Chapter 1

Introduction: Matter and Measurement

OVERVIEW OF THE CHAPTER

1.1, 1.2, 1.3 MATTER: ELEMENTS, COMPOUNDS, AND MIXTURES

Learning Goals: You should be able to:

1. Distinguish between physical and chemical properties and also between simple physical and chemical changes.
2. Differentiate between the three states of matter.
3. Distinguish between elements, compounds, and mixtures.
4. Give the symbols for the elements discussed in this chapter.

1.4 PHYSICAL QUANTITIES AND UNITS

Review: Concept of fraction; exponential notation (see text: Appendix A).

Learning Goal: You should be able to list the basic SI and metric units and the commonly used prefixes in scientific measurements.

1.5 UNCERTAINTY IN MEASUREMENT: SIGNIFICANT FIGURES

Review: Exponential notation (see text Appendix A).

Learning Goals: You should be able to:

1. Determine the number of significant figures in a measured quantity.
2. Express the result of a calculation with the proper number of significant figures.

1.4 TEMPERATURE AND DENSITY: INTENSIVE PROPERTIES

Learning Goals: You should be able to:

1. Convert temperatures among the Fahrenheit, Celsius, and Kelvin scales.
2. Perform calculations involving density.

1.6 DIMENSIONAL ANALYSIS

Review: Concepts of fraction and ratio.

Learning Goal: You should be able to convert between units by using dimensional analysis.

TOPIC SUMMARIES AND EXERCISES

MATTER: ELEMENTS, COMPOUNDS, AND MIXTURES

Matter is any material that occupies space and has mass. Three phases (states) of matter exist: gas, liquid, and solid.

- A sample of matter is either a substance or a mixture.
- **Substances** are either elements or compounds. **Elements** cannot be separated into simpler new substances. **Compounds** consist of two or more elements chemically combined in a definite ratio. A compound can be chemically decomposed into its elements.
- **Mixtures** are physical combinations of two or more substances and are either homogeneous or heterogeneous. A *homogeneous* mixture consists of one phase and a uniform distribution of substances. A *heterogeneous* mixture shows more than one phase and possesses a nonuniform distribution of substances.
- Mixtures can be separated into substances by physical means.

Alterations in matter can involve chemical or physical changes.

- A **chemical change** results in a change in the composition of a substance. A **chemical property** describes the type of chemical change. For example, the property of wood burning is a chemical property.
- A **physical change** does not involve a change in the composition of a substance but rather a change in a **physical property** such as temperature, volume, mass, pressure, or state.

Check your understanding of the new terms you have learned by doing Exercises 1–5.

EXERCISE 1 Identifying characteristics of matter

Match the following characteristics to one or more of the three states of matter: (a) has no shape of its own; (b) definite shape; (c) occupies the total volume of a container; (d) partially takes on the shape of a container; (e) does not take on the shape of a container; (f) readily compressible; (g) slightly compressible; (h) essentially noncompressible.

SOLUTION: Gas—(a), (c), (f); Liquid—(a), (d), (g); Solid—(b), (e), (h)

EXERCISE 2 Identifying characteristics of matter II

Match the term with the best identifying phrase:

Terms

1. Homogeneous mixture
2. Heterogeneous mixture
3. Mixture
4. Substance
5. Element
6. Compound

Phrases

a. Any kind of matter that is pure and has a fixed composition
b. Cannot be decomposed into simpler substances by chemical changes

c. A solution of uniform composition
d. Can be decomposed into simpler substances by chemical changes
e. Any kind of matter that can be separated into simpler substances by physical means
f. Nonuniform composition

SOLUTION: 1-c; 2-f; 3-e; 4-a; 5-b; 6-d

EXERCISE 3 Writing names or symbols of elements

With the help of the periodic table, write the name or the chemical symbol for each of the following elements: (a) F; (b) zinc; (c) potassium; (d) As; (e) Al; (f) iron; (g) helium; (h) barium; (i) Ne.

SOLUTION: (a) fluorine; (b) Zn; (c) K; (d) arsenic; (e) aluminum; (f) Fe; (g) He; (h) Ba; (i) neon

EXERCISE 4 Identifying changes of matter

Are the following changes physical or chemical: (a) the vaporization of solid carbon dioxide; (b) the explosion of solid TNT; (c) the aging of an egg with a resultant unpleasant smell; (d) the formation of a solid when honey is cooled?

SOLUTION: (a) A physical change. The form of carbon dioxide is changed from solid to gas. There is no change in its chemical composition. (b) A chemical and physical change. The explosion results from a change in the chemical composition of TNT and the formation of a gas. (c) A chemical change. A change in the composition of the egg results in the formation of a gas that has an unpleasant smell. (d) A physical change. The solid results from the crystallization of dissolved sugars; no change occurs in the chemical form of the sugars.

EXERCISE 5 Recognizing elements, compounds, and mixtures

Classify each of the following as an element, compound, or mixture: (a) a 100-percent lead bar; (b) wine; (c) gasoline; (d) carbon dioxide (CO_2).

SOLUTION: (a) Lead is an element and cannot be separated by either chemical or physical means into simpler substances. It is listed among the elements in Table 1.2 in the text. (You should know the symbols in Table 1.2.) (b) Wine is a mixture of alcohol, other components, and water. The fact that wines contain varying percentages of alcohol attests to their having different compositions. (c) We know gasoline must be a mixture because it is available with different compositions and properties (no lead, regular, and different brands with different additives). (d) CO_2 is a compound because the ratio of carbon and oxygen atoms is fixed and definite. The name also implies that it is a compound because we do not have such systematic names for mixtures.

PHYSICAL QUANTITIES AND UNITS

A physical property of a sample is measured by comparing it with a standard unit of that property. Measured quantities such as volume, length, mass, and temperature require a number and a reference label, called the unit of measurement. Two systems of unit measurements are shown in Table 1.1 on page 4.

The SI system of units is now the preferred one; however, you will find certain metric units still used. *You must become thoroughly familiar with the units in Table 1.1 before starting the next chapter.*

TABLE 1.1 Metric and SI Units

Physical quantity	Metric unit name	SI unit name[a]
Length	Meter (m)	Meter (m)
Volume	Cubic centimeter (cm^3)[b]	Cubic meter (m^3)
Mass	Gram (g)	Kilogram (kg)
Time	Second (s)	Second (s)
Energy	Calorie (cal)	Joule (J)
Pressure	Atmosphere (atm)	Newton per square meter (N/m^2)

[a] Systeme International d'Unites (SI) or International System of Units.

[b] Chemists commonly use the unit cubic centimeter when dealing with the volume of a solid, but they usually use the unit liter (L) when a substance is a liquid.

A necesary skill requiring proficiency is changing a number with a unit to one with a different unit. We use equivalence relationships between units to do conversions between units. Tables 1.2 and 1.3 give some common equivalences that you will need in this chapter.

Prefixes are used with units to indicate decimal fractions (<1) or multiples (>1) of basic units.

