, electrons in the bonding atoms are assumed to occupy molecular orbitals which spread over the entire molecule rather than atomic orbitals of individual atoms. Advanced students can research MO theory and learn to draw MO diagrams for simple homonuclear diatomic molecules (H_2 , He_2 , Li_2 , etc.)

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 9.4 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 9.4 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

In future chapters, we move away from the microscopic world of chemistry and back to a macroscopic view, focusing on the mathematics involved in the analysis of chemical reactions.

- What are different ways to measure the amount of something?
- Since atoms and molecules are so small, how do chemists keep track of the number of them that are taking part in a given reaction?

9.5 References

- 1. Joy Sheng. CK-12 Foundation . CC BY-NC 3.0
- Ben Mills (User:Benjah-bmm27/Wikimedia Commons). ("Left") http://commons.wikimedia.org/wiki/File:Wa ter-dimensions-from-Greenwood%26Earnshaw-2D.png; ("Right") http://commons.wikimedia.org/wiki/File: Water-3D-balls-A.png. Public Domain

Chapter **10**

The Mole

Chapter Outline

- 10.1 THE MOLE CONCEPT
- 10.2 MASS, VOLUME, AND THE MOLE
- **10.3 CHEMICAL FORMULAS**



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Chapter Overview

The counting unit of the mole is the unit used by chemists to account for the number of representative particles participating in reactions. One mole is always equal to 6.02×10^{23} . The molar mass of a substance is the mass in grams of one mole of particles of that substance. Calculations can be done to convert back and forth between moles and mass, volume of a gas, or the number of particles. Percent composition is the percent by mass of elements in a compound. The empirical and molecular formulas of compounds can be determined from percentage composition data and the molar mass.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 10.1: Pacing Lessons

Lesson	Class Period(s) (60 min)

Lesson	Class Period(s) (60 min)
10.1 The Mole Concept	1-2
10.2 Mass, Volume, and the Mole	2
10.3 Chemical Formulas	2

TABLE 10.1: (continued)

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 10.2	: Chemistry	^v Concepts
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FlexBook® Lesson	Chemistry Concepts
10.1 The Mole Concept	10.1 – 10.3
10.2 Mass, Volume, and the Mole	10.4 -10.9
10.3 Chemical Formulas	10.10-10.13

10.1 The Mole Concept

Key Concept

A mole is the name given to a counting quantity, specifically 6.02×10^{23} of any object. This number is called Avogadro's number. The mole is used by chemists for atoms and molecules because they are so small that physically counting them would be impossible. Dimensional analysis can be used to convert back and forth between the number of particles in a sample of matter and the number of moles of particles. The molar mass of an element is the mass in grams of one mole of atoms of that element and can be found from the periodic table. The molar mass of a compound can be found by adding up the molar masses of all of the atoms in one molecule or formula unit of the compound.

Standards

Lesson Objectives

- Identify three methods for measuring the amount of matter in a sample.
- Define the mole and its relationship to Avogadro's number.
- Use Avogadro's number to convert between moles and the number of representative particles of a substance.
- Relate the atomic mass of an element to its molar mass.
- Calculate the molar mass of a given compound.

Lesson Vocabulary

- Avogadro's number: The number of representative particles in a mole of any substance; the numerical value is 6.02×10^{23} .
- formula mass: The sum of masses of all the atoms represented in a chemical formula.
- molar mass: The mass of one mole of representative particles of a substance.
- mole: The amount of a substance that contains 6.02×10^{23} representative particles of that substance.
- **representative particle:** The smallest unit in which a substance naturally exists, usually an atom, ion, molecule, or formula unit.

Teaching Strategies

Introducing the Lesson

Ask the students what is meant by the term "amount". How is the amount of something measured? They should be able to come up with mass, volume, and number. State that it is the goal of this chapter to be able to convert back and forth between mass, volume, and number in a way that chemists do when performing lab experiments. Ask if they know of any names given to specific numerical quantities (pair, triplet, dozen, etc.) Ask them to recall how small atoms and molecules are and that chemists need a label for the number of particles that is a very large number. Introduce Avogadro's number and the unit of the mole and give the students plenty of practice doing conversion problems.

Common Misconceptions

Avogadro's number can seem very arbitrary to students at first. Why 6.02? Have them recall that when they studied atomic structure, the mass of a single atom was expressed in atomic mass units (amu). Avogadro's number can be thought of as the conversion factor from amu to grams. Now when they look at the atomic masses on the periodic table, they can think of them in the practical unit of grams instead of amu. One atom of hydrogen weighs 1 amu, so 1 mole (6.02×10^{23}) of hydrogen atoms weighs 1 gram. One atom of helium weighs 4 amu, so 1 mole (6.02×10^{23}) of helium atoms weighs 4 grams. This is true for every single element.

Some students will have trouble using exponential notation properly with their calculators. Make sure that they are using the "EE" or "EXP" button correctly. If they perform a division problem with Avogadro's number and input it using " $\times 10^{23}$ ", the order of operations of the calculator will give them the wrong answer. They should input Avogadro's number as "6.02EE23".

Science Inquiry

The activity at the link below is in two parts. Part 1 is an excellent introduction to this lesson, as it introduces students to the idea of counting a very large number of small objects by using mass. Part 2 of the activity would be appropriate at the end of the lesson, once students have been introduced to the mole concept and to molar masses of elements.

http://phs.princetonk12.org/teachers/jgiammanco/Chem%201/Labs/C18-CountByWeigh.pdf

Differentiated Instruction

Get students comfortable with doing conversion problems using familiar quantities like the dozen before moving on to the mole. Give them practice problems and make sure that they show their work and cancel out units. Not canceling units properly is one of the biggest reasons that students run into difficulty later on when doing multi-step problems. If some students have more difficulty with math, pair them up with more advanced students to work on these conversion problems and then on conversion problems using Avogadro's number.

Enrichment

Interested students can research how the number which bears Avogadro's name was actually determined in the laboratory. What chemical reactions were performed that led to 6.02×10^{23} ? The technique for determining the numerical value has changed several times as technology has become more advanced. How was it originally measured? What is the most recent method? Have them present their findings to the class.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 10.1 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 10.1 in *CK-12 Chemistry –Intermediate* FlexBook® resource.

Points to Consider

The molar mass of a compound can be used to convert between the mass of a substance (in grams) and its amount (in moles).

- What would the conversion factors for these calculations look like?
- Is there a way to convert between moles and the volume of a substance?

10.2 Mass, Volume, and the Mole

Key Concept

The molar mass of elements and compounds can be used to make conversions between mass of a sample and the number of moles. To convert between mass and the number of representative particles requires a two-step process. The first step is always to convert the given quantity to moles. Avogadro's hypothesis states that any gas at a given temperature and pressure will contain the same number of particles. The molar volume of a gas is the volume of one mole of gas at standard temperature and pressure (0°C and 1 atm) and is equal to 22.4 L. Molar volume allows conversion calculations between moles and gas volume as well as to determine the density of any gas at STP. The mole road map is a visual aid used to make any calculation involving moles, mass, gas volume, and number of particles.

Standards

Lesson Objectives

- Use molar mass to make conversions between mass and moles of a substance.
- Explain Avogadro's hypothesis and how it relates to the volume of a gas at standard temperature and pressure.
- Convert between moles and volume of a gas at STP.
- Calculate the density of gases at STP.
- Use the mole road map to make two-step conversions between mass, number of particles, and gas volume.

Lesson Vocabulary

- Avogadro's hypothesis: Equal volumes of all gases at the same temperature and pressure contain equal numbers of particles.
- **molar volume:** The volume of one mole of a gas at standard temperature and pressure; the numerical value is 22.4 L.
- standard temperature and pressure (STP): The conditions at which the volumes of gases are measured and compared; it is equal to 0°C (273.15 K) and 1 atm (101.3 kPa and 760 mmHg) pressure.

Teaching Strategies

Introducing the Lesson

This lesson follows directly from the conversion problems done in the last lesson and from the concept of molar mass. Show students how to convert from mass to moles and from moles to mass. Emphasize the canceling of units and dimensional analysis. This is essential so that students will not have to memorize how to do every different calculation. Continue with the two-step problems involving mass and number of particles. Stress that the mole is the central concept of all chemistry calculations and that it is not possible to convert from mass to number of particles (or vice-versa) in just one step.

Common Misconceptions

Students generally can perform most mole calculations with little difficulty when calculations of similar nature are grouped together. When they have difficulty is when different types of calculations are all "mixed up" on a single worksheet or set of problems. It is critical for students to develop the skill of deciding what information is given to them in a problem, what is being asked for, and the step(s) needed to get there. Make sure to provide many practice problems of all types of mole problems together.

Sometimes students will get attached to the 22.4 L molar volume "constant" and overuse it –with gases under nonstandard conditions or with other states of matter. Make sure to define STP and to state that gases are at STP whenever the 22.4 L/mol is needed in a problem. Students will learn how to deal with gases under nonstandard conditions in a later chapter (*The Behavior of Gases*).

Science Inquiry

The molar volume of a gas (22.4 L) can be determined experimentally by a reaction of a metal such as magnesium with hydrochloric acid to produce hydrogen gas. If the mass of the magnesium is known, the volume of the hydrogen gas can be measured. Several calculations will lead to an experimental value for molar volume. Eudiometers are required in order to collect the gas. Students will need assistance in using the combined gas law to calculate the volume that the gas would occupy at STP. This activity would work best after students have become comfortable with the mole calculations in this lesson.

http://www.flinnsci.com/media/960405/ap_chem_3A.pdf

Differentiated Instruction

Some students may learn the mole road map and how to use it by making a type of flashcard set. Have the students write each of the three pairs of conversion factors from the mole road map on to the two sides of a 3×5 card. Have them also make four separate cards with each type of quantity that can be known (or unknown): moles, grams, liters, and representative particles. Students can use the cards for any problem by lining them up in a way that will produce the needed unit for their answer. After they have lined up the cards, they can fill in the numbers on paper and do the math.

Enrichment

Advanced students can be asked to find the molar mass and density of air. They will need to research the percentages of the various gases found in air and perform a weighted average calculation to find the molar mass. Once they know

the density of air, ask them to predict whether balloons filled with certain gases (CO₂, CO, F_2 , Ar, NH₃, N₂, etc.) will float or sink.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 10.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 10.2 in *CK-12 Chemistry –Intermediate* FlexBook® resource.

Points to Consider

The chemical formula of an ionic compound is an empirical formula, the simplest ratio between cations and anions in the crystal. The chemical formula of a molecular compound shows the number of each atom present in the compound.

- How is the composition of a compound related to its formula?
- How can mole calculations be used to analyze chemical formulas?

10.3 Chemical Formulas

Key Concept

Percent composition is the percent by mass of each of the elements in a compound and can be calculated from mass data or from a known chemical formula. A hydrate is an ionic compound that contains water molecules as part of the crystal lattice. The percentage of water in a given hydrate can be calculated from the formula. The empirical formula of an unknown compound can be determined from percent composition data. The molecular formula of a compound may be a multiple of the empirical formula. In order to determine the molecular formula, the molar mass of the compound must be known.

Standards

Lesson Objectives

- Calculate the percent composition of a compound either from mass data or from the chemical formula. Use percent composition to calculate the mass of an element in a certain sample of a compound.
- Calculate the percentage of a hydrate's mass that is due to water.
- Determine the empirical formula of a compound from percent composition data.
- Determine the molecular formula of a compound from the empirical formula and the molar mass.

Lesson Vocabulary

- hydrate: A compound that has one or more water molecules bound to each formula unit.
- percent composition: The percent by mass of each element in a compound.

Teaching Strategies

Introducing the Lesson

Use a simple example to explain percent by mass. If a 200-gram mixture of iron and zinc contained 20 grams of iron, what is the percent by mass of each element? Students should have no trouble with this. Then extend it to percent composition problems with a chemical formula.

Students should know what an empirical formula is from previous chapters, though it is worth quickly reviewing. Explain how the empirical formula problems are somewhat like the reverse of the percent composition problems.

Common Misconceptions

Some students may confuse empirical and molecular formulas. Point out that in many cases the empirical formula of a compound is the same as the molecular compound. This is the case when the subscripts of the molecular formula cannot be "reduced". Students need to make sure that they know what they are being asked when solving the empirical formula and molecular formula problems to avoid confusion.

Science Inquiry

The lab described at the link below allows students to experimentally determine the empirical formula of magnesium oxide. A small sample of magnesium metal is carefully heated in a crucible until it is entirely converted to magnesium oxide. Measurements of the mass of reactants and products followed by conversions to moles will allow students to find the ratio of moles magnesium to moles of oxygen and thus, the empirical formula. This lab works best when students are familiar with empirical formulas and that the subscripts in an empirical formula represent the ratio of moles of the elements in that compound.

http://science.csustan.edu/cstessman/chem3010/Grades/Empirical%20Formula%20of%20Magnesium%20Oxide.pd f

Another lab activity appropriate for this lesson is the determination of the formula of a hydrate. A sample of a hydrate is massed and then heated until all of the water of hydration is driven off. The mass of the anhydrous salt is taken. A conversion to moles allows the students to calculate the mole ratio of water to anhydrous salt and thus determine the formula of the hydrate.

http://gator.gatewayk12.org/~mkohler/Hydrate%20Lab.pdf

Differentiated Instruction

If some students are having difficulty with the concepts of percent composition and empirical formulas, a simple activity may help. Use three different colored poker chips or marbles. Give the students 8 red chips, 12 white chips, and 4 blue chips. The numbers can vary. Ask the students what is the percentage of each color (50% white for example). Ask them to figure out the lowest ratio of red to white to blue chips (2 to 3 to 1). What is the empirical formula of a hypothetical compound of the chips? (R_2W_3B) Use different numbers until the students perfect the exercise.

Enrichment

Advanced students can research the experimental techniques of elemental analysis. There are several different ways that compounds can be analyzed to determine what elements are present as well as the quantities of each element. Students can summarize their findings in a brief report.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 10.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 10.3 in *CK-12 Chemistry –Intermediate* FlexBook® resource.

Points to Consider

Chemical reactions are the essence of chemistry. We will describe chemical reactions both qualitatively and quantitatively.

- What are some ways to classify different types of chemical reactions?
- How are calculations with moles, grams, and volume involved in the analysis of chemical reactions?

CHAPTER

Chemical Reactions

Chapter Outline

- 11.1 **CHEMICAL EQUATIONS**
- 11.2 **TYPES OF CHEMICAL REACTIONS**



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Chapter Overview

All chemical reactions can be described by a chemical equation, which shows how reactants are converted into products. Chemical equations may be word equations or skeleton equations, but must be balanced in order to be useful to a chemist. Balancing a chemical equation means adding coefficients in front of chemical formulas until each element has the same number of atoms on both sides of the equation. Chemical reactions generally fall into one of five types: combination, decomposition, single-replacement, double-replacement, or combustion. Knowledge of the types of reactions allows one to predict the products of many chemical reactions.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 11.1:	Pacing I	Lessons
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Lesson	Class Period(s) (60 min)
11.1 Chemical Equations	1-2
11.2 Types of Chemical Reactions	2-3

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 11.2: Chemistry Concepts

FlexBook® Lesson	Chemistry Concepts
11.1 Chemical Equations	11.1 –11.3
11.2 Types of Chemical Reactions	11.4 –11.9

11.1 Chemical Equations

Key Concept

Chemical reactions are described on paper through the use of chemical equations, where one or more reactants are converted into one or more products. A word equation uses the names of substances, while a skeleton equation uses the chemical formulas. A skeleton equation needs to be balanced in order to satisfy the law of conservation of mass. A balanced chemical equation has the same number of atoms of each element on both the reactant and product side of the equation.

Standards

Lesson Objectives

- Describe chemical reactions using word equations.
- Know the correct symbols to use in order to write skeleton equations for chemical reactions.
- Use coefficients to balance chemical equations so that the law of conservation of mass is followed.

Lesson Vocabulary

- balanced equation: a chemical equation in which mass is conserved and there are equal numbers of atoms of each element on both sides of the equation
- chemical equation: a representation of a chemical reaction that displays the reactants and products with chemical formulas; multiple reactants or products are separated with a plus sign, while the reactants are separated from the products with a yield arrow
- coefficient: a small whole number placed in front of a formula in a chemical equation in order to balance it
- skeleton equation: a chemical equation that shows only the formulas of the reactants and products with nothing to indicate the relative amounts

Teaching Strategies

Introducing the Lesson

Ask the class if anyone can state the law of conservation of mass. Ask if they can relate that law to Dalton's theory of a chemical reaction –that chemical reactions are rearrangements of atoms. Explain the basics between chemical equations, which they have likely seen before. Identify reactants and products, explain the various symbols that are used. Write the skeleton equation for the formation of water from hydrogen and oxygen: $H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(l)}$.

Hold up models of the three molecules and ask if anything seems wrong about this equation regarding the law of conservation of mass and Dalton's theory. This leads into an explanation of how to balance equations.

Common Misconceptions

The majority of students do well with balancing most equations when the skeleton equations are provided. The tougher skill for many to master is writing the correct skeleton equations based on a word equation or a description of the reaction. Many students will incorrectly alter the subscripts on a formula to "fit" the way that makes the equation balancing easier. For example, suppose that the word equation is:

 $sodium + oxygen \rightarrow sodium oxide$

Incorrect equations that you may see are:

 $Na + O \rightarrow NaO$ or $Na + O_2 \rightarrow NaO_2$ In both cases, the formula for sodium oxide is incorrect because the student has not accounted for the fact that it is an ionic compound with only one possible formula because of the charges of the ions. The formula for sodium oxide should be Na₂O. The other mistake with the first equation is that oxygen is not written as a diatomic molecule. It is important to stress that correct formula writing must come **before** balancing the equation. Students will need to practice this.

Science Inquiry

The simulation below is appropriate for students after they have been given a brief introduction on the meaning of a balanced chemical equation, but before they have been given a chance to practice. It has visuals which allow students to see how balancing an equation results in equal numbers of atoms of each element as well as equal mass on both sides of the equations. There are many choices for downloadable activities that can be assigned based on the time and needs of your class.

http://phet.colorado.edu/en/simulation/balancing-chemical-equations

Differentiated Instruction

Students that are having difficulty with equation balancing may benefit from the video at the following link. It goes through writing and balancing equations slowly and deliberately, pointing out many of the pitfalls that students may fall into when learning this material.

http://rmpbs.pbslearningmedia.org/resource/87402e83-2dcc-4618-8571-2047493d6412/chemistry-604-word-equations -and-balancing/

Enrichment

The link below takes students to a list of over 260 equations to balance. The list progresses from easier equations to more difficult ones. Challenge your talented students to see how many they can do, especially the harder ones.

http://www.chembuddy.com/?left=balancing-stoichiometry-questions&right=balancing-questions

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 11.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 11.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

It is important for a chemist to be able to predict the products of chemical reactions. This task is made easier by placing known reactions into specific categories based on the ways in which the substances behave in those reactions.

- What are combination and decomposition reactions?
- What are single replacement and double replacement reactions?
- What is a combustion reaction?

11.2 Types of Chemical Reactions

Key Concept

Chemical reactions can be classified into five basic types: combination, decomposition, single-replacement (or displacement), double-replacement, and combustion. A combination reaction is one in which two simple substances combine to form a single product. A decomposition reaction is one in which a compound breaks down into simpler substances. In a single-replacement reaction a more reactive element replaces a similar element that is part of a compound. Relative reactivities of elements are summarized in the activity series. A double-replacement reaction is one in which two ionic compounds exchange cation-anion partners, producing a precipitate, a gas, or a molecular compound. A combustion reaction is one in which a substance reacts with oxygen.

Lesson Objectives

- Define and give general equations for combination, decomposition, single-replacement, and double-replacement reactions.
- Classify a reaction as combination, decomposition, single-replacement, double-replacement, or combustion.
- Use the activity series to correctly predict whether a given reaction will occur.
- Predict the products of simple reactions, given only the reactants.

Lesson Vocabulary

- · activity series: a list of elements in decreasing order of their reactivity
- combination reaction: a reaction in which two or more substances combine to form a single new substance
- combustion reaction: a reaction in which a substance reacts with oxygen, releasing energy in the form of light and heat
- decomposition reaction: a reaction in which a compound breaks down into two or more simpler substances
- double-replacement reaction: a reaction in which the positive and negative ions of two ionic compounds exchange places to from two new compounds.
- single-replacement reaction: a reaction in which one element replaces a similar element in a compound

Teaching Strategies

Introducing the Lesson

Begin by writing the reactants only of two or three reactions on the board. Ask the students if they know what the products of each reaction will be. They most likely will not know. Ask if it seems like it would be valuable for a chemist to be able to predict the likely products of a chemical reaction. Discuss how all of the many millions of

known chemical reactions generally fall into five main categories and that an understanding of these categories will permit them to be able to predict products of reactions just like the ones written on the board.

Video

The video at the following link shows demonstrations of the types of reactions along with careful explanations of each.

http://rmpbs.pbslearningmedia.org/resource/2a263167-efb4-47b0-98bd-38535347e5b6/chemistry-605-types-of-react ions-and-predicting-products/

Common Misconceptions

Students generally do well with identifying the types of reaction when provided with the entire equation. The more difficult skill is predicting the products of a reaction correctly. It is important to write correct formulas first and balance the equation second. As with the previous lesson in this chapter, students may fall into the trap of writing incorrect formulas because it makes the equation balanced. For example, suppose the reaction is the single-replacement reaction of aluminum reacting with a solution of lead(II) nitrate. You may see this as a "balanced" equation:

 $Al_{(s)} + Pb(NO_3)_{2(aq)} \rightarrow Al(NO_3)_{2(aq)} + Pb_{(s)}$

This is of course incorrect, because the charge of the aluminum ion in a compound is 3+, compared to the 2+ charge of the lead(II) ion. The correct equation is:

 $2Al_{(s)} + 3Pb(NO_3)_{2(aq)} \rightarrow 2Al(NO_3)_{3(aq)} + 3Pb_{(s)}$

Students will need careful instruction and practice to avoid making these kind of mistakes.

Science Inquiry

There are many online labs dealing with the types of reactions. The one below requires students to make observations, test some gaseous products, write balanced equations, and classify the reactions.

http://academic.pgcc.edu/psc/chm101/reactions.pdf

Differentiated Instruction

Have students make cards with the general form of each type of reaction $(A + B \rightarrow AB)$ on one side and the name on the other (combination). When students look at a real chemical reaction on paper, they can compare it to the five cards. Have them identify the elements and/or ions as being A, B, or whichever letter applies. This will help them match up the chemical equation to the correct general equation.

Enrichment

Students can choose a particularly important industrial chemical and research how it is produced. Some choices include sulfuric acid (H_2SO_4), sodium hydroxide (NaOH), titanium(IV) oxide (TiO₂) among many others. They should make a poster showing the relevant chemical reactions involved in its production and identifying the type of reaction for each. Include an explanation and pictures of the end uses of the chemical.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 11.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 11.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Balanced equations allow chemists to control reactions quantitatively. The coefficients in a balanced equation represent the ideal molar ratio for the reactants in the given reaction.

• How can a balanced chemical equation be used to calculate how much of a certain product will be formed in a reaction or how much of a certain reactant will be needed to perform the reaction?



Stoichiometry

Chapter Outline

- 12.1 MOLE RATIOS
- **12.2 STOICHIOMETRIC CALCULATIONS**
- 12.3 LIMITING REACTANT AND PERCENT YIELD



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Chapter Overview

Stoichiometry is the area of chemistry that involves the calculation of amounts of substances used in a chemical reaction from knowledge of the balanced equation. The simplest stoichiometry problems use only mole-mole ratios from the coefficients of the equation. More advanced problems are two or three steps and involve mole-mass and/or mole-volume calculations in addition to the mole ratio. The limiting reactant in a chemical reaction is the reactant that would run out first during the course of a reaction and thus limit the amount of products that could be formed. The other reactant(s) are called excess reactant(s). The percent yield of a reaction is an indication of how successful a reaction has been and can be calculated from the actual yield and the theoretical yield.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 12.1: F	Pacing Lessons
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Lesson	Class Period(s) (60 min)
12.1 Mole Ratios	1
12.2 Stoichiometric Calculations	2
12.3 Limiting Reactant and Percent Yield	2

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 12.2:	Chemistry Concepts
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FlexBook® Lesson	Chemistry Concepts
12.1 Mole Ratios	12.1 –12.2
12.2 Stoichiometric Calculations	12.3 – 12.6
12.3 Limiting Reactant and Percent Yield	12.7 –12.9

12.1 Mole Ratios

Key Concept

A balanced chemical equation is similar to a recipe in that a recipe indicates the relative ratios of ingredients necessary to make a certain dish. The coefficients of the balanced equation are the ratios of moles of the reactants and products that are required for the reaction to proceed as written. Stoichiometry involves using these mole ratios to solve mathematical problems. Moles of any reactant or product can be converted to moles of any other reactant or product by use of a mole ratio conversion factor.

Standards

Lesson Objectives

- Relate balanced chemical equations to everyday analogies, such as a recipe.
- Define stoichiometry.
- Use mole ratios to convert between amounts of substances in a chemical reaction.

Lesson Vocabulary

- mole ratio: a conversion factor that relates the amounts in moles of any two substances in a chemical reaction
- stoichiometry: the area of chemistry that deals with the calculation of amounts of substances in a chemical reaction based on the balanced equation

Teaching Strategies

Introducing the Lesson

Use the ham sandwich example from the FlexBook® student edition or something similar to explain the concept of ratios in terms that students understand. Work your way towards the students being able to determine the number of any part of the sandwich that would be used given any number of another part. For example, how many ham slices to make 5 sandwiches? How many pieces of bread needed to "react" with 8 slices of cheese? How many tomato slices are needed to make the maximum number of sandwiches that can be made with 15 pickles? When you are working through these problems, make sure to write out conversion factors and cancel units every time. Require students to do the same.

Common Misconceptions

Some students may have some trouble with transferring the coefficients from the balanced equation into the conversion factors of the mole ratio. At this point, students have done mass-mole conversions and in those problems the number of moles in the conversion factor is always one. Therefore, some think that either the numerator or the denominator of the conversion factor should likewise by "1 mol". Stress this when you do a problem where neither half of the conversion factor involves the number one. An example with the Haber-Bosch equation is converting back and forth between moles of H2 and moles of NH3.

Science Inquiry

The link below is an introductory activity to mole ratios designed to help students discover the meaning of the coefficients of a balanced equation and how to make basic mole ratio calculations. It has built-in time for class discussion before students move on to practice their skills.

http://www.google.com/url?sa=t&rct=j&q=&esrc=s&source=web&cd=6&ved=0CEUQFjAF&url=http%3A%2F%2Fst affweb.psdschools.org%2Fjjasmann%2FMr_Jasmann%2FChemistry%2520Online%2FUNIT%25207%2520Stoichiome try%2520Calculations%2FWorksheets%2520and%2520Activities%2FHS_POGIL_Stoich_Help.doc&ei=YHPWU8H0A suQyAS_kIKIDg&usg=AFQjCNFaLEj-e8rYDs1rUEuuR7sXrPyt3Q&sig2=IrfC5kDrtxgql07-2oSAgg&bvm=bv.7177 8758,d.aWw&cad=rja

Differentiated Instruction

The tutorial at the site below is a careful step-by-step explanation of coefficients and mole ratios that may help students who struggle with mathematical relationships. Working in pairs, have students work through the tutorial and show their work on paper for the problems at the end of the tutorial.

https://www.wisc-online.com/learn/general-education/general-chemistry/gch7304/mole-ratios-in-chemical-equations

Enrichment

Have advanced students research the Haber-Bosch process, including its historical development, the technology used in the production of ammonia, and the uses of ammonia in modern society. Students can also study how the reaction is performed in a way to maximize the production rate of ammonia. This links to a later lesson on equilibrium processes when they will study LeChâtelier's principle. Another avenue of study is to research how ammonia is present in much larger amounts in the atmospheres of some planets and moons.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 12.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 12.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

In the laboratory, there is no way for a chemist to directly measure out moles of a substance. He or she, instead, must determine amounts either by mass or by volume. Therefore, it is important to be able to perform stoichiometric calculations with mass and volume.

- How is mass converted to moles?
- How are moles converted to mass?
- How many conversion factors are required to convert from the mass of a given substance to the mass of an unknown substance?

12.2 Stoichiometric Calculations

Key Concept

All stoichiometry problems involve a mole ratio between two reactants, two products, or one reactant and one product. However, amounts of reactants and products are normally measured by mass or by volume. Stoichiometry problems can be extended to include mole-mass and/or mole-volume calculations. Because of Avogadro's hypothesis, mole ratios among gases are equal to volume ratios and so volume-volume problems can be done in one step. Mass-mass and mass-volume problems involve a total of three steps. Mass-volume or volume-mass problems make use of the gaseous molar volume of 22.4 L/mol at STP.

Standards

Lesson Objectives

- Calculate the amount in moles of a reactant or product from the mass of another reactant or product. Calculate the mass of a reactant or product from the moles of another reactant or product.
- Calculate the mass of a reactant or product from the mass of another reactant or product.
- Create volume ratios from a balanced chemical equation.
- Use volume ratios and other stoichiometric principles to solve problems involving mass, molar amounts, or volumes of gases.

Teaching Strategies

Introducing the Lesson

Begin the lesson with a refresher of mole ratio stoichiometry problems, emphasizing how all of the problems are one-step calculations. Ask students where in the laboratory they would find the "mole-o-meter" that measures out moles of substances. Hopefully, they will realize there is no such thing. Ask them how amounts of substances are measured in the lab (mass, volume). Ask if they how to calculate moles of a substance from the mass or from the volume. Review how this is done. Now discuss how "real" stoichiometry includes not only the mole ratio problems that they have just learned, but also mass-mole and volume-mole conversion problems, turning the problems into two or three steps.

Common Misconceptions

Once problems have become two or three steps, some students will make the mistake of using the coefficients from the balanced equation in all three steps. Here is an example, from practice problem 3 in the FlexBook® student edition. The equation is:

 $2Fe(OH)_{3(s)} + 3H_2SO_{4(aq)} \rightarrow Fe_2(SO_4)_{3(aq)} + 6H_2O_{(l)}$

The question involves finding the mass of sulfuric acid need to react with 12.72 g of iron(III) hydroxide. The correct solution is:

 $12.72 \ g \ Fe(OH)_3 \times \frac{1 \ mol \ Fe(OH)_3}{106.88 \ g \ Fe(OH)_3} \times \frac{3 \ mol \ H_2SO_4}{2 \ mol \ Fe(OH)_3} \times \frac{98.08 \ g \ H_2SO_4}{1 \ mol \ H_2SO_4} = 17.51 \ g \ H_2SO_4$

Notice that the 3 and the 2 coefficients from the balanced equation are only used in the middle step. Some students may produce this:

 $12.72 \ g \ Fe(OH)_3 \times \frac{2 \ mol \ Fe(OH)_3}{106.88 \ g \ Fe(OH)_3} \times \frac{3 \ mol \ H_2SO_4}{2 \ mol \ Fe(OH)_3} \times \frac{98.08 \ g \ H_2SO_4}{3 \ mol \ H_2SO_4} = 11.67 \ g \ H_2SO_4$

Here, they have incorrectly used the 3 and the 2 coefficients in the mass-mole conversion steps. Make sure to stress that the mole ratio is only for the middle step and to watch out for this type of error.

Science Inquiry

The activity below is called the Fizzy Drink Experiment and allows students to experiment making Kool-Aid drinks with varying amounts of fizz from baking soda and doing stoichiometric calculations involving the amounts of Kool-Aid, sugar, baking soda, and citric acid.

http://www.google.com/url?sa=t&rct=j&q=&esrc=s&source=web&cd=2&cad=rja&uact=8&ved=0CCIQFjAB&url= http%3A%2F%2Fwww.wmich.edu%2Fscience%2Fblast%2FFiles%2FStoichiometry%2FFizzyDrinkStoichiometryLe sson.doc&ei=3XLWU66XA8W0yASUIYDADA&usg=AFQjCNG1VXC1GktY08Ew9k8KZEQaqH9WzQ&sig2=a zIyv3CoxtVp6_Ds39YVUw

Differentiated Instruction

Students that are challenged by the math of stoichiometry problems may benefit by making a set of index cards for each type of stoichiometry problem: (1) mole-mole, (2) mole-mass, (3) mass-mole, (4) mass-mass, (5) mass-volume, (6) volume-mass, and (7) volume-volume. The cards should have the information needed for each conversion factor phrased as "given" or "unknown" along with the correct number of conversion factors. Students can use the cards to help them solve practice stoichiometry problems.

