1 The Foundations of Chemistry

1-2. ***Refer to the Introduction to Chapter 1 and a dictionary.***

(a) Organic chemistry is the study of the chemical compounds of carbon and hydrogen and a few other elements.

(b) Forensic chemistry deals with the chemistry involved in solving crimes, including chemical analyses of crime scene artifacts, such as paint chips, dirt, fluids, blood, and hair.

(c) Physical chemistry is the study of the part of chemistry that applies the mathematical theories and methods of physics to the properties of matter and to the study of chemical processes and the accompanying energy changes.

(d) Medicinal chemistry is the study of the chemistry and biochemistry dealing with all aspects of the medical field.

1-4. ***Refer to the Sections 1-1, 1-4, 1-8, 1-13 and the Key Terms for Chapter 1.***

(a) Weight is a measure of the gravitational attraction of the earth for a body. Although the mass of an object remains constant, its weight will vary depending on its distance from the center of the earth. One kilogram of mass at sea level weighs about 2.2 pounds (9.8 newtons), but that same one kilogram of mass weighs less at the top of Mt. Everest. In more general terms, it is a measure of the gravitational attraction of one body for another. The weight of an object on the moon is about 1/7th that of the same object on the earth.

(b) Potential energy is the energy that matter possesses by virtue of its position, condition, or composition. Your chemistry book lying on a table has potential energy due to its position. Energy is released if it falls from the table.

(c) Temperature is a measurement of the intensity of heat, *i.e.* the "hotness" or "coldness" of an object. The temperature at which water freezes is 0°C or 32°F.

(d) An endothermic process is a process that absorbs heat energy. The boiling of water is a physical process that requires heat and therefore is endothermic.

(e) An extensive property is a property that depends upon the amount of material in a sample. Extensive properties include mass and volume.

1-6. ***Refer to the Section 1-1 and the Key Terms for Chapter 1.***

A reaction or process is exothermic, in general, if heat energy is released, but other energies may be released.

(a) The discharge of a flashlight battery in which chemical energy is converted to electrical energy is referred to as being exothermic the chemical reaction occurring in the battery releases heat.

(b) An activated light stick produces essentially no heat, but is considered to be exothermic because light is emitted.

1-8. ***Refer to Sections 1-1 and 1-5, and the Key Terms for Chapter 1.***

(a) Combustion is an exothermic process in which a chemical reaction releases heat.

(b) The freezing of water is an exothermic process. Heat must be removed from the molecules in the liquid state to cause solidification.

(c) The melting of ice is an endothermic process. The system requires heat to break the attractive forces that hold solid water together.

(d) The boiling of water is an endothermic process. Molecules of liquid water must absorb energy to break away from the attractive forces that hold liquid water together in order to form gaseous molecules.

(e) The condensing of steam is an exothermic process. The heat stored in water vapor must be removed for the vapor to liquefy. The condensation process is the opposite of boiling which requires heat.

(f) The burning of paper is an exothermic process. The heat generated can be used to light the wood in a fireplace.

1-10. ***Refer to Section 1-1.***

Einstein's equation, written as $E = mc^2$, tells us that the amount of energy released when matter is transformed into energy is the product of the mass of matter transformed and the speed of light squared. From this equation, we see that energy and matter are equivalent. Known as the Law of Conservation of Matter and Energy, we can use this equation to calculate the amount of energy released in a nuclear reaction because it is proportional to the difference in mass between the products and the reactants. The energy released (in joules) equals the mass difference (in kilograms) times the square of the speed of light (in m/s).

1-12. ***Refer to Section 1-1.***

Electrical motors are less than 100% efficient in the conversion of electrical energy into useful work, since a part of that energy is converted into frictional heat which radiates away.

However, the Law of Conservation of Energy still applies:

electrical energy = useful work + heat

1-14. ***Refer to Section 1-3 and Figures 1-7 and 1-8.***

Solids: are rigid and have definite shapes;
they occupy a fixed volume and are thus very difficult to compress;
the hardness of a solid is related to the strength of the forces holding the particles of a solid together; the stronger the forces, the harder is the solid object.

