This Instructor's Resource Manual or IRM provides information from and about Nivaldo Tro's *Chemistry: A Molecular Approach*, 2nd edition, from other sources and from the authors' experiences, organized in a manner intended to make the user more productive and effective.

1. Organization of the Chapter Material

Each chapter contains a list of student objectives, organized by section. This portion includes both concepts or ideas and skills or activities.

Chapter 6. Thermochemistry

Student Objectives

6.1 Chemical Hand Warmers

- Define **thermochemistry.**
- Understand the idea of heat exchange as a flow of energy.

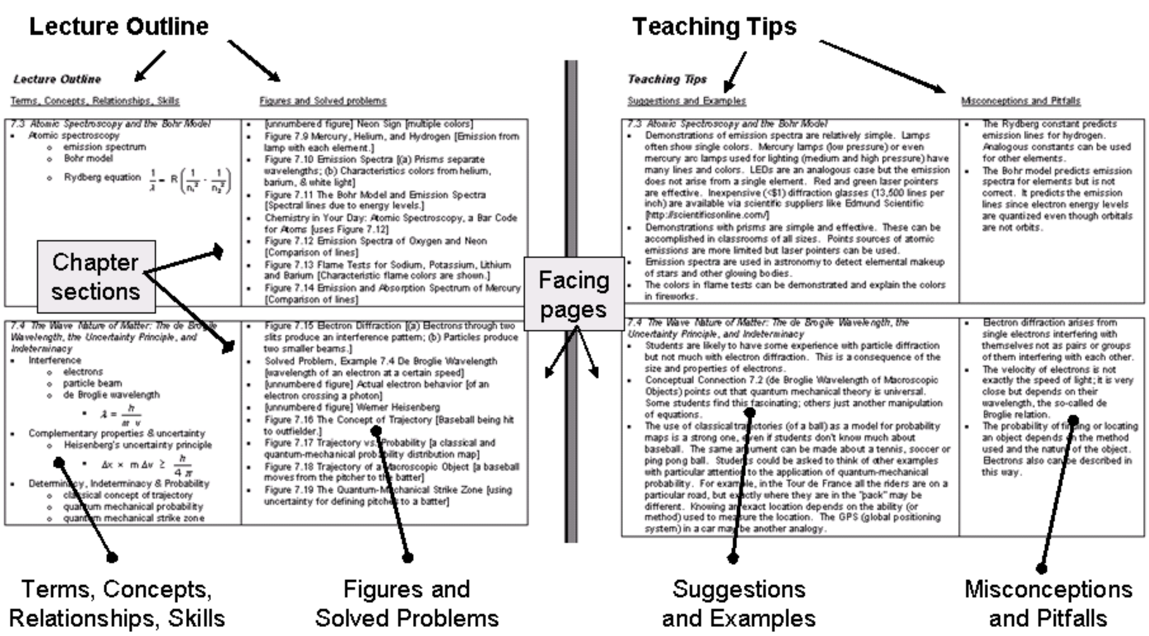
6.2 The Nature of Energy: Key Definitions

- Define **energy**, **work**, **heat**, **kinetic energy**, **thermal energy**, **potential energy**, and **chemical energy**.
- Understand the difference between kinetic and potential energy and know why thermal energy and chemical energy are examples of each, respectively.
- Know the law of conservation of energy.
- Understand the difference between the system and the surroundings.
- Know and convert between the common units of energy: joule, calorie, kilocalorie, kilowatt-hour.

6.3 The First Law of Thermodynamics: There is No Free Lunch

- Define and understand thermodynamics and the **first law of thermodynamics.**
- Understand why a perpetual motion machine violates the first law of thermodynamics.
- Define and understand **internal energy** and **state function.**
- Understand the flow of energy, as work or heat, from the standpoint of the system and the surroundings.
- Understand that energy lost by the surroundings is equal to the energy gained by the system and vice versa.
- Understand the mathematical definition of the first law of thermodynamics in terms of the change in internal energy, heat, and work.
- Understand the sign conventions for heat, work, and the change in internal energy.

For each chapter, section summaries include a four-column organization on facing pages with "Lecture Outline" and "Teaching Tips" portions. One can look across to assess all of the components of a chapter section or look down a column for related items for the entire chapter.



The "Misconceptions and Pitfalls" section is intended to provide or remind instructors of topics that students find challenging. Rather than state what students misunderstand in a negative sense (i.e., "They think an electron orbits around the nucleus like a planet around the sun."), this section contains statements that express the correct concept or idea and not all of the alternative and incorrect versions.

The final part of each chapter includes additional solved problems based on ones from within each chapter. When appropriate, it uses the same problem-solving strategy (i.e., Sort, Strategize, Solve, Check).

Additional Problem for Photon Energy (Example 7.2)	A 1-second pulse of a red laser pointer with a wavelength of 635 nm contains 5.0 mJ of energy. How many photons does it contain?
Sort You are given the wavelength and total energy of a light pulse and asked to find the number of photons it contains.	Given $E_{pulse} = 5.0$ mJ $\lambda = 635$ nm Find number of photons
Strategize In the first part of the conceptual plan, calculate the energy of an individual photon from its wavelength. In the second part, divide the total energy of the pulse by the energy of each photon to get the number of photons in a pulse.	Conceptual Plan $\lambda \rightarrow E_{photon}$ $E = \frac{hc}{\lambda}$ $\frac{E_{pulse}}{E_{photon}}$ = number of photons Relationships Used $E = hc / \lambda$ (Equation 7.3)
Solve Convert wavelength to meters and substitute the values into the energy equation. Concert the energy of the pulse to joules J. Then divide by the energy of a single photon.	Solution $\lambda = 635\ \cancel{nm} \times \frac{10^{-9}\ m}{1\ \cancel{nm}} = 6.35 \times 10^{-7}\ m$ $E_{photon} = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34}\ J \bullet \cancel{s})(3.00 \times 10^{8}\ \frac{\cancel{m}}{\cancel{s}})}{6.35 \times 10^{-7}\ \cancel{m}}$ $= 3.13 \times 10^{-19}\ J$ $5.0\ \cancel{mJ} \times \frac{10^{-3}\ J}{1\ \cancel{mJ}} = 5.0 \times 10^{-3}\ J$ number of photons $= \frac{E_{pulse}}{E_{photon}} = \frac{5.0 \times 10^{-3}\ J}{3.13 \times 10^{-19}\ J} = 1.6 \times 10^{16}$
Check	The magnitude of the answer makes physical sense since the pulse was much larger than the individual photon energy.