Enrichment

Students can take a simple recipe from a cookbook and try to turn it into a chemical equation. This will involve many unit conversions such as cups and tablespoons into metric conversion units. Students will need to find out the approximate densities of ingredients so that they can figure out the masses and then turn those masses into moles of the various reactants. Have them construct the various mole ratios, such as the mole ratio of flour to baking powder in a recipe for cookies. Some approximations may need to be made for ingredients that are mixtures rather than pure substances.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 12.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 12.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

This lesson dealt with ideal stoichiometry, where 100% of the reactants were converted to products. In the real world, many chemical reactions do not proceed entirely in this way.

- How can we calculate the amount of products that could be formed in a reaction when two or more reactants are combined in a ratio other than the mole ratio from the balanced equation?
- How can we express the extent to which a set of reactants is converted to products if it is less than 100%?

12.3 Limiting Reactant and Percent Yield

Key Concept

Chemical reactions can still be performed even when the ideal ratio of reactants is not present. The limiting reactant is the one which will run out first in the course of the reaction. The reactant which remains is called the excess reactant. The theoretical yield of a reaction is the maximum amount of a product that will be formed based upon stoichiometry. Actual yield is the reaction yield that is achieved in the laboratory. Percent yield is the ratio of actual yield to theoretical yield, expressed as a percentage.

Standards

Lesson Objectives

- Analyze a chemical reaction in order to determine which reactant is the limiting reactant and which is the excess reactant.
- Calculate the amount of excess reactant remaining after a reaction is complete.
- Calculate the theoretical yield of a reaction when the available amounts of each reactant are known.
- Calculate the percent yield of a reaction based on the theoretical and actual yields.

Lesson Vocabulary

- actual yield: the amount of product that is formed when a reaction is carried out in the laboratory
- excess reactant (reagent): the reactant that is initially present in a greater amount than will eventually be reacted; some excess reactant is left over after the reaction is complete
- limiting reactant (reagent): the reactant that is used up first in a reaction; the limiting reactant determines the amount of product that can be formed in a chemical reaction
- percent yield: the ratio of the actual yield to the theoretical yield, expressed as a percentage
- theoretical yield: the maximum amount of product that could be formed from the given amounts of reactants; theoretical yield is based on ideal stoichiometry

Teaching Strategies

Introducing the Lesson

Begin the discussion of limiting reactants with a real-world example or a demonstration. The pancake-making discussion in the lesson would work. Perhaps you could say that you are going to mail some letters. Show the

students that you have paper, envelopes, and stamps –but different numbers of each. Ask them which item limits the number of letters you can send. Than change it up by adding in more of whatever was limiting. What is the limiting factor now? Get the students comfortable with the idea of limiting and excess reactants before moving on to chemical equations.

Common Misconceptions

There are several tricky spots for students in the various calculations of this lesson. Make sure that students do not think that the limiting reactant is simply whichever reactant is present in a smaller amount. The masses of reactants must be converted to moles and the mole ratio must be taken into account. As an analogy, suppose you were building bicycles and had 3 seats and 5 wheels. You have fewer seats, but wheels are the limiting reactant because two are needed for each bicycle.

When students are asked to calculate the mass of an excess reactant, some may mistakenly calculate how much more of the limiting reactant would be needed to react in full with the excess reactant.

Stress that the theoretical yield of a product must be based on the starting amount of the limiting reactant. Therefore, the limiting reactant needs to be used in the stoichiometry calculation and not the excess reactant. If they use the starting mass of the excess reactant, they will come up with a theoretical yield that is too large.

Science Inquiry

The inquiry activity below involves students generating a gas by reacting various amounts of baking soda and vinegar. The optimum mole ratio is determined by observing how inflated a balloon will get due to the carbon dioxide gas produced.

http://lorenowicz.weebly.com/uploads/4/6/1/6/4616010/sch3u_limiting_reagent_lab.pdf

Differentiated Instruction

The simulation at the link below may help students who need a more visual approach to limiting reactants. It uses the examples of sandwich making and allows students to see how much of each ingredient is used up as sandwiches are made as well as how much is left over. It also has examples of chemical reactions and a game for students to play when they have mastered the concept. There are many downloadable lesson plans and activities available on the same page.

http://phet.colorado.edu/en/simulation/reactants-products-and-leftovers

Enrichment

When coal is combusted in a power plant, some polluting sulfur dioxide (SO_2) gas is released. Power plants can negate the harmful effects of the SO_2 by adding powdered calcium oxide to the coal to be combusted. The calcium oxide reacts with the sulfur dioxide:

$$2CaO_{(s)} + 2SO_{2(g)} + O_{2(g)} \rightarrow 2CaSO_{4(s)}$$

Have interested students try to figure out what mass of CaO is needed to remove all of the sulfur dioxide from a ton (2000 g) of coal. They will need to research different grades of coal to determine how much sulfur dioxide is typically produced as coal burns. Have them report on the process of how power plants incorporate this reaction into the combustion process and what is done with the calcium sulfate that is produced.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 12.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 12.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Now that you have an understanding of the quantitative relationships that exist between chemical substances in chemical reactions, we will turn our attention to the states of matter and their relationships to one another.

- What are the three states of matter?
- How can changes in temperature and pressure affect the state of a sample of matter?


States of Matter

Chapter Outline

- 13.1 KINETIC-MOLECULAR THEORY AND GASES
- 13.2 LIQUIDS
- 13.3 SOLIDS
- 13.4 CHANGES OF STATE



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Chapter Overview

Common states of matter include solids, liquids, and gases. The kinetic-molecular theory describes the macroscopic behavior of matter by envisioning particles as small spheres in constant and random motion. Gases are distinguished by the very large amounts of space between the particles and by the pressure that gases can exert. Liquids and solids are called the condensed phases because intermolecular forces have forced the particles to be much closer together. Vapor pressure is the pressure exerted by a gas above a liquid or a solid. Most solids have a very ordered structure called a crystal. Changes of state can be described by heating curves and by phase diagrams.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

Lesson	Class Period(s) (60 min)
13.1 Kinetic-Molecular Theory and Gases	2
13.2 Liquids	1-2
13.3 Solids	1 - 2
13.4 Changes of State	1 –2

TABLE 13.1: Pacing Lessons

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

FlexBook® Lesson	Chemistry Concepts
13.1 Kinetic-Molecular Theory and Gases	13.1 –13.5
13.2 Liquids	13.7 –13.11
13.3 Solids	13.12 –13.17
13.4 Changes of State	13.18-13.20

TABLE 13.2: Chemistry Concepts

13.1 Kinetic-Molecular Theory and Gases

Key Concept

The kinetic-molecular theory is a set of assumptions that describes the macroscopic behavior of particles of matter. An ideal gas behaves exactly according to the kinetic-molecular theory. Gas pressure results from the collisions of gas particles with an object. Atmospheric pressure is measured with a barometer, with standard pressure being equal to 760 mmHg. Pressure can be reported in several different units, which can be interconverted using dimensional analysis. Temperature on the Kelvin scale is directly proportional to the average kinetic energy of the particles in a sample of matter. Absolute zero is a theoretical temperature at which all molecular motion ceases.

Standards

Lesson Objectives

- State the main points of the kinetic molecular theory, and describe how it relates to the properties of an ideal gas.
- Define pressure and describe how gases exert pressure.
- Understand the barometer and how it measures atmospheric pressure. Convert between units of gas pressure.
- Relate temperature to average kinetic energy.

Lesson Vocabulary

- absolute zero: The temperature at which the motion of particles theoretically ceases.
- atmospheric pressure: The pressure exerted by gas particles in Earth's atmosphere as those particles collide with objects.
- barometer: An instrument used to measure atmospheric pressure.
- gas pressure: The pressure that results from collisions of gas particles with an object.
- ideal gas: An imaginary gas whose behavior perfectly fits all the assumptions of the kinetic-molecular theory.
- kinetic-molecular theory: A theory that explains the states of matter and is based on the idea that matter is composed of tiny particles that are always in motion.
- pascal: The standard unit of pressure.
- pressure: Applied force per unit area on a surface.

Teaching Strategies

Introducing the Lesson

Students are familiar with the basic descriptions of the states of matter, from earlier science classes as well as Chapter 2 of this FlexBook® student edition. Ask students to make quick sketches of each of the states of matter, showing each particle as a small circle. Focus on the arrangements of the particles and the distances between them. Firmly establish that the particles of a gas are very, very far apart from one another. Ask the students what unique properties of gases are a direct result of this model. See if they can come up with compressibility, fluidity, lack of attractive forces, and what causes gases to exert pressure. This leads directly to a discussion of the points of the kinetic-molecular theory, especially as it relates to gases.

Common Misconceptions

Students may think that mercury is the only liquid that could be used in a barometer. Explain that mercury was traditionally used in a barometer because of its high density and that it evaporates very slowly. It also does not expand or contract a great deal with temperature. Any liquid, including water, could be used in a barometer. However, because of its lower density, a true barometer constructed with water would need to be much, much higher because the atmospheric pressure would push the column of water more easily than the column of mercury.

Some students may mistakenly conclude that temperature on the Celsius scale is also directly proportional to average kinetic energy. An example may clear up why this is not so. Have them convert 300 K and 600 K to degrees Celsius (27°C and 327°C). In which case is the higher temperature double the lower temperature? Doubling the Kelvin temperature means doubling the kinetic energy of the particles because the Kelvin scale is linked to molecular motion.

Science Inquiry

The simulation titled "States of Matter Basics" at the link below allows students to investigate the particle nature of the three states of matter.

http://phet.colorado.edu/en/simulation/states-of-matter-basics

Many downloadable lesson plans are available to accompany the "States of Matter Basics" simulation, though several are aimed at the middle school and even grade school level. The following link is to a lesson plan and other documents to accompany the simulation that would be appropriate for the high school chemistry level. Some introduction to the kinetic-molecular theory would be helpful before students use the simulation.

http://phet.colorado.edu/en/contributions/view/3496

Note that there is another very similar simulation at this same Phet site called "States of Matter" (without the Basics). This simulation is slightly beyond the scope of this particular lesson, as it covers phase changes and phase change diagrams. Save it for the last lesson of this chapter.

Differentiated Instruction

Have groups of students make large detailed illustrations of the five points of the kinetic-molecular theory. If done on butcher paper, they could be displayed around the room. This may help them in further study of the states of matter.

Enrichment

Interested students can build their own barometers, though not with mercury. The easily obtainable materials include a plastic ketchup bottle and a large drinking glass. The link at the site below gives instructions. Have students bring in their finished product, maybe as an extra credit assignment.

http://www.teachengineering.org/view_activity.php?url=collection/cub_/activities/cub_weather/cub_weather_lesson0 2_activity1.xml

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 13.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 13.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The particles of a liquid are much closer together than the particles of a gas, resulting in far different behaviors for liquids as compared to gases.

- How does the kinetic energy of the particles in a liquid relate to the rate of evaporation of the liquid?
- How is the boiling point of a liquid defined?

13.2 Liquids

Key Concept

Liquids differ greatly from gases because the particles of a liquid are much closer together. Properties of a liquid include fluidity and surface tension. Liquids evaporate when particles with the highest kinetic energy escape from the surface and become a gas. The vapor pressure of a liquid is the pressure exerted by a vapor above the liquid's surface and is related to the strength of the intermolecular forces acting on the particles of the liquid. As temperature increases, vapor pressure increases. The relationship can be shown in a vapor pressure curve. The boiling point of a liquid is reached when the temperature of the liquid is equal to the surrounding atmospheric pressure.

Standards

Lesson Objectives

- Describe a liquid according to the kinetic-molecular theory.
- Describe how a liquid exhibits surface tension.
- Describe the evaporation of a liquid and its relationship to the kinetic energy of the evaporating particles.
- Define vapor pressure and understand its relationship to intermolecular forces and to the temperature of the liquid.
- Describe the process of boiling and differentiate between boiling point and normal boiling point.
- Use a vapor pressure curve to determine boiling points at different atmospheric pressures.

Lesson Vocabulary

- boiling point: The temperature at which the vapor pressure of a liquid is equal to the external pressure.
- condensation: The change of state from a gas to a liquid.
- evaporation: The conversion of a liquid to its vapor below the boiling temperature of the liquid.
- fluid: A substance that is capable of flowing from one place to another and takes the shape of its container.
- normal boiling point: The temperature at which the vapor pressure of a liquid is equal to standard pressure.
- surface tension: A measure of the elastic force in the surface of a liquid.
- vapor pressure: A measure of the pressure exerted by a gas above a liquid in a sealed container.
- vaporization: The process in which a liquid is converted to a gas.

Teaching Strategies

Introducing the Lesson

Begin by asking students to write down as many properties of liquids as they can. Discuss how a liquid is very different from a gas because the particles are so much closer together. Liquids are referred to as a condensed phase. The demonstrations below are a good way to show some of the properties of liquids.

Demonstrations

The surface tension of a liquid is easy to demonstrate. Fill a 1-liter beaker with water and place on the overhead projector. Carefully sprinkle a powder (baby powder or sulfur works fine) onto the surface of the water. It will not sink even though the powder is denser than the liquid. Break the surface tension by squirting a drop of soap or detergent into the middle of the surface. The powder will rapidly sink to the bottom.

The following is a simple demonstration of vapor pressure and the equilibrium that exists between a volatile liquid (acetone) and its vapor.

https://www.flinnsci.com/media/620497/91211.pdf

The effect of temperature on vapor pressure can be demonstrated rather easily. Stopper an Erlenmeyer flask that has been partly filled with water. Talk to the students about the vapor pressure exerted by the water vapor. Place the flask on a hot plate and turn up the heat. Go on with the lesson. At some point the stopper will be blown out of the flask, startling the students. Make sure the hot plate is a safe distance away from the class.

If you have excess to a vacuum pump with a bell jar, use it for some entertaining demonstrations. Show how a slightly inflated balloon will expand in a vacuum. Marshmallows will also expand in a vacuum. Finally, show how water will boil at room temperature under vacuum. Many students will think that the water has somehow been heated, so make sure to prove that the temperature hasn't changed by using a thermometer.

Common Misconceptions

Students may believe that evaporation and vaporization mean the exact same thing. Evaporation strictly refers to the liquid to gas state change when it occurs at a temperature below the boiling point. Vaporization is the more general term and refers to either evaporation or boiling.

Students often have difficulty with the relationship of vapor pressure to intermolecular forces and boiling point. A high vapor pressure correlates with weak forces and a low boiling point. Conversely, a low vapor pressure correlates with strong forces and a high boiling point. Use example data and vapor pressure curves to help explain the concepts.

The difference between boiling point and normal boiling point can be confusing. Explain that the boiling point of water varies with elevation because air pressure decreases with increasing elevation, but that the normal boiling point of water is always exactly 100°C.

Science Inquiry

The following link contains a detailed description of vapor pressure that can be used as an introduction to the concept. It includes information on vapor pressure curves, the meaning of boiling point, analysis questions for students to answer, and a brief activity.

https://pogil.org/uploads/media_items/vapor-pressure.original.pdf

Differentiated Instruction

Some students have difficult visualizing the movement of molecules during vaporization, condensation, and the exertion of pressure by the vapor. The links below take students through short tutorials on vapor pressure. The first tutorial illustrates how vapor pressure works, while the second one examines the relationship of boiling points to external pressure.

https://www.wisc-online.com/learn/general-education/general-chemistry/gch4304/vapor-pressure

https://www.wisc-online.com/learn/natural-science/chemistry/ele2107/how-pressure-changes-boiling-temperature

Enrichment

Advanced mathematics students can see that the relationship between vapor pressure and temperature is logarithmic. Have them use the data from Table 13.1 and prepare graphs of log P versus 1/T, where T is the temperature on the Kelvin scale (which will need to be calculated). Prepare one graph with all three sets of data. The graphs will yield straight lines. The most accurate graphs should be made with a graphing program like Microsoft Excel rather than being drawn by hand. What is the relationship of the slope of the line to the strength of intermolecular forces? (The larger the slope relates to stronger forces.)

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 13.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 13.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Solids are the other condensed phase of matter, with particles held closely together by attractive forces.

- What are some of the ways in which the particles of a solid are arranged?
- What happens when the temperature of a solid reaches its melting point?

13.3 Solids

Key Concept

Solids, like liquids, are a condensed state of matter. However, the particles of a solid are held rigidly in place, usually in an ordered, repeating, three-dimensional structure called a crystal. The melting point of a solid is the temperature at which the intermolecular forces holding the particles in place are overcome. Solids have a low vapor pressure, though some solids pass directly from the solid state to the gas state in a process called sublimation. Crystals are categorized into one of seven systems based on their shape. A unit cell is the simple arrangement of particles (ions, atoms, molecules) of a crystal that is repeated over and over again. Crystalline substances include ionic crystals, metallic crystals, covalent network crystals, and molecular crystals. A solid which lacks an ordered structure is called an amorphous solid.

Standards

Lesson Objectives

- Describe a solid according to the kinetic-molecular theory.
- Understand that a solid also has a vapor pressure, and describe the relationship between the vapor pressure of a solid and sublimation.
- Describe the features of the seven basic crystal systems.
- Define a unit cell.
- List the four classes of crystalline solids and describe the properties of each.
- Describe an amorphous solid.

Lesson Vocabulary

- amorphous solid: A solid that lacks an ordered internal structure.
- crystal: A solid substance in which the particles are arranged in an orderly, repeating, three-dimensional pattern.
- deposition: The change of state from a gas to a solid.
- melting point: The temperature at which a solid changes into a liquid; it is the same as the freezing point.
- sublimation: The change of state from a solid to a gas without passing through the liquid state.
- unit cell: The smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire crystal.

Teaching Strategies

Introducing the Lesson

Begin by asking the students about the characteristic properties of solids. Two main points to make are that: (1) the particles of a solid are still in motion, but that motion is restricted to vibrational only, and (2) the majority of solid substances consist of particles in a very ordered, repeating structure. Some examples of crystalline solids would be useful to show the students. An earth science or geology teacher may have some samples you could use. Students should be familiar with the concept of a crystal lattice from earlier chapters about ionic compounds.

Demonstration

The sublimation of iodine is a striking demonstration and easy to do. Place a few crystals of iodine into a large Erlenmeyer flask and cover with a watch glass. Place an ice cube on top to he watch glass and place the flask on low heat on a hot plate. The flask will gradually fill with purple iodine vapor. After some time has passed, carefully move the flask to the fume hood and remove the watch glass. Very small spike-shaped crystals of iodine will have formed on the underside of the watch glass. Leave the flask and the watch glass in the fume hood overnight and most, if not all, of the iodine will have sublimed away by the next day.

Common Misconceptions

Students may think that substances either sublime or melt, but can't do both. Point out that carbon dioxide does have a liquid state, but that it just does not exist at normal atmospheric pressures. Ice certainly melts, but it sublimes as well. That is why snow will gradually disappear even if the temperature remains below freezing. Solids and liquids both have vapor pressures, but that of solids is generally much lower.

The four classes of crystalline solids are very important because an understanding of the differences is key to the physical properties of each substance. Most students are fine with ionic and metallic since those have been covered before. However, they may struggle with the difference between covalent network solids and molecular solids. Make sure to stress that covalent bonds are much stronger than intermolecular forces.

Science Inquiry

The following activity allows students to investigate the shapes of crystals and to identify the seven crystal systems. To do all of the parts of the activity will require microscopes or good hand lenses.

Download the resource at the link below.

http://www.sciencelearn.org.nz/content/download/11607/229990/version/6/file/Crystal+systems.doc.

Differentiated Instruction

Have students make artistic renderings of the seven crystal systems and of the three types of cubic unit cells. These could include colorful drawings or three-dimensional models. Assign different crystals or unit cells to pairs or small groups of students. Display the results around the room.

Enrichment

Liquid crystal displays (LCDs) are used in devices such as calculators and wristwatches. Liquid crystals exhibit the properties of both liquids and solids. Have interested students research LCDs to find out how they have been developed and how they work.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 13.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 13.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

As part of the discussions about the three states of matter, you have seen how matter undergoes changes from one state to another.

- Under what conditions of temperature and pressure do the various changes of state occur?
- What is a phase diagram, and how can it be used to understand the nature of a given substance?

13.4 Changes of State

Key Concept

The changes of state of a substance can be shown with a heating/cooling curve, a graph of temperature of a substance as heat is constantly supplied or removed. The temperature of any substance is constant during a state change. A phase diagram shows the state of a substance as a function of its temperature and pressure. The lines of the diagram represent equilibrium between two states. The triple point is the only temperature/pressure where all three states are in equilibrium. The critical point is the highest temperature at which the substance can possibly exist in the liquid state.

Standards

Lesson Objectives

- Interpret heating and cooling curves.
- Know the terms for the six different changes of state.
- Describe the general features of a phase diagram, including the triple point and the critical point.

Lesson Vocabulary

- critical pressure: The pressure that must be applied to a gas at its critical temperature in order to turn it into a liquid.
- critical temperature: The highest temperature at which a substance can possible exist as a liquid.
- phase diagram: A graph showing the conditions of temperature and pressure under which a substance exists in the solid, liquid, and gas states.
- triple point: The one condition of temperature and pressure where the solid, liquid, and vapor states of a substance all coexist at equilibrium.

Teaching Strategies

Introducing the Lesson

Changes of state are now familiar to students, but make sure to review the vocabulary. The heating curve can be introduced as a lab activity (see Science Inquiry below) or you can simply draw one on the board. Emphasize that during the plateaus on the graph, the temperature is remaining constant. Ask the students about this. If energy is still being supplied to the substance, yet the temperature remains constant, what type of energy is being added? Lead

them to the idea that it can't be kinetic energy (temperature constant), so it must be potential energy. Ask them to rank the states of matter from highest potential energy to lowest (gas >liquid >solid).

Common Misconceptions

Related to the discussion above, students may have a hard time with the notion that adding heat to an ice-water (or water-steam) mixture does not change its temperature. Put another way, an ice water mixture that is truly at equilibrium is always at 0°C, no matter the relative amounts of ice and water that are present. This goes against what we normally may experience because a simple glass of ice water is likely not truly at equilibrium. It would have to be isolated from its surroundings.

Point out how the line between the liquid and gas state on a phase diagram is the same thing as a vapor pressure curve. The students may not realize this.

The critical point can be a difficult concept. With respect to water, make sure to stress that the critical temperature and critical pressure are extremely high and so this is not something easy to experience directly. Ask the students to imagine the molecules moving around so fast that no amount of pressure, no matter how high, can possibly force the gas to condense to a liquid.

Science Inquiry

Students can generate their own heating curve for water with the lab activity found at the link below. This link takes you to a page which describes the lesson, while the lab instructions, titled "Heating Curve of Water" can be downloaded from the link provided in the document.

http://www.cpalms.org/Public/PreviewResource/Preview/51213

The activity at the link below is a good introduction and description of phase diagrams. In the second part of the activity, students construct their own phase diagram from supplied data.

http://pgchemistry.pbworks.com/w/file/fetch/49647910/Student%20Activity%20Phase%20Diagrams.pdf

Differentiated Instruction

Less proficient readers can make a concept map or similar diagram from the extensive vocabulary in this chapter. Have them focus on the three states, the names for the phase changes, and the particular properties of each state. Diagrams should include concepts such as vapor pressure, crystal types, and gas pressure.

Enrichment

Freeze drying is an interesting process in which food is frozen and placed into a chamber that is attached to a vacuum pump. The ice crystals sublime faster at high pressure, leaving freeze-dried foods that can later be reconstituted by simply adding water. If you or one of the students has a freeze drying machine, you can try it out in class. If not, have interested students research the process and make a presentation to the class on the advantages of freeze-dried foods and where they are used most often.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 13.4 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 13.4 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The description of a gas, according to the kinetic-molecular theory, is that gas particles are very far apart from one another and so the particles do not interact with each other. This description leads to a set of mathematical relationships concerning the amount, temperature, pressure, and volume of a gas sample.

- How would the pressure of an enclosed sample of gas be affected if more gas were added?
- How would the pressure of an enclosed sample of gas be affected if the temperature were increased?
- How would the pressure of an enclosed sample of gas be affected if the volume of the container were increased?

CHAPTER **14** The Behavior of Gases

Chapter Outline

- 14.1 GAS PROPERTIES
- 14.2 GAS LAWS
- 14.3 IDEAL GASES
- 14.4 GAS MIXTURES AND MOLECULAR SPEEDS



Garrett Heath. http://www.flickr.com/photos/garrettheath/10262331596/ . CC BY 2.0.

Chapter Overview

Gases are compressible because of the large amount of empty space between the particles. Four variables describe the condition of a gas: amount of gas (moles), volume, temperature, and pressure. The gas laws are simple mathematical relationships between the variables. The simplest laws involve two variables changing while two are held constant. The ideal gas law allows any one of the four variables to be determined if the other three are known. Gases are not truly ideal under all conditions, but are more ideal when the pressure is low and the temperature is high. Dalton's law of partial pressures concerns the pressure of a single gas when it is part of a mixture. The relative speeds of gas molecules are inversely proportional to the square roots of their molar masses according to Graham's law.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

Lesson	Class Period(s) (60 min)
14.1 Gas Properties	1
14.2 Gas Laws	2
14.3 Ideal Gases	2
14.4 Gas Mixtures and Molecular Speeds	1-2

TABLE 14.1: Pacing Lessons

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 14.2:	Chemistry	Concepts
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FlexBook® Lesson	Chemistry Concepts
14.1 Gas Properties	14.1 –14.2
14.2 Gas Laws	14.3 –14.7
14.3 Ideal Gases	14.8 –14.11
14.4 Gas Mixtures and Molecular Speeds	14.12 –14.15

14.1 Gas Properties

Key Concept

Gases are compressible because of the large amount of space between the particles. The four variables that describe any gas are the amount, the volume of the container, the temperature, and the pressure. An increase in the amount of gas in a container will increase the pressure. Increasing the volume of a container will cause gas pressure to decrease. Heating an enclosed sample of gas will result in a pressure increase.

Standards

Lesson Objectives

- Describe how a gas can be compressed.
- Identify three factors that affect gas pressure. Describe the effects according to the kinetic-molecular theory.

Lesson Vocabulary

• compressibility: A measure of how much a given volume of matter decreases when placed under pressure.

Teaching Strategies

Introducing the Lesson

Last chapter's lesson "Kinetic-Molecular Theory and Gases" has given students a good understanding of the particle nature of gases. The focus on this lesson is on the four variables which are used to describe the conditions of a gas. Start the lesson by asking if the students can name them (amount, volume, temperature, pressure). Either use the simulation described in the Science Inquiry section below or draw boxes of gas particles on the board. Ask students to predict the effect of amount, volume and temperature (separately) on the pressure of the gas in the box.

Common Misconceptions

Students generally have little difficulty with the concepts of this lesson. Make sure to stress the difference between a rigid container and a flexible container. Discuss how many containers have the characteristics of both. If you add gas to a completely deflated basketball, the volume of the ball will increase as it inflates. However, the air pressure inside the ball increases as well. Once the ball nears its full size, the addition of more gas primarily effects only the pressure since the ball has reached its maximum volume.

Science Inquiry

The simulation at the link below is a nice review of the properties of a gas and works well for visual learners. Students investigate the effects of gas amount, volume and temperature on the pressure of a gas sample.

http://phet.colorado.edu/en/simulation/gas-properties

Many downloadable activities are available on the page. Choose one that fits with the level of the class and the time allotted to using the simulation. The link below works well if you would like your students to discover the mathematical relationships between the variables (the gas laws). Several files are available from this page.

http://phet.colorado.edu/en/contributions/view/3500

Differentiated Instruction

Have students make a large 4×4 table with amount, volume, temperature, and pressure on each row and column. In each box, have them describe how a change in the row variable will affect the column variable. For example, consider the box in the volume row and the pressure column. Inside, the student would explain how an increase in volume results in a decrease in gas pressure and a decrease in volume results in an increase in pressure. They should draw a large X where the same two variables intersect each other in the table.

Enrichment

Have interested students investigate how an aerosol spray can works. Have them make a detailed diagram on a poster and present to the class.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 14.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 14.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

The relationships between pressure, volume, temperature, and amount of gas can be treated mathematically in a set of equations called the gas laws.

- How many gas laws must there be if each law includes the relationship between two of the variables?
- Can a single gas law relate all of the variables to one another?

14.2 Gas Laws

Key Concept

Gas laws are a set of simple mathematical relationships between the variables of a gas. Boyle's law states that pressure and volume have an inverse relationship. Charles's law states that volume and absolute temperature have a direct relationship. Gay-Lussac's law states that pressure and absolute temperature have a direct relationship. The combined gas law relates pressure, volume and temperature together. Avogadro's law states that the amount of gas in a container is directly related to the volume or to the pressure. Problems for all laws generally involve a set of before and after conditions, where two or more of the variables have been allowed to change.

Standards

Lesson Objectives

- Use Boyle's law to calculate pressure-volume changes at constant temperature.
- Use Charles's law to calculate volume-temperature changes at constant pressure.
- Use Gay-Lussac's law to calculate pressure-temperature changes at constant volume.
- Use the combined gas law to solve problems in which pressure, volume, and temperature all change.
- Use Avogadro's law to calculate volume-number of moles changes at constant temperature and pressure.

Lesson Vocabulary

- Avogadro's law: A gas law which states that the volume of a gas is directly proportional to the number of moles of gas when the temperature and pressure are held constant.
- Boyle's law: A gas law which states that the volume of a given mass of gas varies inversely with the pressure when the temperature is kept constant.
- Charles's law: A gas law which states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when the pressure is kept constant.
- Combined gas law: Gas law which expresses the relationship between the pressure, volume, and absolute temperature of a fixed amount of gas.
- Gay-Lussac's law: A gas law which states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas when the volume is kept constant.

Teaching Strategies

Introducing the Lesson

From the previous lesson, students should know the four variables used to describe a gas (amount, volume, temperature, pressure). Discuss how the gas laws are simple mathematical relationships in which two of the four variables are allowed to change while two are held constant. At first, you will be dealing with the situations in which the amount of gas is always constant, meaning that no gas is allowed in or out of the container. Work your way through each law (Boyle's, Charles's, Gay-Lussac's), emphasizing whether the relationship between the variables is direct or inverse. Provide the students ample practice problems as you work through each law.

Demonstrations

There are many terrific demonstrations for the relationships described by the gas laws.

Put about 400 mL of water in a 1-liter beaker and add some food coloring. Invert a 400-mL beaker and place it inside the larger beaker. Heat the water on a hot place until it boils for a minute. Carefully remove the beaker combination from the hot plate and place on the table. As the steam cools and recondenses, it will create a vacuum, forcing the water up into the inner beaker. The effect can be sped up by squirting cold water onto the outside of the large beaker.

The classic "egg in a bottle" demonstration illustrates Gay-Lussac's law. Find a glass bottle or Erlenmeyer flask whose neck is just the right diameter that a hard-boiled egg will fit loosely into it without falling in. Put a little grease or oil around the neck of the bottle. Set a wooden splint on fire and drop it carefully into the bottle. Wait a few seconds so the air in the bottle heats up. Place the egg on top of the bottle. As the air cools, the egg will be pulled into the bottle.

Some other ideas are provided at the links below.

http://ncsu.edu/project/chemistrydemos/GasLaw/Gas%20Law.pdf

http://www.google.com/url?sa=t&rct=j&q=&esrc=s&source=web&cd=5&ved=0CD8QFjAE&url=http%3A%2F%2Fgr ade11chemistry.wikispaces.com%2Ffile%2Fview%2Fgas%2Blaw%2Bdemo.doc

Common Misconceptions

Some students will forget that all gas law problems that involve temperature (Charles's, Gay-Lussac's, combined) must utilize the Kelvin scale for temperature. So if a problem gives a T1 in Celsius and asks for T2 in Celsius, the T1 must first be converted to Kelvin. After the gas law problem is done, the resulting T2 is then converted from Kelvin back to Celsius. This is not the same thing as doing the gas law problem with the temperatures in Celsius. You may want to prove this to the students with an example.