- Example of a decimal fraction: The prefix centi- means $1/100(= 0.01)$ of a basic unit; thus, $100 \text{ cm} = 100 \times 1/100 \text{ m} = 1 \text{ m}$.
- Example of a multiple: The prefix kilo- means $10^3(= 1000)$; thus, $1 \text{ km} = 1 \times 1000 \text{ m} = 1000 \text{ m}$.

The commonly used prefixes that you must know are shown in Table 1.4 on page 5. *Memorize them.*

TABLE 1.2 Equivalence Relationships between SI and Metric Units

Physical quantity	Metric unit name	SI unit name	Equivalence
Length	Meter	Meter	Same
Mass	Gram	Kilogram	$1000 \text{ g} = 1 \text{ kg}$
Time	Second	Second	Same
Energy	Calorie	Joule	$1 \text{ cal} = 4.184 \text{ J}$
Volume	Cubic centimeter	Cubic meter	$1{,}000{,}000 \text{ cm}^3 = 1 \text{ m}^3$
Volume	Liter	Cubic meter	$1000 \text{ L} = 1 \text{ m}^3$
Pressure	Atmosphere	Newton per square meter	$1 \text{ atm} = 0.1754 \text{ N/m}^2$

TABLE 1.3 Equivalence Relationships between Metric and English Units

Physical quantity	English unit symbol	Metric unit symbol	Equivalence
Mass	lb (= 16 oz)	g	$1 \text{ lb} = 453.6 \text{ g}$
Length	ft (= 12 in.)	m	$3.272 \text{ ft} = 1 \text{ m}$
Length	in.	cm	$1 \text{ in.} = 2.54 \text{ cm}$
Length	mi (= 5280 ft)	m	$1 \text{ mi} = 1609 \text{ m}$
Volume	qt	L	$1.057 \text{ qt} = 1 \text{ L}$

TABLE 1.4 Commonly Used Prefixes for Scientific Measurement in Chemistry

Prefix	Fraction or multiple of base unit	Abbreviation
Deci-	$10^{-1}\left(\frac{1}{10}\right)$	d
Centi-	$10^{-2}\left(\frac{1}{100}\right)$	c
Milli-	$10^{-3}\left(\frac{1}{1000}\right)$	m
Micro-	$10^{-6}\left(\frac{1}{1{,}000{,}000}\right)$	μ
Nano-	$10^{-9}\left(\frac{1}{1{,}000{,}000{,}000}\right)$	n
Pico-	$10^{-12}\left(\frac{1}{1{,}000{,}000{,}000{,}000}\right)$	p
Kilo-	10^{3} (1000)	k
Mega-	10^{6} (1,000,000)	M
Giga-	10^{9} (1,000,000,000)	G

EXERCISE 6 Determining relative magnitudes of quantities

Which quantity of each pair is larger: **(a)** 1 nm or 1 micrometer; **(b)** 1 picogram or 1 cg; **(c)** 1 megagram or 1 milligram?

SOLUTION: Change the pairs so that each quantity is represented by either a fraction or a multiple of the same basic metric unit. Then from their relative magnitudes we can determine which is larger.

(a) $1\text{ nm} = 1\text{ nanometer} = 10^{-9}\text{ meter}$

$1\text{ micrometer} = 1\ \mu\text{m} = 10^{-6}\text{ meter}$

One micrometer is larger in value than one nanometer because the fraction $10^{-6}\left(\frac{1}{1{,}000{,}000}\right)$ is larger in magnitude than the fraction $10^{-9}\left(\frac{1}{1{,}000{,}000{,}000}\right)$.

(b) $1\text{ picogram} = 1\text{ pg} = 10^{-12}\text{ gram}$

$1\text{ cg} = 1\text{ centigram} = 10^{-2}\text{ gram}$

One centigram is larger in value than one picogram because the fraction $10^{-2}\left(\frac{1}{100}\right)$ is larger in magnitude than the fraction $10^{-12}\left(\frac{1}{1{,}000{,}000{,}000{,}000}\right)$.

(c) $1\text{ megagram} = 1\text{ Mg} = 10^{6}\text{ gram}$

$1\text{ mg} = 1\text{ milligram} = 10^{-3}\text{ gram}$

One megagram is larger in value than one milligram because the multiple 10^{6} (1,000,000) is larger in magnitude than the fraction $10^{-3}\left(\frac{1}{1{,}000}\right)$.

EXERCISE 7 Recognizing units with measurements

With what types of measurements are the following units associated?

g, L, m, km, cm, Mg, pg, cm^3

SOLUTION: Mass (g, Mg, pg); volume (L, cm^3); length (cm, m, km). Note that the prefixes such as M- and c- do not change the type of unit. However, the type of unit can be changed if it is raised to some power, as is the case for cm^3. The unit cm^3 means cm × cm × cm, which is a unit for volume ($V = 1 \times w \times h$).

EXERCISE 8 Comparing English to SI System of Units

What is the advantage of the metric system in comparison to the English system?

SOLUTION: In the metric system, all quantities larger or smaller than the basic unit involve multiplication of the basic unit value by some power of 10 (for example, $10^3 = 1000$, $10^{-1} = \frac{1}{10}$, and so on). This is not true of the English system. Smaller or larger quantities of the basic unit in the English system are newly defined units. For example, 4000 qt equals 1000 gal, not 4 "kiloquarts." Many more conversion factors are required in the English unit system than in the metric unit system.

EXERCISE 9 Knowing acceptable SI volume unit

Suggest a reason for the fact that 1 μkL (microkiloliter) is not accepted as an appropriate SI unit for volume.

SOLUTION: The expression 1 μkL involves two prefixes, micro- (μ) and kilo- (k), yielding a compound prefix. This can be confusing, particularly if three or four prefixes are used. Thus, we do not use more than one prefix when expressing numbers. Instead of 1 μkL (microkiloliter), we write 1 mL (milliliter).

UNCERTAINTY IN MEASUREMENTS: SIGNIFICANT FIGURES

Numbers in chemistry are of two types:

- **Exact:** These result from counting objects such as coins or occur as exact numbers in equations or as exact conversion factors.
- **Inexact:** These are obtained from measurements. Uncertainties exist in their values because judgment is required in making measurements.

Measured quantities (inexact numbers) are reported so that the last digit is the first uncertain digit. An uncertain digit is one that requires judgment in determining its value. All certain digits and the first uncertain digit are referred to as **significant figures**. For example:

- 2.86: 2 and 8 are certain and well known. The number 6 is the first that is subject to judgment and is uncertain. The first uncertain digit is assumed to have an uncertainty of $\pm 1 : 2.86 \pm 0.01$. The number 2.86 has three significant figures.
- 0.0020: Zeroes to the left of the first nonzero digit in a number with a decimal point are not significant. The first three zeroes are not significant because they are to the left of the 2 and also define the decimal point. The zero to the right of the 2 is significant. This number has only two significant figures.
- 100: Trailing zeroes that define a decimal point may or may not be significant. Unless stated, assume they are not significant. Therefore, 100 has one significant figure unless otherwise stated; if it is determined from counting objects, it has three significant figures.

Exponential notation is used to remove ambiguity in reporting the number of significant figures a number possesses.