Additional Problem for Wavelength of Light for a Transition in the Hydrogen Atom (Example 7.7)	Determine the wavelength of light emitted when an electron in a hydrogen atom makes a transition from an orbital in n = 5 to n = 4.
Sort Your are given the energy levels of an atomic transition and asked to find the wavelength of emitted light.	Given $n = 5 \rightarrow n = 4$ Find λ
Strategize Calculate the energy of the electron in the n = 5 and n = 4 orbitals using Equation 7.7 and subtract to find the difference. The negative value of the difference indicates that the energy is being emitted. Convert the energy value to a wavelength using Equation 7.3.	Conceptual Plan $n = 5, n = 4 \rightarrow \Delta E_{atom}$ $\Delta E = E_5 - E_4$ $\Delta E_{atom} \rightarrow \Delta E_{photon} \rightarrow \lambda$ $\Delta E_{atom} = -E_{photon}$ $E = \frac{hc}{\lambda}$ Relationships Used $E_n = -2.18 \times 10^{-18}\ J\ (1 / n^2)$ (Equation 7.7) $E = hc / \lambda$ (Equation 7.3)
Solve Follow the conceptual plan to solve the problem. Round the answer to three significant figures to reflect the three significant figures in the least precisely known quantity (4750). These conversion factors are all exact and therefore do not limit the number of significant figures.	Solution $\Delta E_{atom} = E_4 - E_5$ $= -2.18 \times 10^{-18}\ J \left(\frac{1}{4^2}\right) - \left[-2.18 \times 10^{-18}\ J \left(\frac{1}{5^2}\right)\right]$ $= -2.18 \times 10^{-18}\ J \left(\frac{1}{4^2} - \frac{1}{5^2}\right)$ $= -4.90 \times 10^{-20}\ J$ $E_{photon} = -\Delta E_{atom} = +4.90 \times 10^{-20}\ J$ $E = \frac{hc}{\lambda}$ or $\lambda = \frac{hc}{E}$ $= \frac{(6.626 \times 10^{-34}\ \cancel{J} \bullet \cancel{s})(3.00 \times 10^{8}\ m/\cancel{s})}{4.90 \times 10^{-20}\ \cancel{J}}$ $= 4.06 \times 10^{-6}\ m$
Check	The units of the answer are correct (m). Acomparison with the values in Figure 7.21 indicates that the answer should be in the infrared region. Figure 7.5 confirms that the answer is within that region.

A considerable amount has been written about the teaching of chemistry—best practices, pedagogical insights, and research-driven insights. References to some of these materials are provided.

2. Additional Resources

2.1 Pedagogy

Effective teaching strategies improve student learning and their experience.

Monographs and Books

- *Survival Manual for the New Instructor*; Diane Bunce and Cinzia Muzzi (eds); Upper Saddle River (NJ): Prentice Hall Publishing, 2004. [19 chapters: "Meant as a quick read to get an overview of the issues that should be addressed as you prepare to teach or as a reference to answer specific questions that have arisen as you teach…"
- *Chemist's Guide to Effective Teaching*; Norbert J. Pienta, Melanie M. Cooper and Thomas J. Greenbowe (eds); Vol 1; Upper Saddle River (NJ): Prentice Hall Publishing, 2004. [16 chapters: "…this unique book is a collection of information, examples, and references on learning theory, teaching methods, and pedagogical issues related to teaching chemistry to college students"]; Vol 2; Upper Saddle River (NJ): Prentice Hall Publishing, 2008. [18 chapters: in press]
- David K. Gosser, Mark S. Cracolice, J.A. Kampmeier, Vicki Roth, Victor S. Strozak, Pratibha Varma-Nelson; *Peer Led Team Learning: A Guidebook*, Upper Saddle River (NJ): Prentice Hall Publishing, 2001. [9 chapters & 3 appendices: "…this unique book explains the theory behind peer-led team learning, offers suggestions for successful implementation (including how to write effective group problems and how to train peer leaders), discusses how to evaluate the success of the program, and answers frequently asked questions"]
- Additional books in the PLTL series are available with specific guidance for General Chemistry, Organic Chemistry, and General, Organic and Biochemistry courses; information about them

(and the ones above) can be found at the Educational Innovation Series website:
http://vig.prenhall.com/catalog/academic/product/0,1144,0130288055,00.html

Chemistry Education Books

- *Chemical Education: Towards Research-based Practice*; John K. Gilbert, Onno De Jong, Rosaria Justi, David F. Treagust, Jan H. van Driel (eds), Dordrecht (Netherlands): Kluwer Academic Publishers, 2002.
- J Dudley Heron, *The Chemistry Classroom: Formulas for Successful Teaching*, Washington (DC): American Chemical Society, 1996.

2.2 Demonstrations

Live demonstrations or even virtual ones available as multimedia enliven the class, provide motivation and interest, and provide a visual or graphical introduction to a topic:

- Instructor's website and resources for Tro book:
  http://vig.prenhall.com/catalog/academic/product/0,1144,0131000659-IS,00.html
- Journal of Chemical Education and Division of Chemical Education sites:
  - JCE Software: http://jchemed.chem.wisc.edu/JCESoft/Programs/index.html
  - JCE Digi-Demos: http://forums.jce.divched.org:8000/JCE/DigiDemos/
- Bassam Shakhashiri, *Chemical Demonstrations: A Handbook for Teachers of Chemistry*; Vol. 1 (1983); Vol. 2 (1985); Vol. 3 (1989); Vol. 4 (1992); Madison (WI): University of Wisconsin Press.
- Lee R. Summerlin and James L. Ealy. *Chemical Demonstrations: A Sourcebook for Teachers*, Vol. 1, 2nd ed., New York: Oxford University Press, 1988.
- Lee R. Summerlin, Christie L. Borgford, and Julie B. Ealy; *Chemical Demonstrations: A Sourcebook for Teachers*, Vol 2, 2nd ed., New York: Oxford University Press, 1988.
- *Classic Chemistry Experiments: One hundred tried and tested experiments*; Kevin Hutchings (compiler), London: Royal Society of Chemistry, 2000.

2.3 Misconceptions

Misconceptions have been characterized and compiled by several scholars:

- Christopher Horton (Assumption College, Worchester, MA) and members of the Modeling Instruction in High School Action Research Team, Arizona State University. 2001-4. *Students Preconceptions and Misconceptions in Chemistry*.
  [85 page PDF file] http://www.daisley.net/hellevator/misconceptions/misconceptions.pdf
- Queens University (Ontario, Canada)
  http://educ.queensu.ca/~science/main/concept/chem/c07/C07CDTL1.htm
- Royal Society of Chemistry [resources: chemistry misconceptions]
  http://www.chemsoc.org/networks/learnnet/misconceptions.htm
- Vanessa Kind (Durham University, Durham, UK) *Beyond Appearances: Students' Misconceptions about Basic Chemical Knowledge* [84 page PDF]

2.4 Molecular model on-line viewers

Some popular plug-ins for browsers or software can be downloaded:

- Molecule viewer lite: www.axiomdiscovery.com/Downloads.htm
- JMOL browser applet: jmol.sourceforge.net/
- RASMOL / Chime plug-in: www.umass.edu/microbio/rasmol/
- JAVA molecular viewer: www.ks.uiuc.edu/Research/jmv/
- MDL Chime: www.umass.edu/microbio/chime/getchime.htm
- Flash molecular viewer: www.tufat.com/s_3d_molecule_viewer.htm

For additional examples, search "molecular model viewer" on the Internet. Many or most of these tools have a somewhat cyclic history of compatibility with computer operating systems and versions of browsers.

Chapter 1. Matter, Measurement, and Problem Solving

Student Objectives

1.1 Atoms and Molecules

- Define **atoms**, **molecules**, and the science of **chemistry**.
- Represent simple molecules (carbon monoxide, carbon dioxide, water, hydrogen peroxide) using spheres as atoms.