Students generally have little difficulty solving gas law problems when all of the problems relate to one particular gas law. It is essential that students be able to figure out how to solve problems when many different gas laws are all "mixed up" on one worksheet or problem set. After introducing all of the gas laws in this lesson, provide an opportunity for the students to do this.

Science Inquiry

Gas laws are simple enough that they provide a great opportunity for relatively quick inquiry activities.

The link below takes you to a Boyle's law lab, including a video on how the lab works and student instruction page.

www.ck12.org

http://www.flinnsci.com/teacher-resources/teacher-resource-videos/best-practices-for-teaching-chemistry/gas-laws/boy le%27s-law-lab/

The link below takes you to a Charles's law lab and video. The lab also involves the students making an extrapolation of their data to fins absolute zero.

http://www.flinnsci.com/teacher-resources/teacher-resource-videos/best-practices-for-teaching-chemistry/gas-laws/cha rles-law-and-absolute-zero/

Differentiated Instruction

Match up students of high mathematical ability with those who may have more difficulty with algebraically solving for an unknown variable in an equation. Have them work on practice problems together until both can solve them independently.

Enrichment

Interested students can prepare a brief biography on one of the scientists involved in the study of gases and for whom the gas laws are named (Amedeo Avogadro, Robert Boyle, Jacque Charles, Joseph Louis Gay-Lussac, John Dalton, Thomas Graham). Have the students focus on the work that led to one of the gas laws, since some of the scientists did many other types of research in their careers. Students can make a PowerPoint presentation based on what they find.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 14.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 14.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

There is one equation that relates the pressure to the volume, temperature, and number of moles when the gas is assumed to be ideal.

- In what situations is the ideal gas law more useful than the other gas laws?
- Under what conditions can a gas be assumed to be ideal?

14.3 Ideal Gases

Key Concept

The ideal gas law is a single equation that relates the pressure of a gas to the volume, number of moles and temperature. The equation makes use of the variable R, called the ideal gas constant. The ideal gas law can be used to solve for the molar mass of an unknown gas from experimental data or to find the density of a gas under nonstandard conditions. Stoichiometry problems in which the chemical reactions do not take place at STP can be solved with the use of the ideal gas law. Real gases are gases whose behavior does not mathematically conform to the ideal gas law. Gases are most ideal at low pressures and high temperatures.

Standards

Lesson Objectives

- Know the ideal gas law, and know which of the different values for the ideal gas constant to use in a given situation.
- Use the ideal gas law to calculate the pressure, volume, temperature, or number of moles of an ideal gas when the other three quantities are known.
- Use the ideal gas to calculate the molar mass or the density of a gas.
- Use the ideal gas law in stoichiometry problems involving gases that are not at STP.
- Explain the conditions under which real gases are most ideal or least ideal.

Lesson Vocabulary

- ideal gas constant: The variable with the symbol R in the ideal gas equation: value is equal to 8.314 J/K•mol, 0.0806 L•atm/K•mol, and 62.36 L•mmHg/K•mol.
- ideal gas law: A single equation which relates the pressure, volume, temperature, and number of moles of an ideal gas.
- real gas: A gas that does not behave according to the assumptions of the kinetic-molecular theory.

Teaching Strategies

Introducing the Lesson

The ideal gas law is an extension of the combined gas law and Avogadro's law which concluded the Gas Laws lesson. Write the combined gas law as $\frac{PV}{T} = k$ on the board. Ask the students how to add in Avogadro's law, $\frac{V}{n} = k$ to it.

In other words, where in the combined gas law (numerator or denominator) should the n variable go? They should recognize that the n belongs in the denominator since number of moles varies directly with volume (or pressure if the container is rigid). That leaves us with $\frac{PV}{NT} = k$. From here, it is a small rearrangement and a renaming of the constant from k to R and you have the usual form of the ideal gas law.

Common Misconceptions

Sometimes students may become confused about when to use the ideal gas law versus when to use the other gas laws. The other gas laws are used when the initial condition of the gas is changed in some way. When they look at a problem, they will be able to recognize this. For example a Boyle's law problem may have an initial and final pressure given along with an initial volume. The problem is to solve for the final volume. Ideal gas law problems simply relate the four variables to one another for a gas under one particular set of conditions. Three variables are given, while the fourth needs to be solved for to do the problem. There is no change in the conditions. Students should practice doing problems that require the use of the ideal gas law when mixed in with other problems.

Having different values for the constant R troubles some students. Stress that the reason for that is solely because of the different units of pressure. When doing sample problems, cancel out the units, remembering that a joule (J) is equal to a kilopascal-liter (kPa•L). Students need to always choose the correct R value by identifying the pressure unit of the problem.

Science Inquiry

The link below takes you to a set of activities on the ideal gas law. Students can click on simulations that will allow them to explore the gas law variables.

http://butane.chem.uiuc.edu/pshapley/Enlist/Labs/GasLaw/GasLaw.html

Differentiated Instruction

For students that struggle with the many variables possible in gas law problems, have them construct a table similar to the one shown below.

TABLE 14.3:

P ₁	P ₂	
\mathbf{V}_1	V_2	
T ₁	T_2	
n ₁	N ₂	

For each problem, they should fill in the table from the information provided. Compare the filled table with the possible gas laws and choose which one fits. For example, if no volume or amounts are given, Gay-Lussac's law is used. The ideal gas law is used when the entire second column (P_2 , etc.) is empty. Keep in mind that an amount may be given in grams, but that would need to be converted to moles before solving the problem.

Enrichment

Deep sea divers are susceptible to a condition called decompression sickness, often referred to as the bends. The dissolving of nitrogen in the diver's blood is the root cause. Interested students can research decompression sickness, its causes, prevention, and treatment. Have them write a report or prepare an oral presentation on their findings.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 14.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 14.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Dalton's law is a gas law that deals with mixtures of gases. In a gas mixture, all of the particles are still relatively far apart, so the different gases in the mixture are assumed to interact with one another only during brief collisions.

- How do you determine the pressure of one particular gas when that gas is part of a mixture of gases?
- In the laboratory, gases are often collected by bubbling the gas into a water-filled container. What adjustment needs to be made to determine the amount of a gas when it is collected in this way?

14.4 Gas Mixtures and Molecular Speeds

Key Concept

The partial pressure of a gas is the pressure exerts independently of other gases in a mixture. Dalton's law states that the total pressure of a gas mixture is equal to the sum of all the partial pressures. The partial pressure of one gas can also be found by multiplying the total pressure by the mole fraction of that gas. Gaseous products are often collected by water displacement. The pressure of the gas is found by subtracting the vapor pressure of water. Effusion and diffusion are two ways that gas particles move in space. Graham's law relates the relative speeds of gas molecules to their molar masses. As molar mass increases, the speed of the gas particles decreases.

Standards

Lesson Objectives

- Use Dalton's law and mole fraction to calculate the partial pressure of a gas in a mixture.
- Calculate the pressure of a gas that has been collected by eater displacement in order to determine the volume of the dry gas.
- Define diffusion and effusion.
- Use Graham's law to calculate the velocity ratios of two gases based on their molar masses.

Lesson Vocabulary

- Dalton's law of partial pressure: A law that states that the total pressure of a mixture of gases is equal to the sum of all of the partial pressures of the component gases.
- diffusion: The tendency of molecules to move from an area of high concentration to an area of low concentration until the concentration is uniform.
- effusion: The process of a confined gas escaping through a tiny hole in its container.
- mole fraction: The ratio of moles of one substance in a mixture to the total number of moles of all substances.
- partial pressure: The contribution that one gas makes to the total pressure when the gas I part of a mixture.

Teaching Strategies

Introducing the Lesson

Draw two identical boxes on the board that represent containers of gas. Draw a small number of gas particles (say helium) as circles inside the box. Label the pressure to be 1 atm. Repeat for the same number of particles of another

gas (say neon) in the second box. Label the pressure as 1 atm. Discuss how the pressures have to be the same because the volumes and number of particles are the same, as is the presumed temperature. Draw a third box that is the same size as the first two. "Transfer" all of the particles from the first two boxes into the third one. Ask the students what they think the pressure will be. Most will realize that the pressure is now equal to 2 atm. This introduces the fundamental concept of Dalton's law –that gases exert pressure independent of any other gases that may be present. This can also lead into teaching partial pressures ($P_{He} = P_{Ne} = 1$ atm) and mole fractions.

Demonstration

Graham's law can be tested by a demonstration of the reaction of ammonia gas with hydrogen chloride gas to form solid ammonium chloride. The general outline of the demonstration is mentioned in the FlexBook® lesson. Additionally, the mathematics behind Graham's law can be explored by measuring the distances travelled by the two gases before the reaction occurs. The link below gives a detailed description of how to conduct the demonstration.

http://chemmovies.unl.edu/chemistry/dochem/DoChem080.html

Common Misconceptions

Because a mole fraction is a ratio of moles, the quantity has no units. Students may think that the units of a mole fraction should be "mol" or that the mole fraction needs to be multiplied by 100 to make it a percent.

Students may mistakenly believe that the relative speed of gas particles depends on the pressure. The derivation of Graham's law arose from setting the kinetic energies of two gases equal to one another. Therefore, it is inherent that the temperatures of the gases must be the same. In other words, the speed of gas molecules will increase if the temperature is increased. However, increasing the pressure of a gas has no effect on the speed of the particles.

Science Inquiry

Dalton's law can be studied using the Gas Properties simulation. The link below takes you to a page from which you can access a document that gives students instructions on how to use the simulation to study Dalton's law. There is also a link on the page to the simulation.

http://phet.colorado.edu/en/contributions/view/3232

Likewise, Graham's law of effusion can also be studied using the same simulation. The link below gives you access to both a document of student instructions and the link to the Gas Properties simulation.

http://phet.colorado.edu/en/contributions/view/3237

Differentiated Instruction

Graham's law can be challenging to solve for students of lower mathematical ability. Make clear what is meant by the term "velocity ratio" and that it is a single number (with no units) that is equal to v1/v2. Give students ample practice with calculating both possible velocity ratios with pairs of gases. For example, suppose that the gases to compare are nitrogen and chlorine. Have the students calculate the velocity ratio of nitrogen to chlorine and chlorine how to nitrogen. Point out that the two ratios are simply the reciprocal of one another.

Enrichment

Graham's law of effusion played an important historical role during the Manhattan Project, in which the United States built the first atomic bomb. Naturally occurring uranium consists mostly of uranium-238, and isotope which cannot

be used for the nuclear fission process. Have interested students research how Graham's law allowed scientists to enrich naturally occurring uranium with the fissionable isotope of uranium-235.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 14.4 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 14.4 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Many chemical reactions take place between reactants that have been dissolved in water. A thorough understanding of the properties of water is essential to an analysis of reactions that occur in aqueous solution.

- How does the shape of a water molecule affect its polarity and its properties?
- How do the properties of water change when it freezes to ice?





Water

Chapter Outline

- 15.1 **PROPERTIES OF WATER**
- 15.2 **AQUEOUS SOLUTIONS**
- 15.3 **COLLOIDS AND SUSPENSIONS**



NASA/Goddard Space Flight Center/Reto Stöckli. www.flickr.com/photos/gsfc/4426654941/. CC BY 2.0.

Chapter Overview

Water is a very simple substance, but one which is essential to the existence of life as we know it on Earth. Its polarity and hydrogen bonding ability cause water to be able to dissolve many substances and give it unusual properties such as surface tension, low vapor pressure, and a density that is higher than its solid state (ice). Aqueous solutions are encountered frequently in chemistry. Water dissolves most ionic compounds and polar molecular compounds, but cannot dissolve nonpolar molecular compounds. Colloids and suspensions are heterogeneous mixtures rather than true solutions.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 15.1	ŝ	Pacing	Lessons
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Lesson	Class Period(s) (60 min)
15.1 Properties of Water	1
15.2 Aqueous Solutions	2
15.3 Colloids and Suspensions	1 - 2

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

TABLE 15.2:	Chemistry	Concepts
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FlexBook® Lesson	Chemistry Concepts
15.1 Properties of Water	15.1 – 15.3
15.2 Aqueous Solutions	15.4 –15.9
15.3 Colloids and Suspensions	15.10 – 15.11

15.1 Properties of Water

Key Concept

The structure of the water molecule is responsible for many of water's unique properties. The water molecule has a bent shape, making it a highly polar molecule. Its molecules make strong hydrogen bonds, resulting in high surface tension, low vapor pressure, and an unusually high melting and boiling point. When water freezes, the hydrogen bonds create an open framework that causes ice to be less dense than liquid water.

Standards

Lesson Objectives

- Describe the structure and polarity of a water molecule.
- Describe the hydrogen bonding that occurs in water and ice.
- Discuss the unique properties of water and ice.

Teaching Strategies

Introducing the Lesson

The day before the lesson, fill a glass bottle with water and place it in the freezer. As the water freezes, it will expand and break the bottle. Show this to the students when you are discussing the unusual property of water that its solid state is less dense than its liquid state.

Students are largely familiar with the structure and polarity of the water molecule from the Covalent Bonding and States of Matter chapters. These concepts can be quickly reviewed or you can give students warm-up questions to test how much they remember.

Common Misconceptions

Types of crystalline solids were covered in the States of Matter chapter and students may not realize that ice qualifies as a molecular solid with individual water molecules at each lattice point. Emphasize that each oxygen atom in ice is surrounded by four hydrogen atoms in a tetrahedral geometry with two covalent bonds and two hydrogen bonds. If possible, show a molecular model of ice where students can see the open framework that leads to the decrease in density compared to liquid water.

Science Inquiry

The surface tension and cohesion of water can be studied with a quick activity in which students test how many drops of water they can place onto the surface of a penny. Comparisons can be made with other liquids that cannot hydrogen bond. The links below are two options for this introductory lab.

http://www.stevespanglerscience.com/lab/experiments/penny-drops http://sciencespot.net/Media/pennylab.pdf

Differentiated Instruction

Students can create a concept map about water. The map should contain information about the physical properties of the water molecule as well as macroscopic properties of water like surface tension and low vapor pressure.

Enrichment

Cohesion and adhesion are two additional properties of water related to its strong hydrogen bonding. Students can research these two properties and prepare demonstrations for the class illustrating how each works. Have them write up a brief explanation to accompany the demonstrations.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 15.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 15.1 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Water is often referred to as the universal solvent because it is capable of dissolving so many different substances.

- How does the structure of water aid in its ability to dissolve solids?
- What is the nature of solutions made from ionic and molecular compounds?

15.2 Aqueous Solutions

Key Concept

Aqueous solutions are homogeneous mixtures consisting of a solute dissolved in water. Water dissolves many ionic compounds due to an electrostatic attraction between the ions in the crystal and the polar water molecules. Water dissolves polar molecular compounds, but is generally incapable of dissolving molecular compounds that are nonpolar. Electrolytes are compounds that conduct an electric current when in solution. Ionic compounds are electrolytes because the ionic crystal dissociates into ions when it dissolves. Soluble molecular compounds may be either strong or weak electrolytes depending on the extent of ionization that occurs.

Standards

Lesson Objectives

- Define a solution and describe the parts of a solution.
- Describe how an aqueous solution is formed from both ionic compounds and molecular compounds.
- Explain the reasons why some compounds are insoluble in water.
- Write equations for the dissociation of ionic compounds in water.
- Distinguish between strong electrolytes, weak electrolytes, and nonelectrolytes.

Lesson Vocabulary

- aqueous solution: Water that contains one or more dissolved substances.
- dissociation: The separation of ions that occurs when a solid ionic compound dissolves.
- electrolyte: A compound that conducts an electric current when it is in an aqueous solution or melted.
- hydration: The process of solute particles being surrounded by water molecules arranged in a specific manner.
- immiscible: A quality observed by liquids that are insoluble in one another.
- miscible: A quality observed by liquids that dissolve in one another in all proportions.
- nonelectrolyte: A compound that does not conduct an electric current in either aqueous solution or in the molten state.
- solute: The dissolved particles in a solution.
- solution: A homogeneous mixture consisting of a solute dissolved into a solvent.
- solvation: A process in which individual ions are surrounded by solvent particles after dissolving.
- solvent: The dissolving medium of a solution.
- strong electrolyte: A solution in which a large fraction of the dissolved solute exists as ions.
- weak electrolyte: A solution in which only a small fraction of the dissolved solute exists as ions.

Teaching Strategies

Introducing the Lesson

Ask the students to name examples of solutions. Most that they will list will likely be aqueous solutions with solid solutes. Ask if air is a solution. Ask if an alloy such as brass is a solution. Establish that these are also solutions. Try to lead students to the definition that a solution is any homogeneous mixture, regardless of the phases involved. Mix water with an immiscible liquid such as oil and show them the two layers that are formed. Ask if this is a solution (no). Ask why not. Define the parts of a solution (solute, solvent) and move on to describing how the dissolving process works.

Common Misconceptions

Students sometimes fail to realize that dissolved particles are small enough that they will pass unimpeded through a filter. However, if a solution is allowed to sit until the water evaporates (or is heated), only the water disappears and the solute is left behind. Remind students that the dissolving process is a physical change and not a chemical one. Solutions can also contain multiple solutes. A soda, for instance, has dissolved carbon dioxide as well as sugar, flavorings, and other ingredients.

Students may think that oil and water do not mix because they have different densities. Stress that it is the difference in polarity that cause them to be immiscible. Oil then floats on top of water because of its lower density. Another nonpolar liquid may be denser than water and thus sink beneath it.

Students frequently think that insoluble ionic compounds are either nonelectrolytes or weak electrolytes because of their insolubility. However, all ionic compounds are strong electrolytes. Even if a particular ionic compound only dissolves to a very small extent, whatever fraction that does dissolve also dissociates. The complete dissociation of the dissolved solid means that it is a strong electrolyte. Only polar molecular compounds can be weak electrolytes.

Science Inquiry

The process of dissolving can be studied in the inquiry activity at the link below. Students will see how some substances will dissolve in water, but will not dissolve in oil. A student activity page can be downloaded from the main page.

http://www.inquiryinaction.org/classroomactivities/activity.php?id=15

Differentiated Instruction

This lesson contains a large number of vocabulary words. Have less proficient readers or English language learners prepare a list of the vocabulary and to write their definitions. Give a quiz to see if students have learned the words.

Enrichment

Water pollution is an ongoing threat to human health and the survival of natural ecosystems. Have interested students prepare a poster or PowerPoint presentation on water pollution. They can include historical aspects, ways to prevent pollution, methods of treating polluted water, and the outlook for the future. You can also assign different subtopics of water pollution to different groups of students in order to cover the topic more thoroughly.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 15.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 15.2 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Water is capable of interacting with some substances without forming a true solution.

- What are suspensions and colloids?
- How can a colloid be distinguished from a solution?

15.3 Colloids and Suspensions

Key Concept

Suspensions and colloids are heterogeneous mixtures consisting of a dispersed phase and a dispersing medium. The dispersed particles in a suspension settle out due to gravity. The dispersed particles in a colloid are smaller and do not settle, but are larger than the individual ions, atoms, or molecules that constitute the solute of a true solution. Colloids display the Tyndall effect, the scattering of light from the dispersed particles. Types of colloids include gels, aerosols, foams, and emulsions.

Standards

Lesson Objectives

- Describe the properties of a suspension.
- Describe the properties of a colloid and distinguish from a solution or a suspension.

Lesson Vocabulary

- colloid: A heterogeneous mixture whose particle size is intermediate between those of a solution and a suspension.
- emulsion: A colloidal dispersion of a liquid in either a liquid or a solid.
- suspension: A heterogeneous mixture in which some of the particles settle out of the mixture upon standing.
- Tyndall effect: The scattering of visible light by colloidal particles.

Teaching Strategies

Introducing the Lesson

Remind students of the definition of a solution and that the dissolved particles in a solution are so small that they will pass through a filter. Use a laser pointer or other intense light source to show how light that is passed through a solution is not visible (no Tyndall effect) because the wavelength of the light is larger than the particles. Either do the Science Inquiry activity below or show the students ample examples of suspensions and colloids. Demonstrate the Tyndall effect with water in which a few drops of milk have been added.

15.3. Colloids and Suspensions

Common Misconceptions

In many cases, it is difficult to tell the difference between a colloid and a true solution since a colloid appears homogeneous to the naked eye. Stress that the size of the dispersed particles in a colloid is considerably larger, usually consisting of aggregated molecules (molecules that are clumped together into a larger unit). Discuss or show the Brownian motion of a colloid and how this random motion of colloidal particles prevents them from settling to the bottom of the container as with a suspension.

Science Inquiry

The activity below allows students to prepare examples of solutions, colloids, and suspensions and to study their differences. It makes for a good introduction to the lesson.

http://woodstown.org/cms/lib4/NJ01001783/Centricity/Domain/8/Texts/ACS/resources/ac/ch6/act3.pdf

Differentiated Instruction

Students can bring from home some of the colloid examples found in Table 15.3. Experiment with them to see how they are different from true solutions. Explore the concept of an emulsifying agent by making mayonnaise from cooking oil, vinegar, and egg yolk. First have the students look up a recipe for mayonnaise on the internet.

Enrichment

Paint is an interesting example of a colloid. Most students know that paint must first be stirred before using in order to insure that it is uniform. Ask students to investigate different types of paints and the substances that are present in them.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 15.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 15.3 in CK-12 Chemistry –Intermediate FlexBook® resource.

Points to Consider

Solutions have a wide array of properties and uses. While some solutes are very soluble in certain solvents, some other solutes are not.

- How is solubility measured?
- What factors affect solubility?
Chapter **16**

Solutions

Chapter Outline

- 16.1 ELECTROCHEMICAL CELLS
- 16.2 CELL POTENTIALS
- 16.3 ELECTROLYSIS
- 16.4 NET IONIC EQUATIONS



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Chapter Overview

Solutions are an important component of many chemical reactions. The rate of dissolving is affected by temperature, solute particle size, and agitation of the mixture. Solutions may be saturated, unsaturated, or supersaturated. Solubility refers to how much of a solute will dissolve into a solvent at a given temperature. Concentration is a numerical measure of the amount of dissolved solute and can be expressed as molarity, molality, or as a percent. Colligative properties are physical properties of a solvent that are affected by dissolved solute and include vapor pressure lowering, freezing point depression, and boiling point elevation. Chemical reactions that occur between ions in solution are best represented by net ionic equations.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

Lesson	Class Period(s) (60 min)
16.1 Solubility	2
16.2 Solution Concentration	2
16.3 Colligative Properties	2
16.4 Net Ionic Equations	1 - 2

TABLE 16.1: Pacing Lessons

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 16	5.2:	Chemistry	Concepts
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Flexbook Lesson	Chemistry Concepts
16.1 Solubility	16.1 – 16.6
16.2 Solution Concentration	16.7 –16.11
16.3 Colligative Properties	16.12 - 16.16
16.4 Net Ionic Equations	16.17 – 16.19

16.1 Electrochemical Cells

Key Concept

The solutes and solvents of solutions may be solids, liquids, or gases. The rate of dissolving depends on temperature, surface area, and agitation. Solubility is the amount of a substance capable of dissolving and varies for different solutes. Solutions can be saturated, unsaturated, or supersaturated. An increase in temperature generally results in an increase in the solubility of solids, but a decrease in the solubility of gases. The solubility of gases is directly dependent on gas pressure, as described by Henry's law.

Standards

Lesson Objectives

- List examples of solutions made from different solute-solvent combinations.
- List and explain three factors that affect the rate of dissolving of a solid solute in a liquid solvent.
- Explain solution equilibrium and distinguish between saturated, unsaturated, and supersaturated solutions.
- Explain the effects of temperature on the solubility of solids and gases. Use a solubility curve to determine the solubilities of substances at various temperatures.
- Use Henry's law and explain the effect of pressure on the solubility of gases.

Lesson Vocabulary

- Henry's law: A gas law that states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.
- recrystallization: The process of dissolved solute returning to the solid state.
- saturated solution: A solution that contains the maximum amount of solute that is capable of being dissolved.
- solubility: The amount of a substance that is required to form a saturated solution in a given amount of solvent at a specified temperature.
- solution equilibrium: The physical state described by the opposing processes of dissolution and recrystallization occurring at the same rate.
- supersaturated solution: A solution that contains more than the maximum amount of solute that is capable of being dissolved at a given temperature.
- unsaturated solution: A solution that contains less than the maximum amount of solute that is capable of being dissolved.

Teaching Strategies

Introducing the Lesson

Students were introduced to solutions in the previous chapter, Water. Review the definition of a solution and show several examples of solutions, especially those that consist of phase combinations other than the familiar solid dissolved in a liquid. Tell the students that you would like them to prepare a sugar-water solution and to make it as quickly as possible. What should they do to make the dissolving process as rapid as possible? They will most likely be able to come up with the three discussed in the flexbook (particle size, stirring, temperature). Move on to a discussion of the types of solution and either show examples or do the activity below (in Science Inquiry) on degrees of saturation.

Demonstration

Students enjoy seeing the rapid recrystallization that occurs when a seed crystal is added to a supersaturated solution. The link below gives detailed instructions for preparing a supersaturated solution of sodium acetate. Alternatively, do the lab activity described below in the Science Inquiry section.

https://www.flinnsci.com/media/620509/91215.pdf

Common Misconceptions

Students often do not get a clear picture of supersaturated solutions unless they see one or, better yet, make their own. The solubility of sodium chloride in water is about 36 g per 100 g H₂O at 20°C. If one adds 40 g of NaCl to 100 g of water at 20°C, it is not a supersaturated solution. The extra 4 g would not dissolve and the solution is only saturated. A supersaturated solution can only be prepared in the lab by "tricking" the solution into having more dissolved solute than should be possible by heating and cooling.

Science Inquiry

The difference between saturated, unsaturated, and supersaturated solutions is an important distinction and one which the students benefit from seeing directly. The lab below directs students to prepare solutions of sodium acetate with different degrees of saturation and to answer questions based on their observations.

http://hrsbstaff.ednet.ns.ca/dawsonrj/12%20Chem/Lab%20Activities/Degree%20of%20Saturation.htm

Another concept which makes for a good hands-on activity is solubility curves. The link below is for a lab in which students generate their own solubility curve for potassium nitrate and answer questions.

http://www.google.com/url?sa=t&rct=j&q=&esrc=s&source=web&cd=1&ved=0CB0QFjAA&url=http%3A%2F%2Fum anitoba.ca%2Foutreach%2Fcrystal%2Fresources%2520for%2520teachers%2FSolubilityCurveLab.doc&ei=dtPsU5n 8KNOcygTx3YGYAg

Differentiated Instruction

The concept of equilibrium is an important one in chemistry and may be difficult for students to understand at first. You can separate the class into two groups on either side of the room, one representing "dissolved particles" and one representing "solid solute". Count the number in each group –don't make them exactly equal. Discuss that they represent a saturated solution. Allow students to move back and for from one group to the other, but make sure to

maintain the same number in each group at all times. This illustrates the dynamic nature of equilibrium and that the ratio of product to reactant must remain the same.

Enrichment

Have students identify as many solutions as possible in their home. Have them make a list, identifying the solvents and solutes for as many as possible.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 16.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 16.1 in CK-12 Chemistry –Intermediate Flexbook® resource.

Points to Consider

Solutions are prepared by dissolving a certain amount of solute in a certain amount of solvent. The concentration of a solution is a quantitative measure of how much solute has been dissolved.

- What are the ways in which the concentration of a solution can be calculated?
- What happens to the concentration of a solution when additional solvent is added?

16.2 Cell Potentials

Key Concept

The concentration of a solution is a numerical measure of how much solute has been dissolved into a solvent. The simplest concentration units are mass percent and volume percent. Molarity is the moles of solute per liter of solution and is the most widely used unit of concentration for chemical reactions. Solutions of a specific molarity are prepared with the use of a volumetric flask. Dilution is the process of lowering the concentration of a solution by adding by more solvent. Molality, the moles of solute per kilogram of solvent, is a concentration unit used in situations in which the temperature of the solution is altered.

Standards

Lesson Objectives

- Use the terms concentrated and dilute to describe the relative concentration of a solution.
- Calculate the concentration of a solution as either a mass percent or a volume percent.
- Calculate the molarity of a solution. Use molarity to calculate the mass of solute needed to prepare a particular solution.
- Calculate the molarity of a diluted solution. Use the dilution equation to calculate the volume of a concentrated stock solution required for a particular dilution.
- Calculate the molality of a solution and distinguish molality from molarity.

Lesson Vocabulary

- concentrated: A solution that has a relatively large amount of dissolved solute.
- concentration: A measure of the amount of solute that has been dissolved in a given amount of solvent or solution.
- dilute: A solution that has a relatively small amount of dissolved solute.
- molality: The number of moles of solute dissolved in one kilogram of solvent.
- molarity: The number of moles of solute dissolved in one liter of solution.

Teaching Strategies

Introducing the Lesson

This lesson is the mathematics of solutions. Start by making sure that students understand the terms concentrated and dilute, but realize that those are merely general terms. A solution that may be "concentrated" for one particular use might be too "dilute" for another. Review grams to moles and moles to grams calculations as you work through sample problems on calculating molarity and determining the mass of solute needed to make a specific solution.

Common Misconceptions

Students will not necessarily realize that when a solid solute is added to a given volume of water, the volume will increase due to the space taken up by the dissolved solute particles. This can easily be demonstrated with a volumetric flask used incorrectly. Fill the volumetric to the line with water, then add a large portion of solid. Invert several times to dissolve the solid and note how the solution level is now above the line. Because molarity is defined as moles per liter of solution, the dissolving of the solid has to occur before filling to the line. Demonstrate the proper use of a volumetric flask or do the activity described below under Science Inquiry.

It is easy for students to confuse molarity and molality, so it is best not teach the two concepts at the same time. Stress that molarity is by far the most commonly unit of concentration that chemists use in the laboratory. Molality is essentially only used for colligative properties, which are covered in the following lesson.

Science Inquiry

The activity below directs students to make a copper(II) sulfate solution of a specific molarity and then to dilute the solution. Teachers can assign different molarities to different lab groups and compare them. Students also answer questions that require them to be able to do molarity and dilution calculations.

http://www.google.com/url?sa=t&rct=j&q=&esrc=s&source=web&cd=1&sqi=2&ved=0CB0QFjAA&url=http%3A%2 F%2Fteachers.henrico.k12.va.us%2Fdeeprun%2Fnolen_f%2FMolarity%2520and%2520Dilution%2520Lab.doc&ei=09 rsU5ulH8qpyAT5h4KoAw

Differentiated Instruction

Have students who are less skilled at math work with more advanced students to create detailed flowcharts for the different types of problems in this lesson. Have the students use the flowcharts to help with practice problems.

Enrichment

Advanced math students can do problems in which the molarity of a solution can be converted to molality or viceversa. Since one is based on mass of the solvent and the other on volume of solution, the density of the solution must be known. An example is shown below.

Example: The molarity of an aqueous solution of sodium hydrogen carbonate (NaHCO₃) is 5.20 M and the density of the solution is 1.19 g/mL. Calculate the molality.

Answer: Since the molarity is known, begin by assuming exactly 1 L (1000 mL) of solution. The density can be used to determine the mass of the solution.

1000 mL × $\frac{1.19 g}{1 mL}$ = 1190 g solution

Because of the 1-liter assumption, the solution must contain 5.20 mol of NaHCO₃. Convert this to mass.

 $5.20 \ mol \times \frac{84.01 \ g}{1 \ mol} = 437 \ g \ NaHCO_3 \ (solute)$

The mass of the solvent is equal to the mass of the solution minus the mass of the solute.

1190 $g - 437 g = 753 g H_2O$ (solvent)

Now the molality can be calculated.

 $m = \frac{mol \ solute}{kg \ solvent} = \frac{5.20 \ mol}{0.753 \ kg} = 6.91 \ m$

Because the density of the solution is greater than 1 g/mL, the molality is larger than the molarity. If the density of the solution is less than 1 g/mL, the molarity is larger.

To convert from molality to molarity, the easiest starting assumption is exactly 1 kg of solvent. Convert moles solute to grams. Add the mass of the solute to 1 kg to get the mass of the solution. Use the density of the solution to find its volume. Calculate the molarity.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 16.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 16.2 in CK-12 Chemistry –Intermediate Flexbook® resource.

