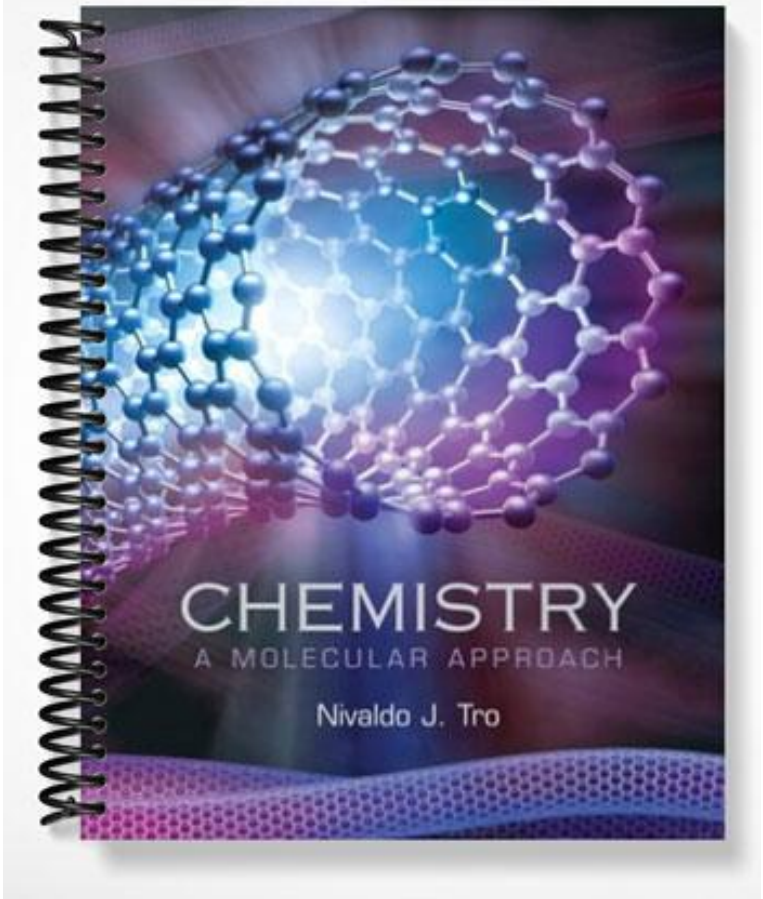
SOLUTIONS MANUAL



Chapter 1. Matter, Measurement, and Problem Solving

Objectives

1.1 Atoms and Molecules

- Define atoms, molecules, and chemistry.
- Draw simple molecules (carbon monoxide, carbon dioxide, water, hydrogen peroxide) using atoms as spheres.

1.2 The Scientific Approach to Knowledge

- Define and distinguish between a hypothesis, theory, and scientific law.
- Define experiment and falsifiable.
- State and understand the law of conservation of mass.
- State and understand Dalton's atomic theory.

1.3 The Classification of Matter

- Define matter, states of matter, gases, liquids, solids.
- Know the properties that differentiate the states of matter.
- Define and understand the difference between crystalline and amorphous solids.
- Define mixture, pure substance, element, compound, heterogeneous, homogeneous.
- Differentiate between pure substance and mixture; element and compound; heterogeneous and homogeneous mixture.
- Classify matter according to the scheme (p 9).
- Define the methods of separating mixtures—decanting, distilling, filtering (decantation, distillation, filtration).

1.4 Physical and Chemical Changes and Chemical Properties

- Define and understand the difference between physical and chemical changes.
- Provide examples of and classify physical and chemical changes.

1.5 Energy: A Fundamental Part of Physical and Chemical Change

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

1.6 The Units of Measurement

- Define system of measurement, English system, metric system, International System (SI)
- Know the SI base units for length, mass, time, temperature, amount.
- Know the temperature scales (Fahrenheit, Celsius, Kelvin), the freezing and boiling points of water on each scale, and the relationships between them.
- Calculate and understand conversions between each scale.
- Know the prefixes that define powers of ten in scientific notation.
- Know and calculate derived units of volume and density.