- Only significant digits are shown. The number 0.0020 becomes 2.0×10^{-3}. The zeroes in front of the two in 0.0020 are not significant whereas the trailing zero is significant.

Calculated numbers must show the correct number of significant figures. The rules for doing this are:

1. Addition and Subtraction: The final answer should have the same uncertainty as the quantity in the calculation with the greatest uncertainty. In the following example, the first uncertain digit in each quantity is in bold.

   325.24 (uncertainty = ±0.01)
   + 21.4 (uncertainty = ±0.1)
   + 145 (uncertainty = ±1)
   491.64 (uncertainty in final answer is ± 1)

   The least precise number is 145 and it controls the number of significant figures in the answer. It has the greatest uncertainty, ±1. Thus 491.64 is rounded to 492 with an uncertainty of ±1.
2. Multiplication and division: When multiplying or dividing numbers, round off the final calculated answer so that it has the same number of significant figures as the least certain number (the one with fewest number of significant figures) in the calculation. A little caution must be used when applying this rule. For example, to divide 101 by 95, you might be tempted to report the final answer to two significant figures because 95 appears to be the least certain number. Yet 95 has almost three significant figures; there is very little difference in error between 1 in 95 and 1 in 101. Thus, in this case it makes more sense to round off the final answer to three significant figures. Use common sense in problems when a number is close in magnitude to 100, 1000, 10,000, and so on.
3. Exact or defined numbers are not used in determining the number of significant figures in a final answer. If you use the equation $A = 4\pi r^2$ to calculate the surface area of a sphere, the number 4 is considered to have an infinite number of significant figures. It therefore not used in determining the uncertainty of the calculated area.

Caution: The final answer of a calculation determined using a calculator often has more digits than any of the numbers in the calculation. You may have to round off the answer to the correct number of significant figures. The rules for rounding off numbers in a calculated answer are:

1. When the number immediately following the last digit to be retained (the first uncertain digit) is less than 5 then the last digit is retained unchanged. If 6.4362 is rounded off to four significant figures it becomes 6.436.
2. When the number immediately following the last digit to be retained is 5* or greater, then increase the last digit by 1. If 6.4366 is to be rounded off to four significant figures it becomes 6.437.

*Note: Your instructor may use an alternative approach: If there are no other numbers or only zeroes beyond the 5, then the last retained digit is increased by 1 if it is odd and left unchanged if it is even. Or if there are numbers other than zero beyond the 5, then the last digit retained is increased by 1. For example, when three significant figures are required, 2.2350 becomes 2.14 (3 is odd and thus it is increased by 1) and 2.1453 becomes 2.15.

Note: Do not round numbers until you have completed your calculation.

EXERCISE 10 Determining uncertain digits

Remembering that measured values are reported to ± 1 uncertainty in the last digit, except for those values determined by counting observable objects, or unless otherwise stated, determine the first uncertain digit in each of the following numbers: (a) 10.03 kg; (b) 5 apples; (c) 5.02 ± 0.02 m.

SOLUTION: (a) The 3 in 10.03 kg is uncertain to ± 1. (b) This is an exact measured value determined by counting. There is no uncertain digit. (c) The 2 in 5.02 m is uncertain to ± 2.

EXERCISE 11 Determining the number of significant figures

The precision of a measurement is indicated by the number of significant figures associated with the reported value. How many significant figures does each number possess: (a) 225; (b) 10,004; (c) 0.0025; (d) 1.0025; (e) 0.002500; (f) 14,100; (g) 14,100.0?

SOLUTION: Try this technique: If a quantity contains a *decimal* point, draw an arrow *starting* at the *left* through all zeroes up to the first nonzero digit; the digits remaining are significant. If the quantity does *not* contain a decimal point, draw an arrow *starting* at the *right* through all zeroes up to the first nonzero digit; the digits remaining are significant.

(a) 225 ⟵ three significant figures (No decimal point—draw arrow to left)
(b) 10,004 ⟵ five significant figures
(c) ~~0.00~~25 two significant figures (Contains a decimal point—draw arrow to the right)
(d) ⟶ 1.0025 five significant figures
(e) ~~0.00~~2500 four significant figures
(f) 14,~~100~~ three significant figures; however, because of our lack of knowledge about the significance of the two trailing zeroes, this number also could have four or five significant figures
(g) ⟶ 14,100.0 six significant figures

EXERCISE 12 Writing numbers in scientific notation

Write the numbers in Exercise 11 using scientific notation.

SOLUTION: Move the decimal in the appropriate direction so that it is to the right of the first nonzero digit reported in the number. If the decimal is moved to the left, multiply the resulting quantity by 10 raised to a power that equals the number of digits the decimal is moved past. If the decimal is moved to the right, the power of 10 is again the number of digits the decimal is moved past, but with a negative sign. That is, a number that is greater than 1 will appear as $A.BC \times 10^x$, while one that is less than 1 will appear as $A.BC \times 10^{-x}$. (a) $225 = 2.25 \times 10^2$. The decimal is moved two digits to the left, thus the power of 10 is 2. (b) 10,004 = 10,004 = 1.0004×10^4. The power of 10 is 4 because the decimal is moved four places to the left. (c) $0.0025 = 0002.5 \times 10^{-3} = 2.5 \times 10^{-3}$. The power of 10 is -3

because the decimal is moved three places to the right. Note that the nonsignificant zeros are omitted. (d) 1.0025. We do not write 1.0025×10^0. (e) $0.002500 = 0002.500 \times 10^{-3} = 2.500 \times 10^{-3}$. The zeros after the 2.5 are written because they define the number of significant figures. (f) $14,100 = 1.4100 \times 10^4 = 1.41 \times 10^4$ if the zeros in 14,100 are not significant. (g) $14,100.0 = 1.41000 \times 10^4$.

EXERCISE 13 Rounding answers in calculations

Round the answers in the following problems to the correct number of significant figures:

(a) $12.25 + 1.32 + 1.2 = 14.770$
(b) $13.7325 - 14.21 = -0.4775$
(c) $12300 + 2.11 = 12302.11$

SOLUTION: In each problem, identify the quantity with the greatest uncertainty and use this uncertainty to determine the correct number of significant figures for the answer. (a) The 1.2 has the greatest uncertainty, ± 0.1. Therefore, the answer must be rounded to one digit to the right of the decimal point: 14.8. (b) 14.21 has the greatest uncertainty, ± 0.01. Therefore, the answer must be rounded to two digits to the right of the decimal point: -0.48. *Note*: An answer obtained by subtraction may have fewer significant figures than either number used. (c) 12300 has an uncertainty of ± 100. The trailing zeros are not identified as being significant; therefore, we normally assume that they are not significant. (If the trailing zeros are significant, they should have been identified as 12300., or preferably 1.2300×10^4.) The answer must be rounded at the hundreds place: 12300. Any digit to the right of the hundreds place is assigned a zero value.