1.2 The Scientific Approach to Knowledge

- Define and distinguish between a **hypothesis**, a **scientific law**, and a **theory**.
- Understand the role of experiments in testing hypotheses.
- State and understand the law of mass conservation as an example of scientific law.
- Understand that scientific theories are built from strong experimental evidence and that the term "theory" in science is used much differently than in pop culture.

1.3 The Classification of Matter

- Define **matter** and distinguish between the three main states of matter: solid, liquid, gas.
- Define and understand the difference between **crystalline** and **amorphous** solids.
- Define **mixture**, **pure substance**, **element**, **compound**, **heterogeneous**, and **homogeneous**.
- Differentiate between mixtures and pure substances; elements and compounds; and heterogeneous and homogeneous mixtures.
- Use the scheme on page 7 to classify matter.
- Define and understand the methods of separating mixtures: decantation, distillation, and filtration.

1.4 Physical and Chemical Changes and Physical and Chemical Properties

- Define, recognize, and understand the difference between physical and chemical changes.

1.5 Energy: A Fundamental Part of Physical and Chemical Change

- Define **energy**, **work**, **kinetic energy**, **potential energy**, and **thermal energy**.
- State and understand the law of conservation of energy.

1.6 The Units of Measurement

- Understand the importance of reporting correct units with measurements.
- Know the differences between the three most common sets of units: English system, metric system, and International System (SI).
- Know the SI base units for length, mass, time, and temperature.
- Know the three most common temperature scales (Fahrenheit, Celsius, and Kelvin), the freezing and boiling points of water on each scale, and the relationships between the scales.
- Calculate temperature conversions between each scale.
- Know and use the SI prefix multipliers for powers of ten.
- Know and calculate using the derived units of volume and density.

1.7 The Reliability of a Measurement

- Understand that all measurements have some degree of uncertainty and that the last digit in a measurement is estimated.
- Know how to determine the number of significant figures in a measurement using a set of rules.
- Know how to determine the number of significant figures after calculations.
- Distinguish between accuracy and precision.

1.8 Solving Chemical Problems

- Understand dimensional analysis and know how to use conversion factors.
- Understand the problem-solving strategy: sort, strategize, solve, and check.
- Convert from one unit to another.
- Make order-of-magnitude estimations without using a calculator.
- Rearrange algebraic equations to solve for unknown variables.

Section Summaries

Lecture Outline

- Terms, Concepts, Relationships, Skills
- Figures, Tables, and Solved Examples

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

Chapter 1. Matter, Measurement, and Problem Solving

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples
1.1 Atoms and Molecules • Definitions of atoms, molecules • Interactions of CO and CO_2 with hemoglobin • Composition of water and hydrogen peroxide • Definition of chemistry	• Intro figure: crystal structure of hemoglobin surrounded by CO molecules • Figure 1.1 Binding of Oxygen and Carbon Monoxide to Hemoglobin • unnumbered figures: models of CO_2, H_2O, H_2O_2

1.2 The Scientific Approach to Knowledge
• Definitions of hypothesis, falsifiable, experiments, scientific law, theory
• Scientific method:
  o Observations and experiments lead to hypotheses.
  o More experiments may lead to a law and a theory.
  o A theory explains observations and laws.
	• unnumbered figure: painting of Antoine Lavoisier
	• Figure 1.2 The Scientific Method
	• The Nature of Science: Thomas S. Kuhn and Scientific Revolutions

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls

1.1 Atoms and Molecules
• Chemistry involves a great deal of what can't be seen directly, requiring representations and models.
  o The intro figure shows hemoglobin, but the actual molecule is not a green and blue ribbon.
  o Chemists look at microscopic, macroscopic, and symbolic representations of atoms and molecules interchangeably. If you say "water", you might mean the formula H_2O or a molecular model or a large collection of molecules (e.g., a glass of water). Students need help recognizing which representation to think about when a chemical name is used.	

1.2 The Scientific Approach to Knowledge • Experiments test ideas. They are designed to support a hypothesis or to disprove it. Good scientific hypotheses must be testable or falsifiable. • Theories are developed only through considerable evidence and understanding, even though theories often are cited in popular culture as unproven or untested. • Figure 1.2 shows how the scientific method is cyclic and allows for the refining of ideas. • Conceptual Connection 1.1 Laws and Theories • The box about Thomas Kuhn can help to clear misconceptions of science being completely objective and immutable.	• Theories are *not* as easily dismissible as pop culture suggests. • Scientific knowledge constantly evolves as new information and evidence are gathered.

Chapter 1. Matter, Measurement, and Problem Solving

Lecture Outline

<u>Terms, Concepts, Relationships, Skills</u> — <u>Figures, Tables, and Solved Examples</u>

1.3 The Classification of Matter
• States of matter: their definitions and some of their characteristics
  o gas
  o liquid
  o solid
    ➢ crystalline
    ➢ amorphous
• Classification of Matter
  o pure substance
    ➢ element
    ➢ compound
  o mixture
    ➢ heterogeneous
    ➢ homogeneous
• Separating mixtures
  o decantation
  o distillation
  o filtration
	• Figure 1.3 Crystalline Solid
	• unnumbered figure: illustrations of solid, liquid, and gas phases
	• Figure 1.4 The Compressibility of Gases
	• unnumbered figure: classification of matter
	• Figure 1.5 Separating Substances by Distillation
	• Figure 1.6 Separating Substances by Filtration

1.4 Physical and Chemical Changes and Physical and Chemical Properties • Differences between physical and chemical changes • Examples and classifying changes	• Figure 1.7 Boiling, a Physical Change • Figure 1.8 Rusting, a Chemical Change • Figure 1.9 Physical and Chemical Changes • Example 1.1 Physical and Chemical Changes and Properties

1.5 Energy: A Fundamental Part of Physical and Chemical Change
• Definitions of work and energy
• Classification and types of energy
  o kinetic
    ➢ thermal
  o potential
• Definition and examples of the law of conservation of energy
	• unnumbered figure: illustration of work (physical definition)
	• Figure 1.10 Energy Conversions
	• Figure 1.11 Using Chemical Energy to Do Work

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.3 The Classification of Matter • Properties of matter define its state: gas, liquid, or solid. Temperature is one example, and everyone recognizes steam, water, and ice. Ask for additional examples such as dry ice or liquid nitrogen. • Compressibility is a property that differentiates especially gases from liquids and solids. • The thickened glass at the bottoms of old windows helps students appreciate the amorphous nature of glass. • Conceptual Connection 1.2 The Mass of a Gas • Classifying additional examples of matter, e.g. mayonnaise, Jell-O, and milk, according to the scheme demonstrates some of the challenges. • Students are likely to have varying personal experience with distillation and filtration. Kitchen analogies may be useful: steam condenses on the inside of a pot lid; macaroni and water are poured into a colander; wine is often decanted.	• The differences between the space-filling models from Section 1.1 and the ball-and-stick model of diamond may be missed by some students. • Students may not have experience with elemental forms other than diamond and charcoal.
1.4 Physical and Chemical Changes and Physical and Chemical Properties • Conceptual Connection 1.3 Chemical and Physical Changes	• Boiling (especially) does *not* change a substance's chemical identity. • Confront the confusion that can occur when a physical change accompanies a chemical one: burning liquid gasoline produces gases. (physical or chemical or both?)
1.5 Energy: A Fundamental Part of Physical and Chemical Change • The examples of work being done by a person moving a box and chemical energy ultimately moving the car are consistent and simple. Additional examples using gravitation (very familiar) are straightforward. • Several examples are cited for the law of conservation of energy; ask students to name and describe other forms of energy (solar, mechanical, chemical, electrical) and devices that convert between these forms.	• Work is a form of energy and thus has the same units as energy.