Liquids: occupy essentially constant volume but have variable shape;
they are difficult to compress;
particles can pass freely over each other;
their boiling points increase with increasing forces of attraction among the particles.

Gases: expand to fill the entire volume of their containers;
they are very compressible with relatively large separations between particles.

The three states are alike in that they all exhibit definite mass and volume under a given set of conditions. All consist of some combination of atoms, molecules or ions. The differences are stated above. Additional differences occur in their relative densities:

gases <<< liquids < solids.

Molecular representations of these three phases can be seen in Figure 1-8. Note that water is an exceptional compound. The density of the liquid is greater than the solid phase. That is why solid ice floats in liquid water

1-16. ***Refer to Section 1-6 and the Key Terms for Chapter 1.***

(a) A substance is a kind of matter in which all samples have identical chemical composition and physical properties, e.g., iron (Fe) and water (H_2O).

(b) A mixture is a sample of matter composed of two or more substances in variable composition, each substance retaining its identity and properties, e.g., soil (minerals, water, organic matter, living organisms, etc.) and seawater (water, different salts, dissolved gases, organic compounds, living organisms, etc.).

(c) An element is a substance that cannot be decomposed into simpler substances by chemical means, e.g., nickel (Ni) and nitrogen (N).

(d) A compound is a substance composed of two or more elements in fixed proportions. Compounds can be decomposed into their constituent elements by chemical means. Examples include water (H_2O) and sodium chloride (NaCl).

1-18. ***Refer to Section 1-6.***

(a) Gasoline is a homogeneous liquid mixture of organic compounds distilled from oil.

(b) Tap water is a homogeneous liquid mixture, called an aqueous solution, containing water, dissolved salts, and gases such as chlorine and oxygen.

(c) Calcium carbonate is a compound, $CaCO_3$, consisting of the elements Ca, C and O in the fixed atomic ratio, 1:1:3.

(d) Ink from a ball-point pen is a homogeneous mixture of solvent, water and dyes.

(e) Vegetable soup is a heterogeneous mixture of water, vegetables and the compound, NaCl (table salt), depending on the recipe.

(f) Aluminum foil is composed of the metallic element, Al.

1-20. ***Refer to Section 1-6.***

The coin is a heterogeneous mixture of gold and copper because it consists of two distinguishable elements that can be recognized on sight.

1-22. ***Refer to Section 1-4.***

(a) Striking a match, causing it to burst into flames, is a chemical property, since a change in composition is occurring of the substances in the match head and new substances including carbon dioxide gas and water vapor, are being formed.

(b) The hardness of steel is a physical property. It can be determined without a composition change.

(c) The density of gold is a physical property, since it can be observed without any change in the composition of the gold.

(d) The ability of baking soda to dissolve in water with the evolution of carbon dioxide gas is a chemical property of baking soda, since during the reaction, its composition is changing and a new substance is being formed.

(e) The ability of fine steel wool to burn in air is a chemical property of steel wool since a compositional change in the steel wool occurs and heat is released.

(f) The ripening of fruit is a chemical property. When the temperature of the fruit decreases when put into a refrigerator, the rate of the chemical reaction slows. So, the lowering of the fruit's temperature is a physical change, but temperature has a definite effect on the chemical properties of the fruit.

1-24. ***Refer to Section 1-5.***

The observations that identify chemical properties are: (c) ultraviolet light converts ozone into oxygen, (e) sodium metal reacts violently with water, and (f) CO_2 does not support combustion.

Some chemists think that dissolution is a chemical process, since it is actually very complex, so some chemists would include (a).

1-26. ***Refer to Section 1-1 and the Key Terms for Chapter 1.***

(b), (d) and (e) are examples of potential energy. An inflated balloon (b) possesses energy which will be released if it is popped. The stored chemical energy in a flashlight battery (d) will convert to electrical energy, then into kinetic energy once it is put to use. A frozen lake (e) is stored energy. Once spring comes, the water molecules will be free to move, the lake will be circulating and the energy will convert to kinetic energy. However, a lake can also be a source of potential energy that can be converted into kinetic energy if the water is released via a dam.