Points to Consider

Some of the physical properties of a solvent are altered by the process of dissolving a solute into that solvent. The vapor pressure, freezing point, and boiling point of a solution are different than that of a pure solvent.

- How is the vapor pressure of a solvent affected when a solute is dissolved?
- How can the freezing and boiling point of a solution be calculated?

16.3 Electrolysis

Key Concept

Colligative properties of a solution depend only on the number of dissolved solute particles and not on their chemical nature. The addition of a solute lowers the vapor pressure of a solvent. The freezing point of a solution is lower than that of the pure solvent while the boiling point of a solution is higher than that of the pure solvent. Changes in freezing and boiling points are dependent upon the molality of the solution, the molal freezing-point or boiling-point constant, and the number of particles produced when the solute dissolves. The measurement of a freezing point or boiling point or boiling point change can be used to calculate the molar mass of an unknown solute.

Standards

Lesson Objectives

- Define a colligative property and identify three colligative properties of solutions.
- Differentiate between the effects that an electrolyte has on the colligative properties of a solution compared to a nonelectrolyte.
- Calculate the freezing and boiling points of a solution of known molality.
- Use freezing or boiling point information to calculate the molar mass of an unknown solute.

Lesson Vocabulary

- boiling point elevation: The difference in temperature between the boiling point of the pure solvent and that of the solution.
- colligative property: A property of a solution that depends only upon the number of solute particles dissolved in the solution and not on their identity.
- freezing point depression: The difference in temperature between the freezing point of the pure solvent and that of the solution.
- molal boiling-point elevation constant: A constant that is equal to the change in the boiling point for a 1-molal solution of a nonvolatile molecular solute.
- molal freezing-point depression constant: A constant that is equal to the change in the freezing point for a 1-molal solution of a nonvolatile molecular solute.

Teaching Strategies

Introducing the Lesson

After introducing the term colligative property, discuss how they are properties of the solvent and not of the solute. Both freezing point depression and boiling point elevation depend on the lowering of the vapor pressure. Make sure to show and discuss the graphs shown in Figures 16.11 and 16.13. Emphasize that the degree to which the freezing and boiling points are affected is dependent only on the number of solute particles. Show with equations why a solution of NaCl would have twice the effect of a solution of $C_6H_{12}O_6$, since the latter is a nonelectrolyte. Show that MgCl₂ would have three times the effect. Discuss how cities and towns where it snows take advantage of freezing point depression by salting the roads.

Demonstration

A demonstration of freezing point depression is easy. Prepare a large beaker of ice water and allow it to cool to 0°C. Pour in a generous amount of salt and continually stir, watching the temperature drop below the normal freezing point of water.

Common Misconceptions

When students learn about boiling point elevation, they may mistakenly assume that the reason cooks add salt to boiling water (for pasta) is to raise the boiling point and decrease the cooking time. Discuss how the effect on the boiling point is extremely minimal if a couple teaspoons of salt are added to a pot of water. It is done for taste and not for any effect on the boiling temperature.

Students may have difficulty distinguishing electrolytes from nonelectrolytes. Remind them that the easiest way to do so is to remember that electrolytes are ionic compounds and so must be a compound consisting of a metal and a nonmetal. If the solute is a compound consisting only of nonmetals, it is likely to be a nonelectrolyte. The main exceptions are salts of the ammonium ion, such as ammonium chloride (NH4Cl), which is an electrolyte.

Science Inquiry

The lab described below allows students to determine the molar mass of an unknown solute from freezing point depression data. The lab skills needed to complete this lab are somewhat advanced, so judge your class and decide if this is a lab that they can do. The necessary calculations are described in the lesson.

http://www.ulm.edu/chemistry/courses/manuals/chem1010/experiment_01.pdf

A much more relaxed activity in which students can see colligative properties in action is to make homemade ice cream! The link below contains an ice cream recipe along with pertinent questions on why the traditional ice cream making process is dependent on the colligative property of freezing point depression. http://www.google.com/url?sa =t&rct=j&q=&esrc=s&source=web&cd=3&sqi=2&ved=0CC0QFjAC&url=http%3A%2F%2Ffc.dekalb.k12.ga.us% 2F~john_r_snyder%2FFOV1-00261F3B%2FFOV1-00263AB2%2FFOV1-0026932F%2FLab%2520%2520Colligat ive%2520Properties%2520%2526%2520%2520Cream.doc

Differentiated Instruction

The tutorials at the links below may help students with colligative properties. Have them watch the tutorials and discuss what they've learned in small groups.

http://rmpbs.pbslearningmedia.org/resource/398e7a47-9d70-48cd-856a-e9b8df987ca1/chemistry-1003-molarity-and-colligative-properties/

http://rmpbs.pbslearningmedia.org/resource/90899a5d-0cef-4e60-bbd2-fbfcf6825638/teachable-moment-freezing-po int-depression/

Enrichment

Another colligative property of solutions is osmotic pressure. Students can research osmotic pressure and how to calculate its value. They can produce a brief description of osmotic pressure, in what situations it is a concern, and example problems showing how it is calculated.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 16.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 16.3 in CK-12 Chemistry –Intermediate Flexbook® resource.

Points to Consider

Many chemical reactions take place in aqueous solutions, and many of those reactions involve electrolytes that dissociate in solution.

- Which types of reactions most often involve ionic substances?
- How can reaction between ionic substances be shown most succinctly?

16.4 Net Ionic Equations

Key Concept

Chemical reactions which take place in aqueous solution are more accurately described by equations in which dissolved ionic substances are represented as ions. Most double-replacement and single-replacement reactions involve ions in solution. A net ionic equation is on which shows only the ions which actively participate in the reaction. Ions which remain unchanged are called spectator ions and are omitted from the equation. The solubility rules can be used to predict the formation of a precipitate from a reaction mixture of ions.

Standards

Lesson Objectives

- Write net ionic equations for double-replacement reactions that produce precipitates, gases, or molecular compounds.
- Write net ionic equations for single-replacement reactions.
- Use the solubility rules to predict precipitate formation.

Lesson Vocabulary

- ionic equation: An equation in which dissolved ionic compounds are shown as free ions.
- molecular equation: An equation in which the formulas of the compounds are written as though all substances exist as molecules.
- net ionic equation: A chemical equation that shows only those elements, compounds, and ions that are directly involved in the chemical reaction.
- spectator ion: An ion that does not take part in the chemical reaction and is found in solution both before and after the reaction.

Teaching Strategies

Introducing the Lesson

Start the lesson with a quick review of types of reactions from the Chemical Reactions chapter. When you have examples of each type on the board, ask the students which type or types tend to occur in aqueous solution. Focus on double-replacement reactions first, since all common double-replacement reactions have all substances in solution. Go step-by-step through the molecular equation, full ionic equation, and net ionic equations for each example that

you show. While discussing a certain equation, demonstrate the reaction if possible (see below), showing how double-replacement reactions produce a precipitate, a gas, or a molecular compound (often water). When students have mastered double-replacement reactions, move on to single-replacement.

Demonstrations

There are many precipitate-producing reactions to show, but the one below is interesting because the lead(II) iodide precipitate is bright yellow.

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow 2KNO_{3(aq)} + PbI_{2(s)}$$

When metal carbonates react with an acid, the double replacement reaction produces unstable carbonic acid, which immediately decomposes to form carbon dioxide and water. This is an example of a double-replacement reaction that produces both a gas and a molecular compound.

$$NaCO_{3(aq)} + 2HCl \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

A copper wire suspended in a solution of silver nitrate is a good single-replacement reaction to show. The wire "grows" strands of silver on it while the solution gradually turns blue from the copper(II) ion. The reaction is slow and you will want to leave it overnight.

$$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$$

All of the equations shown above are molecular equations.

Common Misconceptions

Students seem to have more difficulty with the net ionic equations for single-replacement reactions. Because they have learned the charges of common ions, some students will want to "change" the element in the reaction to a charged ion. For example, consider the molecular equation for the copper reacting with silver nitrate.

$$Cu_{(s)} + 2AgNO_{3(aq)} \rightarrow Cu(NO_3)_{2(aq)} + 2Ag_{(s)}$$

Students may write the following for the ionic equation.

$$Cu_{(s)}^{2+} + 2Ag_{(aq)}^{+} + 2NO_{3(aq)}^{-} \rightarrow Cu_{(aq)}^{2+} + 2NO_{3}^{-}(aq) + 2Ag_{(s)}^{+}$$

This is clearly incorrect because the solid elements should not have been written as ions. Stress that ions are only present in aqueous solutions and that solids, liquids, and gases should not be changed when writing the ionic equations.

Science Inquiry

The lab activities at the links below require students to perform simple reactions by mixing pairs of aqueous solutions. In cases where a reaction occurs, they must write molecular equations and net ionic equations. It is therefore necessary to have an understanding of the material from this lesson before doing this lab. It would be relatively easy to shorten the time required by cutting down on the number of reactions.

http://www.hschem.org/Honors/CH04/Labs/solruleslab.pdf http://mhchem.org/221/pdfLabs221/NetIonicReactions.pdf

Differentiated Instruction

Some students may encounter difficulty turning molecular equations into ionic equations. Reteach how to write dissociation equations for ionic compounds and give the students practice. Discuss how that is exactly what they

16.4. Net Ionic Equations

need to do to every substance in a molecular equation that is in solution (has the (aq) following the formula). Have them work through a few examples in which there are no coefficients to deal with, then work your way up to harder examples.

Enrichment

Cave formations such as stalactites and stalagmites form as a result of some interesting precipitation reactions. Ask students to investigate the chemical reactions involved and produce a PowerPoint presentation or poster on their findings.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 16.4 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 16.4 in CK-12 Chemistry –Intermediate Flexbook® resource.

Points to Consider

Heat and energy are important concepts in chemistry, as most chemical reactions are accompanied by a transfer of energy.

- What is heat, and how does a transfer of heat occur?
- How do different substances respond to an input or loss of heat?

Chapter **17**

Thermochemistry

Chapter Outline

17.1	HEAT FLOW
17.2	THERMOCHEMICAL EQUATIONS
17.3	HEAT AND CHANGES OF STATE
17.4	HESS'S LAW

Chapter Overview

Heat is a type of energy that is transferred from an object at higher temperature to an object at lower temperature. Chemical reactions involve a transfer of heat energy and the study of those heat changes is called thermochemistry. Every substance has a specific heat, the amount of heat required to change the temperature of a given mass of that substance. The heat change for a reaction at constant pressure is called enthalpy and enthalpy changes are measured by calorimetry experiments. Heat changes also occur for substances undergoing a change of state and for substances that dissolve in water. Hess's law allows enthalpy changes to be calculated for any reaction, including those for which the enthalpy change cannot be measured directly.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 17.1: Pacing Lessons

Lesson	Class Period(s) (60 min)
17.1 Heat Flow	2
17.2 Thermochemical Equations	2
17.3 Heat and Changes of State	2
17.4 Hess's Law	2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

Flexbook Lesson	Chemistry Concepts
17.1 Heat Flow	17.1 –17.5
17.2 Thermochemical Equations	17.6 –17.9
17.3 Heat and Changes of State	17.10 – 17.13
17.4 Hess's Law	17.14 – 17.17

TABLE 17.2: Chemistry Concepts

17.1 Heat Flow

Key Concept

Chemical potential energy is the energy that is stored within the chemical bonds of a substance. Heat is energy that is transferred between objects due to a difference in temperature of the objects. An exothermic reaction is one which releases heat to the surroundings, while an endothermic reaction is one which absorbs heat from the surroundings. The specific heat of a substance is a measure of how much heat must be supplied to a given amount of that substance in order to raise its temperature. Water has a very high specific heat, while metals generally have very low specific heats.

Standards

Lesson Objectives

- Describe how chemical potential energy relates to heat and work.
- Describe the law of conservation of energy and how heat flows between system and surroundings during both endothermic and exothermic processes.
- Use the specific heat equation to perform calculations that relate mass, specific heat, change in temperature, and the amount of heat absorbed or released.

Lesson Vocabulary

- calorie: The quantity of heat required to raise the temperature of 1 gram of water by 1°C.
- chemical potential energy: The energy stored in the chemical bonds of a substance.
- endothermic: The condition of a process where heat is absorbed by the system from the surroundings.
- exothermic: The condition of a process where heat is released by the system into the surroundings.
- heat: Energy that is transferred from one object or substance to another because of a difference in temperature between them.
- heat capacity: The amount of heat required to raise the temperature of an object by 1°C.
- law of conservation of energy: A law that states that in any physical or chemical process, energy is neither created nor destroyed.
- specific heat: The amount of energy required to raise the temperature of 1 gram of a substance by 1°C.
- surroundings: Everything in the universe that is not part of the system.
- system: The specific portion of matter in a given space that is being studied during an experiment or an observation.
- thermochemistry: The study of energy changes that occur during chemical reactions and changes of state.

Teaching Strategies

Introducing the Lesson

Students are most likely familiar with the terms exothermic and endothermic from earlier science classes. Begin by asking the class for definitions of the two terms. Ask how a chemist can tell if a certain chemical reaction is exothermic or endothermic. This leads to a discussion of the difference between the system and the surroundings. A simple and dramatic demonstration that can be done is to show the highly exothermic process of diluting sulfuric acid. Place about 200 mL of water in a 600-mL beaker. Place a thermometer in the water and check the temperature. Carefully add about 20 mL of concentrated sulfuric acid to the beaker. The temperature will rise rapidly and steam will be released. The water will likely boil when it first contacts the acid. Emphasize to the students that an exothermic reaction is one in which heat flows from the system to the surroundings and that the temperature of the surroundings increases.

Endothermic reactions are not quite as common, but a commercial cold pack can be shown to the students. When activated, a chemical reaction causes the temperature of the cold pack to decrease by absorbing heat energy from the surroundings. The cold pack can then be applied to an injured body part in order to diminish swelling.

Common Misconceptions

Students may easily confuse the concepts of heat capacity and specific heat. Specific heat is a much more useful quantity because it allows for direct comparisons between different substances. Make sure to explain that the specific heat of a substance does not depend on how much of that substance is present. The specific heat of a tiny copper pellet is the same as a large block of copper. Heat capacity, on the other hand, is dependent on mass and so is not as useful for comparisons.

Make sure that students understand that in many laboratory chemical reactions, it is the surroundings that are being observed. Frequently this involves reactions that take place in aqueous solution, where the water is part of the surroundings. Therefore, an exothermic reaction which releases heat to the surroundings causes the water temperature to rise. Conversely, an endothermic reaction absorbs heat from the surroundings, causing the water temperature to decrease.

Science Inquiry

The specific heat of a metal can easily be determined by experiment. The process involves heating a sample of the metal in a boiling water bath, then transferring the heated metal to a given volume of room temperature water in a Styrofoam cup calorimeter. Measurement of the final temperature of the metal/water mixture allows the student to calculate the specific heat of the metal. The links below lead to two example handouts that can be printed and given to students in order to do this lab exercise.

http://cmhsscience.pbworks.com/f/spec+heat+metals+%2703+CP1.pdf

http://www.kwanga.net/chemnotes/specific-heat-lab.pdf

Differentiated Instruction

The equation that relates heat to specific heat, mass, and temperature change can be challenging for students with limited math ability. Have students work out separate equations that allow them to solve for each of the four variables $(q, c_p, m, \Delta T)$. Provide several examples of problems that students can practice with and which require them to solve for each of those variables. Students can make the problems clearer by listing out all of the known variables that are given in the problem and then choosing the correct equation that allows them to solve for the one unknown variable.

Enrichment

Specific heats can be converted to molar heat capacities (C_p) by multiplying by the molar mass of the substance. As an example, the molar heat capacity of water is given by: $4.18 J/g \cdot C \times \frac{18.02 g}{mol} = 75.3 J/mol \cdot C$

Have advanced students calculate the molar heat capacities of various substances. They may notice that all metals have very similar molar heat capacities. Have them speculate as to why this is the case. Provide more problems which require students to use the equation below to relate molar heat capacities to heat, moles, and temperature change.

 $q = C_p \times n \times \Delta T$

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 17.1 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 17.1 in CK-12 Chemistry –Intermediate Flexbook® resource.

Points to Consider

A complete understanding of a chemical reaction requires knowledge of the amount of heat that is absorbed or released during the reaction.

- How can heat be shown in the chemical equation?
- How can the heat involved in a reaction be used in stoichiometry calculations?

17.2 Thermochemical Equations

Key Concept

The enthalpy change of a chemical reaction or physical process is equal to the heat absorbed or released when the change occurs at constant pressure. Calorimetry is the experimental procedure of measuring heat changes in the laboratory. A thermochemical equation is a chemical equation which also shows the enthalpy change of the reaction. Exothermic reactions have a negative enthalpy change, while endothermic reactions have a positive enthalpy change.

Standards

Lesson Objectives

- Define enthalpy, and know the conditions under which the enthalpy change in a reaction is equal to the heat absorbed or released.
- Describe the principles behind calorimetry, and be able to calculate the heat absorbed or released during a process that occurs in a calorimeter.
- Write and solve problems with thermochemical equations.

Lesson Vocabulary

- calorimeter: An insulated container that is used to measure heat changes during a chemical or physical process.
- calorimetry: The measurement of the transfer of heat into or out of a system during a chemical reaction or physical process.
- enthalpy: The heat content of a system at constant pressure.
- heat of reaction: The enthalpy change for a chemical reaction.
- thermochemical equation: A chemical equation that includes the enthalpy change of the reaction.

Teaching Strategies

Introducing the Lesson

In the previous lesson, the concept of heat was introduced and heat changes were calculated for substances whose temperature was changed. This lesson extends the idea of heat changes to chemical reactions. Define enthalpy and discuss how the majority of chemical reactions encountered in the lab are under constant pressure conditions and so for most situations heat change (q) and enthalpy change (ΔH) are the same. Use the Science Inquiry activity below

to introduce the concept of calorimetry. Either do the activity as a lab or as a demonstration, working through the calculations together.

Common Misconceptions

Students have some difficulty with the sign of enthalpy changes as it relates to calorimetry experiments. In the majority of calorimetry experiments, the water is the surroundings. Therefore, when ΔT is positive (a temperature increase), the q_{surr} is also positive. This means that q_{sys} is negative since the heat gained by the surroundings was lost by the system. This is why exothermic reactions always have a negative value for ΔH –because the direction of heat flow is from the system to the surroundings. The opposite is true for endothermic reactions. The surroundings decreases in temperature and the sign of ΔH for the system is positive.

Science Inquiry

The following link is for an experiment which allows students to determine the enthalpy change for the reaction of magnesium metal with hydrochloric acid. Students will use a simple foam cup calorimeter for the experiment.

http://fcw.needham.k12.ma.us/~annap/FOV1-0010C014/FOV1-0010D34F/Molar%20Enthalpy%20of%20Mg%2BHCl .pdf?Plugin=Loft

Differentiated Instruction

Review mass and mole stoichiometry problems (Chapter 12) before bringing in thermochemical equations and enthalpy changes into the calculations. Help students see that the ΔH of a reaction can be thought of in the same way as other reactants and products when doing stoichiometry calculations with energy. Give students plenty of guided practice problems.

Enrichment

Combustion reactions produce carbon dioxide, a greenhouse gas which has been linked to global warming due to an ever-increasing amount of the gas present in the atmosphere. One of the most important contributors to the increase in CO_2 is the transportation industry, or the burning of gasoline by cars and trucks. Have students research how much gasoline is consumed worldwide each year and do stoichiometry calculations to determine how much CO<sub2 is produced and how many kilojoules of heat energy are released when the gasoline is burned.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 17.2 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 17.2 in CK-12 Chemistry –Intermediate Flexbook® resource.

17.2. Thermochemical Equations

Points to Consider

Enthalpy changes always accompany changes in state.

- What are the heat of fusion and heat of vaporization, and how can they be used to solve problems?
- What is the heat of solution?

17.3 Heat and Changes of State

Key Concept

Enthalpy changes accompany the changes of state of matter. Molar heat of fusion is the heat absorbed when one mole of a substance melts, while molar heat of solidification is the heat released when one mole of a substance freezes. Molar heat of vaporization is the heat absorbed when one mole of a substance boils, while molar heat of condensation is the heat released when one mole of a substance condenses. The enthalpy change that occurs when one mole of a solute dissolves in water is called the molar heat of solution, a process that can be either exothermic or endothermic.

Standards

Lesson Objectives

- Describe the enthalpy change that occurs as a substance changes between the solid and liquid states.
- Describe the enthalpy change that occurs as a substance changes between the liquid and gas states.
- Calculate the enthalpy change involved in the change of state for any amount of a given substance.
- Calculate the enthalpy changes involved as substances dissolve in water.

Lesson Vocabulary

- molar heat of condensation: The heat released by one mole of a substance as it is converted from a gas to a liquid.
- molar heat of fusion: The heat absorbed by one mole of a substance as it is converted from a solid to a liquid.
- molar heat of solidification: The heat released by one mole of a substance as it is converted from a liquid to a solid.
- molar heat of solution: The heat absorbed or released when one mole of a substance is dissolved in water.
- molar heat of vaporization: The heat absorbed by one mole of a substance as it is converted from a liquid to a gas.

Teaching Strategies

Introducing the Lesson

Students are certainly familiar with changes of state, especially for water. Discuss the direction of heat flow as ice melts, as water freezes, as water boils, and as steam condenses. Introduce the concepts of molar heat of fusion,

solidification, vaporization, and condensation. If time permits, do the activity described below under the Science Inquiry in order to determine experimentally the molar heat of fusion of ice. Do practice calculations using each of the values.

Common Misconceptions

Students are familiar with the notion that heat must be supplied to ice to make it melt or to water to make it boil. However, it is not quite so intuitive that the freezing and condensation processes release heat and are exothermic. Make sure to write the equations out so that students can see that solidification and fusion are the exact opposite processes (as are condensation and vaporization). If the equation is reversed, the numerical value of the enthalpy change is the same, but only the sign changes.

Science Inquiry

The molar heat of fusion of ice can be easily measured in the laboratory by adding a known quantity of ice to a known volume of warm water and measuring the temperature change that occurs as the ice melts. The link below describes the procedure for such a lab activity.

http://www.google.com/url?sa=t&rct=j&q=&esrc=s&source=web&cd=7&sqi=2&ved=0CEMQFjAG&url=http%3A%2 F%2Fwww.greenwichschools.org%2Fuploaded%2Ffaculty%2Farleene_ferko%2FUnit_02_LAB_Heat_of_Fusion_of_Ice.doc&ei=uOOeVKeFDonroAS-0oLQBg&usg=AFQjCNFCT1uGsuy3pihGZRUmOg3rhhTlOQ&sig2=lLV9zcDo tkpRG8bn0Hb5Sg&cad=rja

Differentiated Instruction

Students with less mathematical ability will struggle with the 5-step type of problems described in the text: heating ice from below 0°C to steam above 100°C. Have students work with a partner and break these problems down into discrete steps. In other words, first calculate the heat required to warm a given amount of ice up to 0°C. Make sure that students remember how to do that calculation. In a separate question, have the students calculate the heat required to melt that amount of ice. Continue until all five calculations are mastered. Lastly, add the enthalpy changes together.

Enrichment

Have students design a procedure for measuring the molar heat of solution for a soluble ionic compound such as calcium chloride. Different groups can compare their procedures to work out the best way to do the experiment and test it out in the laboratory. Compare the results to the literature value for the molar heat of solution.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 17.3 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 17.3 in CK-12 Chemistry –Intermediate Flexbook® resource.

Points to Consider

A formation reaction is a reaction in which elements in their standard states are combined to form a compound.

- How can knowledge of heats of formation be used to determine the enthalpy change for any reaction?
- What does it mean to say that heats of reaction are additive?

17.4 Hess's Law

Key Concept

Hess's law of heat summation states that when two or more thermochemical equations are added together to give a final equation, the heats of reaction for those equations can also be added together to give the heat of reaction for that final equation. The enthalpy change for any reaction can also be calculated by summing the standard heats of formation of the products and subtracting the sum of the standard heats of formation of the reactants.

Standards

Lesson Objectives

- Use Hess's law of heat summation to add chemical reactions together in such a way as to produce a desired final equation. Calculate the enthalpy change for that final reaction.
- Define the standard heat of formation of a compound.
- Use known values for standard heats of formation to calculate a previously unknown standard heat of reaction.

Lesson Vocabulary

- heat of combustion: The heat released when one mole of a substance is completely burned.
- Hess's law of heat summation: A law that states that if two or more thermochemical equations can be added together to give a final equation, then the heats of reaction can also be added to give a heat of reaction for the final equation.
- standard heat of formation: The enthalpy change associated with the formation of one mole of a compound from its elements in their standard states.

Teaching Strategies

Introducing the Lesson

Ask the class why enthalpy changes for reactions cannot always be determined directly by an experiment. What might prevent it from being possible? Describe Hess's law and carefully work through an example of adding multiple equations together to give a final equation, as in the acetylene example from the text. Make sure that students understand how to cancel out substances if they appear in equal amounts as a reactant in one equation and as a product in another.

Common Misconceptions

When working with heats of formation, students may have trouble understanding that it is only an element in its standard state that has a standard heat of formation that is equal to zero. They must be aware of elements whose standard state is diatomic, such as N2 and H2, or unusual elements such as P4, which is the standard state for phosphorus. Some elements have different allotropes, and only one of those forms is the standard state. For example, graphite is the standard state of carbon and the rhombic form of sulfur is the standard state for that element.

Students will often forget to multiply the standard heats of formation by the coefficient from the balanced chemical equation when using that form of Hess's law (see Sample Problem 17.8).

Science Inquiry

Hess's Law can be investigated through the following lab activity. Students measure the enthalpy change associated with three simple reactions. The first two reactions can be added together to give the third reaction and so students can test to see if the enthalpy changes (Δ H) add up in the same way to give the Δ H for the third reaction.

http://www2.ucdsb.on.ca/tiss/stretton/chem2/enthlab2.htm

Differentiated Instruction

Some students may have difficulty visualizing how substances can cancel out when adding reactions together with Hess's law. Use the following simple math equations, written in the style of a chemical equation, to illustrate.

 $1\!+\!3\to4$

 $4+5 \rightarrow 6+3$

If the 4 and the 3 are canceled out from the two equations, the result is:

 $1 + 5 \rightarrow 6$

Since that is a correct equation, it shows how the canceling of "reactants" and "products" works.

Enrichment

Many more standard heats of formation are known than what is shown in Table 17.3. Have students choose an element and research to find the heats of formation of as many compounds of that element as they can find. Have them write the chemical equations showing each of the formation reactions.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 17.4 worksheets in CK-12 Chemistry –Intermediate Workbook. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 17.4 in CK-12 Chemistry –Intermediate Flexbook® resource.

Points to Consider

Chemical reactions proceed at a wide variety of speeds or rates. Some reactions happen almost instantly, as soon as the reactants come into contact with one another, while other reactions may take years to reach completion.

- How do chemical reactions occur on the molecular level?
- What factors influence the rate of a chemical reaction?

CHAPTER 18

Kinetics

Chapter Outline

- 18.1 **RATES OF REACTIONS**
- 18.2 **RATE LAWS**
- 18.3 **REACTION MECHANISMS**

Chapter Overview

Chemical kinetics is the study of the rates of chemical reactions. Collision theory is a set of principles which states that the rate of a reaction is governed by the frequency of collisions between reacting particles as well as the force of the collisions and the orientation of the particles. The rate of any chemical reaction is mathematically described by the rate law for that reaction. A rate law includes a specific rate constant and the dependency of the rate on each of the reactant concentrations. A reaction mechanism is a sequence of elementary reactions which together comprise the overall chemical reaction.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

Lesson	Class Period(s) (60 min)
18.1 Rates of Reactions	2
18.2 Rate Laws	2
18 3 Reaction Mechanisms	1-2

TABLE 18.1: Class Period(s) (60 min)

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 18.2: Chemistry Concepts

Flexbook Lesson	Chemistry Concepts
18.1 Rates of Reactions	18.1 – 18.7
18.2 Rate Laws	18.8 - 18.10

TABLE 1	8.2:	(continued)
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Flexbook Lesson	Chemistry Concepts
18.3 Reaction Mechanisms	18.11 – 18.15

Keywords

- catalyst
- intermediate
- kinetics
- rate law
- reaction mechanism
- reaction rate

18.1 Rates of Reactions

Key Concept

The rate of a chemical reaction is measured as the change in concentration of a reactant or product as a function of time. Reaction rates are governed by collision theory, which states that reacting particles must collide with sufficient energy and in the correct orientation. The energy required to initiate a reaction is called activation energy. Activation energy is a barrier or hill in a potential energy diagram. Reaction rates generally increase with increases in concentration, gas pressure, temperature, or surface area of reactants. The use of a catalyst increases the rate by lowering the required activation energy.

Standards

Lesson Objectives

- Be able to express the rate of a chemical reaction.
- Describe the collision theory as it relates to chemical reactions.
- Draw and analyze a potential energy diagram for a reaction, including heat of reaction, activation energy, and the activated complex.
- Describe and explain various factors that influence the rates of reactions.

Lesson Vocabulary

- activated complex: An unstable arrangement of atoms that exists momentarily at the peak of the activation energy barrier.
- activation energy: The minimum energy that colliding particles must have in order to undergo a reaction.
- catalyst: A substance that increases the rate of a chemical reaction by lowering the activation energy without being used up in the reaction.
- Collision theory: A set of principles that states that reacting particles can form products when they collide with one another, provided those collisions have enough kinetic energy and the correct orientation.
- Potential energy diagram: A diagram that shows the change in potential energy of a system as reactants are converted into products.
- Reaction rate: The change in concentration of a reactant or product with time.

Teaching Strategies

Introducing the Lesson

Begin by discussing the concept of reaction rate and emphasize that there is an extremely wide variation in rates possible. Put three columns on the board, titled "fast", "medium", and "slow". Ask the students to think of everyday chemical reactions and to place them in the correct category. Make sure that they understand that slow reactions can even mean hundreds or thousands of years. Once you have some examples, discuss why some reactions are fast, while others are slow. See if the students can come up with ways to make a slow reaction occur more quickly.

Common Misconceptions

Students may mix up the effects of temperature and of a catalyst on reaction rate, especially as it relates to a potential energy diagram. Describe it with the following analogy. Suppose that you are riding your bicycle and see a big hill coming up. What do you do? You speed up as much as you can before you get to the hill in order to make it easier to get over it. This is what an increase in temperature does to a chemical reaction. It gives the reacting particles more energy so that it is more likely that they will be able to get over the hill. What about a catalyst? Another way to get over the hill on your bicycle is to have a friend with some heavy explosives blow off the top of the hill (not recommended). A catalyst changes the way in which the reaction happens so that the activation energy is lower. It's not as big of a hill to climb. A catalyst does not give the reacting particles any more energy than what they had before. Only temperature does that.

Science Inquiry

The various factors that affect reaction rates are a great topic for inquiry. The lab at the link below allows students to investigate the effects of temperature, concentration, particle size, and catalysts on some simple reactions. Do this before discussing collision theory.

http://www.haspi.org/curriculum-library/Med-Chem-lessons/09%20Standard%208%20Reactions%20Rates/Labs%20an d%20Activities/rate_factors.pdf

Demonstration

The demonstration described at the link below is a good way to show students the effect of a catalyst on a colorful reaction, as well as a brief view of an activated complex. The reaction involves multiple oxidation states of cobalt ion, which of which is a different color. The cobalt(III) chloride catalyst is pink. When it is added to the reaction, the color gradually changes to green because the cobalt changes its oxidation state from 3+ to 6+. As with any catalyst, the original Co³⁺ is regenerated at the end and the solution turns back to pink.

https://www.flinnsci.com/media/622180/95020.pdf

Differentiated Instruction

Understanding of potential energy diagrams is essential. In small groups, have students draw some potential energy diagrams based on information that you give to them. Give groups various values for the following: (1) PE of reactants, (2) PE of products, (3) enthalpy change (ΔH), (4) activation energy without a catalyst, and (5) activation energy with a catalyst. Have the groups draw colorful potential energy diagrams on butcher paper.

Enrichment

Catalysts are very important in the industrial production of common chemicals as well as in cutting-edge chemical research. Have interested students research the many varied types of catalysts that are commonly used today and to make a brief report on how these catalysts work. The end result of their findings could be a PowerPoint presentation or a poster.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 18.1 worksheets in *CK-12 Chemistry - Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 18.1 in *CK-12 Chemistry - Intermediate* Flexbook[®] resource.