Chapter 1. Matter, Measurement, and Problem Solving

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

1.6 The Units of Measurement
• Loss of Mars Climate Orbiter because of inconsistent units
• Systems of measurement and units
  o English system
  o metric system
  o International System (SI)
• SI base units
  o length: meter
  o mass: kilogram
  o time: second
  o temperature: Kelvin
• Temperature scales and conversions
  o Fahrenheit to Celsius and vice versa
  o Celsius to Kelvin and vice versa
• Derived units
  o volume (cubic meter, cubic centimeter, liter, milliliter)
  o density, mass per unit volume (g/mL, g/cm^3)
	• unnumbered figure: Mars Climate Orbiter
	• unnumbered figures: heights in meters of Empire State Building and basketball player
	• Table 1.1 SI Base Units
	• unnumbered figure: electronic balance
	• Figure 1.12 Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales
	• unnumbered figure: The Celsius Temperature Scale
	• Example 1.2 Converting between Temperature Scales
	• Table 1.2 SI Prefix Multipliers
	• Figure 1.13 The Relationship between Length and Volume
	• Table 1.3 Some Common Units and Their Equivalents
	• Table 1.4 The Density of Some Common Substances at 20 $^{\circ}$C
	• Example 1.3 Calculating Density
	• Chemistry and Medicine: Bone Density

1.7 The Reliability of a Measurement
• Significance and reporting of numerical values
  o estimating measurements
• Counting significant figures or digits
  o nonzero digits
  o interior zeroes
  o leading zeroes
  o trailing zeroes
  o exact numbers
• Significant figures in calculations
  o multiplication and division (fewest significant figures)
  o addition and subtraction (fewest decimal places)
  o rounding (best only after the final step)
• Precision *vs.* accuracy
• Scientific integrity and data reporting
	• unnumbered figures: CO concentration in L.A. county; two tables with different significant figures for the data
	• Figure 1.14 Estimation in Weighing
	• Example 1.4 Reporting the Correct Number of Digits
	• Example 1.5 Determining the Number of Significant Figures in a Number
	• Example 1.6 Significant Figures in Calculations
	• unnumbered figure: accuracy and precision
	• Chemistry in Your Day: Integrity in Data Gathering

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.6 The Units of Measurement • Students are amazed and horrified that NASA could lose an expensive spacecraft because of inconsistent units. • Metric and SI units are unfamiliar to most Americans. That a nickel has a mass of 5 g and that a yard is nearly as long as a meter gives a good frame of reference. • The practical examples of different temperatures on the Celsius scale (unnumbered figure) provide practical reference points. • Several of the large SI unit prefixes (mega, giga, tera) are already familiar from memory capacity in computers. • Conceptual Connection 1.4 Density • The Chemistry and Medicine box on bone density provides an open-ended conceptual question about designing an experiment to measure bone density; this may be good for a brief in-class discussion.	• A common misconception is that 100 cm^3 is equal to 1 m^3. • Some students initially are confused that density can be used as a conversion factor even when the units are inverted.

1.7 The Reliability of a Measurement • Use a 400-mL beaker and a 100-mL graduated cylinder to measure quantities of water. Make the point about the importance of estimating measurements. Add the quantities of water together and ask the students to calculate the final volume...to the correct precision. • Two tables present air quality data (with different precision) that might appear in a newspaper or other publication. Initiate a discussion of the certainty of digits in reported data. • Water-quality standards have evolved substantially since the advent of instrumental methods for quantitative analysis. Ask the question: Does zero mean that a particular analyte is not present? • The number on a calculator display requires interpretation; only the user knows the certainty of the values entered. • A discussion about why integrity in data reporting is particularly important in science is appropriate. It should point out that scientists report how they did the experiments so others can try to repeat and verify the work. Use recent examples from the media.	• Students presume that calculators are flawless but forget that calculators do only what the user dictates.

Chapter 1. Matter, Measurement, and Problem Solving

Lecture Outline

<u>Terms, Concepts, Relationships, Skills</u>	<u>Figures, Tables, and Solved Examples</u>

1.8 Solving Chemical Problems
• Converting from one unit to another
  o dimensional analysis
  o multiple approaches to any problem
• General problem-solving strategy
  o sort
  o strategize
  o solve
  o check
• Calculations using units raised to a power
• Order-of-magnitude estimations
• Using equations
	• Example 1.7 Unit Conversion
	• Example 1.8 Unit Conversion
	• Example 1.9 Unit Conversions Involving Units Raised to a Power
	• Example 1.10 Density as a Conversion Factor
	• Example 1.11 Problems with Equations
	• Example 1.12 Problems with Equations

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.8 Solving Chemical Problems • General chemistry classes at most schools have students with a wide range of math skills. A quick review of algebra may be useful. • Emphasize that watching an instructor work problems is not nearly as effective as working those same problems on one's own. Give students time to work a problem or two in class; allow them to work in small groups. • Emphasize the good practice of writing units and keeping track of units in every calculation. Simple dimensional analysis prevents many headaches throughout the year of general chemistry. • Promote estimation as part of the problem solving model. Tell the students to ask themselves, "Does this answer make sense?" Reduce the reliance on blindly entering numbers into a calculator and transcribing whatever answer comes up. • Cognitive load theory says that a person can remember 7–9 items in short-term memory. A problem loaded with unit conversions, spurious facts, and many steps does not test a person's understanding of an underlying idea or concept. It becomes a measure of cognitive ability outside the realm of chemistry.	• Students often want to follow one particular "recipe" to solve one particular kind of problem.

Procedure for Solving Unit Conversion Problems	**Additional Problem (Example 1.7 Unit Conversion)**
	Convert 1.76 miles to meters.
Sort Begin by sorting the information in the problem into *Given* and *Find*.	**Given** 1.76 mi **Find** m
Strategize Devise a *conceptual plan* for the problem. Begin with the *given* quantity and symbolize each conversion step with an arrow. Below each arrow, write the appropriate conversion factor for that step. Focus on the units. The conceptual plan should end at the *find* quantity and its units. In these examples, the other information needed consists of relationships between the various units as shown.	**Conceptual Plan** $\text{mi} \xrightarrow[\frac{1\text{ km}}{0.6214\text{ mi}}]{} \text{km} \xrightarrow[\frac{1000\text{ m}}{1\text{ km}}]{} \text{m}$ **Relationships Used** 1 km = 0.6214 mi 1 km = 1000 m (These conversion factors are from Tables 1.2 and 1.3.)
Solve Follow the conceptual plan. Begin with the *given* quantity and its units. Multiply by the appropriate conversion factor(s), cancelling units, to arrive at the *find* quantity. Round the answer to the correct number of significant figures by following the rules in Section 1.7. Remember that exact conversion factors do not limit significant figures.	**Solution** $1.76\ \cancel{\text{mi}} \times \frac{1\ \cancel{\text{km}}}{0.6214\ \cancel{\text{mi}}} \times \frac{1000\text{ m}}{1\ \cancel{\text{km}}} = 2832.31\text{ m}$ 2832.31 m = 2830 m
Check Check your answer. Are the units correct? Does the answer make physical sense?	The units (m) are correct. The magnitude of the answer (2830) makes physical sense since a meter is a much smaller unit than a mile.