EXERCISE 14 Rounding answers in calculations II

Round the final answer in each of the following calculations:

(a) $(1.256)(2.42) = 3.03952$
(b) $\dfrac{16.231}{2.20750} = 7.352661$
(c) $\dfrac{(1.1)(2.62)(13.5278)}{2.650} = 14.712121$

SOLUTION: (a) The least precise number in the calculation is 2.42 (three significant figures). The final answer must be rounded to three significant figures: 3.04. (b) The least precise number in the calculation is 16.231 (five significant figures). The final answer must be rounded to five significant figures: 7.3527. (c) The least precise number in the calculation is 1.1 (two significant figures). The final answer must be rounded to two significant figures: 15.

TEMPERATURE AND DENSITY

Two important concepts discussed in Chapter 1 are temperature and density.

- They are both **intensive properties** as their values are independent of the amount of substance.
- This contrasts with **extensive properties** such as volume and mass, which depend on the amount of substance.

Temperature is a measure of the intensity of heat—the "hotness" or "coldness" of a body.

- Heat is a form of energy. Heat flows from a hot object to a colder one.
- When there is no heat flow between two objects in contact, they have the same temperature.
- Three temperature scales are used: Celsius (°C), Fahrenheit (°F), and Kelvin (K). You need to know how their reference points differ and how to change between them.

Reference Points			
	Fahrenheit	Celsius	Kelvin
Freezing point of water	32 °F	0 °C	273.15 K
Boiling point of water	212 °F	100 °C	373.15 K

- *Note that a 1-degree interval is the same on both the Celsius and Kelvin scales, but a 1 °C interval equals a 1.8 °F interval.* The only temperature at which both the Fahrenheit and Celsius scales are equivalent is $-40°(-40\ °C = -40\ °F)$. This fact enables us to make conversions between the two scales using the following approach, which is different from the one given in the text.

	°F ⟶ °C	°C ⟶ °F
(a)	Add 40° to °F = (1)	Add 40° to °C = (1)
(b)	$(1) \times \frac{1\ °C}{1.8\ °F} = (2)$**	$(1) \times \frac{1.8\ °F}{1\ °C} = (2)$**1
(c)	Subtract 40° from (2) = °C	Subtract 40° from (2) = °F

**Notice that only step (b) is different. Step (b) converts Fahrenheit to Celsius or Celsius to Fahrenheit using the relationship that a 1-degree Celsius interval equals a 1.8-degree Fahrenheit interval. Also, do not round off until the calculation is finished.

- The following relationship is used to convert between Celsius and Kelvin temperatures.

$$K = \left(\frac{1\ K}{1\ °C}\right)(°C) + 273.15\ K$$

Density (d) measures the amount of a substance (m) in a given volume (V):

- $d = \frac{\text{mass}}{\text{volume}} = \frac{m}{V}$
- Density varies with temperature because volume changes with temperature.
- Density can be used to change mass to volume and vice versa for the same substance.
- Chemists commonly use the following units for density: g/mL for liquids, g/cm^3 for solids, and g/L for gases.

Exercise 15 Converting temperature to a different scale

The temperature on a spring day is around 22 °C. What is this temperature in degrees Fahrenheit and degrees Kelvin?

SOLUTION: To change 22 °C to °F, first add 40°:

$$22\ °C + 40° = 62\ °C$$

Then multiply by 1.8 °F/1 ° C:

$$(62\ °C)\left(\frac{1.8\ °F}{1\ °C}\right) = 112\ °F$$

Finally, subtract 40° from 112 °F:

$$112° - 40° = 72\ °\text{F}$$

To calculate degrees Kelvin, write the relationship between the two degrees and substitute for °C:

$$\text{K} = \left(\frac{1\ \text{K}}{1\ °\text{C}}\right)(°\text{C}) + 273.15\ \text{K} = \left(\frac{1\ \text{K}}{1\ °\text{C}}\right)(22\ °\text{C}) + 273.15\ \text{K} = 295\ \text{K (rounded)}$$

Note: The method presented in the text gives the same answer as follows.

$$°\text{F} = \left(\frac{1.8\ °\text{F}}{1\ °\text{C}}\right)(°\text{C}) + 32\ °\text{F}$$

Substituting 22 °C for °C yields:

$$(°\text{F}) = \left(\frac{1.8\ °\text{F}}{1\ °\text{C}}\right)(22\ °\text{C}) + 32\ °\text{F} = 39.6\ °\text{F} + 32\ °\text{F} = 72\ °\text{F (rounded)}$$

EXERCISE 16 Comparing densities

At 20 °C, carbon tetrachloride and water have densities of 1.60 g/mL and 1.00 g/mL respectively. When water and carbon tetrachloride are poured into the same container, two layers form, one being water and the other carbon tetrachloride. Based on their densities, which one will occupy the lower layer in a container?

SOLUTION: Since carbon tetrachloride and water do not mix together permanently, the heavier substance per unit volume will fall to the bottom of the container. Carbon tetrachloride will be that substance because it has a higher mass per unit volume (density), 1.60 g/mL, than does water, 1.00 g/mL.

EXERCISE 17 Using density in a calculation

Which has the greater mass, 2.0 cm^3 of iron (d = 7.9 g/cm^3) or 1.0 cm^3 of gold (d = 19.32 g/cm^3)?

SOLUTION: The mass of a substance is related to its density by the equation $d = m/V$. Multiplying both sides of the equation by V yields $m = \text{d} \times V$. The mass of 2.0 cm^3 of iron is calculated as follows:

$$m = d \times V = \left(7.9\frac{\text{g}}{\text{cm}^3}\right)(2.0\ \text{cm}^3) = 16\ \text{g}$$

The mass of 1.0 cm^3 of gold is calculated similarly:

$$m = d \times V = \left(19.32\frac{\text{g}}{\text{cm}^3}\right)(1.0\ \text{cm}^3) = 19\ \text{g}$$

Thus 1.0 cm^3 of gold has a greater mass than 2.0 cm^3 of iron.

DIMENSIONAL ANALYSIS

You need to develop the habit of including units with all measurements in calculations. *Units are handled in calculations as any algebraic symbol:*

- Numbers added or subtracted must have the same units.
- Units are multiplied as algebraic symbols: (2 L)(1 atm) = 2 L-atm
- Units are cancelled in division if they are identical. Otherwise, they are left unchanged: (3.0 m)/(2.0 mL) = 1.5 m/mL.

Dimensional analysis is the algebraic process of changing from one system of units to another. A fraction, called a unit conversion factor, is used to make the conversion. These fractions are obtained from an equivalence between two units. For example, consider the equality 1 in. = 2.54 cm. This equality yields two conversion factors.

$$\frac{1\text{ in.}}{1\text{ in.}} = \frac{2.54\text{ cm}}{1\text{ in.}} \quad \text{and} \quad \frac{1\text{ in.}}{2.54\text{ cm}} = \frac{2.54\text{ cm}}{2.54\text{ cm}}$$

$$1 = \frac{2.54\text{ cm}}{1\text{ in.}} \quad \text{and} \quad \frac{1\text{ in.}}{2.54\text{ cm}} = 1$$

Note that the two conversion factors each equal one and are the inverse of one another. They enable us to convert between units in the equality. For example, to convert from centimeters to inches or vice versa:

$$5.08\ \cancel{\text{cm}} \times \frac{1\text{ in.}}{2.54\ \cancel{\text{cm}}} = 2.00\text{ in.}$$

$$4.00\ \cancel{\text{in.}} \times \frac{2.54\text{ cm}}{1\ \cancel{\text{in.}}} = 10.2\text{ cm}$$

Note: $\text{given unit} \times \dfrac{\text{new unit}}{\text{given unit}} = \text{new unit}$

EXERCISE 18 Converting units of volume

Convert 10.5 L to milliliters.