Points to Consider

Rate laws provide a quantitative relationship between the rate of a reaction and the concentrations of its reactants.

- How can experiments be designed in order to determine the rate law for a reaction?
- What is a specific rate constant?

18.2 Rate Laws

Key Concept

The rate law for a reaction shows the dependence of the reaction rate on the concentration of each of the reactants in the reaction. The rate is equal to a specific rate constant (k) multiplied by the concentration of each reactant, raised to some power. A rate law can only be determined experimentally. A large value for the rate constant is indicative of a fast reaction, while a small value is indicative of a slow reaction. Reaction order refers to the power to which the reactant concentration is raised in the rate law.

Standards

Lesson Objectives

- Use a rate law to describe the concentration dependence of a reaction rate.
- Determine the rate law for a reaction by analysis of experimental data.
- Calculate the specific rate constant from experimental data.

Lesson Vocabulary

- first-order reaction: A reaction in which the rate is proportional to the concentration of only one reactant.
- rate law: An expression showing the relationship of the reaction rate to the concentrations of each reactant.
- specific rate constant: A proportionality constant relating the rate of a reaction to the concentrations of reactants.

Teaching Strategies

Introducing the Lesson

Start by drawing a simple reaction on the board showing a molecule of A reacting with a molecule of B. With just one molecule of each, there is just one possible collision. Then draw a second reaction with two molecules of A and still one molecule of B. Now there are two possible collisions. With twice as many possible collisions, the rate would double. Repeat this with two molecules of B and one of A. Finally, show how there are four possible collisions if there are two molecules of A and two molecules of B. The reaction would now be four times as fast as the original reaction. This allows the students to relate rate laws to collision theory. The rate law for this "reaction" would be: rate = k[A][B].

Common Misconceptions

Students may think that a rate law can be written simply by using the coefficients of the balanced equation. Stress that rate laws can only be determined by experiment. Do several examples of the problems where concentrations are systematically changed and the initial rates are used to figure out the rate law. When mechanisms of a reaction are known, the rate law can indeed be written simply from knowledge of the rate-determining step (the slowest step) of the reaction. However, that is beyond the scope of this lesson and appears in the next one. In reality, knowledge of the experimental rate law is one part of deducting the correct mechanism.

Science Inquiry

The determination of a rate law is a rather time-consuming experiment and may prove to be a bit complex for a beginning level class. The lab activity below is an example of an iodine clock reaction, where the time for the reaction to be completed is detected by the sudden appearance of a blue color due to the appearance of iodine in the presence of a starch indicator. The concentrations of various reactants are systematically changed, the reaction is timed, and the data is analyzed to determine the rate law. If you have an advanced class and ample time, the experiment will work well. Alternatively, several of the reactions could be done as a demonstration to illustrate the concept of how kinetics experiments are performed.

http://www2.ohlone.edu/people/jklent/labs/101B_labs/Rate%20Law.pdf

Differentiated Instruction

Practice determining the rate of a reaction at different points by calculating the slope of tangents to a curved line (Figure 18.8). Draw these types of curves on graph paper or generate them with a computer program. Have students use a ruler to figure out the "rise over run" or $\frac{\Delta y}{\Delta x}$ at several points on the curve. They should find that the slope is higher early in the reaction and lower later in the reaction. This emphasizes how the majority of chemical reactions slow down over the course of the reaction. Discuss that the reason for this slowing down is because there are fewer reactant molecules available and so collision frequency decreases with time.

Enrichment

Advanced math students (who have studied logarithms) can perform calculations with a form of a rate law called an integrated rate law. The integrated rate law for a 1st-order reaction is shown below.

$$\ln\left(\frac{[A]}{[A]_o}\right) = -kt$$

In this equation, $[A]_o$ is the initial concentration of reactant A, k is the specific rate constant, and t is the time. It allows the student to calculate the concentration of a reactant at any point during a reaction, as long as the rate constant is known. Alternatively, if the concentration of the reactant is measured at time "t", the rate constant can be calculated.

Example: The rate constant for a first-order reaction is 0.0025 s^{-1} . If the starting concentration in a certain reaction is 1.60 M, what will be the concentration after 5.0 minutes have passed?

Solution: 5.0 minutes is equal to 300 seconds. To solve the equation, take e^x (antilogarithm) of both sides of the equation. Then multiply by $[A]_o$.

$$\ln\left(\frac{[A]}{1.60}\right) = -0.0025 \, s^{-1}(300 \, s)$$

А

$$= 1.60 M \times e^{(-0.0025 s^{-1} \times 300 s)} = 0.76 M$$

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 18.2 worksheets in *CK-12 Chemistry - Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 18.2 in *CK-12 Chemistry - Intermediate* Flexbook[®] resource.

Points to Consider

Chemical reactions can generally be broken down into a series of simple steps that convert the reactants into the products. This set of steps is referred to as the reaction mechanism.

- What is the relationship between a reaction mechanism and the rate law for that reaction?
- What is an intermediate in a chemical reaction?
18.3 Reaction Mechanisms

Key Concept

A reaction mechanism consists of a sequence of elementary steps which combine together to produce the entire reaction. The molecularity of an elementary step refers to the number of reacting particles in that step. Most elementary reactions are unimolecular or bimolecular, but can occasionally be termolecular. An intermediate is a substance that is produced in an early step of a mechanism, then consumed in a later step. The rate law for a given chemical reaction is equal to the rate law for the slowest step of the mechanism, called the rate-determining step.

Standards

Lesson Objectives

- Describe how the elementary steps of a chemical reaction combine to produce the overall reaction mechanism.
- · Be able to identify intermediates and catalysts when they appear in a reaction mechanism.
- Write a rate law for an elementary step, and identify the molecularity of the reaction.
- Identify the rate-determining step of a reaction.
- Analyze the potential energy diagram for a multi-step reaction.

Lesson Vocabulary

- elementary step: One step in a series of simple reactions that show the progress of a reaction at the molecular level.
- intermediate: A species which appears in the mechanism of a reaction, but not in the overall balanced equation.
- molecularity: The number of molecules reacting in an elementary step of a reaction.
- rate-determining step: The slowest step in the sequence of steps in a reaction mechanism.
- reaction mechanism: The sequence of elementary steps that together comprise an entire chemical reaction.

Teaching Strategies

Introducing the Lesson

The concept of a reaction mechanism is best understood by relating it to something familiar like building a birdhouse. Ask students what some of the steps would be if one were to build a birdhouse. Each individual step (measuring, sawing wood, gluing or nailing together) are analogous to the elementary steps of the overall reaction (building the birdhouse). Further, the overall rate of a reaction can be no faster than the slowest step. If the birdhouse building

18.3. Reaction Mechanisms

involves waiting for glue to dry at one point, that is likely going to be the slowest step. It doesn't really matter how fast you can measure. The rate of birdhouse building is determined by the time it takes for the glue to dry. The concept of rate-determining step is further illustrated by the activity found under the Science Inquiry section below.

Common Misconceptions

It is easy for students to be confused at the difference between a catalyst and an intermediate as both relate to the sequence of steps in a reaction mechanism. Stress that an intermediate is generated in an early step and consumed in a later one. Catalysts are sort of the opposite. They are consumed in an early step, then are regenerated in a later step.

Science Inquiry

The concept of rate-determining step and reaction mechanism is brought out in this classroom activity below. Students work in groups to perform a task of stuffing envelopes, with each person in the group performing a certain task. Some tasks, by their nature, are slower than others. Make sure to copy the accompanying worksheet for after the activity is completed.

http://mste.illinois.edu/courses/educ362sp04/folders/thompson/Lesson%20Plan%202.htm

Differentiated Instruction

This chapter uses a great deal of new vocabulary that students use sparingly or not at all in any other topics in chemistry. Have students write out a list of all of the vocabulary terms for the chapter along with definitions. They can also include examples of each that they have encountered in the text or in class activities.

Enrichment

The destruction of the stratospheric ozone layer by a class of chemicals called chlorofluorocarbons (CFCs) is a good example of a real-life chemical reaction that proceeds by a multi-step mechanism. The reaction produces chlorine atoms which act as a catalyst in the reaction, meaning that a single chlorine atom can bring about the decomposition of many, many ozone molecules. Have students prepare a poster or report discussing the importance of the ozone layer, the chemical reactions involved when it reacts with CFCs, the history and ultimate banning of CFCs, and the slow recovery of the ozone layer.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 18.3 worksheets in *CK-12 Chemistry - Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 18.3 in *CK-12 Chemistry - Intermediate* Flexbook[®] resource.

Points to Consider

Many chemical reactions do not proceed entirely to completion. In other words, after some of the products are formed in the reaction, they can recombine to form the reactants again. Such reactions are said to be reversible.

- How can a reversible reaction be shown in a chemical equation?
- What is the relationship between the rate of the forward reaction and the rate of the reverse reaction?



Equilibrium

Chapter Outline

- **19.1 THE NATURE OF EQUILIBRIUM**
- **19.2 CELL POTENTIALS**
- **19.3 SOLUBILITY EQUILIBRIUM**

Chapter Overview

Equilibrium is a state of balance that occurs in a reversible chemical reaction when the rate of the forward reaction is exactly equal to the rate of the reverse reaction. Since the concentrations of reactants and products are unchanging at equilibrium, a particular ratio of product concentrations to reactant concentrations can be calculated and is called the equilibrium constant. A system that is at equilibrium will respond to a disturbance in a way that brings the system back to equilibrium again. Solubility equilibrium is the equilibrium that exists between dissolved and undissolved solute in a saturated solution of an ionic compound.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 19. 1	: Class Period(s) (60 min)
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Lesson	Class Period(s) (60 min)
19.1 The Nature of Equilibrium	2
19.2 Le Châtelier's Principle	2
19.3 Solubility Equilibrium	2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 19.2:	Chemistry Concepts
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Flexbook Lesson	Chemistry Concepts
19.1 The Nature of Equilibrium	19.1 – 19.4
19.2 Le Châtelier's Principle	19.5 – 19.10

TABLE 19.2: (continued)

Flexbook Lesson	Chemistry Concepts
19.3 Solubility Equilibrium	19.11 – 19.15

Keywords

- common ion effect
- equilibrium
- equilibrium constant
- LeChâtelier's principle
- molar solubility

19.1 The Nature of Equilibrium

Key Concept

In a reversible reaction, products are capable of being converted back into reactants. Chemical equilibrium is achieved when the rate of the forward reaction is equal to the rate of the reverse reaction. At equilibrium, the concentrations of reactants and products remain constant. For every reaction that reaches equilibrium, a specific ratio of product concentrations to reactant concentrations is maintained and is called the equilibrium constant (K_{eq}) for that reaction. An equilibrium constant that is much larger than 1 means that the products are present in larger concentrations than the reactants, while an equilibrium constant that is much less than 1 means that the reactant concentrations are greater than the product concentrations.

Standards

Lesson Objectives

- Describe the nature of a reversible reaction.
- Define chemical equilibrium.
- Write chemical equilibrium expressions from balanced chemical equations.
- Calculate the equilibrium constant for a reaction, and perform calculations involving that constant.

Lesson Vocabulary

- chemical equilibrium: The state of a system in which the rate of the forward reaction is equal to the rate of the reverse reaction.
- equilibrium constant (K_{eq}) : The ratio of the mathematical product of the concentrations of the products of a reaction to the mathematical product of the concentrations of the reactants of the reaction, with each concentration raised to the power of its coefficient in the balanced chemical equation.
- reversible reaction: A reaction in which the conversion of reactants to products and the conversion of products to reactants occur simultaneously.

Teaching Strategies

Introducing the Lesson

Equilibrium has been introduced before, with respect to the vaporization and condensation that occurs in a stoppered flask of a liquid. Hold up such a flask and ask the class if the level of liquid in the flask will change over time. Write

an equilibrium equation on the board and discuss how the level remains the same because the rate of vaporization is exactly equal to the rate of condensation.

Take an empty flask and pour some water into it. Place a stopper on the flask. Is the system at equilibrium right away? Discuss how at first, only the "forward" reaction can take place because no water vapor was originally present. It takes some time before equilibrium can be attained.

Common Misconceptions

Students need practice writing and manipulating equilibrium constant (K_{eq}) expressions. Students need to remember that the product(s) of the reaction are written as the numerator, while the reactant(s) are written as the denominator. Further, when there are multiple reactants and/or products, their concentrations are multiplied together and not added, as it looks like from the chemical equation. Finally, they must remember to raise each concentration to the power of its coefficient from the balanced equation. When calculating, make sure to follow the order of operations and apply the exponents before multiplying and dividing.

Science Inquiry

To accurately determine an equilibrium constant in the laboratory takes great care, a substantial amount of time, and specific instrumentation. The lab below describes a classic examination of the equilibrium between iron(III) ions and thiocyanate ions to produce a complex ion:

$$Fe^{3+}(aq) + SCN^{-}(aq) \rightarrow FeSCN^{2+}(aq)$$

The experiment can be performed with an advanced class, but required the use of a spectrophotometer. The typical spectrophotometer that is used in the high school lab is a simple Spectronic 20, available through most chemical supply companies.

http://staff.buffalostate.edu/nazareay/che112/ex6.htm

Differentiated Instruction

The concept of reversible reactions and equilibrium can be introduced simply by a whole-class activity. In a large open space, divide the class roughly, but not exactly, in half. Count the number in each group. Designate one group as the reactants and another as the products. Initiate the reaction at a "low temperature" by having one "reactant" class member move to the product group. Simultaneously, direct a "product" class member to move to the reactant group. Raise the temperature by having more class members move to the other group, but always keep the numbers in each group the same. This illustrates the dynamic nature of equilibrium.

Enrichment

The Haber-Bosch process for the production of ammonia is one of the most industrially important reactions and is a good example of a reaction that naturally attains equilibrium. Have students research the history of the Haber-Bosch process, including how it was developed and the importance of ammonia as one of the most-produced chemicals in the world. They can also study how the reaction has been refined over the years and how the equilibrium is manipulated in order to maximize the production of ammonia.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 19.1 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 19.1 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Disruptions to a system at equilibrium can occur in the form of a change in temperature, pressure, or concentrations of one or more of the substances in the equilibrium.

- How does an equilibrium respond to a change in conditions?
- Is the value of the equilibrium constant affected by such changes?

19.2 Cell Potentials

Key Concept
Standards
Lesson Objectives
Lesson Vocabulary
Teaching Strategies
Introducing the Lesson
Common Misconceptions
Science Inquiry
Differentiated Instruction
Enrichment

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 19.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 19.2 in *CK-12 Chemistry –Intermediate* Flexbook® resource.

Points to Consider

19.3 Solubility Equilibrium

Key Concept

The equilibrium that exists between undissolved solute and dissociated ions in a saturated solution of an ionic compound is called a solubility equilibrium. The solubility product constant, K_{sp} , is equal to the mathematical product of the molar concentrations of the ions in solution, each raised to the power of their coefficient in the dissociation equation. Experimental solubility data can be used to determine K_{sp} values, or known K_{sp} values can be used to calculate solubility. The value of K_{sp} can be used to predict whether a precipitate will form when two different solutions are mixed. The common ion effect results in a lowering of the solubility of an ionic compound.

Standards

Lesson Objectives

- Write solubility product constant expressions for nearly insoluble ionic compounds.
- Calculate the value of K_{sp} for a compound from its solubility and determine the solubility of a compound with a known K_{sp} .
- Use the K_{sp} of a compound to predict whether a precipitate will form when two solutions are mixed together.
- Describe the common ion effect and its relationship to solubility.

Lesson Vocabulary

- common ion: An ion that is in common to both salts in a solution.
- common ion effect: A decrease in the solubility of an ionic compound as a result of an ion that is in common to one of the ions already in the saturated solution.
- molar solubility: The number of moles of solute in one liter of saturated solution.
- solubility product constant (K_{sp}) : The mathematical product of the ions that come from the dissociation of a nearly insoluble salt, each raised to the power of the coefficient of the ion in the equation.

Teaching Strategies

Introducing the Lesson

Show the students a saturated solution of an ionic compound. Make sure the solution has visible solid on the bottom of the container. Choose a colorful one such as the brilliant yellow of lead(II) iodide. Ask them what would happen if you added more solid PbI_2 to the solution. They should predict that nothing will happen because the solution

is already saturated and thus cannot dissolve any more solute. Discuss the equilibrium that exists and that the rate of dissolving is equal to the rate of recrystallization. Write the solubility product constant expression (K_{sp}) on the board. Introduce the concept of molar solubility as the highest possible molarity solution that can be attained and that whatever happens to be the molarity of the *PbI*₂ solution, that is its molar solubility.

Common Misconceptions

When doing ICE problems in order to find molar solubility from K_{sp} , students can get confused about the effect of a coefficient for one or both of the ions in the dissociation equation. For example, suppose that one of the ions has a coefficient of two in the balanced dissociation equation. That means that the change in concentration of that ion in the ICE table is +2s rather than just +s. It also means that in the K_{sp} expression the coefficient of two necessitates that the term for that ion be squared. In other words, the term for that particular ion is $(2s)^2$. So the coefficient in the equation affects the mathematics of the problem twice.

Science Inquiry

The K_{sp} of a slightly soluble salt can be determined experimentally. The procedure involves mixing two solutions and observing if a precipitate forms or not. If so, the original solutions are diluted and mixed again. Eventually, the solutions will become so dilute, that no visible precipitate will form. These concentrations can then be used to calculate the K_{sp} .

http://www.delsearegional.us/academic/classes/highschool/science/chemistry/firstyear/labs/labs/ksp/ksp.pdf

Differentiated Instruction

ICE tables will be a challenge for students of lower mathematical ability. Pair such students up those of higher math ability and assign plenty of practice problems. Have each of the students take turns writing up the solutions to the problems as they work them out.

Enrichment

Kidneys stones are often comprised of the mostly insoluble compound calcium oxalate. For students that are interested in biology, have them research kidney stones, including causes, risk factors, and treatment options. Have them prepare a brief report of their findings.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 19.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 19.3 in *CK-12 Chemistry –Intermediate* Flexbook® resource.

Points to Consider

In the course of an exothermic reaction, heat is released from the system into the surroundings, resulting in a decrease in the enthalpy of the system. This is a favorable reaction because nature prefers a state of lower energy.

- What is meant by the term "driving force" as it relates to chemical reactions?
- What other force is responsible for the occurrence of endothermic reactions, which absorb heat into the system?

Sample Answers

- A driving force is a particular change in the system that causes the reaction to occur
- Entropy change is the other driving force for reactions (other than enthalpy change), and reactions tend to occur so that the entropy of the system increases.

CHAPTER 20 Entropy and Free Energy

Chapter Outline

- **20.1 ENTROPY**
- 20.2 CELL POTENTIALS
- 20.3 ELECTROLYSIS

Chapter Overview

The driving forces for all chemical reactions and physical processes are the drive towards lower enthalpy and higher entropy. Entropy is a measure of the randomness or disorder of a system. Reactions in which the driving forces favor the formation of products are said to be spontaneous. Spontaneous reactions release free energy. The temperature is the controlling factor in the spontaneity of reactions in which one driving force favors the reaction while the other does not. A reversible reaction that is spontaneous in the forward direction has an equilibrium constant that is larger than one, indicating that the concentrations of products are greater than the concentrations of reactants.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 20. 1	: Class Period(s)	(60 min)
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Lesson	Class Period(s) (60 min)
20.1 Entropy	1-2
20.2 Spontaneous Reactions and Free Energy	2
20.3 Free Energy and Equilibrium	2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 20.2: Chemistry Concepts

Flexbook Lesson	Chemistry Concepts
20.1 Entropy	20.1 - 20.2
20.2 Spontaneous Reactions and Free Energy	20.3 - 20.5

TABLE 20.2: (continued)

Flexbook Lesson	Chemistry Concepts
20.3 Free Energy and Equilibrium	20.6 - 20.8

Keywords

- driving force
- entropy
- free energy
- nonspontaneous reaction
- spontaneous reaction

20.1 Entropy

Key Concept

All chemical reactions and physical processes are governed by two driving forces. One is the drive towards a decrease in enthalpy of the system. The second driving force is towards an increase in entropy. Entropy is a measure of the disorder or randomness of a system. The entropy change for any chemical reaction can be determined from the standard entropies of all the reactants and products.

Standards

Lesson Objectives

- Identify the two driving forces behind all chemical reactions and physical processes.
- Describe entropy, and be able to predict whether the entropy change for a reaction is increasing or decreasing.
- Calculate the standard entropy change for a reaction from the standard entropies of all substances in the reaction.

Lesson Vocabulary

• entropy: A measure of the degree of randomness or disorder of a system.

Teaching Strategies

Introducing the Lesson

Discuss how the energy change of a reaction is one of the controlling factors of a reaction. The stability of a chemical system is increased when a reaction is exothermic, releasing enthalpy to the surroundings. However, endothermic reactions do exist. Why? The reason is that there is another controlling factor for reactions and that factor is called entropy. Use everyday examples like the Rubik's cube or a messy desk to help the students understand the concept of entropy.

Common Misconceptions

Most students grasp the concept of entropy as disorder as it pertains to everyday situations. The biggest misconception revolves around the meaning of driving force. Explain that reactions don't simply occur every time two reactants are mixed together. A driving force refers to a cause of a potential reaction. The driving forces for chemical reactions are the drive towards lower enthalpy and the drive towards higher entropy. If a proposed reaction results in an increase in enthalpy and a decrease in entropy, then there is no possible way that reaction can occur.

Science Inquiry

The simulation below acts as an introduction to the concept of entropy from the perspective of probabilities.

http://www.presentingscience.com/quantumcasino/simulations/entropyanddisorder.html

Differentiated Instruction

Students can draw particle diagrams to help them see how entropy is related to the particular state of matter. Since gas particles are so far apart, the entropy of a gas is much greater than the entropy of the condensed states, liquids and solids. When examining a chemical reaction, use the particle diagrams to help decide whether the entropy of the system is increasing or decreasing.

Enrichment

The second law of thermodynamics concerns the concept of entropy. Have students look up the second law and summarize its meaning as it relates to the current lesson.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 20.1 worksheets in *CK-12 Chemistry - Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 20.1 in *CK-12 Chemistry - Intermediate* Flexbook[®] resource.

Points to Consider

The drives toward lower enthalpy and greater entropy determine whether a chemical reaction is likely to occur under a given set of conditions.

- How can enthalpy and entropy changes be combined into one quantity?
- If a reaction is not spontaneous under one set of conditions, is it possible that it will be spontaneous under some other set of conditions?

20.2 Cell Potentials

Key Concept

A spontaneous process or reaction is one in which the products are favored over the reactants under a given set of conditions. A spontaneous reaction releases free energy, which can be used to do work. The change in free energy of a reaction is a combination of enthalpy change and entropy change, the two driving forces for reactions. The sign of the free energy change, ΔG , is negative for a spontaneous reaction and positive for a nonspontaneous reaction. The temperature at which a reaction takes place can affect the spontaneity of the reaction.

Standards

Lesson Objectives

- Describe the meaning of a spontaneous reaction in terms of enthalpy and entropy changes.
- Define free energy and calculate the change in free energy for a reaction using known values of the changes in enthalpy and entropy.
- Determine the spontaneity of a reaction based on the value of its change in free energy at various temperatures.

Lesson Vocabulary

- free energy: Energy that is available to do work; free energy is a combination of enthalpy and entropy and is used to determine spontaneity of a reaction.
- nonspontaneous reaction: A reaction that does not favor the formation of products at the given set of conditions.
- spontaneous reaction: A reaction that favors the formation of products at the conditions under which the reaction is occurring.

Teaching Strategies

Introducing the Lesson

Review the concept of driving forces and make sure that the class understands that favorable reactions are ones in which the enthalpy of the system decreases and the entropy of the system increases. Ask how many possibilities there are for a given reaction with these two variables. There are four, since both ΔH and ΔS can either increase or decrease. Provide examples that fit into each category. Introduce free energy as a new variable that will roll

enthalpy and entropy into one value so that chemists will be able to predict whether a given reaction is favorable or unfavorable.

Common Misconceptions

The most common misconception about spontaneity as it pertains to chemical reactions is that students tend to assume that a spontaneous reaction is fast. Use the combustion of a fuel such as wood or gasoline as an example. Combustion reactions are exothermic and result in an increase in entropy due to the production of gases. Therefore, they are spontaneous. However, most combustion reactions have a high enough activation energy that the reactions don't occur unless heated to a high temperature. Discuss the conceptual difference between kinetics and thermodynamics. Thermodynamics allows one to predict if a reaction will occur, but says absolutely nothing about how fast the reaction might be.

Science Inquiry

The activity below can be run as a demonstration or a short lab activity. It involves the modeling of a spontaneous reaction by using small straws and observing the ordering process of the straws as they float in a beaker of water.

https://www.flinnsci.com/media/622087/91846.pdf

Differentiated Instruction

Some students may have difficulty understanding that the reported values for entropy change, ΔS , must be converted from $J/K \cdot mol$ to $kJ/K \cdot mol$ when using the free energy equation. Either provide the ΔS values in the correct units already or specifically direct students to make the unit conversion first, before calculating the free energy change.

Enrichment

Free energy changes are very important in biological processes. Often, a necessary biochemical reaction will have a positive free energy change, called being endergonic. Such a reaction can happen if the reaction is linked to another process which is highly exergonic, releasing free energy. In this way, the sum change of the free energy of the two (or more) reactions can be negative and the entire linked process is spontaneous. Provide an opportunity for interested students to pursue the concept of free energy changes in biochemical reactions to find examples to show to the class.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 20.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 20.2 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

When a reversible reaction is at equilibrium, the concentrations of the reactants and products are constant, so neither the forward nor reverse reaction is spontaneous (favored).

- How can we determine the conditions under which a certain reaction will be at equilibrium?
- What is the relationship of the free energy change, ΔG , to the equilibrium constant, K_{eq} ?

20.3 Electrolysis

Key Concept

When a reaction is at equilibrium, the value of the free energy change, ΔG , is zero. One can solve the free energy equation for the temperature at which a reaction changes from being spontaneous to being nonspontaneous. The reactants and products in a change of state process are in equilibrium. The equilibrium constant (K_{eq}) of a reversible reaction can be calculated if the ΔG is known, or vice-versa.

Standards

Lesson Objectives

- Determine the temperature at which a reversible reaction will achieve equilibrium by using the Gibbs free energy equation.
- Describe the relationship between standard free energy change (ΔG°) and the equilibrium constant (K_{eq}) for reversible reactions.
- Convert between K_{eq} and ΔG° for a reaction at a given temperature.

Teaching Strategies

Introducing the Lesson

Ask the students how they can tell if a certain reaction is likely to be spontaneous. They should be able to remember that a reaction is spontaneous if it releases free energy, meaning that the sign of ΔG is negative. Now ask which side of an equation, the reactant side or the product side, would be favored if a reaction is spontaneous. In other words, if the reversible reaction $A \leftrightarrow B$ is spontaneous, is the concentration of A or B going to be higher at equilibrium? Lead them into seeing that a spontaneous forward reaction leads to more products while a spontaneous reverse reaction leads to more reactants.

Common Misconceptions

Students tend to think that spontaneity is an all-or-nothing concept. Describe how a reaction that is nonspontaneous is one in which the reactants are simply favored over the products. It is not the case that no products at all are formed in the reaction. If the forward reaction is endothermic, then an increase in temperature will begin to drive the reaction in the forward direction according to LeChâtelier's principle. This will cause the K_{eq} value to increase. Because of the relationship described in this lesson, the ΔG will decrease (become less positive). Eventually, a temperature will be reached when K_{eq} will be equal to 1 and ΔG will be equal to 0. At that point, the system has reached equilibrium.

Science Inquiry

The activity below allows students to see the effect of temperature on an equilibrium involving different colored cobalt complexes. Though qualitative in nature, the lab describes the effect of changing conditions on the value of the equilibrium constant. This could be taken further to discuss the effect on the free energy change for the reactions.

https://www.flinnsci.com/media/622072/91838.pdf

Differentiated Instruction

The mathematics behind the relationship between ΔG and K_{eq} may be beyond some students at this point. If necessary, teach the concept behind this lesson without including the math. That is, emphasize that a spontaneous reaction is one in which the products of the reaction are favored ($K_{eq} > 1$) as well as free energy is released ($\Delta G < 0$). Conversely, a nonspontaneous reaction has a $K_{eq} < 1$ and a $\Delta G > 0$, corresponding to the reactants being favored and free energy being absorbed.

Enrichment

The text shows how to calculate the entropy change that occurs during phase changes such as melting or vaporization. Find a table of enthalpies of fusion and vaporization for various substances along with their melting and boiling temperatures. Have students calculate the entropy changes (ΔS_{fus} and ΔS_{vap}) for the substances. See if they can relate the values that they find to variations in the strength of the intermolecular forces of the substances. Generally, substances with strong intermolecular forces (such as the hydrogen bonding in water) have higher entropies of fusion and vaporization than those substances with weaker forces.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 20.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 20.3 in *CK-12 Chemistry –Intermediate* Flexbook® resource.

Points to Consider

Acids and bases are extremely important types of chemical compounds that we use virtually every day.

- What are the physical and chemical properties of acids and bases?
- What types of reactions do acids and bases undergo?

CHAPTER **21**

Acids and Bases

Chapter Outline

- 21.1 ACID-BASE DEFINITIONS
- 21.2 THE PH CONCEPT
- 21.3 ACID AND BASE STRENGTH
- 21.4 ACID-BASE NEUTRALIZATION
- 21.5 SALT SOLUTIONS

Chapter Overview

Acids and bases are substances that take part in chemical reactions in which a hydrogen ion is transferred. Acid solutions have a pH lower than 7, while base solutions have a pH higher than 7. The pH of a solution is the negative logarithm of the hydrogen ion concentration. Acids and bases can be either strong or weak, depending on the extent of ionization in aqueous solution. Weak acids and bases can be described by acid or base ionization constants. Acids and bases combine chemically in a neutralization reaction and the controlled neutralization experiment is called a titration. The product of a neutralization reaction is a salt and water. When one of the ions of a salt reacts with water in a process called hydrolysis, the resulting solution can be either acidic or basic.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 21.1:

Lesson	Class Period(s) (60 min)
21.1 Acid-Base Definitions	1-2
21.2 The pH Concept	1-2
21.3 Acid and Base Strength	2
21.4 Acid-Base Neutralization	2
21.5 Salt Solutions	2

Chemistry Concepts

The table below matches each lesson from the FlexBook® student edition to the Chemistry Concepts.

FlexBook Lesson	Chemistry Concepts
21.1 Acid-Base Definitions	12.1–12.7
21.2 The pH Concept	21.8-21.11
21.3 Acid and Base Strength	21.12-21.15
21.4 Acid-Base Neutralization	21.16-21.20
21.5 Salt Solutions	21.21-21.23

TABLE 21.2: Chemistry Concepts

Keywords

- acids
- Arrhenius theory
- bases
- Brønsted-Lowry theory
- hydrolysis
- ionization constants
- Lewis theory
- neutralization
- titration

21.1 Acid-Base Definitions

Key Concept

Acids and bases can be described by their observable properties. Both are electrolytes and they react with each other to produce water and a salt. According to the Arrhenius definition, an acid is a substance that produces hydrogen ion in aqueous solution, while a base produces hydroxide ion. Under the Brønsted-Lowry definition, an acid donates a hydrogen ion in a reaction, while a base accepts a hydrogen ion. A conjugate acid-base pair is a pair of substance that are linked by the presence or absence of a hydrogen ion. A Lewis acid is a substance that accepts a pair of electrons in a reaction, while a base is a substance that donates a pair of electrons.

Standards

Lesson Objectives

- Describe the properties of acids and bases.
- Define an acid and a base according to the Arrhenius theory.
- Define and acid and a base according to the Brønsted-Lowry theory. Be able to identify the conjugate acid-base pairs in a Brønsted-Lowry acid-base reaction.
- Define an acid and a base according to the Lewis theory.