Additional Problem for Unit Conversion Involving Units Raised to a Power (Example 1.9)	Calculate the number of cubic meters of concrete necessary to support a deck if each of 14 concrete piers require 4750 cubic inches.
Sort Begin by sorting the information in the problem into *Given* and *Find*.	**Given** 14 piers, 4750 in^3 **Find** m^3
Strategize Write a *conceptual plan* for the problem. Begin with the *given* information and devise a path to the information that you are asked to find. Notice that for cubic units, the conversion factors must be cubed.	**Conceptual Plan** piers → in^3 → m^3 14 piers $\left(\frac{1\ m}{39.37\ in}\right)^3$ **Relationships Used** 1 m = 39.37 in (Conversion factor from Table 1.3) 1 pier = 4750 in^3 (Given)
Solve Follow the conceptual plan to solve the problem. Round the answer to three significant figures to reflect the three significant figures in the least precisely known quantity (4750). These conversion factors are all exact and therefore do not limit the number of significant figures.	**Solution** $14\ \cancel{piers} \times \frac{4750\ \cancel{in^3}}{1\ \cancel{pier}} \times \frac{(1\ m)^3}{(39.37\ \cancel{in})^3} = 1.0897\ m^3$ $1.0897\ m^3 = 1.09\ m^3$
Check	The units of the answer are correct and the magnitude makes sense. The unit meters is larger than inches, so cubic meters are much larger than cubic inches.

Additional Problem for Density as a Conversion Factor (Example 1.10)	An experimental automobile has a 100.0 liter fuel tank filled with ethanol. How many pounds does the fuel add to the mass of the car?
Sort Begin by *sorting* the information in the problem into *Given* and *Find*.	**Given** 100.0 L **Find** lb
Strategize Devise a *conceptual plan* by beginning with the *given* quantity, in this case the volume in liters (L). The overall goal of this problem is to find the mass. You can convert between volume and mass using density (g/cm^3). However, you must first convert the volume to cm^3. Once you have converted the volume to cm^3, use the density to convert to g. Finally, convert g to lb.	**Conceptual Plan** $L \rightarrow mL \rightarrow cm^3 \rightarrow g \rightarrow lb$ $\frac{1000\ mL}{1\ L}$ $\frac{1\ cm^3}{1\ mL}$ $\frac{0.789\ g}{1\ cm^3}$ $\frac{1\ lb}{453.59\ g}$ **Relationships Used** 1000 mL = 1 L 1 mL = 1 cm^3 *d* (ethanol) = 0.789 g/cm^3 1 lb = 453.59 g (These conversion factors are from Tables 1.2, 1.3 & 1.4.)
Solve Follow the conceptual plan to solve the problem. Round the answer to three significant figures to reflect the three significant figures in the density.	**Solution** $100\ \cancel{L} \times \frac{1000\ \cancel{mL}}{1\ \cancel{L}} \times \frac{1\ \cancel{cm^3}}{1\ \cancel{mL}} \times \frac{0.789\ \cancel{g}}{1\ \cancel{cm^3}} \times \frac{1\ lb}{453.59\ \cancel{g}}$ $= 173.94\ lb$ 173.94 lb = 174 lb
Check	The units of the answer (lb) are correct. The magnitude of the answer (174) makes physical sense since a liter of water has a mass of 1 kilogram or about 2.2 pounds; 100 liters of water is about 220 lbs. Ethanol has a lower density than water (about 80% or 8/10).

Additional Problem for Solving Problems Involving Equations (Example 1.12)	What is the mass in grams of an ice cube that is 1.1 inches per side?
Sort Begin by sorting the information in the problem into *Given* and *Find*.	**Given** $l = 1.1$ in **Find** g
Strategize Write a *conceptual plan* for the problem. Focus on the equation(s). The conceptual plan shows how the equation takes you from the *given* quantity (or quantities) to the *find* quantity. The conceptual plan may have several parts, involving other equations or required conversions. In these problems, you must use the geometrical relationships given in the problem as well as the definition of density.	**Conceptual Plan** $l \rightarrow V$ $V = l^3$ $\text{in}^3 \xrightarrow{\left(\frac{2.54\text{ cm}}{1\text{ in}}\right)^3} \text{cm}^3 \xrightarrow{\frac{0.917\text{ g}}{1\text{ cm}^3}} \text{g}$ **Relationships Used** $V = l^3$ [volume of a cube with a length of l] 2.54 cm = 1 in d (ice) = 0.917 g/cm^3 (These conversion factors are from Tables 1.3 and 1.4.)
Solve Follow the conceptual plan. Solve the equation(s) for the find quantity. Gather each of the quantities that must go into the equation in the correct units. (Convert to the correct units if necessary.) Substitute the numerical values and their units into the equation(s) and compute the answer. Round the answer to the correct number of significant figures.	**Solution** $V = (1.1\text{ in})^3 = 1.331\text{ in}^3$ $1.331\ \cancel{\text{in}^3} \times \left(\frac{2.54\ \cancel{\text{cm}}}{1\ \cancel{\text{in}}}\right)^3 \times \frac{0.917\text{ g}}{1\ \cancel{\text{cm}^3}} = 20.0008\text{ g}$ 20.0008 g = 20. g
Check	The units (g) are correct. The magnitude of the answer (20.) seems to make physical sense.

Chapter 2. Atoms and Elements

Student Objectives

2.1 Imaging and Moving Individual Atoms

- Describe scanning tunneling microscopy (STM) and how atoms are imaged on surfaces.
- Define **atom** and **element**.

2.2 Early Ideas about the Building Blocks of Matter

- Describe the earliest definitions of atoms and matter (Greeks).
- Know that greater emphasis on observation and the development of the scientific method led to the scientific revolution.

2.3 Modern Atomic Theory and the Laws That Led to It

- State and understand the law of conservation of mass (also from Section 1.2).
- State and understand the law of definite proportions.
- State and understand the law of multiple proportions.
- Know the four postulates of Dalton's atomic theory.

2.4 The Discovery of the Electron

- Describe J. J. Thomson's experiments with the cathode ray tube and understand how they provide evidence for the electron.
- Describe Robert Millikan's oil-drop experiment and understand how it enables measurement of the charge of an electron.

2.5 The Structure of the Atom

- Define **radioactivity**, **nucleus**, **proton**, and **neutron**.
- Understand Thomson's plum-pudding model and how Ernest Rutherford's gold-foil experiment refuted it by giving evidence for a nuclear structure of the atom.