SOLUTION: The required operations for converting 10.5 L to milliliters are as follows:

1. State the general relation required to convert units:

$$?\text{ mL} = (10.5\text{ L})(\text{conversion factor that changes L to mL})$$

2. Find the conversion factor (or factors) that converts L to mL. In this case it is 1000 mL = 1 L. Using this equivalence relation, determine the appropriate ratio of units that converts L to mL. Because we want to change liters to milliliters, we will have to divide by 1 L:

$$\frac{1000\text{ mL}}{1\text{ L}}$$

3. Substitute 1000 mL/1 L for the conversion factor in the equation in step 1 and solve the problem.

$$?\text{ mL} = (10.5\ \cancel{\text{L}})\left(\frac{1000\text{ mL}}{1\ \cancel{\text{L}}}\right) = 10{,}500\text{ mL} = 1.05 \times 10^4\text{ mL}$$

Change the final answer to scientific notation, if necessary. It is sometimes more convenient to change all numbers to scientific notation before doing the mathematics of the problem:

$$?\text{ mL} = (1.05 \times 10^1\ \cancel{\text{L}})\left(\frac{1 \times 10^3\text{ mL}}{1\ \cancel{\text{L}}}\right) = 1.05 \times 10^4\text{ mL}$$

4. Check that the units properly cancel to yield the desired unit. In this problem, if we had used the conversion factor 1 L/1000 mL instead of 1000 mL/1 L, the result would have been:

$$? \text{ mL} \neq (1.05 \times 10^1 \text{ L})\left(\frac{1 \text{ L}}{1 \times 10^3 \text{ mL}}\right) \neq 1.05 \times 10^{-2} \frac{\text{L}^2}{\text{mL}}$$

The unit L^2/mL does not equal mL; therefore, we know we have used the wrong conversion factor.

EXERCISE 19 Converting SI units

Convert 6.23 ft^3 to the appropriate SI unit.

SOLUTION: The appropriate SI unit is meter^3. The unit ft^3 is based on an English unit of length, the foot. Use the same method shown in Exercise 18 to convert ft^3 to m^3.

1. $? \text{ m}^3 = (6.23 \text{ ft}^3)$ (conversion factor that changes ft^3 to m^3)
2. From Table 1.3 we find that 3.272 ft = 1 m, but there is no unit conversion given for ft^3 to m^3. What we must recognize is that if

$$1 = \frac{1 \text{ m}}{3.272 \text{ ft}}$$

then we can cube both sides of the expression

$$(1)^3 = \left(\frac{1 \text{ m}}{3.272 \text{ ft}}\right)^3 = \frac{1 \text{ m}^3}{(3.272)^3 \text{ ft}^3} = 1$$

3. Converting units yields:

$$? \text{ m}^3 = (6.23 \cancel{\text{ft}^3})\left(\frac{1 \text{ m}^3}{(3.272)^3 \cancel{\text{ft}^3}}\right) = 0.178 \text{ m}^3$$

EXERCISE 20 Using multiple conversion factors

The Empire State Building in New York City was for many years the tallest building in the world. Its height is 484.7 yd to the top of the lightning rod. Convert this distance to meters.

SOLUTION: The conversion of 484.7 yd to meters is accomplished by the method used in Exercise 18:

1. State the general relation required to convert units:

   ? m = (484.7 yd)(conversion factor or series of factors that changes yards to meters)

2. Look up the required equivalences in appropriate tables. From the tables in this chapter of the *Student's Guide*, we find that 3.272 ft = 1 m. Because no conversion between yards and feet is given, we will have to go to another source or remember from your past experience that 1 yd = 3 ft. To convert yards to meters we will have to make a series of unit conversions that will look like

$$(\cancel{\text{yd}})\left(\frac{\cancel{\text{ft}}}{\cancel{\text{yd}}}\right)\left(\frac{\text{m}}{\cancel{\text{ft}}}\right) = \text{m}$$

Thus the required conversion factors are

$$\frac{3\text{ ft}}{1\text{ yd}} \text{ and } \frac{1\text{ m}}{3.272\text{ ft.}}$$

3. We could solve the problem in two steps:

$$(484.7\ \cancel{\text{yd}})\left(\frac{3\text{ ft}}{1\ \cancel{\text{yd}}}\right) = 1454\text{ ft}$$

$$(1454\ \cancel{\text{ft}})\left(\frac{1\text{ m}}{3.272\ \cancel{\text{ft}}}\right) = 444.3\text{ m}$$

However, it is usually simpler to do it all in one step:

$$?\text{ m} = (484.7\ \cancel{\text{yd}})\left(\frac{3\ \cancel{\text{ft}}}{1\ \cancel{\text{yd}}}\right)\left(\frac{1\text{ m}}{3.272\ \cancel{\text{ft}}}\right) = 444.3\text{ m}$$

4. Check that the units properly cancel. Inspection of the previous equation shows that they do.

Exercise 21 Converting units of measurements having a ratio of units

Florence Griffith Joyner (USA) set a world record in the women's 100 m dash on July 16, 1988, running the distance in 10.49 s. This record has not been broken as of May 30, 2007. Assuming that the distance is exactly 100 m, what is her average speed in miles per hour?

SOLUTION: Joyner's average speed in meters per second is

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{100\text{ m}}{10.49\text{ s}} = 9.533\ \frac{\text{m}}{\text{s}}$$

The answer has four significant figures because the problem says assume the distance is exact. The units are not requested in the problem; the ratio m/s has to be converted to mi/hr.

1. The conversion of units can be viewed as follows:

$$\frac{\text{m} \rightarrow \text{mi}}{\text{s} \rightarrow \text{min} \rightarrow \text{hr}}$$

Table 1.3 gives us the necessary equivalences between meters and miles and we should know the equivalences between seconds and minutes and minutes and hours:

1 mi = 1609 m
60 s = 1 min
60 min = 1 hr

2. Using these equivalences we can now make the unit conversions by converting meters to miles in the numerator and also seconds to minutes to hours in the denominator:

$$9.533\frac{\text{m}\left(\dfrac{1\text{ mi}}{1609\text{ m}}\right)}{\text{s}\left(\dfrac{1\text{ min}}{60\text{ s}}\right)\left(\dfrac{1\text{ hr}}{60\text{ min}}\right)} = 21.33\ \frac{\text{mi}}{\text{hr}}$$

(3) Alternatively we can do the conversion in a long sequence:

$$\text{speed} = \left(9.533\frac{\text{m}}{\text{s}}\right)\left(\frac{1\ \text{mi}}{1609\ \text{m}}\right)\left(\frac{60\ \text{s}}{1\ \text{min}}\right)\left(\frac{60\ \text{min}}{1\ \text{hr}}\right) = 21.33\ \frac{\text{mi}}{\text{hr}}$$

SELF-TEST QUESTIONS

Having reviewed key terms in Chapter 1, match key terms with phrases and identify statements as true or false. If a statement is false, indicate why it is incorrect.