(a), (c) and (f) are all examples of kinetic energy due to their motion.

1-28. ***Refer to Section 1-5.***

When the sulfur is heated, some of it obviously became a gas. However, there is not enough information to tell whether or not this was the result of a physical or a chemical change.

Hypothesis 1: Solid sulfur could be changing directly into gaseous sulfur. This is a physical change called sublimation.

Hypothesis 2: Solid sulfur could be reacting with oxygen in the air to form a gaseous compound consisting of sulfur and oxygen. This would be a chemical change. The sharp odor may indicate the presence of SO_2, but the smell test is not conclusive.

To verify which hypothesis is correct, we need to identify the gas that is produced.

1-30. ***Refer to Appendix A.***

(a) $423.\underline{00}6$ mL = 4.23006×10^2 mL (6 significant figures)

(b) $0.001\underline{0}73\underline{0}4\underline{0}$ g = 1.073040×10^{-3} g (7 significant figures)

(c) $1\underline{0}81.\underline{0}2$ pounds = 1.08102×10^3 pounds (6 significant figures)

1-32. ***Refer to Appendix A.***

(a) 50600 (c) 0.1610 (e) 90000.

(b) 0.0004060 (d) 0.000206 (f) 0.0009000

1-34. ***Refer to Appendix A.***

? volume (cm^3) = 252.56 cm x 18.23 cm x 6.5 cm = 29927 = $\mathbf{3.0 \times 10^4\ cm^3}$ (2 significant figures based on 6.5 cm)

1-36. ***Refer to Section 1-9, the conversion factors from Tables 1-6 and 1-8, and Examples 1-3 and 1-4.***

(a) $? \text{ km} = 453.4 \text{ m} \times \frac{1 \text{ km}}{1000 \text{ m}} = \mathbf{0.4534 \text{ km}}$

(b) $? \text{ m} = 36.3 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} = \mathbf{3.63 \times 10^4 \text{ m}}$

(c) $? \text{ g} = 487 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = \mathbf{4.87 \times 10^5 \text{ g}}$

(d) $? \text{ mL} = 1.32 \text{ L} \times \frac{1000 \text{ mL}}{1\text{L}} = \mathbf{1.32 \times 10^3 \text{ mL}}$

(e) $? \text{ L} = 55.9 \text{ dL} \times \frac{1 \text{ L}}{10 \text{ dL}} = \mathbf{5.59 \text{ L}}$

(f) $? \text{ cm}^3 = 6251 \text{ L} \times \frac{1000 \text{ cm}^3}{1 \text{ L}} = \mathbf{6.251 \times 10^6 \text{ cm}^3}$ (Note: $1 \text{ cm}^3 = 1 \text{ mL}$)

1-38. ***Refer to Section 1-9, the conversion factors listed in Table 1-8, and Example 1-9.***

$? \text{ cents/L} = \frac{\$3.119}{1 \text{ gal}} \times \frac{1 \text{ gal}}{4 \text{ qt}} \times \frac{1.057 \text{ qt}}{1 \text{ L}} \times \frac{100 \text{ cents}}{\$1} = \mathbf{82.42 \text{ cents/L}}$

1-40. ***Refer to Section 1-10, the conversion factors from Table 1-8, and Examples 1-7 and 1-9.***

(a) $? \text{ L} = 0.750 \text{ ft}^3 \times \frac{(12 \text{ in})^3}{(1 \text{ ft})^3} \times \frac{(2.54 \text{ cm})^3}{(1 \text{ in})^3} \times \frac{1 \text{ L}}{1000 \text{ cm}^3} = \mathbf{21.2 \text{ L}}$

(b) $? \text{ pints} = 1.00 \text{ L} \frac{1.057 \text{ qt}}{1 \text{ L}} \times \frac{2 \text{ pt}}{1 \text{ qt}} = \mathbf{2.11 \text{ pt}}$

(c) $? \frac{\text{km}}{\text{L}} = \frac{1 \text{ mile}}{1 \text{ gal}} \times \frac{1.609 \text{ km}}{1 \text{ mile}} \times \frac{1 \text{ gal}}{4 \text{ qt}} \times \frac{1.057 \text{ qt}}{1 \text{ L}} = \mathbf{0.4252 \frac{km}{L}}$

Therefore, to convert miles per gallon to kilometers per liter, one multiplies the miles per gallon by the factor, 0.4252.