1.7 The Reliability of a Measurement

- Know the significance of how numerical values are reported.
- Determine the number of significant figures or digits using a set of rules.
- Determine and predict significant figures in calculations.
- Define and understand the difference between precision and accuracy.

1.8 Solving Chemical Problems

- Know and understand the process of dimensional analysis.
- Understand the textbook problem solving plan—sort, strategize, solve, check.
- Calculate and understand unit conversions.
- Calculate with and understand the use of powers.
- Calculate with and understand the use of equations.

Section Summaries

Lecture Outline

- Terms, Concepts, Relationships, Skills
- Figures and Solved Problems

Teaching Tips

- Suggestions and Examples
- Misconceptions and Pitfalls

Chapter 1. Matter, Measurement, and Problem Solving

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures and Solved problems
1.1 Atoms and Molecules • Definitions of atoms, molecules, chemistry • Representations of atoms and molecules as spheres • Models of carbon monoxide, carbon dioxide, water, hydrogen peroxide	• Figure 1.1 Binding of Oxygen and Carbon Dioxide to Hemoglobin [An example for representing atoms and molecules in a practical example.] • Models of molecules (CO, 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 & experiments lead to a hypothesis; o More experiments may lead to a theory or law; o Theory explains observations and a law.	• Antoine & Marie Lavoisier painting and history [Science is discovered by people.] • Scanning tunneling microscopy STM image of gallium arsenide and cesium atoms. • Figure 1.2 The Scientific Method [A concept map that describes the relationship among observation, hypothesis, theory, and law.] • The Nature of Science [a description of the writing of philosopher Thomas S. Kuhn]
1.3 The Classification of Matter • States of matter, their definitions, and some of their defining properties (e.g., temperature, compressibility) o gas o liquid o solid ▪ crystalline ▪ amorphous • Classification of matter o pure substance ▪ element ▪ compound o mixture ▪ heterogeneous ▪ homogeneous • Separating mixtures o decanting (decantation) o distilling (distillation) o filtering (filtration)	• A particulate (microscopic) and macroscopic view of solid, liquid and gas [Green spheres are used to represent a view and properties of each phase.] • Figure 1.3 Crystalline and Amorphous Solid [A microscopic and macroscopic view of diamond and charcoal.] • Figure 1.4 The Compressibility of Gases [A solid and gas in a piston.] • Classification scheme for matter [A flowchart with the categories and both a macroscopic and microscopic view of an example of each: helium, pure water, wet sand and tea with sugar.] • Figure 1.5 Separating Substances by Distillation [A laboratory setup] • Figure 1.6 Separating Substances by Filtration [A laboratory apparatus.]

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls

1.1 Atoms and Molecules
• Chemistry involves a great deal of what can't be seen. This requires representations and models.
  o The opening image shows hemoglobin. The "real" molecule is not a green and blue ribbon.
  o Chemists look at microscopic, macroscopic, and symbolic representations of atoms and molecules interchangeably. Students need help recognizing which representation to think about when a name is used.
  o If you say "water" you might mean the characters "H_2O" or the spheres (i.e., a picture), the molecule, or a large collection of molecules (e.g., a glass of water). This may need to be repeated in the next two chapters.
	• Atoms and molecules have the macroscopic properties exhibited by a collection of them.

1.2 The Scientific Approach to Knowledge
• Experiments test ideas. They are contrived to prove a hypothesis wrong or to provide support. Hypotheses have to be testable (i.e., falsifiable) or it's not science.
• Theories actually are accompanied by considerable understanding and evidence even though theories are often cited as unproven or untested.
• Figure 1.2 will help clarify the components of the scientific method and their relationships.
• Conceptual Connection 1.1 (Difference between laws and theories) can be used as a 1-minute discussion within small student groups.
• The box about Thomas Kuhn and his writings provide information about the history and philosophy of science; it provides an opportunity to reflect about how chemistry moves forward.
	• Theories are easily dismissible.
	• Science is a set of objective, inarguable facts.
	• The STM image may be confusing:
	  o The physical and electronic properties of the atoms are being measured and plotted in the image;
	  o Atoms and molecules are not colored.