Match each phrase with the best term:

1.1 Increases in value with decreasing volume.

1.2 0.01200 contains four of these.

1.3 Heat emitted from burning wood is not an example of this.

1.4 The kg unit in 1.00 kg of silver tells us about this measurement.

1.5 This property does not change with amount of material.

1.6 This property changes with amount of material.

1.7 The temperature of water at 75 °C is an example.

1.8 The freezing of water is an example.

1.9 A chemical reaction is an example.

1.10 The ability of carbon to form carbon dioxide is an example.

1.11 Makes up the composition of a compound.

1.12 HF is an example.

1.13 A temperature scale with the divisions between the freezing point and melting point of water.

Terms:

(a) Celsius	(h) Intensive
(b) Chemical change	(i) Mass
(c) Chemical property	(j) Matter
(d) Compound	(k) Physical property
(e) Density	(l) Physical change
(f) Elements	(m) Significant figures
(g) Extensive	

True-False Statements:

1.14 When ice completely melts in a glass of water, there is a change from a heterogeneous *mixture* to a homogeneous one.

1.15 The *SI unit* for mass is the gram.

1.16 A *conversion factor* contains a ratio of units.

1.17 Coke is a *substance.*

1.18 −273.15 °C is equivalent to 0 K.

1.19 A student measures the mass of an object three times: 3.60 g, 3.90 g, and 3.75 g. The object actually has a mass of 3.75 g. The student's measurements show good *precision.*

1.20 In problem 1.19, the average value is 3.75 g. Therefore, the student's measurements gave good *accuracy.*

1.21 A *solution* is a homogeneous mixture of two or more substances.

1.22 The basic unit of mass in the *metric system* is the same as in the SI system of units.

1.23 A *chemical reaction* involves a change in the chemical composition of substances.

1.24 A *change of state* of $N_2(l)$ to $N_2(g)$ is a chemical change.

1.25 10 g of CO_2 *gas* occupies less space than 10 g of CO_2 solid in a one liter flask.

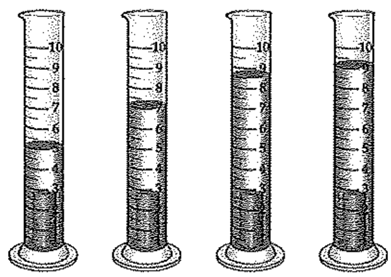
1.26 A *solid* consists of particles closer together than in the gas state.

1.27 A *liquid* is slightly compressible.

1.28 According to the *law of constant composition (definite proportions)*, H_2O and H_2O_2 represent the same substance.

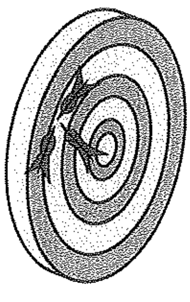
Problems and Short-Answer Questions

1.29 Shown below is a 10 mL graduate cylinder. It is initially filled with 5.0 mL of water. 4.0 g of a solid object with a density of 2.0 g/mL is added. Which figure below best represents the volume of water in the graduate cylinder after the solid is added?

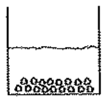


(a) (b) (c) (d)

1.30 Characterize the following dartboard in terms of precision and accuracy of the results.



1.31 Identify the type of matter described by the following figure:



1.32 You calculate your income for last year as \$42,125.00. However, when you file your income tax form, you report an income of \$42,000.00. You do this because you want to report a rounded-off income figure to two significant figures. Why would an IRS agent be unhappy with your use of significant figures?

1.33 How many significant figures do the following numbers possess?

(a) 20.03 kg
(b) 1.90×10^3 L
(c) 120 m
(d) 10 dollar bills
(e) 0.00067 cm^3

1.34 Convert the following numbers to scientific notation form with the correct number of significant figures.

(a) 0.00067 cm^3
(b) 210.0 m
(c) 0.040 L
(d) 46900.0 g
(e) 200 rattlesnakes

1.35 Complete the following calculations and round off answers to the correct number of significant figures:

(a) $11.020 + 300.0 + 2.0030 =$
(b) $1211 + 2.205 - 1.70 =$
(c) $\dfrac{(1.425 \times 10^2)(2.61 \times 10^3)}{2.89 \times 10^5} =$
(d) $\dfrac{(0.012)(0.100)}{11.0265} =$

1.36 Make the following conversions.

(a) 1 ML to cubic centimeters
(b) $\dfrac{1}{1000}$ g to kg
(c) 8.00 qt to milliliters
(d) 16 lb to grams

1.37 Make the following temperature conversions.

(a) −70 °F in the Antarctic (it gets cold there) to °C
(b) 480 °C at the surface of Venus to °F
(c) 0.95 K, the freezing point of helium gas, to °C

1.38 Calculate the density of bromine, given that a 125.0 mL sample weighs 375.0 g.

1.39 A gold bar has the following dimensions: 2.50 cm × 2.00 cm × 1.50 cm. Assuming that gold can be sold for \$400/oz, what is the value of the gold bar? The density of gold is 19.32 g/cm^3, and 1 lb = 453.6 g.

1.40 A column of mercury is contained in a cylindrical tube. This tube has a diameter of 8.0 mm, and the height of the mercury column is 1.20 m. Given that the density of mercury is 13.6 g/cm^3 and that the volume of mercury in the tube can be calculated from the relation $V = \pi r^2 h$ (r radius of the tube and h is the height of mercury column), calculate the mass of mercury present in the cylindrical tube.

1.41 Substances A, B, and C are all liquids and immiscible (do not mix) in each other. It is observed that A lies on top of C and C lies on top of B. Do you have sufficient information to determine the relative densities of A, B, and C? Explain.

1.42 The density of water at 3 °C is greater than its density at 0 °C. What does this information tell you about change in the volume of water from 3 °C to 0 °C? Are water molecules coming closer together or further apart? Explain.

Multiple-Choice Questions

1.43 Which of the following is not correct?

(a) There are 1000 mg in a gram.
(b) There are 100 cm in a meter.
(c) There are 1000 mL in a liter.
(d) There are 100 mm in a centimeter.
(e) There are 1000 μJ in a millijoule.

1.44 Which is the standard unit of volume in the SI system?

(a) meter
(b) meter^3
(c) liter
(d) milliliter
(e) gallon

1.45 Which prefix means $\dfrac{1}{1{,}000{,}000}$ of a unit?

(a) kilo-
(b) centi-
(c) milli-
(d) micro-
(e) nano-

1.46 2.5 nm equals:

(a) 2.5×10^{-9} m
(b) 2.5×10^{-4} mm
(c) 2.5×10^{-7} cm
(d) (a) and (b)
(e) (a) and (c)

1.47 Which number has exactly four significant figures?

(a) 0.020
(b) 2000
(c) 0.2000
(d) 0.020
(e) 2210

1.48 When 1210.42 is rounded to three significant figures, it becomes:

(a) 1210
(b) 121
(c) 1000
(d) 1210.
(e) 210.42

1.49 Solid carbon dioxide, dry ice, changes directly from a solid to a vapor at 195 K if left in an open container. What is this temperature in degrees Celsius and Fahrenheit?