1-42. ***Refer to Appendix A.***

$\text{Average} = \frac{58.2 + 56.474}{2} = 57.337 = \mathbf{57.3\ \%}$ since the answer must be rounded to the tenths place

1-44. ***Refer to Section 1-9, Appendix A, the conversion factors from Table 1-8 and Example 1-9.***

(a) $18 \text{ pints} \times \frac{1 \text{ qt}}{2 \text{ pints}} = \mathbf{9.0 \text{ qt}}$

(b) $\frac{55.0 \text{ miles}}{\text{hr}} \times \frac{1.609 \text{ km}}{1 \text{ mile}} = \mathbf{88.5 \text{ km/hr}}$

(c) $15.45 \text{ s} + 2.2 \text{ s} + 55 \text{ s} = 72.65 \text{ s} = \mathbf{73 \text{ s}}$ since the answer must be rounded to the one's place.

1-46. ***Refer to Section 1-11, and Examples 1-11 and 1-12.***

Density (mg/mm^3) $= \dfrac{m}{V} = \dfrac{6.080 \text{ mg}}{(2.20 \text{ mm} \times 1.36 \text{ mm} \times 1.23 \text{ mm})} = 1.65 \text{ mg/mm}^3$

Density (g/cm^3) $= \dfrac{1.65 \text{ mg}}{1 \text{ mm}^3} \times \dfrac{1 \text{ g}}{1000 \text{ mg}} \times \dfrac{(10 \text{ mm})^3}{(1 \text{ cm})^3} = \mathbf{1.65\ g/cm^3}$

1-48. ***Refer to Section 1-11 and Example 1-12.***

(a) Method 1: $D = \dfrac{m}{V}$; $V\,(\text{cm}^3) = \dfrac{m\,(\text{g})}{D\,(\text{g/cm}^3)} = \dfrac{443 \text{ g}}{10.5 \text{ g/cm}^3} = \mathbf{42.2\ cm^3}$ since 0.443 kg $\equiv$ 443 g

Method 2: Dimensional Analysis

$? \text{ cm}^3 \text{ silver} = 0.443 \text{ kg} \times \dfrac{1000 \text{ g}}{1 \text{ kg}} \times \dfrac{1 \text{ cm}^3}{10.5 \text{ g}} = \mathbf{42.2\ cm^3}$

(b) length of each edge (cm) $= \sqrt[3]{V} = \sqrt[3]{42.2 \text{ cm}^3} = \mathbf{3.48\ cm}$

(c) length of each edge (in.) $= 3.48 \text{ cm} \times \dfrac{1 \text{ in.}}{2.54 \text{ cm}} = \mathbf{1.37\ in.}$

1-50. ***Refer to Section 1-11.***

Plan: (1) Find the volume of the aluminum wire, assuming that 10-lb spool contains 10.0 lb of aluminum
(2) Calculate the radius of the wire in meters.
(3) Solve for the length of wire in meters, using $V = \pi r^2 \ell$

(1) $? V = 10.0 \text{ lb Al} \times \dfrac{453.6 \text{ g Al}}{1 \text{ lb Al}} \times \dfrac{1 \text{ cm}^3 \text{ Al}}{2.70 \text{ g Al}} \times \dfrac{1 \text{ m}^3 \text{ Al}}{(100 \text{ cm})^3 \text{ Al}} = 1.68 \times 10^{-3} \text{ m}^3 \text{ Al}$

(2) $? \text{ radius, } r = \text{diameter}/2 = \dfrac{0.0808 \text{ in.}}{2} \times \dfrac{2.54 \text{ cm}}{1 \text{ in.}} \times \dfrac{1 \text{ m}}{100 \text{ cm}} = 1.03 \times 10^{-3} \text{ m}$