Lesson Vocabulary

- amphoteric: A substance that is capable of acting as either an acid or a base by donating or accepting hydrogen ions.
- Arrhenius acid: A compound which ionizes to yield hydrogen ions (H⁺) in aqueous solution.
- Arrhenius base: A compound which ionizes to yield hydroxide ions (OH⁻) in aqueous solution.
- Brønsted-Lowry acid: A molecule or ion that donates a hydrogen ion in a reaction.
- Brønsted-Lowry base: A molecule or ion that accepts a hydrogen ion in a reaction.
- conjugate acid: The particle produced when a base accepts a proton.
- conjugate acid-base pair: A pair of substances related by the loss or gain of a single hydrogen ion.
- conjugate base: The particle produced when an acid donates a proton.
- hydronium ion (H_3O^+) : An ion consisting of a water molecule with an attached hydrogen ion.
- Lewis acid: A substance that accepts a pair of electrons to form a covalent bond.
- Lewis base: A substance that donates a pair of electrons to form a covalent bond.
- monoprotic acid: An acid that contains one ionizable hydrogen.
- polyprotic acid: An acid that contains multiple ionizable hydrogens.

Teaching Strategies

Introducing the Lesson

Students generally have some knowledge about acids and bases since many products encountered around the home are either acids or bases. Start with an open-ended discussion to see what the students know. Bring in some products such as fruits, fruit juices, vinegar, sour candy, ammonia, bleach, and soaps as examples of household items that are acids or bases.

Demonstration

When discussing the concept of electrolytes, demonstrate the conductivity of various solutions using a simple apparatus consisting of two electrodes connected to a light bulb. Use distilled water to show how it does not conduct a current. Compare to solutions of a strong acid (HCl), a weak acid (CH₃COOH), a strong base (NaOH), and a weak base (NH₃). Emphasize that various substances can be nonelectrolytes, weak electrolytes, or strong electrolytes.

Common Misconceptions

Students frequently get confused about the presence of three distinct definitions of acids and bases: Arrhenius, Brønsted-Lowry, and Lewis. Emphasize that each of the three definitions is simply an expansion. All Arrhenius acids and bases also qualify under the other two definitions. Brønsted-Lowry acids and bases also qualify under the Lewis definition, but may include some substances (such as NH₃) which would not qualify by the Arrhenius definition. There are no substances which would be classified as an acid under one definition, but as a base under a different definition. The following Venn diagram may help cement the concept.



Science Inquiry

The link below is an introductory lab activity where students investigate the acid-base properties of some common household substances such as orange juice and vinegar. Students will use indicators such as litmus paper and red cabbage juice to test the substances, comparing their results with readings from a pH meter.

http://www.nclark.net/acidbase_lab.doc

The link below is for a similar activity, focusing on the preparation and use of red cabbage juice as an acid-base indicator. Students will enjoy the preparation of the cabbage juice.

http://web.stanford.edu/~ajspakow/downloads/outreach/ph-student-9-30-09.pdf

Differentiated Instruction

The link below is for a computer-based simulation that illustrates some of the primary properties of acids and bases, including pH and electrical conductivity.

https://phet.colorado.edu/en/simulation/acid-base-solutions

There are several companion activity choices on this page that can be used as a handout for an introductory lesson.

Enrichment

Universal indicator is an indicator which turns all of the different colors of the rainbow; from red (very low pH) to violet (very high pH). Have students research materials with various pH values so that each color of the rainbow is represented. They can practice their technique of adding small amounts of the universal indicator to each substance and finish off with a classroom demonstration.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 21.1 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 21.1 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

The relative acidity or basicity of an aqueous solution is measured using a scale called the pH scale.

- How are the hydrogen and hydroxide ion concentrations related in all aqueous solutions?
- How is the pH of a solution calculated?

21.2 The pH Concept

Key Concept

Water is a very weak electrolyte, undergoing self-ionization into a hydronium ion and a hydroxide ion. The equilibrium constant for the self-ionization process is called the ion product of water (K_w) and has a value of 1.0×10^{-14} at 25°C. Consequently, the concentration of hydrogen ions and hydroxide ions in any aqueous solution are inversely proportional. The pH of a solution is equal to the negative logarithm of the hydrogen ion concentration. A pH of 7 represents a neutral solution, while a pH of less than 7 is acidic and a pH of greater than 7 is basic. The pOH of a solution is defined as the negative logarithm of the hydroxide ion concentration. The sum of the pH and the pOH of any solution is 14. The pH and pOH of any strong acid or strong base solution can be calculated if the concentration is known or vice-versa.

Standards

Lesson Objectives

- Describe the equilibrium process of water ionizing into hydrogen and hydroxide ions, and explain the ionproduct constant.
- Calculate the concentration of either the hydrogen ion or the hydroxide ion in water when the other value is known.
- Calculate the pH of an aqueous solution of an acid or a base.
- Calculate the hydrogen-ion concentration of a solution with a known pH.
- Calculate the pOH of acids and bases and know the relationships between pH, pOH, hydrogen-ion concentration, and hydroxide-ion concentration.

Lesson Vocabulary

- acidic solution: A solution in which the concentration of hydrogen ions is greater than the concentration of hydroxide ions.
- basic solution: A solution in which the concentration of hydroxide ions is greater than the concentration of hydrogen ions.
- ion-product of water (K_w): The mathematical product of the concentration of hydrogen ions and hydroxide ions in aqueous solution.
- pH: The negative logarithm of the hydrogen-ion concentration in a solution.
- pOH: The negative logarithm of the hydroxide-ion concentration in a solution.
- self-ionization: The process in which water ionizes to hydronium ions and hydroxide ions.

Teaching Strategies

Introducing the Lesson

The concepts of pH and the pH scale are likely known to students and/or were touched on in the previous lesson. Tell the students that the emphasis in this lesson is on the mathematics of pH and how the pH of a solution can be determined from its concentration. Discuss how the pH scale is what is called a logarithmic scale. Ask them if they are familiar with the Richter scale and what it measures. Compare the two scales in this way. An earthquake that is a 7 on the Richter scale is 10 times as damaging as an earthquake that measures a 6 on the scale. Similarly, an acid with a pH of 4 is 10 times as acidic as one with a pH of 5.

Common Misconceptions

Using both the hydronium ion (H_33O^+) and the hydrogen ion (H^+) to represent an acid can be problematic to some students. Stress that the H+ ion does not actually exist in an aqueous solution. Because of its positive charge, it will always be attracted to, and ultimately bond with one of the lone pairs of electrons on the oxygen atom of a water molecule. However, chemists frequently use H⁺ in chemical equations involving acids because it is more convenient. Some may wonder why H_4O^{2+} doesn't exist. That is because another H⁺ ion will not be attracted to the positively charged hydronium ion. They are both positive and would repel each other.

Students may have trouble understanding how a logarithmic scale works. Some will conclude that if an acid of pH = 4 is 10 times as acidic as one with a pH = 5, then it will be 20 times as acidic as one with a pH = 6. Stress that each unit on the scale is a power of 10 and $10^2 = 100$. This means the solution with a pH of 4 is 100 times as acidic as one with a pH of 6. This is worth making clear before teaching the rigorous math of calculating pH.

It may be confusing to some students that a <u>lower</u> number on the pH scale is more acidic. Explain how this is because of the negative sign involved when calculating the pH. That is: $pH = -\log[H^+]$. Without the presence of the negative sign, the entire scale would consist of negative numbers which would be confusing.

Science Inquiry

The simulation below works well as a visual introduction to the intricacies of pH. Solutions of various well known substances can be "tested" with the simulation. The effect of concentration on pH can be seen as well as the relative amounts of hydronium ion and hydroxide ion in each solution. Companion lesson plans to use with the simulation are available to download on the same page.

https://phet.colorado.edu/en/simulation/ph-scale

Differentiated Instruction

Have students of lower math ability work with higher ability students to master the various pH calculations. Have them work through multiple examples of each type of calculation:

- 1. given the H^+ ion concentration \rightarrow calculate the concentration of OH^- (and vice-versa)
- 2. given the concentration of an acid \rightarrow calculate the pH
- 3. given the concentration of a base \rightarrow calculate the pH (2 steps)
- 4. given the pH of an acid \rightarrow calculate the concentrations of H⁺ and of OH⁻
- 5. given the pH of a base \rightarrow calculate the concentrations of H⁺ and of OH⁻

Make copies of the flowchart at the end of the lesson as that will help students visualize how to solve each type of problem, especially the ones that require two steps.

Enrichment

The pH of natural bodies of water such as ponds, lakes, and rivers is very important. Have interested students research how acid rain has lowered the pH of bodies of water and the harmful effects on plants and animals that have been the result. Have them also discover how scientists combat this problem by treating the water with various chemicals that can raise the pH back to normal levels.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 21.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 21.2 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Acids and bases can be characterized as either strong or weak, based on the extent to which they ionize in water.

- What are acid and base ionization constants?
- How is the pH of a solution containing a weak acid or a weak base calculated?

21.3 Acid and Base Strength

Key Concept

Strong acids and bases ionize completely in aqueous solution, while weak acids and bases ionize only to a small extent. The stronger a particular acid, the weaker is its conjugate base. The acid ionization constant (K_a) is the equilibrium constant for the ionization of a weak acid. The base ionization constant (K_b) is the equilibrium constant for the reaction in which a base removes a hydrogen ion from a water molecule, yielding hydroxide ion as a product. Stronger acids have larger K_a values, while stronger bases have larger K_b values. The value of K_a or K_b for a particular acid or base can be calculated if the concentration of the solution and the pH are known. Tables of known K_a and K_b constants can be used to determine the pH of an acid or base solution of a given concentration.

Standards

Lesson Objectives

- Explain the difference between a strong acid or base and a weak acid or base.
- Write equilibrium expressions for the ionizations of weak acids and weak bases. Explain how the value of K_a or K_b relates to the strength of the acid or base.
- Calculate an unknown K_a or K_b from the solution concentration and the pH.
- Calculate the expected pH of a solution containing a given concentration of a weak acid or base that has a known K_a or K_b value.

Lesson Vocabulary

- acid ionization constant: The equilibrium constant for the ionization of an acid.
- base ionization constant: The equilibrium constant for the ionization of an base.
- strong acid: An acid which ionizes completely in aqueous solution.
- strong base: A base which ionizes completely in aqueous solution.
- weak acid: An acid that ionizes only slightly in aqueous solution.
- weak base: A base that ionizes only slightly in aqueous solution.

Teaching Strategies

Introducing the Lesson

Begin by reviewing how strong acids are 100% ionized in solution. For this reason, the concentration of H^+ in a particular solution of a strong acid is equal to the concentration of the acid. This makes the pH of the solution easy to calculate. Write an equation on the board, emphasizing the "one-way" arrow.

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Now use a pH meter to compare the pH values of two acids that are the same molarity. A solution of 0.10 M HCl will have a pH of 1.0, while a solution of 0.10 M CH₃COOH will have a pH of 2.9. Ask the class why there is a difference and the acetic acid is nearly 100 times less acidic. When someone has realized that it is because the acid is weak and thus not fully ionized, write the ionization equation on the board, emphasizing the double arrow of equilibrium.

$$CH_3COOH(aq) \rightarrow H^+(aq) + CH_3COO^-(aq)$$

Because fewer H^+ ions are produced by the acid, the pH is higher.

Common Misconceptions

Students may expect that polyprotic acids ionize all at once, with one ionization constant. Stress that the ionization of a polyprotic acid such as $H_2C_2O_4$ (oxalic acid) occurs in two distinct steps. The K_a value for the first ionization is always significantly higher than the K_a value for the second ionization (and likewise for any subsequent ionizations). This is because the loss of a positively charged hydrogen ion from a negative ion ($HC_2O_4^-$), happens to a lesser extent. For the purpose of calculating the pH of the solution of a polyprotic acid, only the first ionization needs to be considered since the vast majority of the hydrogen ions are produced in the first ionization step.

Science Inquiry

The link below is for a lab activity which requires students to test the pH of a variety of solutions: strong acids, weak acids, strong bases, weak bases, and some salt solutions. The hydrolysis of salts is covered in the last lesson in this chapter, so that part of the lab could be either cut out or delayed until later.

http://www.smccd.edu/accounts/batesa/chem220/lab/labmanual-sum12/7-SASBWAWB-SUM12.pdf

Differentiated Instruction

Have students prepare a series of index cards that describes each step that needs to be taken for the calculations in this lesson. There are four different calculations:

- 1. given the concentration of an acid and the pH \rightarrow calculate the K_a
- 2. given the concentration of a base and the pH \rightarrow calculate the K_b
- 3. given the concentration of an acid solution and the $K_a \rightarrow$ calculate the pH
- 4. given the concentration of a base solution and the $K_b \rightarrow$ calculate the pH

Enrichment

Mathematically adept students can try using the quadratic equation to solve for "x" in the ICE problems where the pH of a solution needs to be calculated. The approximation discussed in the text does not work well for acids or bases that have relatively high K values or when the concentration of the acid or base is very small. Consider the example below.

Find the pH of a 0.05 M solution of oxalic acid. Only the first ionization constant needs to be considered and the K_a value is 0.065. The equation from the ICE tale becomes:

$$K_a = \frac{[H^+][HC_2O_4^-]}{[H_2C_2O_4]} = 0.065 = \frac{x^2}{0.05 - x}$$

Use the quadratic equation to solve for x (solution is that x = 0.0331). The pH of the solution is then 1.48.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 21.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 21.3 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

TABLE 21.3:

	HClO	H^+	ClO ⁻
I	0.250	0	0
С	-8.5×10^{-5}	$+8.5 \times 10^{-5}$	$+8.5 imes 10^{-5}$
Е	$0.250 - 8.5 \times 10^{-5}$	8.5×10^{-5}	8.5×10^{-5}

TABLE 21.4:

	HCN	H^+	CN^{-}
Ι	0.20	0	0
С	-x	+x	+x
Е	0.20 - x	x	x

TABLE 21.5:

	HA	H^+	A^-
Ι	1.25	0	0
С	-1.9×10^{-5}	$+1.9 \times 10^{-5}$	$+1.9 \times 10^{-5}$
Ε	$1.25 - 1.9 imes 10^{-5}$	1.9×10^{-5}	1.9×10^{-5}

TABLE 21.6:

	В	HB^+	OH^-
Ι	0.350	0	0
С	-1.66×10^{-3}	$+1.66 \times 10^{-3}$	$+1.66 \times 10^{-3}$
Е	$0.350 - 1.66 \times 10^{-3}$	1.66×10^{-3}	1.66×10^{-3}

TABLE 21.7:

	C_6H_5COOH	H^+	$C_6H_5COO^-$
Ι	2.40	0	0
C	-x	+x	+x
Е	2.40 - x	x	x

TABLE 21.8:

	C_5H_5N	$C_5H_5NH^+$	OH^-
Ι	0.745	0	0
C	-x	+x	+x
Е	0.745 - x	x	x

TABLE 21.9:

	HF	H^+	F^{-}
Ι	Μ	0	0
С	-0.0100	+0.0100	+0.0100
Е	M - 0.0100	0.0100	0.0100

Points to Consider

Acids and bases react with each other in a reaction called a neutralization.

- What are the products of a neutralization reaction?
- How can a neutralization reaction be used to determine the unknown concentration of an acid or base in an aqueous solution?

21.4 Acid-Base Neutralization

Key Concept

A neutralization reaction is one in which equal moles of acid and base are combined, producing a salt and water. The net ionic equation for the neutralization of a strong acid by a strong base is always the same: hydrogen ion combing with hydroxide ion to form water. Neutralizations also take place when one or both of the solutions is weak. A titration is an experiment where the unknown concentration of an acid or base is determined in a controlled neutralization reaction where the concentration of the other reactant is known. A titration curve is a graphical representation, showing how the pH of the solution changes during the course of a titration. An indicator is a substance that is used to see the end point of a titration because the indicator displays a different color in an acidic or basic solution.

Standards

Lesson Objectives

- Write balanced molecular and net ionic equations for acid-base neutralization reactions.
- Know the steps of an acid-base titration experiment.
- Calculate the concentration of an acid or a base from data obtained in a titration.
- Describe the appearance of a titration curve.
- Explain the chemistry of acid-base indicators.

Lesson Vocabulary

- end point: The point during a titration experiment when the indicator changes color.
- equivalence point: The point in a neutralization reaction where the number of moles of hydrogen ions is equal to the number of moles of hydroxide ions.
- indicator: A substance that has a distinctly different color when in an acidic or basic solution.
- neutralization reaction: A reaction in which an acid and a base react in an aqueous solution to produce a salt and water.
- salt: An ionic compound composed of a cation from a base and an anion from an acid.
- standard solution: The solution in a titration whose concentration is known.
- titration: An experiment where a volume of a solution of known concentration is added to a volume of another solution in order to determine its concentration.
- titration curve: A graphical representation of the pH of a solution during a titration.

Teaching Strategies

Introducing the Lesson

Students should remember double replacement reactions from an earlier chapter. Write the reactants HCl and NaOH on the board and ask them to predict the products. Then, review how to turn the molecular equation into a net ionic equation. The end result is the reaction that occurs when an acid (HCl) is neutralized by a base (NaOH). You may wish to demonstrate the reaction by using phenolphthalein as an indicator. As the NaOH solution is slowly added to the HCl solution, the color of the solution changes from colorless to pink. This tells us that the acid has been neutralized.

Common Misconceptions

Once students have learned the $M_A V_A = M_B V_B$ equation, they may want to use it for problems for which it will not work. When the number of H⁺ ions in the acid is not equal to the number of OH⁻ ions in the base, the problem must be solved by accounting for the stoichiometry of the balanced equation. See Sample Problem 21.6 in the text.

Even though an acid may be weak, it is still fully neutralized by reaction with a strong base. The reaction below is an example: the weak acid HCN is neutralized by NaOH, a strong base.

molecular: $HCN(aq) + NaOH(aq) \rightarrow NaCN(aq) + H_2O(l)$

net ionic: $HCN(aq) + OH^{-}(aq) \rightarrow CN^{-}(aq) + H_2O(l)$

Hydrocyanic acid by itself is weak and not fully ionized. However, the strong hydroxide ion completely removes all of the hydrogen ions from the HCN. A numerical problem is solved the same way as for a strong acid –strong base titration.

Science Inquiry

Titration is an essential experiment that all chemistry students should be able to do. The easiest type of titration involves a strong acid with a strong base, using phenolphthalein as an indicator. The link below describes the procedure for a simple strong acid –strong base titration. It is best if the students are already familiar with neutralization reactions and have done some practice titration calculations.

https://www.apsu.edu/files/chemistry/CHEM_1011_Titration_of_an_acid_with_a_base_lab.pdf

Differentiated Instruction

Students may have difficulty writing accurate net ionic equations when one of the components of the neutralization reaction is weak. Provide numerous practice problems of weak acid –strong base neutralizations where the acid is monoprotic. Have students compare all of the net ionic equations to see how similar they are. Gradually move up to equations with a polyprotic weak acid such as H_2CO_3 . The same technique can be used for weak base –strong acid neutralizations. However, the only weak bases should be NH₃ or its related compounds such as methylamine, CH_3NH_2 .

Enrichment

Have advanced students attempt to construct a titration curve from calculations of the pH of a solution at various points during a titration. As an example, suppose that 20.0 mL of 0.500 M HCl were to be titrated with 0.500 M NaOH. What would be the pH after 5.00 mL of NaOH had been added?
Calculate the moles of both acid and base present at that point.

 $HCl: 0.500 M \times 0.0200 L = 0.0100 mol HCl$

 $NaOH: 0.500 M \times 0.00500 L = 0.002500 mol NaOH$

The HCl would neutralize all of the NaOH since the HCl is present in a greater amount. The moles of HCl left over can be determined by subtraction: 0.0100 - 0.002500 = 0.007500 mol HCl.

The total solution volume is now 25.00 mL, so the molarity of the HCl has become:

 $\frac{0.007500 \ mol}{0.02500 \ L} = 0.300 \ M \ HCl$

Since the HCl is a strong acid, the $[H^+] = 0.300 M$. This makes the pH of the solution easy to calculate:

 $pH = -\log(0.300) = 0.52$

By repeating this calculation at regular intervals, a titration curve can be created. Intervals of about 2 mL would work well, though more calculations should be done near the equivalence point to get a more accurate curve.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 21.4 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 21.4 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Salt solutions can be acidic, basic, or neutral.

- How can the acidity or basicity of a salt solution be determined?
- What is a hydrolysis reaction, and how does it affect the pH?

21.5 Salt Solutions

Key Concept

Hydrolysis of a salt solution occurs when either the cation or the anion (or both) from the salt reacts with water, producing either hydronium ions or hydroxide ions. Salts that are derived from the neutralization of a strong acid by a weak base will yield acidic solutions. Salts that are derived from the neutralization of a weak acid by a strong base will yield basic solutions. Salts derived from the neutralization of a strong acid by a strong base will yield basic solutions. Salts derived from the neutralization of a strong acid by a strong base do not undergo hydrolysis and produce a neutral solution. The pH of a solution resulting from salt hydrolysis can be determined if the K_a or K_b of the hydrolyzing ion is known. Buffers are solutions of a weak acid or base combined with a salt of its conjugate. Buffers are capable of resisting changes in pH brought about by the addition of external acid or base.

Standards

Lesson Objectives

- Predict whether a certain salt solution is acidic, basic, or neutral.
- Write balanced equations for hydrolysis reactions.
- Calculate the pH of a salt solution when the K_a or the K_b of the hydrolyzing ion is known.
- Use equations to show how buffers resist changes in pH by using up additional hydrogen or hydroxide ions.

Lesson Vocabulary

- buffer: A solution composed of a weak acid or a base and its salt. A buffer resists changes in pH when either an acid or a base is added to it.
- buffer capacity: The amount of acid or base that can be added to a buffer solution before a large change in pH occurs.
- salt hydrolysis: A reaction in which one of the ions from a salt reacts with water, forming either an acidic or basic solution.

Teaching Strategies

Introducing the Lesson

Show the class three samples of salts: sodium chloride, ammonium chloride, and sodium carbonate. All are similar looking white solids. Ask them what they think the pH of the resulting solutions will be when each of these salts is dissolved in water. They will likely think that the solutions will be neutral with a pH of near 7. Dissolve each of

the solids in distilled water and add a few drops of universal indicator. The sodium chloride solution will indeed be neutral. However, the ammonium chloride solution will be acidic, while the sodium carbonate solution will be basic. Why is this true? Remind the students that a salt can be thought of as the product of an acid-base neutralization reaction and that acids and bases can be either strong or weak. Show the equations for the acid-base neutralizations that produce each of the salts.

> $HCl + NaOH \rightarrow NaCl + H_2O$ strong acid + strong base $HCl + NH_4OH \rightarrow NH_4Cl + H_2O$ strong acid + weak base $H_2CO_3 + 2NaOH \rightarrow Na_2CO_3 + 2H_2O$ weak acid + strong base

This leads to a discussion of how different salt solutions can be either neutral, acidic, or basic.

Common Misconceptions

Students can get confused about whether hydrolysis occurs for a salt solution. Salts that hydrolyze must have either the cation or the anion be the conjugate of a weak acid or base. For example, KCl does not hydrolyze because the K^+ ion is an alkali metal ion and the Cl⁻ ion is the conjugate base of a strong acid (HCl). However, KCN does hydrolyze because the CN⁻ ion is the conjugate base of a weak acid (HCN). The cyanide ion will react with water, accepting a proton and producing hydroxide ion. The result is a basic solution. Salts that hydrolyze to produce acidic solutions typically have ammonium as the cation or cations derived from ammonium. The ammonium ion hydrolyzes and produces an acidic solution because ammonia is a weak base. The ammonium ion donates a proton to water, producing the hydronium ion.

Science Inquiry

The link below is for a very rigorous lab in which students investigate the properties of various salt solutions and conclude by making a buffer solution. It would work best as a concluding activity for the lesson.

http://www.academia.edu/4302677/Hydrolysis_of_salts

Differentiated Instruction

The following video helps explain the concept of slat hydrolysis. Have students view the video and write a short summary of what they have learned.

http://www.flinnsci.com/teacher-resources/teacher-resource-videos/best-practices-for-teaching-chemistry/acids-and-bases/hydrolysis-of-salts/

Enrichment

The buffers that occur in blood are critical to maintaining a constant pH. Students can research the buffering chemical reactions that occur in blood. Prepare a poster illustrating their findings.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 21.5 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 21.5 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

TABLE 21.10:

	CH ₃ COO ⁻	СН ₃ СООН	OH^-
Ι	0.200	0	0
С	-x	+x	+x
E	0.200 - x	x	x

TABLE 21.11:

	CO_{3}^{2-}	HCO_3^-	OH^-
Ι	0.400	0	0
С	- <i>x</i>	+x	+x
Е	0.400 - x	x	x

TABLE 21.12:

	$CH_3NH_3^+$	H_3O^+	CH ₃ NH ₃
Ι	1.50	0	0
С	-x	+x	+x
Е	1.50 - x	x	x

TABLE 21.13:

	<i>B</i> ⁻	НВ	OH^-
Ι	1.00	0	0
С	$-6.46 imes 10^{-5}$	$+6.46 \times 10^{-5}$	$+6.46 \times 10^{-5}$
Е	$1.00 - 6.46 \times 10^{-5}$	6.46×10^{-5}	6.46×10^{-5}

TABLE 21.14:

	ClO ⁻	HClO	OH^-
Ι	0.81	0	0
С	- <i>x</i>	+x	+x
E	0.81 - x	x	x

Points to Consider

Acid-base reactions involve the transfer of a hydrogen ion from one substance to another. An oxidation-reduction reaction is another common type of reaction that involves the transfer of one or more electrons from one substance to another.

- How can an oxidation-reduction reaction be recognized?
- What is corrosion, and how does it relate to a transfer of electrons?



Oxidation-Reduction Reactions

Chapter Outline

- 22.1 THE NATURE OF OXIDATION AND REDUCTION
- 22.2 OXIDATION NUMBERS
- 22.3 BALANCING REDOX REACTIONS

Chapter Overview

Oxidation-reduction (redox) reactions are a class of chemical reactions in which electrons are transferred from one substance to another. The substance that loses electrons is said to be oxidized, while the substance that gains electrons is reduced. Oxidation numbers can be assigned to every atom in a chemical reaction and allow one to track how electrons are being transferred. An oxidation number increase means that atom is being oxidized, while an oxidation number decrease means that atom is being reduced. Special techniques are employed to balance redox equations, including the oxidation-number change method and the half-reaction method.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 22.1	Class Period(s) (60 min)
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Lesson	Class Period(s) (60 min)
22.1 The Nature of Oxidation and Reduction	2
22.2 Oxidation Numbers	1-2
22.3 Balancing Redox Reactions	1-2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 22.2:	Chemistry	Concepts
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Flexbook Lesson	Chemistry Concepts
22.1 The Nature of Oxidation and Reduction	22.1 – 22.5

TABLE 22.2:	(continued)
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Flexbook Lesson	Chemistry Concepts
22.2 Oxidation Numbers	22.6 - 22.8
22.3 Balancing Redox Reactions	22.9 - 22.11

Keywords

- corrosion
- oxidation
- oxidation number
- redox
- reduction

22.1 The Nature of Oxidation and Reduction

Key Concept

The chemical process of oxidation was originally defined as the addition of oxygen to a substance during a reaction. Reduction was defined as the removal of oxygen. The modern definition of oxidation is the loss of electrons, while reduction is the gain of electrons. Both processes occur simultaneously in an oxidation-reduction or redox reaction. Redox reactions may involve complete transfers of electrons as in ionic reactions, or the partial transfer of electrons as in molecular reactions. An oxidizing agent is a substance which causes oxidation, while a reducing agent is a substance that causes reduction. Corrosion is the deterioration of metals by a redox process.

Standards

Lesson Objectives

- Define oxidation and reduction in terms of a gain or loss of oxygen atoms.
- Define oxidation and reduction in terms of a gain or loss of electrons.
- Describe oxidation and reduction as it occurs in both ionic and molecular reactions.
- Identify the oxidizing and reducing agents in a redox reaction.
- Describe corrosion and some of the ways to prevent it or slow it down.

Lesson Vocabulary

- corrosion: The deterioration of metals by redox processes.
- half-reaction: An equation that shows either the oxidation or the reduction reaction that occurs during a redox reaction.
- oxidation: The full or partial loss of electrons, or the gain of oxygen.
- oxidation-reduction (redox) reaction: A reaction that involves the full or partial transfer of electrons from one reactant to another.
- oxidizing agent: A substance that causes oxidation by accepting electrons.
- reducing agent: A substance that causes reduction by losing electrons.
- reduction: The full or partial gain of electrons or the loss of oxygen.

Teaching Strategies

Introducing the Lesson

Demonstrate the reaction between magnesium and oxygen by inserting a small piece of magnesium ribbon into the flame of a Bunsen burner. The reaction produces a grayish powder as the magnesium is consumed in the flame. Ask students what has happened. The magnesium has undergone a reaction with oxygen. What is the product? Write the equation on the board.

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

Ask what type of compound is magnesium oxide. It is an ionic compound and so can be written as a combination of magnesium ions and oxide ions. So in the course of the reaction, what happened to the magnesium metal? It lost two electrons to become magnesium ions. What happened to the oxygen? Each oxygen atom in the oxygen molecule gained two electrons to become oxide ions. So as a whole, the oxygen molecule gained four electrons. We have arrived at a new kind of chemical reaction.

$$Mg \rightarrow Mg^{2+} + 2e^{-}$$

 $O_2 + 4e^{-} \rightarrow 2O^{2-}$

Common Misconceptions

Students may have some difficulty grasping how a reaction involving only molecular compounds can be considered a redox reaction since no ions are formed. Stress that it is the density of electrons around the atoms in a molecule that changes during the course of this kind of reaction. When oxidation numbers are introduced in the next lesson, these reactions will be easier to recognize as redox reactions.

Science Inquiry

The link below is an activity that is designed to be an introduction to redox chemistry. The reaction involves the oxidation of aluminum foil pieces by a copper(II) chloride solution.

https://www.google.com/url?q=http://newyorkscienceteacher.com/sci/files/user-submitted/Green_Chunks_and_Foil .doc&sa=U&ei=POskVa7FBYrtsAW2zYCgDA&ved=0CAMQFjAA&client=internal-uds-cse&usg=AFQjCNEF4d2tao tqj-auBz9gJ87-6Zup9Q

Differentiated Instruction

Provide students with butcher paper in order to diagram the electron transfer that occurs during redox reactions. Provide example reactions including combination reactions of ionic compounds, decomposition reactions, and single replacement reactions. Students can draw out the reactions using Lewis electron dot symbols and drawing arrows to show the transfers of electrons.

Enrichment

Students can make a PowerPoint presentation on all types of corrosion processes and the lengths to which humans go to combat corrosion. Different groups can be assigned different metals such as iron, copper, or aluminum.

Presentations can be made to the rest of the class.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 22.1 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 22.1 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

An oxidation number is a number that is assigned to an atom that indicates its degree of oxidation.

- How are oxidation numbers determined?
- How can oxidation numbers assist in determining if a given reaction is a redox reaction?

22.2 Oxidation Numbers

Key Concept

An oxidation number is a positive or negative number that is assigned to an atom to indicate its degree of oxidation or reduction. Specific rules are followed to assign oxidation numbers to each of the elements in a compound. The sum of all the oxidation numbers in a neutral compound is zero, while the sum of all the oxidation numbers for a polyatomic ion is equal to the charge of the ion. During a redox reaction, the oxidation number of the element that is oxidized increases, while the oxidation number of the element that is reduced decreases. Reactions in which no oxidation numbers change are not redox reactions.

Standards

Lesson Objectives

- Use the oxidation number rules to determine the oxidation number of an atom of any element in a pure substance.
- Define oxidation and reduction in terms of a change in oxidation number.
- Identify which atoms are being oxidized and which atoms are being reduced in a redox reaction.
- Know which types of reactions are usually redox reactions and which types are usually not. Use oxidation number changes to decide whether a given reaction involves oxidation and reduction.

Lesson Vocabulary

• oxidation number: A positive or negative number that is assigned to an atom to indicate its degree of oxidation or reduction.

Teaching Strategies

Introducing the Lesson

Students have long been familiar with the common charges of ions: 1+ for alkali metals, 1– for halogens, etc. Introduce oxidation numbers as a way to take this idea to atoms that are in any type of compound, that is ionic or molecular. Use water as an example. Draw the structure on the board. Ask which element is more electronegative. Remind the students that when they studied hydrogen bonding, they placed partial charges on the hydrogen (δ^+) and oxygen (δ^-) atoms. The concept of oxidation numbers now takes that and extends it to a complete electron transfer. Hydrogen's oxidation number is +1, while oxygen's oxidation number is -2. Go from there to the rules.