2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms

- Define **atomic mass unit**, **atomic number**, and **chemical symbol**.
- Recognize chemical symbols and atomic numbers on the periodic table.
- Define **isotope**, **mass number**, and **natural abundance**.
- Determine the number of protons and neutrons in an isotope using the chemical symbol and the mass number.
- Define **ion**, **anion**, and **cation**.
- Understand how ions are formed from elements.

2.7 Finding Patterns: The Periodic Law and the Periodic Table

- Define the **periodic law**.
- Know that elements with similar properties are placed into columns (called groups) in the periodic table.
- Define and distinguish between metals, nonmetals, and metalloids.
- Identify main-group and transition elements on the periodic table.
- Know the general properties of elements in some specific groups: noble gases, alkali metals, alkaline earth metals, and halogens.
- Know and understand the rationale for elements that form ions with predictable charges.

2.8 Atomic Mass: The Average Mass of an Element's Atoms

- Calculate atomic mass from isotope masses and natural abundances.
- Define **mass spectrometry** and understand how it can be used to measure mass and relative abundance.

2.9 Molar Mass: Counting Atoms by Weighing Them

- Understand the relationship between mass and count of objects such as atoms.
- Define **mole** and **Avogadro's number**.
- Calculate and interconvert between number of moles and atoms.
- Calculate and interconvert between number of moles and mass.

Section Summaries

Lecture Outline

- Terms, Concepts, Relationships, Skills
- Figures, Tables, and Solved Examples

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

Chapter 2. Atoms and Elements

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples
2.1 Imaging and Moving Individual Atoms • Description of scanning tunneling microscopy (STM) • Introduction to macroscopic and microscopic perspectives. • Definitions of atom and element.	• Intro figure: tip of an STM moving across a surface • Figure 2.1 Scanning Tunneling Microscopy • Figure 2.2 Imaging Atoms
2.2 Early Ideas about the Building Blocks of Matter • History of chemistry from antiquity (~450 bc) • Scientific revolution (1400s-1600s)	

2.3 Modern Atomic Theory and the Laws That Led to It
• Law of conservation of mass
  o Matter is neither created nor destroyed.
  o Atoms at the start of a reaction may recombine to form different compounds, but all atoms are accounted for at the end.
  o Mass of reactants = mass of products.
• Law of definite proportions
  o Different samples of the same compound have the same proportions of constituent elements independent of sample source or size.
• Law of multiple proportions
• John Dalton's atomic theory
	• unnumbered figure: models and photos of Na and Cl_2 forming NaCl
	• Example 2.1 Law of Definite Proportions
	• unnumbered figure: models of CO and CO_2 illustrating the law of multiple proportions
	• Example 2.2 Law of Multiple Proportions
	• Chemistry in Your Day: Atoms and Humans

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.1 Imaging and Moving Individual Atoms • Other STM images can be found readily on the Internet. • It is useful to reiterate the analogies about size; the one used in the chapter compares an atom to a grain of sand and a grain of sand to a large mountain range.	• STM is not actually showing images of atoms like one might imagine seeing with a light microscope. • Atoms are not colored spheres; the images use color to distinguish different atoms.

2.2 Early Ideas about the Building Blocks of Matter • The view of matter as made up of small, indestructible particles was ignored because more popular philosophers like Aristotle and Socrates had different views. • Leucippus and Democritus may have been proven correct, but they had no more evidence for their ideas than Aristotle did. • Observations and data led scientists to question models; the scientific method promotes the use of a cycle of such inquiry.	• Theories are not automatically accepted and may be unpopular for long periods of time. • Philosophy and religion can be supported by arguments; science requires that theories be testable and therefore falsifiable.

2.3 Modern Atomic Theory and the Laws That Led to It • That matter is composed of atoms grew from experiments and observations. • Conceptual Connection 2.1 The Law of Conservation of Mass • Investigating the law of definite proportions requires preparing or decomposing a set of pure samples of a compound like water. • Investigating the law of multiple proportions requires preparing or decomposing sets of pure samples from related compounds like NO, NO_2, and N_2O_5. • Conceptual Connection 2.2 The Laws of Definite and Multiple Proportions	• Measurements to establish early atomic theories were performed at the macroscopic level. The scientists observed properties for which they could collect data (e.g., mass or volume).

Chapter 2. Atoms and Elements

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

2.4 The Discovery of the Electron
• Thomson's cathode-ray tube experiments
   o High voltage produced a stream of particles that traveled in straight lines.
   o Each particle possessed a negative charge.
   o Thomson measured the charge-to-mass ratio of the electron.
• Millikan's oil-drop experiments
   o Oil droplets received charge from ionizing radiation.
   o Charged droplets were suspended in an electric field.
   o The mass and charge of each oil drop was used to calculate the mass and charge of a single electron.
	• Figure 2.3 Cathode Ray Tube
	• unnumbered figure: properties of electrical charge
	• Figure 2.4 Thomson's Measurement of the Charge-to-Mass Ratio of the Electron
	• Figure 2.5 Millikan's Measurement of the Electron's Charge

2.5 The Structure of the Atom
• Thomson's plum-pudding model: negatively charged electrons in a sea of positive charge
• Radioactivity
   o Alpha decay provides the alpha particles for Rutherford's experiment.
• Rutherford's experiment
   o Alpha particles directed at a thin gold film deflect in all directions, including back at the alpha source.
   o Only a concentrated positive charge could cause the alpha particles to bounce back.
• Rutherford's nuclear theory
   o most mass and all positive charge contained in a small nucleus
   o most of atom by volume is empty space
   o protons: positively charged particles
   o neutral particles with substantial mass also in nucleus
	• unnumbered figure: plum-pudding model
	• Figure 2.6 Rutherford's Gold Foil Experiment
	• Figure 2.7 The Nuclear Atom
	• unnumbered figure: scaffolding and empty space

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.4 The Discovery of the Electron • Review the attraction, repulsion, and additivity of charges. • Discuss the physics of electric fields generated by metal plates. • A demonstration of a cathode ray tube will help students better understand Thomson's experiments. • Demonstrate how Millikan's calculation works and why he could determine the charge of a single electron.	• Millikan did not measure the charge of a single electron; he measured the charge of a number of electrons and deduced the charge of a single electron.

2.5 The Structure of the Atom • It may be useful to give a brief description of radioactivity. Rutherford's experiment makes more sense if one knows some properties of the alpha particle and from where it comes. • Thomson identified electrons and surmised the existence of positive charge necessary to form a neutral atom. The plum-pudding model is the simplest way to account for the observations. • The figure about scaffolding supports discussion about an atom being mostly empty space but still having rigidity and strength in the macroscopic view. This is another example of apparent differences between the microscopic and macroscopic properties.	• Students often don't understand the *source* of alpha particles in Rutherford's experiments.