1.3 The Classification of Matter • Properties of matter define its state—gas, liquid, or solid. The temperature is one example, and everyone recognizes steam, water and ice. Ask for additional examples. Dry ice and liquid air or nitrogen are ones they have seen. • Compressibility is another property that differentiates states of matter. They can try it with air in a syringe or balloon, water in a bag, or an ice cube. The water changes shape but is not compressible. • Conceptual Connection 1.2 (Mass of water going from a liquid to a gas) can be used as a 1-minute discussion within small student groups and serves as an introduction to physical vs. chemical changes. • Classifying additional examples of matter according to the scheme demonstrates some of the pitfalls and challenges: mayonnaise, Jell-O, milk. • Students are likely to have varying personal experience with distillation and filtration. Kitchen analogies may be useful—steam condenses on the inside cover of a pot; macaroni and water are poured into a colander for a reason.	• The presence or absence of sticks (i.e., bonds) in the diamond or charcoal structures may confuse students. • Students may not have experience with elemental forms beyond diamond and charcoal to appreciate other examples of crystalline or amorphous solids. • When considering changes in state for a substance, some students are likely to confuse physical and chemical changes; this is covered in the next section.

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures and Solved problems

1.4 Physical and Chemical Changes and Chemical Properties
• Properties and Changes
  o Physical
  o Chemical
• Examples and classifying changes
	• Figure 1.7 Boiling, a Physical Change [with a microscopic and macroscopic view]
	• Figure 1.8 Rusting, a Chemical Change [with a microscopic and macroscopic view]
	• Figure 1.9 Physical and Chemical Changes [Additional examples with microscopic and macroscopic representations]
	• Solved Problem, Example 1.1 [Determining whether a change is physical or chemical.]

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 statements (i.e., forms) of the law of conservation of energy
• Examples of the law of conservation of energy
	• An example of the physics definition of work [A person pushing a box.]
	• Figure 1.10 Energy Conversions [a mass falling from a building]
	• Figure 1.11 Using Chemical Energy to Do Work [conversion of chemical energy stored in molecules]

1.6 The Units of Measurement
• 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
  o amount of substance = mole
• Temperature scales
  o Kelvin $K = {}^{o}C + 273.15$
  o Celsius $^{o}C = \frac{(^{o}F - 32)}{1.8}$
  o Fahrenheit
• Derived units
• Volume (cubic meter, cubic centimeter, liter, milliliter)
• Density (mass per volume) $d = \frac{m}{V}$
	• Mars Climate Orbiter image and story [a calculation disaster]
	• Heights of a skyscraper and athlete
	• Mass of a kilogram and a coin.
	• Figure 1.12 Comparison of ^{o}F, ^{o}C and K Scales
	• Practical examples on the Celsius temperature scale [water freezing, a fall day, room temperature, Death Valley]
	• Solved Problem, Example 1.2 Converting between Temperature Scales [a body temperature on all three scales]
	• Table 1.2 SI Prefixes and Multipliers
	• Figure 1.13 Relationship between Length and Volume
	• Table 1.3 Common Units and Equivalents [for length, mass, volume]
	• Table 1.4 Density of Some Common Substances
	• Solved Problem, Example 1.3 Calculating Density [of a ring to determine its metal content].
	• Chemistry and Medicine [a box describing bone density].