Common Misconceptions

Students generally do well with oxidation numbers. However, remind them that for ionic compounds, especially those with a polyatomic anion, the oxidation number of the metal ion is simply its charge. Students may look at a formula like $CuCO_3$ and think that they cannot figure out the oxidation numbers because there is no rule for either copper or carbon. However, since it is an ionic compound, the charge of the copper ion is 2+. So the oxidation number of carbon +4.

When they are analyzing chemical equations for oxidation number changes, remind students that they should not pay any attention to the coefficients of the balanced equation. They have no effect on oxidation numbers.

Science Inquiry

There are many available practice worksheets online to allow students to master assigning oxidation numbers. The link below is one which also includes solutions and explanations.

http://gsereda.wikis.hsd.ca/file/detail/07+Finding+Oxidation+Numbers+Worksheet.doc

Differentiated Instruction

Students can make creative and colorful posters out of the oxidation number rules. For each rule, have them provide specific examples, adding in drawings or pictures of the compounds involved. Display the best posters in the room.

Enrichment

Carbon is the most versatile element of the periodic table in that carbon makes more different compounds than any element. One main reason is that carbon can have oxidation numbers from -4 to +4 in its compounds. Have students try to find examples of carbon compounds in which the carbon atoms have each of the possible oxidation numbers.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 22.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 22.2 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Balancing equations can often be done by inspection, simply counting atoms. However, many redox reactions are quite complex, and a more systematic approach can be helpful.

- How can oxidation numbers be used to help balance redox equations?
- How can redox reactions be broken into half-reactions to make balancing them easier?

22.3 Balancing Redox Reactions

Key Concept

Balancing some redox reactions is often too difficult to do simply by inspection, so specific techniques are described. In the oxidation number change method, the increase in oxidation number by one atom is compared to the decrease in oxidation number by another atom. The half-reaction method of balancing redox reactions works better for reactions occurring in aqueous solution. In this method, separate half-reactions are written and balanced, including the addition of water and either hydrogen ions or hydroxide ions. The number of electrons lost is made equal to the number of electrons gained and the reactions are added together.

Standards

Lesson Objectives

- Balance a redox equation using the oxidation-number-change method.
- Balance a redox equation by breaking it down into oxidation and reduction half-reactions and using the half-reaction method.

Teaching Strategies

Introducing the Lesson

Students have likely mastered balancing equations by inspection and may be surprised that there are equations that are too difficult to balance in that way. Put an example of an unbalanced redox reaction on the board and give students a minute or so to try to balance by inspection.

$$Cr_2O_7^{2-} + C_2O_4^{2-} \to Cr^{3+} + CO_2$$

It is highly unlikely that anyone will succeed. Then, start to introduce the specific techniques for balancing redox reactions.

Common Misconceptions

Students should make sure to follow the steps for balancing very carefully. One small error in an early step of the process will result in an incorrectly balanced equation. They should make sure to check the balancing when they are finished, making sure to check both the number of atoms and the total charge. If the equation is incorrectly balanced, it will be easiest to completely start the process over again.

Science Inquiry

There are many good tutorial videos for balancing redox reactions. The links below describe the half-reaction method in both acidic solution and basic solution.



MEDIA Click image to the left or use the URL below.

https://www.youtube.com/watch?v=EtOrcKw6wso





https://www.youtube.com/watch?v=imqYliG5BCU

Differentiated Instruction

Pair students together to have them practice balancing by the half-reaction method. Have each student do one step at a time, comparing their work with each other as they go. This may reduce frustration and careless errors by trying to go too fast.

Enrichment

There are many practice problems for balancing redox reactions available online. An example is linked below.

http://pages.towson.edu/ladon/redoxprac.htm

Some research may provide students with some extremely challenging examples that diligent students may enjoy.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 22.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 22.3 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

A battery is a device that uses a redox reaction to turn chemical energy into electrical energy.

- Why can't a redox reaction directly generate electricity?
- What is an electrochemical cell?



Electrochemistry

Chapter Outline

- 23.1 ELECTROCHEMICAL CELLS
- 23.2 CELL POTENTIALS
- 23.3 ELECTROLYSIS

Chapter Overview

An electrochemical cell is a device that converts chemical energy to electrical energy or vice-versa. In a voltaic cell, a spontaneous redox reaction produces an electrical current at a certain voltage. The cell potential for any half-cell is determined by measuring the potential versus the standard hydrogen electrode. The result is a listing of reduction potentials for all possible half-reactions. The overall cell potential for any combination of half-cells can be calculated. The voltage must be positive for the reaction to be spontaneous. Electrolysis is the process of using a direct electric current to force a nonspontaneous redox reaction to occur. Electrolysis is used in the process of electroplating of metals for decorative purposes.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 23.1	: Class Period	(s) (60 min)
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Lesson	Class Period(s) (60 min)
23.1 Electrochemical Cells	1-2
23.2 Cell Potentials	2
23.3 Electrolysis	1-2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

T/	B	LE	2	23	.2:	Chemistry	Concepts
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Flexbook Lesson	Chemistry Concepts
23.1 Electrochemical Cells	23.1 - 23.3, 23.7

Flexbook Lesson	Chemistry Concepts
23.2 Cell Potentials	23.4 - 23.6
23.3 Electrolysis	23.8 - 23.11

TABLE 23.2: (continued)

Keywords

- cell potential
- electrochemical cell
- electrochemistry
- electrolysis
- voltaic cells

23.1 Electrochemical Cells

Key Concept

Electrochemistry is the branch of chemistry concerned with the interconversion of chemical energy and electrical energy. In a voltaic cell, a spontaneous redox reaction is used to generate an electrical current. The voltaic cell consists of two half-cells in which a metal electrode is immersed into a solution of its own cation. The two half-cells are connected by an external wire and by a salt bridge. Oxidation occurs at the anode, while reduction occurs at the cathode. Electrons move through the external wire, while ions move through the salt bridge. Dry cells, lead storage batteries, and hydrogen fuel cells are examples of commonly used electrochemical cells.

Standards

Lesson Objectives

- Use the activity series to identify elements that are more easily oxidized than others, and write oxidation and reduction half-reactions.
- Describe the parts of a voltaic cell and explain how redox reactions are used to generate an electric current.
- Describe the general features of a dry cell, a lead storage battery, and a fuel cell.

Lesson Vocabulary

- anode: The electrode at which oxidation occurs.
- battery: A group of electrochemical cells combined together as a source of direct electric current at a constant voltage.
- cathode: The electrode at which reduction occurs.
- electrochemical cell: Any device that converts chemical energy into electrical energy or electrical energy into chemical energy.
- electrochemistry: A branch of chemistry that deals with the interconversion of chemical energy and electrical energy
- electrode: A conductor in a circuit that is used to carry electrons to a nonmetallic part of the circuit.
- fuel cell: An electrochemical cell that requires a continuous supply of reactants to keep functioning.
- half-cell: One part of a voltaic cell in which either the oxidation or reduction half-reaction takes place.
- salt bridge: A tube containing an inert electrolyte that allows the passage of ions between the two half-cells.
- voltaic cell: An electrochemical cell that uses a spontaneous redox reaction to produce electrical energy.

Teaching Strategies

Introducing the Lesson

Be sure to be ready to construct a zinc-copper electrochemical cell during class. Discuss how the zinc is more easily oxidized than the copper because it is higher on the activity series. Place the electrodes into the corresponding solutions and connect to a voltmeter. The reading will not show a voltage. Ask the students why it does not. Someone will realize that the circuit is not complete. Add in the salt bridge, which can be a paper towel soaked in a concentrated solution of KCl. The circuit is complete and the voltmeter will register a voltage.

Common Misconceptions

Students may get the direct redox process confused with the electrochemical cell. In the direct process, a strip of metal is placed into a solution of another ion (one above it on the activity series). In an electrochemical cell, each metal strip is immersed into a solution of its own ions.

Stress the direction of electron flow and of ion flow. Electrons always move from the anode to the cathode. Positive ions move toward the cathode, while anions move toward the anode.

Science Inquiry

The link below can be used to help students explore the basic parts of an electrochemical cell. Have students view the page as an introduction to the topic, especially if the materials to construct an actual cell are not available.

http://www.chemistry.uoguelph.ca/educmat/chm19105/galvanic/galvanic1.htm

Differentiated Instruction

It is essential for students to be able to write correct half-cell reactions for both oxidation and reduction.Provide students a practice worksheet in which they must write balanced half-reactions. As examples:

- 1. Oxidation of Al
- 2. Reduction of Mn^{2+}

Once students have mastered writing the half-reactions, continue by having them combine them together and balance the number of electrons. Add the half-reactions together. The final result for the examples above would be:

 $2Al + 3Mn^{2+} \rightarrow 2Al^{3+} + 3Mn$

Enrichment

Students can research the many different types of electrochemical cells in common use. There are many other besides the one described in the text. The link below describes some of them. Place the students into small groups and assign each group a different type of battery. Have the students create a poster, written report, or PowerPoint presentation.

http://batteryuniversity.com/learn/article/whats_the_best_battery

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 23.1 worksheets in *CK-12 Chemistry –Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 23.1 in CK-12 Chemistry - Intermediate Flexbook[®] resource.

Points to Consider

The electrical potential of an electrochemical cell is the voltage that the cell produces. It is dependent on the particular oxidation and reduction reactions that take place in the cell.

- What is the standard for measuring half-cell electrical potentials?
- How can the electrochemical cell potential be calculated for any cell?

23.2 Cell Potentials

Key Concept

Electrical potential is a measurement of the ability of a voltaic cell to produce an electric current. The standard hydrogen electrode is used as a reference electrode in order to determine the standard reduction potential of all other half-cells. A table of standard reduction potentials allows one to calculate the expected voltage for any combination of half cells. Reduction takes place in the half cell with the higher reduction potential, while oxidation takes place in the half cell with the lower reduction potential. The calculated voltage of a voltaic cell must be positive in order to be spontaneous.

Standards

Lesson Objectives

- Describe how an electrical potential is generated in an electrochemical cell.
- Describe the standard hydrogen electrode and how it is used to determine the standard cell potentials of other half-cells.
- Calculate the standard cell potentials from a table of standard reduction potentials.
- Predict the behavior of oxidizing and reducing agents based on their position in the table of standard reduction potentials.

Lesson Vocabulary

- cell potential: The difference in reduction potential between the two half-cells in an electrochemical cell.
- electrical potential: A measurement of the ability of a voltaic cell to produce an electric current.
- reduction potential: A measure of the tendency of a given half-reaction to occur as a reduction in an electrochemical cell.
- standard cell potential (E°_{cell}): The cell potential when the temperature is 25°C, the solution concentrations are all 1 M, and the pressure of any gases is standard (1 atm).
- standard hydrogen electrode: A reference electrode that is used with another electrode (half-cell) to determine its standard reduction potential.

Teaching Strategies

Introducing the Lesson

Show some examples of simple batteries, pointing out the various voltages of each. Ask the class what is meant by voltage. Describe electrical potential as the ability of a certain reaction to generate an electric current. A reaction with a larger electrical potential will be more efficient at generating a current and will yield a higher voltage. The goal of the lesson is to be able to calculate the voltage for any pair of oxidation-reduction reactions when combine together in a voltaic cell.

Common Misconceptions

Students may mistakenly believe that a voltaic cell cannot have one of the half-reactions with a negative reduction potential. Make sure that they understand that it is the sum of the two electrode potentials which has to be positive. The half-cell with the negative reduction must simply be paired with an oxidation of a substance that is below it on the reduction potential table. In that way, the overall cell potential will have a positive voltage.

Science Inquiry

The link below is for a lab activity that involves constructing several electrochemical cells in a micro scale procedure and measuring the resulting voltages. It will work best after students have an understanding of electrochemical cells and how to calculate the overall cell potential from the potentials of each half-cell. It is recommended to only do part 1 of the experiment.

http://hhs.heard.k12.ga.us/common/pages/DisplayFile.aspx?itemId=3030785

Differentiated Instruction

Assign pairs of students different electrochemical cell combinations to analyze. Students need to use the reduction potential chart to determine which will be the anode and which will be the cathode. Use butcher paper to make colorful drawings of the voltaic cell. Make sure to include the balanced half-cell reactions and the overall reaction, as well as the voltages for each.

Enrichment

Standard reduction potentials are the potentials when solutions in half-cells are at exactly 1 M and pressures of any gases are at 1 atm. Have interested students study the Nernst equation, which describes the dependency of cell potentials on reaction conditions. Have the students write up their findings, complete with examples of problems.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 23.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

23.2. Cell Potentials

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 23.2 in *CK-12 Chemistry* - *Intermediate* Flexbook[®] resource.

Points to Consider

Nonspontaneous redox reactions can be driven to completion by the application of an electric current in a process called electrolysis.

- How does an electrolytic cell work?
- What is electroplating?

23.3 Electrolysis

Key Concept

Electrolysis is a process in which a nonspontaneous redox reaction is forced to occur by supplying electrical energy via direct current. The voltage provided by a battery must be at least as great as the negative cell potential. Examples of electrolytic cells include a Down's cell, which produces sodium metal and chlorine gas from molten sodium chloride. The electrolysis of water produces hydrogen and oxygen gases. Electroplating is a process where metal ions are reduced to the metal and plated on to the surface of another object.

Standards

Lesson Objectives

- Distinguish between voltaic and electrolytic cells.
- Describe a Down's cell, and identify the products of the electrolysis of molten sodium chloride.
- Describe the reactions that occur during the electrolysis of water.
- Identify the products that would be generated during the electrolysis of an aqueous solution of sodium chloride.
- Describe the process of electroplating.

Lesson Vocabulary

- electrolysis: The process in which electrical energy is used to cause a nonspontaneous chemical reaction to occur.
- electrolytic cell: The apparatus used for carrying out an electrolysis reaction.
- electroplating: A process in which a metal ion is reduced in an electrolytic cell and the solid metal is deposited onto a surface.

Teaching Strategies

Introducing the Lesson

Discuss the opening story to the chapter about how aluminum was at one time considered a precious metal because of the difficulty in reducing aluminum ions to aluminum metal. Point out aluminum's position on the reduction potential chart and see how there are very few metals capable of reducing aluminum ions and that none of those metals occur naturally either. The only way to easily reduce aluminum ions to aluminum metal is to force a nonspontaneous redox

reaction to occur and the application of direct electric current can do that. Either demonstrate an electrolytic process or do one of the science inquiry experiments described below.

Common Misconceptions

Students should understand that in some instances there are multiple possibilities for reactions that can occur at the electrodes in an electrolytic cell. When aqueous sodium chloride is electrolyzed, sodium ions are not produced at the cathode because water is preferentially reduced. The reason for this is that the potential for the reduction of water is less negative than is the potential for the reduction of sodium ions. The only possible way to make elemental sodium is to electrolyze molten sodium chloride because there is no water present.

Science Inquiry

The link below is to a relatively easy electroplating experiment where students decoratively plate nickel on to a copper surface. No calculations are involved.

http://micschem.hkbu.edu.hk/emanual/Expt%2010-Electroplating-english.pdf

This link is for a more advanced electrolysis experiment in which students will be required to write the equations for the reactions involved.

http://www.walnuths.net/apps/download/Puc1kgz3ttL71BE9Cxld5uIG6VakjiL2StXjmz5o6Qq1ZVmQ.pdf/Electroplating%20Lab.pdf

Differentiated Instruction

Students can use the work they've done in the last lesson to assist them in writing the half-reactions and overall reactions for electrolysis. Have them use their completed work to write reactions that would take place in an electrolytic cell. The reactions should be reversed, switching what occurs at the anode and the cathode. The cell potential changes sign to negative.

Enrichment

Direct interested students to research the chrome plating process that is used on car bumpers and other objects. If possible, have the students gather the materials and perform the process on a small scale. If this is not feasible, have them prepare a PowerPoint presentation on the chrome plating process and how it is performed on an industrial scale.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 23.3 worksheets in *CK-12 Chemistry - Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 23.3 in *CK-12 Chemistry - Intermediate* Flexbook[®] resource.

Points to Consider

Aluminum is widely used in all sorts of modern materials from beverage cans to airplanes.

- Why was pure aluminum metal so difficult and expensive to obtain prior to the development of electrolysis on an industrial scale?
- How much energy savings comes from recycling aluminum instead of producing it from aluminum ore?

Chapter **24**

Nuclear Chemistry

Chapter Outline

- 24.1 NUCLEAR RADIATION
- 24.2 HALF-LIFE
- 24.3 FISSION AND FUSION

Chapter Overview

Unstable atomic nuclei emit radiation in the form of particles and/or energy. Natural decay processes include alpha decay, beta decay, positron emission and electron capture. The half-life of a particular radioisotope is the amount of time it takes for one half of a given amount of that isotope to decay. Radioactive decay may require many steps before a stable isotope is produced. Fission is the splitting of a heavy nucleus into lighter nuclei, while fusion is the combining together of very light nuclei. Both processes result in the release of very large amounts of energy. Nuclear power plants and the atomic bomb take advantage of the energy released during fission.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 24.1:

Lesson	Class Period(s) (60 min)
24.1 Nuclear Radiation	2
24.2 Half-Life	1-2
24.3 Fission and Fusion	1-2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 24.2:

FlexBook Lesson	Chemistry Concepts
24.1 Nuclear Radiation	24.1 - 24.2, 24.9
24.2 Half-Life	24.4

TABLE 24.2:	(continued)
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FlexBook Lesson	Chemistry Concepts
24.3 Fission and Fusion	24.3, 24.6 - 24.8, 24.10

Keywords

- alpha decay
- beta decay
- fission
- fusion
- half-life
- radioactivity
- transmutation

24.1 Nuclear Radiation

Key Concept

Radioactivity is the spontaneous breakdown of an atom's nucleus by the emission of particles and/or energy. The nuclear binding energy of a nucleus is the energy that is released when a nucleus forms from its individual nucleons. Nuclei can be unstable and radioactive if they are either very large or have a proton to neutron ratio that does not fall within a band of stability. Radioactive decay produces a change in the identity of the nuclide, called transmutation. Types of radioactive decay include alpha decay, beta decay, positron emission, electron capture and gamma ray emission. Radioactive decay gives a nucleus greater stability.

Standards

Lesson Objectives

- Explain how radioactivity involves a change in the nucleus of a radioisotope.
- Describe mass defect and calculate the conversion of mass to energy according to Einstein's equation.
- Describe the band of stability, the odd-even effect and magic numbers in terms of their influence on the stability of nuclei.
- Describe and write equations for the primary types of radioactive decay.

Lesson Vocabulary

- alpha particle (α): A helium nucleus, two protons and two neutrons, which is emitted during some types of radioactive decay.
- band of stability: The range of stable nuclei on a graph of number of neutrons to protons.
- beta particle (β): A high-speed electron emitted from the nucleus of an atom during some kinds of radioactive decay.
- gamma ray (γ): Very high energy electromagnetic waves emitted from a nucleus.
- mass defect: The difference between the mass of an atom and the sum of the masses of its protons, neutrons and electrons.
- nuclear binding energy: The energy released when a nucleus is formed from its nucleons.
- nuclear reaction: A reaction that affects the nucleus of an atom.
- nucleon: A nuclear subatomic particle (proton or neutron).
- nuclide: The nucleus of a given isotope of an element; an atom that is identified by the number of protons and neutrons in its nucleus.
- positron: A particle with the same mass as an electron, but with a positive charge.
- radiation: The emission of energy through space in the form of particles and/or waves.

- radioactive decay: A reaction in which a nucleus spontaneously disintegrates into a slightly lighter nucleus, accompanied by the emission of particles, energy, or both.
- radioisotope: An isotope of an element that is unstable and undergoes radioactive decay.
- transmutation: A change in the identity of a nucleus as a result of a change in the number of protons.

Teaching Strategies

Introducing the Lesson

Ask the class to give examples of chemical reactions. They will likely offer up some of the types of reactions such as combustion or double replacement. Ask them what all chemical reactions have in common. To lead them in the right direction, ask them to think of how John Dalton defined a chemical reaction –as a rearrangement of atoms. When chemical bonds are broken and formed, what subatomic particles are involved? What particles are responsible for ionic and covalent bonds? The answer to both questions is the electron. Now say that you are about to introduce a completely different type of reaction, a nuclear reaction, in which the electrons of the atoms are essentially irrelevant. Ask them if they can think of two main applications of nuclear reactions in the "real world". The answer is nuclear power plants and the atomic bomb.

Common Misconceptions

Nuclear binding energy and mass defect can be difficult concepts for students to understand. It is likely that they have all seen Einstein's $E = mc^2$ equation, but may not really know what it means. Discuss how when a large radioactive nucleus like uranium or plutonium is split apart into smaller particles, that some of the mass is "lost", but really converted into energy. Stress that the law of conservation of mass still holds, but that mass can be converted to energy in this case. In fact, tremendous amounts of energy are released and it is that energy that is responsible for atomic bombs and nuclear power plants.

Beta decay involves the ejection of a high speed electron from the nucleus and that may trouble students who know that electrons are not normally located in the nucleus. Emphasize that the conversion of a neutron to a proton and an electron precedes the ejection of the electron, which is simply called a beta particle in this instance.

Science Inquiry

Hands-on experiments with radioactivity are understandably problematic. The two links below are for simulations showing the processes of alpha decay and beta decay. They can be used as an introduction to the topic or after the concepts of radioactivity and balancing of nuclear equations have been established. Each page has many options for activities and worksheets to go along with the simulations.

- https://phet.colorado.edu/en/simulation/alpha-decay
- https://phet.colorado.edu/en/simulation/beta-decay

Differentiated Instruction

This lesson contains a great deal of new vocabulary terms. Have students create a list of the terms and write out the definitions. Create a word search using the definitions as clues. There are many on-line sites which can create custom word search puzzles. The link below is to one such site.

http://puzzlemaker.discoveryeducation.com/WordSearchSetupForm.asp?campaign=flyout_teachers_puzzle_wordcross

Enrichment

Assign students or pairs of students an element that has many different isotopes. It can be an element that has only radioactive isotopes (such as plutonium) or an element in which some of the isotopes are radioactive and some are stable (such as mercury). Have them go online to discover all or most of the isotopes and prepare a colorful poster showing which ones are radioactive and writing balanced nuclear equations based on the decay mode for each of the radioisotopes. If desired, limit their search to only the naturally occurring isotopes as opposed to artificially produced ones, which can be numerous. They can also add additional information such as pictures of the elements and brief descriptions of any common uses for the radioisotopes.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 24.1 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 24.1 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

The stability of a radioisotope is indicated by its half-life, which is the amount of time required for half of the nuclei in a given sample to undergo a decay process.

- What is the range of half-lives for known isotopes?
- How can half-life be used to determine the age of objects, such as fossils or certain rocks and minerals?

24.2 Half-Life

Key Concept

The half-life of a radioisotope is the time that it takes for one half of the atoms in any sample of the isotope to decay. The half-life of certain radioisotopes can be used to determine the ages of fossils and other very old objects. Often, a certain radioactive nuclide will undergo many decays steps before a stable nuclide is produced. The bombardment of nuclides with small particles can produce nuclear reactions in a process called artificial transmutation. The transuranium elements were all formed in this fashion.

Standards

Lesson Objectives

- Define half-life as it relates to radioactive nuclides and solve half-life problems.
- Describe the general process by which radioactive dating is used to determine the age of various objects.
- Explain the mechanism of a decay series.
- Define and write equations for artificial transmutation processes.

Lesson Vocabulary

- artificial transmutation: The bombardment of stable nuclei with charged or uncharged particles in order to cause a nuclear reaction.
- decay series: A sequence of successive radioactive decays that proceeds until a stable nuclide is reached.
- half-life $(t_{\frac{1}{2}})$: The time required for one half of the nuclei of a radioactive nuclide to decay.
- radioactive dating: A process by which the approximate age of an object is determined through the use of certain radioactive nuclides.

Teaching Strategies

Introducing the Lesson

Ask the students if they are familiar with the process of radioactive carbon dating. What is it used for? What isotope of carbon is radioactive? How much of naturally occurring carbon is carbon-14? Explain how, once an organism dies, it no longer produces new carbon and so the amount of carbon-14 in its body slowly diminishes over time. Introduced the concept of half-life and work some problems using the carbon-14 half-life of 5730 years.

Common Misconceptions

Students sometimes mistakenly think that when one half of a certain sample of a radioisotope has decayed that the matter it represents has completely disappeared. Stress that the law of conservation of mass cannot be violated and that the original isotope has just been transformed into something else.

Science Inquiry

The process of using radioactive isotopes of carbon and uranium for dating of fossils and rocks can be studied through the use of the simulation below. The page also contains many activities for students that can be downloaded and used as needed.

https://phet.colorado.edu/en/simulation/radioactive-dating-game

Differentiated Instruction

Some students may have trouble with half-life problems. You can teach it by using a chart to keep track of time and amount. Suppose that a 20-gram sample of phosphorus-32 (half-life of 14.3 days) was allowed to decay until 2.5 grams remained. How much time has passed? Make a two-column table of time on the left and amount on the right. The completed table would look like this:

TABLE 24.3:

Time (days)	Amount (g)
0	20
14.3	10
28.6	5
42.9	2.5

The answer is 42.9 days. Make sure to stress that for each row of the table, you must <u>add</u> a new time equal to the half-life. Some students may mistakenly double the time for each row; in which case the fourth row in the above table would be 57.2 days instead of 42.9.

Enrichment

Advanced math students can learn to use natural logarithms to solve half-life problems in which the number of half-lives that have passed is not a whole number. The relevant equations are shown below:

$$\ln \frac{N_t}{N_0} = -kt \quad \text{and} \quad t_{\frac{1}{2}} = \frac{\ln 2}{k}$$

The variable k is a first-order rate constant, measured in reciprocal time (s^{-1} for example). N_t and N_0 are the number of nuclei and time t and the number of nuclei at time = 0, respectively. Mass of the sample can be used in place of the number of nuclei.

For example, suppose that a 10.0-gram sample of a certain radioisotope has decayed for 15.2 hours and there is 1.94 g of the sample remaining. To find the half-life, use the first equation to solve for k (in hr⁻¹).

$$\ln \frac{1.94 g}{10.0 g} = -k(15.2 hr) \qquad k = -\frac{\ln(0.194)}{15.2 hr} = 0.108 hr^{-1}$$

Now, use the second equation to solve for the half-life.

$$t_{\frac{1}{2}} = \frac{\ln 2}{k} = \frac{\ln 2}{0.108 \ hr^{-1}} = 6.42 \ hr$$

Similar problems can be constructed where the half-life is known and the amount remaining after a certain amount of time has passed (N_t) can be calculated.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 24.2 worksheets in *CK-12 Chemistry –Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 24.2 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Nuclear fission and nuclear fusion are two processes that can occur naturally or through bombardment. Both release tremendous amounts of energy.

- What are two modern applications of nuclear fission?
- Where does nuclear fusion occur naturally?

24.3 Fission and Fusion

Key Concept

The process of nuclear fission involved the splitting of very large nuclei into nuclei of intermediated mass, along with the release of large amounts of energy. A chain reaction occurs when a critical mass of fissionable material undergoes the fission process very rapidly. Fission is used in atomic bombs and nuclear power plants. Fusion is the combining of very small nuclei and is the energy produced powers the Sun and other stars. Ionizing radiation can do damage to living tissues and so care must be taken to limit exposure. Radiation can be used in cancer treatments, to study chemical reactions and to prolong the shelf life of some foods.

Standards

Lesson Objectives

- Define fission and explain why energy is released during the fission process.
- Describe a nuclear chain reaction and how it is applied in both a fission bomb and in a nuclear power plant.
- Define fusion and explain the difficulty in using fusion as a controlled energy source.
- Explain how ionizing radiation is measured and how it is detected.
- Describe some uses of radiation in medicine and agriculture.

Lesson Vocabulary

- chain reaction: A reaction in which the material that starts the reaction is also one of the products and can start another reaction.
- control rod: A rod made of cadmium or boron that is used in a nuclear reactor to limit the available neutrons by absorbing some of them and preventing the fission reaction from proceeding too rapidly.
- critical mass: The minimum amount of fissionable material needed to sustain a chain reaction.
- film badge: A device that consists of several layers of photographic film that can measure the exposure, of the wearer, to radiation.
- Geiger counter: A device that uses a gas-filled metal tube to detect radiation.
- ionizing radiation: Radiation that has the required energy to knock electrons off the atoms of a bombarded substance, and produce ions.
- moderator: A material which slows down high-speed neutrons in a nuclear reactor.
- nuclear fission: A process in which a very heavy nucleus (mass >200) splits into smaller nuclei of intermediate mass.
- nuclear fusion: A process in which light-mass nuclei combine to form a heavier and more stable nucleus.
- radioactive tracer: Radioactive atoms that are incorporated into substances so that the movement of these substances can be tracked by a radiation detector.
- rem: The amount of ionizing radiation that does as much damage to human tissue as is done by 1 roentgen of high-voltage x-rays.
- roentgen: A unit that measures nuclear radiation and is equal to the amount of radiation that produces 2×10^9 ion pairs when it passes through 1 cm³ of air.
- scintillation counter: A device that uses a phosphor-coated surface to detect radiation by the emission of bright bursts of light.

Teaching Strategies

Introducing the Lesson

Begin by asking the students if they know what is in atomic bomb. How does it work? Why are atomic bombs so much more destructive than other bombs? Ask if they are familiar with the Manhattan Project. Describe nuclear fission and how the energy obtained from a fission reaction is millions of times larger than for ordinary chemical reactions.

Common Misconceptions

Students may think that if a fission reaction releases energy, then a fusion reaction must absorb energy. While this seems reasonable, both reactions release energy and fusion reactions release a great deal more than fission. In both cases, mass is lost and converted to energy according to Einstein's equation. The mass of the products in a fission reaction is less than the mass of the reactants. In a fusion reaction, the mass of the product nucleus is also slightly less than the mass of the original nuclei that were fused.

Science Inquiry

Nuclear fission, the chain reaction and a nuclear power plant can all be investigated using the simulation below. It allows students to see how a certain critical number (or mass) of fissionable nuclei must be present in order for a chain reaction to occur. Student activities can be downloaded from this page as well.

https://phet.colorado.edu/en/simulation/nuclear-fission

Differentiated Instruction

Students can make a list of some jobs and careers that involve increased risk of exposure to radiation. Examples are workers at a nuclear power plant or doctors who treat cancer patients with radiation. Create posters or collages of these careers and descriptions of the precautions that people must take in order to enter these fields.

Enrichment

Nuclear power is a controversial subject and can be the subject of a class debate. Obtain a few articles that provide arguments both for and against the expanded use of nuclear power. Assign the students to read the articles and decide on a stance: either for or against nuclear power. Hold a debate in class, allowing students to ask each other questions, make arguments and counterarguments. Act as the moderator without giving your own opinion. Have the students write up a brief conclusion after the debate, outlining their current position on the use of nuclear power.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 24.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 24.3 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Fission and fusion are promising as energy sources, but they are not without difficulties and controversy.

- Do the advantages of nuclear power justify the risk of accidents and the problems associated with the disposal of nuclear waste?
- What methods are being investigated as a way to use controlled nuclear fusion as an energy source?