Chapter 2. Atoms and Elements

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms
• Properties of subatomic particles
    o atomic mass units (amu)
        ➢ proton, neutron: ~1 amu
        ➢ electron: ~0.006 amu
    o charge
        ➢ relative value: –1 for electron, +1 for proton
        ➢ absolute value: 1.6×10^{-19} C
• Atomic number (number of protons): defining characteristic of an element
• Isotope: same element, different mass (different number of neutrons)
• Ion: atom with nonzero charge
    o anion: negatively charged (more electrons)
    o cation: positively charged (fewer electrons)
	• unnumbered figure: baseball
	• Table 2.1 Subatomic Particles
	• unnumbered figure: lightning and charge imbalance
	• Figure 2.8 How Elements Differ
	• Figure 2.9 The Periodic Table
	• unnumbered figure: portrait of Marie Curie
	• Example 2.3 Atomic Numbers, Mass Numbers, and Isotope Symbols
	• Chemistry in Your Day: Where Did Elements Come From?

2.7 Finding Patterns: The Periodic Law and the Periodic Table
• Periodic law and the periodic table
    o generally arranged by ascending mass
    o recurring, periodic properties; elements with similar properties arranged into columns: groups (or families)
• Major divisions of the periodic table
    o metals, nonmetals, metalloids
    o main-group elements, transition elements
• Groups (families)
    o noble gases (group 8A)
    o alkali metals (group 1A)
    o alkaline earth metals (group 2A)
    o halogens (group 7A)
• Ions with predictable charges: based on stability of noble-gas electron count
    o group 1A: 1+
    o group 2A: 2+
    o group 3A: 3+
    o group 5A: 3–
    o group 6A: 2–
    o group 7A: 1–
	• unnumbered figure: discovery of the elements
	• Figure 2.10 Recurring Properties
	• Figure 2.11 Making a Periodic Table
	• unnumbered figure: stamp featuring Dmitri Mendeleev
	• Figure 2.12 Metals, Nonmetals, and Metalloids
	• Figure 2.13 The Periodic Table: Main-Group and Transition Elements
	• unnumbered figure: the alkali metals
	• unnumbered figure: the halogens
	• Figure 2.14 Elements That Form Ions with Predictable Charges
	• Example 2.4 Predicting the Charge of Ions
	• Chemistry and Medicine: The Elements of Life
	• Figure 2.15 Elemental Composition of Humans (by Mass)

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.6 Subatomic Particles: Protons, Neutrons, and Electrons in Atoms • The analogy of the baseball and a grain of rice to a proton and an electron is meant to illustrate the difference in mass but not size. • Electrical charge can be demonstrated with static electricity. Two balloons charged with wool or human hair will repel each other. • Names of elements come from various sources. Tom Lehrer's "Element Song" can be found on the Internet. • Isotopic abundances are invariant in typical lab-sized samples because of such large numbers of atoms. • Conceptual Connection 2.3 The Nuclear Atom, Isotopes, and Ions • The history of chemistry involves considerable cultural and gender diversity. Examples include both Lavoisiers (French), Dalton (English), Thomson (English), Marie Curie (Polish/French), Mendeleev (Russian), Millikan (American), Robert Boyle (Irish), Amedeo Avogadro (Italian). • The Chemistry in Your Day box gives a broad description of the origin of atoms.	• Students sometimes confuse the mass number as being equal to the number of neutrons, not the number of neutrons plus the number of protons. • Students logically (but mistakenly) presume that the mass of an isotope is equal to the sum of the masses of the protons and neutrons in that isotope.
2.7 Finding Patterns: The Periodic Law and the Periodic Table • Other displays of the periodic table can be found in journals (Schwartz, *J. Chem. Educ*. **2006**, *83*, 849; Moore, *J. Chem. Educ*. **2003**, *80*, 847; Bouma, *J. Chem. Educ.* **1989**, *66*, 741), books, and on the Internet. • Periodic tables are arranged according to the periodic law but can compare many features, e.g. phases of matter, sizes of atoms, and common ions. These are presented as a series of figures in the text. • Chemistry and Medicine: The Elements of Life provides an opportunity to relate the topics to everyday life. Some of the other elements in the figure and table represent trace minerals that are part of good nutrition. The periodic law accounts for why some are necessary and others are toxic.	• The periodic table is better at predicting microscopic properties, though macroscopic properties are also often illustrated.

Chapter 2. Atoms and Elements

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples

2.8 Atomic Mass: The Average Mass of an Element's Atoms
• Average atomic mass is based on natural abundance and isotopic masses.
• Mass spectrometry
   o atoms converted to ions and deflected by magnetic fields to separate by mass
   o output data: relative mass *vs*. relative abundance
	• unnumbered figure: periodic table box for Cl
	• Example 2.5 Atomic Mass
	• Figure 2.16 The Mass Spectrometer
	• Figure 2.17 The Mass Spectrum of Chlorine

Terms, Concepts, Relationships, Skills	Figures, Tables, and Solved Examples
2.9 Molar Mass: Counting Atoms by Weighing Them • Mole concept and Avogadro's number • Converting between moles and number of atoms • Converting between mass and number of moles	• unnumbered figure: pennies containing ~1 mol of Cu • unnumbered figure: 1 tbsp of water contains ~1 mol of water • Example 2.6 Converting between Number of Moles and Number of Atoms • unnumbered figure: relative sizes of Al, C, He • unnumbered figure: balance with marbles and peas • Example 2.7 Converting between Mass and Amount (Number of Moles) • Example 2.8 The Mole Concept–Converting between Mass and Number of Atoms • Example 2.9 The Mole Concept

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
2.8 Atomic Mass: The Average Mass of an Element's Atoms • The masses of isotopes must be reconciled with an element having only whole number quantities of protons and neutrons; the values should be nearly integral since the mass of electrons is so small. • Mass spectrometry is an effective way to demonstrate where values of natural abundance are obtained.	• Students are tempted to calculate average atomic mass by adding together isotopic masses and dividing by the number of isotopes. • Atomic mass on the periodic table is usually not integral even though elements have only whole numbers of protons and neutrons.

2.9 Molar Mass: Counting Atoms by Weighing Them • Review the strategy for solving numerical problems: sort, strategize, solve, check. • Estimating answers is an important skill; the number of atoms will be very large (i.e. some large power of ten) even from a small mass or small number of moles. • Conceptual Connection 2.4 Avogadro's Number • Conceptual Connection 2.5 The Mole	• Many students are intimidated by estimating answers in calculations involving powers of ten.

Additional Problem for Converting between Number of Moles and Number of Atoms (Example 2.6)	Calculate the number of moles of iron in a sample that has 3.83 x 10^{23} atoms of iron.
Sort You are given a number of iron atoms and asked to find the amount of iron in moles.	**Given** 3.83 x 10^{23} Fe atoms **Find** mol Fe
Strategize Convert between number of atoms and number of moles using Avogadro's number.	**Conceptual Plan** atoms → mol $\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ **Relationships Used** 6.022 x 10^{23} = 1 mol (Avogadro's number)
Solve Follow the conceptual plan. Begin with 3.83 x 10^{23} Fe atoms and multiply by the ratio that equates moles and Avogadro's number.	**Solution** $3.83 \times 10^{23} \text{ Fe atoms} \times \frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}} = 0.636 \text{ mol Fe}$
Check	The sample was smaller than Avogadro's number so the answer should be a fraction of a mole. The value of the sample has 3 significant figures, and the answer is provided in that form.