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.4 Physical and Chemical Changes and Chemical Properties • Conceptual Connection 1.3 (Difference between chemical and physical changes) can be used as a 1-minute discussion within small student groups. • Ask the students in small groups to contrive examples that will stump their colleagues.	• Confront the confusion when physical change accompanies a chemical one: burning liquid gasoline produces gaseous fumes. [physical or chemical or both?]
1.5 Energy: A Fundamental Part of Physical and Chemical Change • The example of work being done by a person (moving the box) and chemical energy ultimately moving the car are consistent and simple. Chemical work is defined later. Additional examples could match that theme or the gravitational example. • Several examples are cited for the conservation of energy; ask students to name and describe other forms of energy (solar, mechanical, thermal, chemical, electrical) and devices that convert between these forms.	• Experience in physics may be limited or uneven: the physics definition of work is used; units of work and energy are the same.
1.6 The Units of Measurement • The SI base units may be easier to remember when compared to English equivalents or to common objects. The yardstick and meter stick appear in a graphic. • The practical examples (water freezing, fall day, room temperature, body temperature, Death Valley) calculated on all three scales provide practical reference points and in-class practice. • Several of the large SI unit prefixes are already familiar since they measure the amount of memory in computer storage devices. • A volume unit conversion makes a good in-class group assignment. Just setting it up only takes a few minutes. • An additional solved density calculation can start with the density; Conceptual Connection 1.4 (temperature dependence of a metal) can be used as a 1-minute discussion within small student groups. • The Chemistry and Medicine box provides an open-ended conceptual question about designing an experiment to measure bone density.	• Converting volume units is often a challenge for students (1 cubic meter = ?? cubic centimeters). • Density calculations are deceptively simple.

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures and Solved problems

1.7 The Reliability of a Measurement
• Significance and method of reporting numerical values
• 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 in the final step)
• Precision vs. accuracy
• Scientific integrity and data reporting
	• Tables with carbon monoxide data for Los Angeles County [demonstrate certainty in digits within a number]
	• Figure 1.14 Estimation in weighing [reading a scale and balance]
	• Volume in a burette and values on a thermometer [reading a liquid level with a meniscus and a mercury thermometer]
	• Significant figures rules [with examples]
	• Solved Problem, Example 1.5 [Practice in determining sig figs with explanations]
	• Significant figures in calculations [with examples]
	• Solved Problem, Example 1.6 [Practice in determining sig figs in calculations]
	• Table and graphs of a set of data [detailed graphics]
	• Integrity in Data Gathering box [about scientific data and reporting]

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.7 The Reliability of a Measurement • Two tables present air quality data that might appear in a newspaper or publication initiate the discussion of the certainty of digits in a number. This could be reinforced by weather or air quality numbers from a local newspaper. We can measure temperature very accurately—why doesn't the meteorologist report the temperature as 87.45^{o}F? Students have had very little experience reflecting on the meaning of numerical values. • Reading scales takes practice; digital readouts could be misleading unless you know something about the quality of the data. • Scientists take significant figures for granted. They worry about them when they have to and follow the simple rules to determine them. Coverage in the course is intended to introduce novices to the idea and practice although students seem to ascribe incredible importance to the issue. • The number on the display of a calculator requires interpretation; only the user knows the certainty of the values entered. • The table and graphics for accuracy and precision clearly and succinctly illustrate the ideas. A demonstration involving the collection of an actual data set could involve a set of four coins (e.g., quarters) and an electronic (top-loading) or triple-beam balance. A second example entails measuring the dimensions of the book. Provide some student groups with 12-inch rulers and some with only 8.5 x 11 inch unmarked sheets of paper. Collect and analyze the data. • A discussion about why integrity in data reporting is particularly important in science is appropriate. It can and should include examples of the checks and balances. It should point out that scientists report how they did the experiments so others can try to repeat their work.	• Any answer that appears on the display of a calculator is "sacred" and must be the "right" one.