(a) −78 °C, 468 °F
(b) −108 °C, 468 °F
(c) 468 °C, −108 °F
(d) −108 °C, −78 °F
(e) −78 °C, −108 °F

1.50 Density can be thought of as a conversion factor. For example, the density of aluminum is 2.70 g/cm^3, which can be interpreted to mean 2.70 g of aluminum = 1 cm^3. Using this equivalence, what is the volume occupied by 223.5 g of aluminum?

(a) 603 cm^3
(b) 82.8 cm^3
(c) 0.0121 cm^3
(d) 0.0270 cm^3
(e) 223.5 cm^3

1.51 When a solid substance undergoes a physical change to a liquid, which of the following is always true?

(a) A new substance is formed.
(b) Heat is given off.
(c) A gas is given off.
(d) It vaporizes.
(e) It melts.

1.52 An empty container weighs 15.230 g. When filled with water (density = 1.00 g/mL), it weighs 35.920 g. When filled with an unknown liquid to the same mark as it was filled to with the water, it weighs 36.261 g. What is the density of the unknown liquid?

(a) 1.02 g/mL
(b) 1.02 g/m^3
(c) 1.20 g/mL
(d) 1.20 g/m^3
(e) none of the above

1.53 The maximum speed limit on many interstate highways is 70 mi/hr. How many kilometers can you travel in 4.5 hours at this speed?

(a) 75
(b) 150
(c) 245
(d) 390
(e) 500

1.54 Which is *not* a characteristic property of a compound useful in its identification?

(a) chemical formula
(b) density
(c) melting point temperature
(d) mass
(e) elemental composition

1.55 The copper content of a normal healthy human person is approximately 1.1×10^{-4} percent by mass. How many grams of copper would exist in a person weighing 1.00×10^3 lb (1.0 kg = 2.2 lb)

(a) 0.00050 g
(b) 0.050 g
(c) 0.50 g
(d) 5.0 g
(e) 50.0 g

1.56 A pure solid is heated and it decomposes into two substances, one a liquid and the other a gas. One can conclude with certainty that:

(a) The two products are elements.
(b) One of the two products is an element.
(c) The original solid is not an element.
(d) The liquid is a compound and the gas is an element.
(e) Both products are compounds.

1.57 When 125 mg, 1.2 dg and 1.2223 g are added, how many significant figures does the answer have?

(a) two
(b) three
(c) four
(d) five
(e) six

1.58 An intensive property of matter

(a) depends on the size of the sample.
(b) may depend on a ratio of two extensive properties.
(c) is a property that cannot be easily measured.
(d) does not depend on temperature.
(e) cannot be used to characterize matter.

1.59 A student determines the mass of silver in a sample and does four determinations. The results are 1.75 g, 1.71 g, 1.85 g, and 1.93 g. The true value is 1.81 g. Which statement concerning the results is correct?

(a) High precision and accurate results.
(b) High precision and poor accuracy.
(c) Poor precision and poor accuracy.
(d) Poor precision and accurate results.
(e) Reasonable precision and poor accuracy.

SELF-TEST SOLUTIONS

1.1 (e). **1.2** (m). **1.3** (j). **1.4** (i). **1.5** (h). **1.6** (g). **1.7** (k). **1.8** (l). **1.9** (b). **1.10** (c). **1.11** (f). **1.12** (d). **1.13** (a). **1.14** True.

1.15 False. It is the kilogram. **1.16** True. **1.17** False. It is a mixture of substances. **1.18** True. **1.19** False. If the student's measurements had shown good precision, the masses would have been closer in value, not showing a range of 0.30 g between low and high values, but a much smaller range. **1.20** True. **1.21** True. **1.22** False. In the metric system it is the gram, and in the SI system it is the kilogram. **1.23** True. **1.24** False. A physical change—its composition does not change. **1.25** False. A gas expands to fill the entire vessel holding it. A solid only occupies a limited and lesser amount of space. **1.26** True. **1.27** True. **1.28** False. In the case of water the ratio of hydrogen to oxygen is 2:1, whereas in the case of hydrogen peroxide (H_2O_2) the ratio is 2:2 or 1:1. Thus, they are different substances.

1.29 (b) V = m/d = 4.0 g/2.0 g/mL = 2.0 mL. The object will displace 2.0 mL of water. When added to the original 5.0 mL of water, the total volume becomes 7.0 mL.

1.30 The darts are close to one another; thus the precision is very good. However, the darts are far from the center; thus the throws result in dart positions that are not accurate.

1.31 The figure show two different phases, solid and liquid; thus, it is a heterogeneous mixture.

1.32 The IRS requires that all digits to the left of the decimal be significant when you report your income. You may round off the cents to the nearest dollar. Your actual income of $42,125.00 has five significant figures to the left of the decimal. Your reported income of $42,000.00 implies that the three zeros to the right of the 2 are significant and certain. The IRS agent is unhappy because you underreported your income and underpaid your taxes.

1.33 (a) four; (b) three; (c) two; (d) infinite, because this is an exact number; (e) two, because the zeros immediately to the right of the decimal are not significant.

1.34 (a) 0.00067 $cm^3 = 6.7 \times 10^{-4}$ cm^3;
(b) 210.0 = 2.100×10^2 m;
(c) 0.040 L = 4.0×10^{-2} L;
(d) 46900.0 g = 4.69000×10^4 g;
(e) 200 rattlesnakes = 2.00×10^2 rattlesnakes. Because this is an exact quantity, we need only write 2×10^2 rattlesnakes.

1.35 (a) $11.020 + 300.0 + 2.0030 = 313.0230$. The answer can have only one digit to the right of the decimal because 300.0 has the fewest number of digits to the right of the decimal—that is, one. The answer is rounded off to 313.0.

(b) $1211 + 2.205 - 1.70 = 1211.505$. The answer can have no digits to the right of the decimal because 1211 has no digits to the right of the decimal. The answer is rounded off to 1212.

(c) $$\frac{(1.425 \times 10^2)(2.61 \times 10^3)}{2.89 \times 10^5} = 1.2869377$$

The numbers with the fewest significant figures in the calculation are 2.61×10^3 and 2.89×10^5 with three significant figures. The final answer is rounded off to three significant figures: 1.29.

(d) First convert all numbers to exponential notation and then solve:

$$\frac{(0.012)(0.100)}{11.0265} = \frac{(1.2 \times 10^{-2})(1.00 \times 10^{-1})}{1.10265 \times 10^1}$$
$$= 1.08828 \times 10^{-4}$$

The number with the fewest significant figures in the calculation is 1.2×10^{-2}, with two significant figures. The final answer is rounded off to two significant figures: 1.1×10^{-4}.