Additional Problem for Converting between Mass and Number of Moles (Example 2.7)	Calculate the number of grams of silver in an American Silver Eagle coin that contains 0.288 moles of silver.
Sort You are given the amount of silver in moles and asked to find the mass of silver.	**Given** 0.288 mol Ag **Find** g Ag
Strategize Convert amount (in moles) to mass using the molar mass of the element.	**Conceptual Plan** $\text{mol Ag} \xrightarrow[\text{1 mol Ag}]{\text{107.87 g Ag}} \text{g Ag}$ **Relationships Used** 107.87 g Ag = 1 mol Ag
Solve Follow the conceptual plan to solve the problem. Start with 0.288 mol, the given number, and multiply by the molar mass of silver.	**Solution** $0.288\ \cancel{\text{mol Ag}} \times \frac{107.87\text{ g Ag}}{1\ \cancel{\text{mol Ag}}} = 31.07\text{ g Ag}$ 31.07 g = 31.1 g Ag
Check	The magnitude of the answer makes sense since we started with an amount smaller than a mole. The molar amount and answer both have 3 significant figures.

Additional Problem for the Mole Concept—Converting between Mass and Number of Atoms (Example 2.8)	What mass of iron (in grams) contains 1.20×10^{22} atoms of Fe? A paperclip contains about that number of iron atoms.
Sort You are given a number of iron atoms and asked to find the mass of Fe.	**Given** 1.20×10^{22} Fe atoms **Find** g Fe
Strategize Convert the number of Fe atoms to moles using Avogadro's number. Then convert moles Fe into grams of iron using the molar mass of Fe.	**Conceptual Plan** Fe atoms → mol Fe → g Fe $\frac{1 \text{ mol Fe}}{6.022 \times 10^{23} \text{ Fe atoms}}$ $\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ **Relationships Used** 6.022×10^{23} = 1 mol (Avogadro's number) 55.85 g Fe = 1 mol Fe
Solve Follow the conceptual plan to solve the problem. Begin with 1.20 x 10^{22} atoms of Fe, multiply by the ratio derived from Avogadro's number, and finally multiply by the atomic mass of Fe.	**Solution** $1.20 \times 10^{22}\ \cancel{\text{Fe atoms}} \times \frac{1\ \cancel{\text{mol Fe}}}{6.022 \times 10^{23}\ \cancel{\text{Fe atoms}}} \times \frac{55.85 \text{ g Fe}}{1\ \cancel{\text{mol Fe}}}$ = 1.11 g Fe
Check	The units and magnitude of the answer make sense. The sample is smaller than a mole. The number of atoms and mass both have 3 significant figures.

Additional Problem for the Mole Concept (Example 2.9)	An iron sphere contains 8.55×10^{22} iron atoms. What is the radius of the sphere in centimeters? The density of iron is 7.87 g/cm^3.
Sort You are given the number of iron atoms in a sphere and the density of iron. You are asked to find the radius of the sphere.	**Given** 8.55×10^{22} Fe atoms d = 7.87 g/cm^3 **Find** radius (r) of a sphere
Strategize The critical parts of this problem are density, which relates mass to volume, and the mole, which relates number of atoms to mass: (1) Convert from the number of atoms to the number of moles using Avogadro's number; (2) Convert from the number of moles to the number of grams using the molar mass of iron; (3) Convert from mass to volume using the density of iron; (4) Find the radius using the formula for the volume of a sphere.	**Conceptual Plan** Fe atoms → mol Fe → g Fe → V (cm^3) $\frac{1\text{ mol Fe}}{6.022\times10^{23}\text{ Fe atoms}}$ $\frac{55.85\text{ g Fe}}{1\text{ mol Fe}}$ $\frac{1\text{ cm}^3}{7.87\text{ g Fe}}$ V (cm^3) → r (cm) $V = \frac{4}{3}\pi r^3$ **Relationships Used** 6.022 x 10^{23} = 1 mol (Avogadro's number) 55.85 g Fe = 1 mol Fe d (density of Fe) = 7.87 g/cm^3 V = 4/3 πr^3 [volume of a sphere with a radius of r]
Solve Follow the conceptual plan to solve the problem. Begin with 8.55 x 10^{22} Fe atoms and convert to moles, then to grams and finally to a volume in cm^3. Solve for the radius using the rearranged equation.	**Solution** $8.55\times10^{22}\ \cancel{\text{atoms}} \times \frac{1\ \cancel{\text{mol Fe}}}{6.022\times10^{23}\ \cancel{\text{atoms}}} \times \frac{55.85\ \cancel{\text{g Fe}}}{1\ \cancel{\text{mol Fe}}} \times \frac{1\text{ cm}^3}{7.87\ \cancel{\text{g Fe}}} = 1.00757\text{ cm}^3$ $r = \sqrt[3]{\frac{3\text{ V}}{4\ \pi}} = \sqrt[3]{\frac{3\times1.00757\text{ cm}^3}{4\ \pi}} = 0.622\text{ cm}$
Check	The units (cm) are correct and the magnitude of the answer makes sense compared with previous problems.

Chapter 3. Molecules, Compounds, and Chemical Equations

Student Objectives

3.1 Hydrogen, Oxygen, and Water

- Know some chemical and physical properties of H_2, O_2, and H_2O.
- Know and understand that compounds, e.g. NaCl, are different from the elements, e.g. Na and Cl_2, from which they are composed.

3.2 Chemical Bonds

- Define and understand the difference between **ionic** and **covalent bonds**.
- Describe and understand the formation of an ionic compound from its elements.
- Describe and understand the sharing of electrons in a covalent bond.

3.3 Representing Compounds: Chemical Formulas and Molecular Models

- Define and understand **empirical formula**, **molecular formula**, and **structural formula**.
- Write the empirical formula, molecular formula, and structural formula for simple molecules.
- Recognize and understand the differences between ball-and-stick models and space-filling models.
- Recognize and identify characteristic colors for elements in molecular models.

3.4 An Atomic-Level View of Elements and Compounds

- Identify elements as atomic or molecular.
- Differentiate between atomic or molecular elements and ionic or molecular compounds.
- Know and understand that ionic compounds are composed of formula units and not discrete molecules.
- Know and understand that covalent compounds tend to exist as discrete molecules.
- Know and understand that a polyatomic ion is composed of atoms that are covalently bound to each other.

3.5 Ionic Compounds: Formulas and Names

- Know that ionic compounds are ubiquitous in the Earth's crust as minerals.
- Know and understand the rules for writing formulas for ionic compounds.
- Write formulas for ionic compounds using the charges of the ions and the principle of electrical neutrality.
- Know and understand the rules for naming ionic compounds.
- Write names from formulas and formulas from names of ionic compounds.

3.6 Molecular Compounds: Formulas and Names

- Know and understand the rules for naming molecular compounds.
- Write names from formulas and formulas from names of molecular compounds.
- Write names and formulas for binary acids and oxyacids.

3.7 Formula Mass and the Mole Concept for Compounds

- Define **formula mass** (a.k.a. molecular weight, molecular mass) and **molar mass** for a compound.
- Understand and calculate the molar mass of a compound.
- Calculate and interconvert between mass, moles, and molecules of a compound.