Lecture Outline

Terms, Concepts, Relationships, Skills	Figures and Solved problems

1.8 Solving Chemical Problems
• Converting from one unit to another
  o dimensional analysis
  o multiple approaches to any problem
• General problem solving strategy (utilized in the text)
  o Sort [What is given and what to find?]
  o Strategize [Conceptual plan]
  o Solve [Follow your plan]
  o Check [Does the answer make sense?]
• Calculations using powers
• Estimates using orders of magnitude
• Calculations using equations
	• Dimension analysis examples [including a review of the algebraic identity]
	• Review of a strategic plan including the model used in the text for solved problems.
	• Solved Problems, Examples 1.7 and 1.8 [unit conversion using the text model]
	• Solved Problem, Example 1.9 [unit conversion involving a power]
	• Solved Problem, Example 1.10 [density as a conversion factor]
	• Solved Problems, Examples 1.11 and 1.12 [calculations involving the use of equations for the volume of a sphere and density]

Teaching Tips

Suggestions and Examples	Misconceptions and Pitfalls
1.8 Solving Chemical Problems • Students will appreciate a quick review of algebra. Two key skills to refresh are: o Multiplying by 1 (multiplication identity) is allowed under any circumstances • Two items equal to each other can be used to create the identity: • 12 eggs = 1 dozen, so $\frac{12\text{ eggs}}{1\text{ dozen}} = 1 = \frac{1\text{ dozen}}{12\text{ eggs}}$ o Performing the same mathematical operation to both sides of an equation are always allowed. $\frac{17}{4} = 4.25$ $\cancel{4} \times \frac{17}{\cancel{4}} = 4 \times 4.25$ $17 = 4 \times 4.25$ • Watching an instructor working problems repeatedly is not as effective as forcing students to try it themselves in class. Provide a set of questions and allow students to work in groups. Escalate the complexity. Promote correct process over correct answers. Show what happens when the student follows a wrong approach (i.e., uses the wrong ratio as the identity) • Promote the use of estimation as part of the problem solving model. Checking oneself prevents careless errors. Reduce reliance on blindly entering values into a calculator. • 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 the underlying idea or concept. It becomes a measure of cognitive ability, outside the realm of chemistry.	• There is only one correct way to solve a problem. • If there is more than one way to do a problem, show them all to me so I can "know" them too.

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** mi → km → m $\frac{1\text{ km}}{0.6214\text{ mi}}$ $\frac{1000\text{ m}}{1\text{ km}}$ **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 $\left(\frac{1\text{ m}}{39.37\text{ in}}\right)^3$ **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** $14\ \cancel{\text{piers}} \times \frac{4750\ \cancel{\text{in}^3}}{1\ \cancel{\text{pier}}} \times \frac{(1\text{ m})^3}{(39.37\ \cancel{\text{in}})^3} = 1.0897\ \text{m}^3$ $1.0897\ \text{m}^3 = 1.09\ \text{m}^3$
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 → mL → cm^3 → g → lb $\frac{1000\ \text{mL}}{1\ \text{L}}$ $\frac{1\ \text{cm}^3}{1\ \text{mL}}$ $\frac{0.789\ \text{g}}{1\ \text{cm}^3}$ $\frac{1\ \text{lb}}{453.59\ \text{g}}$ **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** $100\ \cancel{\text{L}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mL}}} \times \frac{0.789\ \cancel{\text{g}}}{1\ \cancel{\text{cm}^3}} \times \frac{1\ \text{lb}}{453.59\ \cancel{\text{g}}}$ = 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** $\ell = 1.1$ 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** $\ell \rightarrow V$ $V = \ell^3$ $in^3 \rightarrow cm^3 \rightarrow g$ $\left(\frac{2.54\ cm}{1\ in}\right)^3 \qquad \frac{0.917\ g}{1\ cm^3}$ **Relationships Used** $V = \ell^3$ [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** $V = (1.1\ in)^3 = 1.331\ in^3$ $1.331\ \cancel{in^3} \times \left(\frac{2.54\ \cancel{cm}}{1\ \cancel{in}}\right)^3 \times \frac{0.917\ g}{1\ \cancel{cm^3}} = 20.0008\ g$ 20.0008 g = 20. g
Check	The units (g) are correct. The magnitude of the answer (20.) seems to make physical sense.