1.36 (a) $$(1\ \text{ML})\left(\frac{10^6\ \text{L}}{1\ \text{ML}}\right)\left(\frac{10^3\ \text{mL}}{1\ \text{L}}\right)\left(\frac{1\ \text{cm}^3}{1\ \text{mL}}\right) = 1 \times 10^9\ \text{cm}^3$$

(b) $$\left(\frac{1}{1000}\ \text{g}\right)\left(\frac{1\ \text{kg}}{1000\ \text{g}}\right)$$
$$= (1 \times 10^{-3}\ \text{g})\left(\frac{1 \times 10^{-3}\ \text{kg}}{1\ \text{g}}\right) = 1 \times 10^{-6}\ \text{kg}$$

(c) $$(8.00\ \text{qt})\left(\frac{1\ \text{L}}{1.057\ \text{qt}}\right)\left(\frac{1000\ \text{mL}}{1\ \text{L}}\right) = 7.57 \times 10^3\ \text{mL}$$

(d) $$(16\ \text{lb})\left(\frac{453.6\ \text{g}}{\text{lb}}\right) = 7.3 \times 10^3\ \text{g}$$

1.37 (a) $$°\text{C} = (-70\ °\text{F} + 40\ °)\left(\frac{1\ °\text{C}}{1.8\ °\text{F}}\right) - 40\ ° = -57\ °\text{C}$$

(b) $$°\text{F} = (480\ °\text{C} + 40\ °)\left(\frac{1.8\ °\text{F}}{1\ °\text{C}}\right) - 40\ ° = 896\ °\text{F}$$

(c) $$\text{K} = \left(\frac{1\ \text{K}}{1\ °\text{C}}\right)(°\text{C}) + 273.15\ \text{K} \quad \text{or}$$

$$°\text{C} = \left(\frac{1\ °\text{C}}{1\ \text{K}}\right)(\text{K} - 273.15\ \text{K})$$

$$= \left(\frac{1\ °\text{C}}{1\ \text{K}}\right)(0.95\ \text{K} - 273.15\ \text{K}) = -272.20\ °\text{C}$$

1.38 Density of bromine = mass/volume = $375.0g/125.0\ \text{mL} = 3.000\ \text{g/mL}$.

1.39 The mass of the gold bar is calculated from its density using the relation mass = density × volume. The volume of the gold bar equals the product of its length times width times height: $V = (2.50\ \text{cm})(2.00\ \text{cm})(1.50\ \text{cm}) = 7.50\ \text{cm}^3$. Calculating the mass of the gold bar with this volume: m = density × volume = $(19.32\ \text{g/cm}^3)(7.50\ \text{cm}^3) = 145\ \text{g}$. The value of the gold bar is $\$400/\text{oz} \times 16\ \text{oz/lb} \times 1\ \text{lb}/453.6\ \text{g} \times 145\ \text{g} = \$2046$. (Note: The gold dealer normally does not adhere to significant figure rules!)

1.40 Because the density is expressed in units of cm^3, it is convenient to change all length measurements into that unit. For this problem, the required equivalences are 10 mm = 1 cm and 100 cm = 1 m. Calculating the volume of the tube the mercury occupies in the unit of cm^3:

$$V = \pi r^2 h = \pi\left(\frac{8.00\ \text{mm}}{2}\right)^2\left(\frac{1\ \text{cm}}{10\ \text{mm}}\right)^2$$
$$\times (1.20\ \text{m})\left(\frac{100\ \text{cm}}{1\ \text{m}}\right)$$
$$= 60.3\ \text{cm}^3$$

Mass of mercury = density × volume

$$= \left(13.6\frac{\text{g}}{\text{cm}^3}\right)(60.3\ \text{cm}^3)$$
$$= 820\ \text{g}$$

1.41 The information tells you that C is more dense than A as A lies on top of C. Furthermore you know that B is more dense than C as C lies on top of B. If density(A) < density(C) and density(C) < density(B), it must also be true that density(A) < density(B). Thus, the problem provides sufficient information to conclude that the trend in relative densities is density(A) < density(B) < density(C).

1.42 Density = mass/volume or volume = mass/density. If the density of water decreases, then the volume increases because mass is divided by a smaller number as the temperature decreases. An increase in volume suggests that the molecules must be further apart from one another as the temperature decreases. If they were more closely associated there would be less space between molecules and the volume would decrease.

1.43 (d) 100 mm = (100 mm)(1 m/1000 mm) = 1/10 m; 1 cm = 1/100 m. Thus the two quantities are not equal.

1.44 (b)

1.45 (d)

1.46 (e)

$$(2.5\ \text{nm})\left(\frac{10^{-9}\ \text{m}}{1\ \text{nm}}\right) = 2.5 \times 10^{-9}\ \text{m}$$

$$(2.5\ \text{nm})\left(\frac{10^{-9}\ \text{m}}{1\ \text{nm}}\right)\left(\frac{100\ \text{cm}}{1\ \text{m}}\right) = 250 \times 10^{-9}\ \text{cm}$$
$$= 2.5 \times 10^{-7}\ \text{cm}$$

1.47 (c) The zeroes after the two in 0.2000 are significant because they do not define the decimal point: 2.000×10^{-1}.

1.48 (a) The number of significant figures is counted from the left and trailing zeroes in a number without a decimal point are considered not significant.

1.49 (e)

1.50 (b) $(223.5\ \text{g})(1\ \text{cm}^3/2.70\ \text{g}) = 82.8\ \text{cm}^3$

1.51 (e)

1.52 (a) Mass of water = 35.920 g − 15.230 g = 20.690 g; mass of liquid = 36.261 g − 15.230 g = 21.031 g; volume occupied by liquid = (mass water) (1.00 mL/1 g) = 20.7 mL; density of liquid = m/V = 21.031 g/20.7 mL = 1.02 g/mL.

1.53 (d) $$\left(70\ \frac{\text{mi}}{\text{hr}}\right)(3.5\ \text{hr})\left(1.609\ \frac{\text{km}}{\text{mi}}\right)$$
$$= 390\ \text{km (rounded to two significant figures)}$$

1.54 (d) A compound has a specific formula and elemental composition, density, and melting point. Its mass depends on the amount present and is not useful for identification.

1.55 (c)
$$(1.00 \times 10^3\ \text{lb})\left(\frac{1000\ \text{g}}{2.2\ \text{lb}}\right)(1.1 \times 10^{-4}/100) = 0.50\ \text{g}$$

1.56 (c) The original solid is pure and it must be a compound because it can be chemically decomposed into simpler substances. You are not given specific information about the composition of the products and thus you cannot make any definite conclusions about them.

1.57 (b) First convert all numbers to the same units: 125 mg is 0.125 g; 1.2 dg is 0.12 g; and 1.2223 g is unchanged. The sum of these numbers is 1.4673 g. The final answer must be rounded to two digits to the right of the decimal: 1.47 g. Thus, the answer has three significant figures.

1.58 (b) Density is an example of an intensive property that depends on the ratio of two extensive properties, mass and volume. An intensive property does not depend on the amount of substance present.

1.59 (d) The average of the measurements is 1.81 g, which is the same as the true value. Therefore, the accuracy is high. The range of measured values is 1.75 g − 1.93 g, or a difference of 0.18 g, or 10% of the average value, which is poor precision. Precision is a measure of how close the four masses are to each other.