Organic Chemistry

Chapter Outline

- 25.2 HYDROCARBON RINGS
- 25.3 FUNCTIONAL GROUPS
- 25.4 ORGANIC REACTIONS

Chapter Overview

Organic chemistry is the study of compounds containing the element carbon. Hydrocarbons contain only carbon and hydrogen, and include alkanes, alkenes, alkynes, as well as cyclic hydrocarbons. Aromatic hydrocarbons contain one or more benzene rings. Functional groups are atoms or groups of atoms that give a particular compound its unique chemical and physical properties. Common functional groups include alcohols, ethers, aldehydes, ketones, carboxylic acids, esters, and amines. Chemical reactions in organic chemistry include substitution, addition, oxidation, condensation, and polymerization.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 25.1:

Lesson	Class Period(s) (60 min)
25.1 Hydrocarbons	2
25.2 Hydrocarbon Rings	1-2
25.3 Functional Groups	2
25.4 Organic Reactions	2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 25.2:

Flexbook Lesson	Chemistry Concepts
25.1 Hydrocarbons	25.1 – 25.5
25.2 Hydrocarbon Rings	25.6 - 25.7
25.3 Functional Groups	25.8 - 25.14
25.4 Organic Reactions	25.15 - 25.20

Keywords

- aliphatic
- aromatic
- functional group
- hydrocarbon
- isomer
- organic chemistry
- organic reactions

25.1 Hydrocarbons

Key Concept

Organic chemistry is the study of compounds that contain carbon. Hydrocarbons are molecules that contain only carbon and hydrogen. Alkanes are saturated hydrocarbons, meaning that any carbon-carbon bonds are single bonds and there is the maximum number of hydrogen atoms possible. Alkanes may be either straight-chain or branched. Alkenes and alkynes are unsaturated hydrocarbons. Alkenes have at least one carbon-carbon double bond, while alkynes have at least one carbon-carbon triple bond. Two or more molecules with the same chemical formula but different structures are called isomers. Isomers may be structural, geometric, or optical.

Standards

Lesson Objectives

- Describe the bonding characteristics of carbon that lead to such a large number of organic compounds.
- Differentiate between aliphatic and aromatic hydrocarbons.
- Describe, name, and draw structures for straight-chain and branched alkanes.
- Differentiate between saturated and unsaturated hydrocarbons.
- Describe alkenes and alkynes.
- Describe the characteristics of structural isomers, geometric isomers, and optical isomers.

Lesson Vocabulary

- aliphatic hydrocarbon: A hydrocarbon that does not contain the benzene group or benzene ring.
- alkane: A hydrocarbon in which there are only single covalent bonds.
- alkene: A hydrocarbon with one or more carbon-carbon double covalent bonds.
- alkyl group: A hydrocarbon substituent.
- alkyne: A hydrocarbon with one or more carbon-carbon triple covalent bonds.
- aromatic hydrocarbon: A hydrocarbon that contains one or more benzene rings.
- geometric isomer: Molecules in which the order of atom bonding is the same but the arrangement of atoms in space is different.
- hydrocarbon: An organic compound that is made up of only carbon and hydrogen.
- optical isomer: Molecules that are non-superimposable mirror images of each other.
- organic compound: Covalently bonded compound containing carbon, excluding carbonates and oxides.
- saturated hydrocarbon: A hydrocarbon that contains the maximum number of hydrogen atoms that can possibly bond with the number of carbon atoms present.
- structural isomer: One of multiple molecules that have the same molecular formula, but different structural formulas.

• unsaturated hydrocarbon: A hydrocarbon that contains less than the maximum number of hydrogen atoms that can possibly bond with the number of carbon atoms present.

Teaching Strategies

Introducing the Lesson

Ask the students how many valence electrons are in an atom of carbon. How many covalent bonds does a carbon atom typically form? Draw the Lewis structure of a carbon atom on the board to emphasize that carbon can make bonds to atoms in four different directions. Show a molecular model of methane to the class. Ask about the shape of the molecule. Discuss how carbon's versatile bonding ability means that an entire field of chemistry is devoted solely to carbon and its compounds. While discussing other molecules throughout the course of this chapter, have models on hand to show the class. Begin with hydrocarbons, extensions of methane with longer and longer carbon chains.

Common Misconceptions

It is very important to emphasize the geometry of the organic compounds discussed throughout the chapter. Students will likely see structural formulas on the board or on a book page and think that the molecules represented are planar. Show models and remind the students of the approximate 109° angle that exists around a tetrahedral carbon atom. When drawing branched structure, be sure to point out that a single alkyl branch be drawn either "above" or "below" a carbon atom in a chain (see below). Because of the free rotation of single bonds, the two structural drawings are the exact same molecule.



The two structures of 3-methylpentane are the same.

Science Inquiry

The activity below can be used for students to further enhance their skills with creating structures of hydrocarbons and properly naming them. It works best after students have had a basic introduction to the nomenclature system, including straight chain and branched structures.

http://www.rsc.org/education/teachers/resources/aflchem/resources/50/index.htm

Differentiated Instruction

The link below is for a very basic activity involving drawing and naming straight-chain hydrocarbons. It could be used as an introductory activity or as reinforcement for students who may be having difficulty with the concepts of this lesson.

http://www.esrl.noaa.gov/gmd/infodata/lesson_plans/Naming%20and%20Creating%20Hydrocarbons.pdf

Enrichment

Students often enjoy the challenge of trying to draw all of the possible structural isomers for hydrocarbons of longer and longer carbon chains. The link below is a worksheet with some formulas for students to work with and try to come up with all of the possible isomers. This type of activity can easily be turned into a competition between small groups of students within the class.

http://legacy.jefferson.kctcs.edu/users/kaya.muller/che120/supplements/structures/worksheets/isomers_worksheet/w orksheet.html

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 25.1 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 25.1 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Carbon is capable of making ringed structures called cyclic hydrocarbons. Benzene is an unusually stable type of cyclic hydrocarbon that is referred to as an aromatic ring.

- What are cycloalkanes, and how are they named?
- What is the system for naming aromatic compounds with substituents attached to a benzene ring?

25.2 Hydrocarbon Rings

Key Concept

Cyclic hydrocarbons are hydrocarbons composed of rings of carbon atoms. The most common and stable cycloalkanes are cyclopentane and cyclohexane. Aromatic hydrocarbons contain one or more benzene rings. The pibonding electrons in benzene are delocalized around the six-membered ring, making benzene a very stable molecule. Substituted aromatic compounds are those in which other atoms or groups of atoms have replaced one or more of the hydrogen atoms around the ring.

Standards

Lesson Objectives

- Describe and name cyclic hydrocarbons.
- Explain how ring strain contributes to the shape of cyclic hydrocarbons. Describe the two primary conformations of cyclohexane.
- Relate the concept of electron delocalization to the structure of benzene.
- Name and write structural formulas for substituted benzene compounds.

Lesson Vocabulary

- cyclic hydrocarbon: A hydrocarbon in which the carbon chain joins to itself in a ring.
- cycloalkane: A cyclic hydrocarbon in which all of the carbon-carbon bonds are single bonds.
- cycloalkene: A cyclic hydrocarbon with at least one carbon-carbon double bond.
- cycloalkyne: A cyclic hydrocarbon with at least one carbon-carbon triple bond.
- delocalized electrons: Electrons that are not confined to the bond between two atoms but are, instead, allowed to move between three or more.

Teaching Strategies

Introducing the Lesson

Show images of some common molecules that have biological significance such as DNA bases or a glucose molecule in its ringed from. Ask the students how the carbon atoms of these molecules appear different than the hydrocarbons of the last chapter. Discuss how rings of carbon atoms are very prevalent in many, many organic compounds in the natural world. Emphasize that the rules for constructing and naming molecules will be largely the same as what they

learned in the last lesson. Start with simple cycloalkanes, from cyclopropane up through cyclohexane. Make sure to explain how the ideal bond angle of 109° for tetrahedral carbon atoms dictates that 5 and 6-membered carbon rings are by far the most common.

Common Misconceptions

Students may easily confuse cycloalkanes with the aromatic compounds derived from benzene. Stress that while the structure look very similar on the board, the benzene ring is planar while cyclohexane is not. This is because of the need for the carbon atoms to come as close as possible to the ideal bond angle. Because the carbon atoms in benzene are trigonal planar, the ideal bond angle is 120° and a planar structure is adopted.

Science Inquiry

The link below serves as a capstone to the study of nomenclature of alkanes, alkenes, alkynes, and cyclic hydrocarbons. Use it at the end of the lesson to gauge students' mastery of the material so far.

http://www.sophiebwrightschool.com/ourpages/auto/2012/4/25/48559588/Drawing%20and%20Naming%20Alkanes .pdf

Differentiated Instruction

Working in groups, assign students a few different compound names. Have them draw the structures on butcher paper and construct models using a molecular model kit. Each student in the group gets assigned a different molecule. When finished, he or she should show the models and pictures to the other members of the group and the group can agree or disagree as to whether the structure is correct.

Enrichment

The link below has a nice selection of practice problems for students who have mastered the basics of nomenclature of aromatic compound

http://2012books.lardbucket.org/books/introduction-to-chemistry-general-organic-and-biological/s16-08-structure-and-nomenclature-of-.html

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 25.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 25.2 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

There is a wide variation in the reactivity of organic compounds due to specific groups in the molecule, most of which contain atoms other than hydrogen and carbon.

- What are some of the most common functional groups?
- How are functional groups indicated in the name of an organic compound?

25.3 Functional Groups

Key Concept

A functional group is an atom or group of atoms that is primarily responsible for the properties of the organic compound in which it is present. Alkyl halides contain one or more halogen atoms in place of a hydrogen atom. Alcohols contain the hydroxyl functional group. Ethers are compounds with an oxygen atom bonded to two alkyl groups. Aldehydes and ketones contain the carbonyl functional group. Carboxylic acids contain the carboxyl functional group and are weak acids. An ester is a compound in which the hydrogen of a carboxyl group has been replaced by an alkyl group. Amines are nitrogen-containing compounds related to ammonia and are weak bases.

Standards

Lesson Objectives

- Define a functional group and explain the importance of functional groups to organic chemistry.
- Be able to identify the following functional groups in an organic molecule: alkyl halide, alcohol, ether, aldehyde, ketone, carboxylic acid, ester, and amine.
- Know how to name simple molecules containing one of the basic functional groups listed above.
- Relate the properties of organic compounds to their structures and the functional groups they contain, and describe the uses of some of these compounds.

Lesson Vocabulary

- alcohol: An organic compound that contains one or more hydroxyl (-OH) groups.
- aldehyde: An organic compound in which a carbonyl group (C=O) is attached to a carbon atom at the end of a carbon chain.
- alkyl halide: An organic compound in which one or more halogen atoms are substituted for one or more hydrogen atoms in a hydrocarbon.
- amine: An organic compound that can be considered to be a derivative of ammonia (NH₃).
- carboxylic acid: An organic compound that contains the carboxyl functional group (-COOH).
- ester: An organic compound that is a derivative of a carboxylic acid in which the hydrogen atom of the hydroxyl group has been replaced with an alkyl group.
- ether: An organic compound in which two hydrocarbon groups are bonded to the same atom of oxygen.
- functional group: An atom or group of atoms that is primarily responsible for the properties and reactions of a given organic compound.
- ketone: An organic compound in which a carbonyl group (C=O) is attached to a carbon atom within the carbon chain.

Teaching Strategies

Introducing the Lesson

Thus far, students have learned about organic compounds that contain carbon and hydrogen only. Ask the class with what other elements carbon may easily from covalent bonds. If anyone names metals, remind them that covalent bonds must be between nonmetal atoms. They will most likely come up with elements such as oxygen, nitrogen, sulfur, and the halogens. Discuss how a key aspect of organic chemistry is the study of compounds which contain these elements, in addition to carbon and hydrogen. In fact, it is these other elements which usually give a particular compound its distinctive chemical and physical properties.

Common Misconceptions

One misconception students may have is that alcohols are bases due to the presence of the –OH group. Stress that in a base, the hydroxyl group must be ionically bonded to a metal, typically from Groups 1 or 2. Because of the large difference in electronegativity between the metal atom and the oxygen of the –OH, that bond breaks, with the bonding electrons remaining on the oxygen. When the –OH is bonded to a carbon atom instead, the electronegativities are closer and the covalent bond does not easily break. Alcohols are definitely not bases.

Students can easily confuse the four different functional groups that contain the carbonyl group (C=O): aldehydes, ketones, carboxylic acids, and esters. Make sure to stress how each is different from the others.

Science Inquiry

The following worksheet is a good summary of all the different functional groups and gives the student additional practice on identifying them.

https://www.chem.wisc.edu/areas/clc/general/104/practsheet_1_104.pdf

Differentiated Instruction

Students that have difficulty visualizing the various organic molecules from drawings should use molecular models. Have a group build a simple example of a molecule containing each of the different functional groups. From the models, make three-dimensional sketches and display them in the room.

Enrichment

Assign different functional groups to students or pairs of students. Have them research different compounds of importance that contain these groups. They can make a poster or PowerPoint presentation summarizing their findings.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 25.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 25.3 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Common organic chemical reactions can be classified into several broad categories.

- What are some of the most common types of organic reactions?
- How do functional groups give organic compounds their chemical properties?

Sample Answers

- Common organic reactions include addition reactions, substitution reactions, oxidations, condensation reactions, esterifications, and polymerizations.
- A particular functional group is typically the most reactive portion of the molecule. Each functional group has its own unique type of chemical reactions.

25.4 Organic Reactions

Key Concept

Basic organic reactions can be broken down into several categories. A substitution reaction involves one atom in a molecule being replaced by a different atom. In an addition reaction, a molecule is added across a double or triple bond and includes the processes of hydrogenation and hydration. Hydrocarbons and other organic molecules can undergo oxidation, either by the gain of oxygen or the loss of hydrogen. Condensation reactions involve the combining together of two molecules, eliminating a molecule of water in the process. Polymers are very large molecules made of small repeating units and polymers can be produced by addition or condensation reactions.

Standards

Lesson Objectives

- Describe and distinguish between substitution reactions and addition reactions, and give examples of each.
- Relate the concepts of oxidation and reduction to organic reactions.
- Describe and give examples of condensation reactions, including esterification.
- Define "polymer" and describe how polymers are formed by an addition reaction or by a condensation reaction.

Lesson Vocabulary

- addition polymer: A polymer formed by chain addition reactions between monomers that contain a double bond.
- addition reaction: A reaction in which an atom or molecule is added to an unsaturated molecule, making a single product.
- condensation polymer: A polymer formed by condensation reactions.
- condensation reaction: A reaction in which two molecules combine to form a single molecule.
- esterification: A reaction in which an ester is formed from an alcohol and a carboxylic acid.
- hydration reaction: A reaction in which water is added to an alkene.
- hydrogenation: A reaction that occurs when molecular hydrogen is added to an alkene to produce an alkane.
- monomer: The small molecules which are repeated over and over and make up a polymer.
- polymer: A large molecule formed of many smaller molecules covalently bonded in a repeating pattern.
- saponification: The alkaline hydrolysis reaction of an ester; a reaction that makes soap molecules.
- substitution reaction: A reaction in which one or more atoms replace another atom or group of atoms in a molecule.

Teaching Strategies

Introducing the Lesson

Ask the students if they heard of the terms saturated and unsaturated in regards to fats that are contained in foods. Which is better? Unsaturated fats are healthier. Is there a way to turn a saturated hydrocarbon into a saturated one? Draw two molecules on the board such as butene and butane. Ask for the difference between the two. If a molecule of hydrogen can be added to butene, it can be converted to butane. This is an example of a hydrogenation reaction. Many food labels that they have seen may contain an ingredient called partially hydrogenated vegetable oil. An oil that contains multiple double bonds is reacted with hydrogen gas, causing some of the double bonds to become hydrogenated.

Common Misconceptions

Addition reactions and substitution reactions can be confused by the students. One way to remember them is that for a molecule to take part in an addition reaction, it must have a double or triple bond. Substitution reactions can occur for molecules with all single bonds like alkanes.

Students will have trouble remembering the sequence of oxidations from alcohol to aldehyde to carboxylic acid to carbon dioxide. Make sure to draw out the structures on the board and/or use models so that they can see how each step is either a gain of oxygen or a loss of hydrogen.

Science Inquiry

The preparation of esters is an excellent way to introduce students to organic reactions in the laboratory. Esters are relatively easy to make and have all different sorts of recognizable fragrances from bananas to grapes to wintergreen. There are many examples of ester labs that can be found online, with a couple of links below.

http://www.kentschools.net/ccarman/files/2009/08/4-6-Making-Artificial-Fragrances-Lab-fy11.pdf

http://www.scribd.com/doc/30755499/Organic-Chemistry-Esters-Lab-Lab-Report-Making-Scents-of-Esters#scribd

Differentiated Instruction

Students will enjoy making a slimy polymer from borax and Elmer's glue. A set of directions is shown at the link below. If desired, have students look up other possible polymers that can easily be made in the lab. Compare their properties.

http://www.stevespanglerscience.com/lab/experiments/glue-borax-gak

Enrichment

Polymers are a very important class of compounds that are used in all sorts of products in everyday life. Assign different types of polymers (nylon, polyesters, vinyl chloride, etc.) to groups of students and have them research products that their polymer is used to make. They can display their findings on a poster. complete with drawings of the molecular structure and pictures of the products that have been cut out from articles or advertisements. If possible, they can also bring in examples of products made from their polymer.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 25.4 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 25.4 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Chemistry and chemical reactions are key factors in living things.

- How can a basic understanding of organic chemistry provide a foundation for studying biological molecules?
- What are the main categories of biomolecules?



Biochemistry

Chapter Outline

26.1	CARBOHYDRATES
26.2	AMINO ACIDS AND PROTEINS
26.3	LIPIDS

26.4 NUCLEIC ACIDS

Chapter Overview

Carbohydrates serve as energy storage in plants and animals. They consist of organic molecules called monosaccharides, which can be linked together into disaccharides as well as longer structures called starches and cellulose. Amino acids are small molecules with an amino group and a carboxyl group. Amino acids link together to form peptides, polypeptides and proteins. Proteins adopt specific structures based on their amino acid sequence. Lipids are water-insoluble molecules and include triglycerides, phospholipids and waxes. Nucleic acids are the genetic material of the cell. DNA and RNA are very long double-helical strands composed of four different nitrogen-containing bases. The base sequence of DNA is decoded by the cell to produce specific proteins.

Online Resources

See the following Web sites for appropriate activities:

Pacing the Lessons

TABLE 26.1:

Lesson	Class Period(s) (60 min)
26.1 Carbohydrates	2
26.2 Amino Acids and Proteins	2
26.3 Lipids	2
26.4 Nucleic Acids	2

Chemistry Concepts

The table below matches each lesson from the flexbook to the chemistry concepts.

TABLE 26.2:

Flexbook Lesson	Chemistry Concepts
26.1 Carbohydrates	26.1 - 26.3
26.2 Amino Acids and Proteins	26.4 - 26.7
26.3 Lipids	26.8 - 26.10
26.4 Nucleic Acids	26.11 - 26.14

250-Character Chapter Description

This chapter covers biomolecules including carbohydrates, amino acids, peptides, proteins, enzymes, lipids, triglycerides, phospholipids, waxes and nucleic acids.

Keywords

- amino acids
- carbohydrates
- enzymes
- genetic code
- lipids
- nucleic acids
- peptides
- proteins

26.1 Carbohydrates

Key Concept

Carbohydrates are a class of biomolecules that store energy in plants and animals and are found in many foods. Monosaccharides, also called simple sugars, consist of six carbon atoms along with multiple hydroxyl groups and a carbonyl group. Glucose and fructose are common monosaccharides. Disaccharides include maltose, sucrose and lactose. Starches and glycogen are long chains of glucose monomers and are used as longer term energy storage. Cellulose is a monosaccharide that makes up many plant fibers and is indigestible by most animals.

Standards

Lesson Objectives

- Describe the general structures of the simple sugars, glucose and fructose.
- Describe how two simple sugars can be linked to form a disaccharide.
- Compare the simple sugar components of sucrose, lactose and maltose.
- Identify the structure and function of the polysaccharides starch, glycogen and cellulose.

Lesson Vocabulary

- carbohydrates: Monomers and polymers of aldehydes and ketones that have multiple hydroxyl groups attached.
- disaccharide: A carbohydrate formed by the joining of two monosaccharides.
- monosaccharide: A carbohydrate consisting of one sugar unit.
- polysaccharide: A complex carbohydrate polymer formed from the linkage of many monosaccharide monomers.

Teaching Strategies

Introducing the Lesson

Bring in samples of high-carbohydrate foods such as pasta, bread and potatoes. Ask the class if they know what nutrients these foods are mostly known for. Who typically eats a large amount of these types of foods and why? Show models of simple sugars and discuss that they are called carbohydrates because of their empirical formulas.

26.1. Carbohydrates

Common Misconceptions

Many resources show the structural formulas for monosaccharides as open chain molecules. While that can be instructive for seeing the functional groups involved, emphasize that in solution these molecules adopt a cyclic structure.

Science Inquiry

Benedict's solution is a chemical test used to detect simple sugars. Iodine is used to test for starches. The lab activity below is written to test some components of a typical meal for both monosaccharides and for starches. It can be done as an introductory class activity or as a teacher demonstration. The lab also includes a test for proteins, appropriate for the following lesson.

http://www.napavalley.edu/people/sburns/Documents/Biol%20105%20Labs/Lab%203%20Biological%20Molecules .pdf

Differentiated Instruction

Keeping track of the different types of biomolecules can be challenging for students because of their size and complexity. Have students draw a large circle on a sheet of butcher paper. Put in dividing lines to turn the circle into a pie chart with four equal areas. Each of these areas will eventually contain drawings and other information related to the four lessons of this chapter. Students can start by labeling one of the areas as "Carbohydrates". In the area they can put drawings, describe the uses by the body, and attach pictures cut out from articles or advertisements. As they work through the other lessons, add to the diagram.

Enrichment

Have interested students study the chemical reactions that occur in the body when various carbohydrates undergo the process of digestion. Alternatively, other reactions take place when starches are used as the long-term energy storage. Finally, some animals like cows are capable of digesting cellulose. Students can present their findings to the class.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 26.1 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 26.1 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Proteins are polymers of another class of biomolecules called amino acids.

- How many naturally occurring amino acids are used to form proteins?
- What is the function of an enzyme in biological reactions?

26.2 Amino Acids and Proteins

Key Concept

An amino acid is an organic molecule that consists of a carboxyl group, an amino group, and a variable side chain that are all covalently bonded to a central carbon atom. A reaction between two amino acids is a condensation reaction and the resulting C-N bond is called a peptide bond. The twenty fundamental amino acids present in all living organisms link together in unique sequences to form polypeptides and proteins. Proteins have four levels of structure called primary, secondary, tertiary and quaternary. Enzymes are proteins that catalyze biochemical reactions by binding a specific substrate at its active site.

Standards

Lesson Objectives

- Describe the general structure of an amino acid.
- List the twenty common amino acids found in living organisms.
- Describe how a peptide bond forms.
- Define a protein and differentiate between the four levels of structure of a protein.
- Describe the effect of an enzyme on a biochemical reaction and the general features of its reaction with a substrate.

Lesson Vocabulary

- amino acid: A compound that contains both an amino group $(-NH_2)$ and a carboxyl group (-COOH) in the same molecule.
- enzyme: A protein that acts as a biological catalyst.
- peptide: A combination of amino acids in which the amino group of one amino acid has undergone a reaction with the carboxyl group of another amino acid.
- peptide bond: The amide bond that occurs between the amino nitrogen of one amino acid and the carboxyl carbon of another amino acid.
- polypeptide: A sequence of amino acids between ten and one hundred in length.
- primary structure: The amino acid sequence of a protein.
- protein: A peptide that is greater than one hundred amino acids in length.
- quaternary structure: The specific interaction and orientation of the subunits of a protein.
- secondary structure: A highly regular sub-structure of a protein.
- tertiary structure: The overall three-dimensional structure of a protein.

Teaching Strategies

Introducing the Lesson

Ask the class to remember the functional groups of carboxylic acids and amines. Draw the structures of each on the board. Describe how there is a certain important class of compounds that have both of these functional groups contained in the same molecule. Draw the basic structure of an amino acid on the board and show a model. Stress that all biologically relevant amino acids have this same structure, but that there are twenty different R groups. Amino acids join together one after the other to form proteins. Ask what foods are high in protein. What is the function of protein in the human body?

Common Misconceptions

Stress the directionality of a peptide sequence by drawing two different dipeptides on the board. Use two simple R groups such as glycine and alanine. One dipeptide should have the glycine on the left with the free amino group and the alanine on the right with the free carboxyl group. The other dipeptide should be reversed.

The levels of structure of a protein can be difficult for students to differentiate. The difference between the secondary structures of alpha helix can and beta sheet should be shown with the aid of a model or an animation.

Science Inquiry

The presence of protein in a sample can be tested by reacting with a solution called a biuret solution. The link below describes a detailed activity in which students conduct tests for protein, as well as for carbohydrates and lipids.

http://seplessons.ucsf.edu/node/362

The activity at the link below uses a model kit specifically designed to show amino acids and to demonstrate how they link together to form peptide sequences. The specific models would have to be purchased in order to make use of the activity.

http://www.3dmoleculardesigns.com/3DMD-Files/AASK/PDFs/Student-Handouts/AASKStudentHandout1Key.pdf

Differentiated Instruction

Students have likely studied proteins before at some level –either in biology or a middle school science class. As a start to the lesson, pair students together and have them brainstorm as many things about proteins that they can. After a set amount of time, ask the pairs for their thoughts and make a list of the contributions of the class on the board.

Enrichment

Just a single mutation in the amino acid sequence of a protein can cause dramatic changes in the structure and function of the protein. For example, one amino acid change in the large hemoglobin molecule is responsible for a debilitating disease called sickle cell anemia where the normal shape of the red blood cell is altered and its oxygen carrying capability is diminished. Have students research other mutations in biologically important proteins and the problems that those mutations cause.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 26.2 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 26.2 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

The class of biomolecules called lipids is comprised of fats, oils and other water-insoluble compounds.

- What is the structure of a triglyceride?
- What are phospholipids?

26.3 Lipids

Key Concept

Lipids are a class of water-insoluble biomolecules that includes oils, fats and waxes. A triglyceride is derived from the esterification of glycerol with three long-chain fatty acids. Phospholipids consist of long hydrocarbon chains with a phosphate group on the end. They form a bilayer which functions as the cell membrane. Waxes are soft, low-melting solids formed from long-chain fatty acids and long-chain alcohols.

Standards

Lesson Objectives

- Describe the structure of a triglyceride.
- Distinguish between saturated and unsaturated fats.
- Describe the structure of a phospholipid and a lipid bilayer.
- Describe the structure and properties of waxes.

Lesson Vocabulary

- lipid: A member of a class of water-insoluble compounds that includes fats, oils, waxes.
- phospholipid: A lipid that contains a phosphate group and is a major component of cell membranes.
- saturated fat: A fat that consists of triglycerides whose carbon chains consist entirely of carbon-carbon single bonds.
- triglyceride: An ester derived from glycerol combined with three fatty acid molecules.
- unsaturated fat: A fat that consists of triglycerides whose carbon chains contain one or more carbon-carbon double bonds.
- wax: An ester of a long-chain fatty acid and long-chain alcohol.

Teaching Strategies

Introducing the Lesson

Ask the class if they know what biomolecule makes up butter, lard and the coating of plant leaves. They all belong to a class of compounds called lipids. What is significant about the behavior of lipids? If necessary, list some other examples. The distinguishing characteristic of lipids is their insolubility in water. Why is this the case? Remind students of the "like dissolves like" rule that they learned in the chapter on solutions. Is water polar or nonpolar?

26.3. Lipids

Since water is polar, lipids must be nonpolar. What is it about the chemical structure of organic molecules that allows them to be nonpolar? Hopefully, they will be able to reason that it is the long length of the hydrocarbon chains that gives lipids their insolubility.

Common Misconceptions

Students have difficulty understanding the difference between fats and oils. Most fats are animal-derived and solid, while oils are plant-derived and liquid. Fats consist of mostly unsaturated carbon chains which are able to stack rather tightly next to one another. This increases the intermolecular interaction and results in fats being solid. Oils are generally comprised of carbon chains that are more unsaturated. Recall that means there are double bonds between carbon atoms. This structure causes bends in the chains and they do not stack together nearly as efficiently as the saturated chains do. As a result, oils are liquid rather than solid.

Science Inquiry

The activity at the link below contains two tests for the presence of lipids: a Sudan red test and a grease spot test. This activity can be used in conjunction with the protein and carbohydrate tests from the two previous lessons.

http://seplessons.ucsf.edu/node/362

Differentiated Instruction

Have students build models of a phospholipid bilayer, complete with embedded proteins. Have them brainstorm what materials could be used for the hydrophilic heads and the hydrophobic tails. Display their creations in the room.

Enrichment

If time allows, a saponification reaction is an interesting lab activity for the students. Saponification is the process of making soap by hydrolyzing oils and fats by boiling with an aqueous solution of an alkali-metal hydroxide. The link below is one of many online that provides instruction for a saponification experiment.

http://chemlab.truman.edu/CHEM100Labs/SAPONIFICATION.pdf

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 26.3 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 26.3 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

Nucleic acids form the genetic material that is passed from generation to generation in the form of deoxyribonucleic acid (DNA).

- What are the structures of nucleic acids?
- How does the nucleic acid sequence of DNA lead to the production of different proteins?

26.4 Nucleic Acids

Key Concept

Deoxyribonucleic acid (DNA) and ribonucleic acid (RNA) are biopolymers found in the nuclei of cells. DNA carries generic material, while RNA takes part in the translation process that produces proteins. Each nucleotide of a nucleic acid contains a 5-membered sugar, a phosphate group and a nitrogen-containing base. DNA contains the bases adenine, thymine, cytosine and guanine. In RNA, uracil is substituted for thymine. DNA adopts a double helical structure, with bases on opposite strands pairing up via hydrogen bonds. The genetic code of DNA is read by the cell, with each sequence of three base pairs corresponding to a specific amino acid. These amino acids are linked together to make proteins.

Standards

Lesson Objectives

- Identify the functions of DNA and RNA.
- Describe the structure of a single nucleotide and how that structure leads to the overall structure of DNA and RNA.
- Describe how the base sequence of DNA codes for specific amino acids during protein synthesis.

Lesson Vocabulary

- gene: A segment of DNA that carries a code for making a specific polypeptide chain.
- nucleic acid: A biopolymer of deoxyribonucleotides (DNA) or ribonucleotides (RNA) found in the nuclei of cells.
- nucleotide: A molecule that contains a five-carbon sugar, a phosphate group and a nitrogen-containing base called a nucleobase.

Teaching Strategies

Introducing the Lesson

Ask the class what they know about genetics. How are certain characteristics, like blue eyes or big feet passed from one generation to the next? What chemical is responsible for each and every living thing having its own special set of features? They have likely heard of DNA before. This lesson will delve into the specific biochemical structure of DNA and a related molecule called RNA. The understanding of how DNA works was one of the greatest

advancements in science during the twentieth century, but even more is being learned about it every day in labs all over the world.

Common Misconceptions

Students underestimate how incredibly large and massive a DNA strand can be. Have them calculate the molar mass of one nucleotide. Multiply by two to account for the double stranded nature of DNA. Multiply by three to get the "piece" of DNA that is necessary to code for just one amino acid. Assuming a certain protein has 100 amino acids, calculate the mass of just the piece of DNA required to produce one protein. The actual molar masses of DNA molecules reaches well into the millions.

Science Inquiry

Extraction of DNA is a relatively straightforward process that students will enjoy. Strawberries are a convenient source of the genetic material and the link below describes a process for extracting their DNA.

http://www.gs.washington.edu/outreach/dhillon_dnaprocedure.pdf

Differentiated Instruction

Have students practice their decoding skills. Write out some short DNA sequences and have pairs of students decode them by using the table found in the text. Test out the effect of random mutations in the DNA strand by changing out one particular nucleotide for another. How does it affect the protein that would be produced? Other mutations can also result in additions or deletions in the nucleotide sequence.

Enrichment

DNA fingerprinting is a laboratory technique that allows scientists to determine the DNA sequence that is unique to each individual. It is used increasingly in forensics labs to help determine the guilt or innocence of persons that may have been part of a crime scene. Have students research the process of DNA fingerprinting and how it is used by police departments in the analysis of crimes.

Reinforce and Review

Lesson Worksheets

Copy and distribute the Lesson 26.4 worksheets in *CK-12 Chemistry* –*Intermediate Workbook*. Ask students to complete the worksheets to reinforce lesson content.

Lesson Review Questions

Have students answer the Lesson Review Questions at the end of Lesson 26.4 in *CK-12 Chemistry –Intermediate* Flexbook[®] resource.

Points to Consider

DNA must be able to reproduce itself as cells divide so that the genetic material can be passed on. That process is called replication. Next, DNA is transcribed into RNA. Finally a form of RNA is used to make proteins in a process called translation.

- How does DNA replicate itself during cell division?
- What is the mechanism by which proteins are built in the cell from the genetic code of the DNA?