(3) $? \text{ length, } \ell = \dfrac{V}{\pi r^2} = \dfrac{1.68 \times 10^{-3} \text{ m}^3}{3.1416(1.03 \times 10^{-3} \text{ m})^2} = \mathbf{504\ m}$

1-52. ***Refer to Sections 1-10 and 1-11.***

Plan: L solution $\xRightarrow{(1)}$ mL solution $\xRightarrow{(2)}$ g solution $\xRightarrow{(3)}$ g iron(III) chloride

Using 3 unit factors,

(1) Convert liters to milliliters using 1000 mL = 1 liter,
(2) Convert mL of solution to mass of solution using density, then
(3) Convert mass of solution to mass of iron(III) chloride using the definition of % by mass.

$? \text{ g iron(III) chloride} = 2.50 \text{ L soln} \times \dfrac{1000 \text{ mL soln}}{1 \text{ L soln}} \times \dfrac{1.149 \text{ g soln}}{1 \text{ mL soln}} \times \dfrac{11 \text{ g iron(III) chloride}}{100 \text{ g soln}} = \mathbf{3.2 \times 10^2\ g}$

1-54. ***Refer to Appendix A, Section 1-12, and Examples 1-16 and 1-17.***

In determining the correct number of significant figures, note that the following values are exact: 32°F, 1°C/1.8°F, and 1°C/1 K and have an infinite number of significant figures.

(a) $? \text{ °C} = \dfrac{1\text{°C}}{1.8\text{°F}} \times (15\text{°F} - 32\text{°F}) = \mathbf{-9.4\text{°C}}$

(b) $? \text{°C} = \frac{1\text{°C}}{1.8\text{°F}} \times (32.6\text{°F} - 32.0\text{°F}) = 0.6\text{°C}$ (1 sig. fig. due to subtraction rules)

$? \text{K} = \frac{1 \text{K}}{1\text{°C}} \times (0.6\text{°C} + 273.2\text{°C}) = \mathbf{273.8\ K}$ since 0°C = 273.15 K

(c) $? \text{°C} = \frac{1\text{°C}}{1 \text{K}} \times (328 \text{K} - 273 \text{K}) = 55\text{°C}$

$? \text{°F} = \left(55\text{°C} \times \frac{1.8\text{°F}}{1\text{°C}}\right) + 32\text{°F} = \mathbf{130°F}$ (2 sig. figs.)

(d) $? \text{°F} = \left(11.3\text{°C} \times \frac{1.8\text{°F}}{1\text{°C}}\right) + 32\text{°F} = \mathbf{52.3°F}$

1-56. ***Refer to Section 1-12.***

	Freezing Point of Water (FP)	Boiling Point of Water (BP)
Celsius Scale	0°C	100°C
Fahrenheit Scale	32°F	212°F
Rėamur Scale	0°R	80°R

(a) $\frac{\text{BP}_{\text{water}} - \text{FP}_{\text{water}} \text{ on Celsius Scale}}{\text{BP}_{\text{water}} - \text{FP}_{\text{water}} \text{ on Rėamur Scale}} = \frac{100\text{°C} - 0\text{°C}}{80\text{°R} - 0\text{°R}} = \frac{100\text{°C}}{80\text{°R}} = \frac{1.0\text{°C}}{0.8\text{°R}} = \frac{5\text{°C}}{4\text{°R}}$

Therefore, since both scales set the freezing point of water = 0°, then $? \text{°C} = \left(x\text{°R} \times \frac{5\text{°C}}{4\text{°R}}\right)$

(b) $\frac{\text{BP}_{\text{water}} - \text{FP}_{\text{water}} \text{ on Fahrenheit Scale}}{\text{BP}_{\text{water}} - \text{FP}_{\text{water}} \text{ on Rėamur Scale}} = \frac{212\text{°F} - 32\text{°F}}{80\text{°R} - 0\text{°R}} = \frac{180\text{°F}}{80\text{°R}} = \frac{9\text{°F}}{4\text{°R}}$

Therefore, $? \text{°F} = \left(x\text{°R} \times \frac{9\text{°F}}{4\text{°R}}\right) + 32\text{°F}$

Note that we must add 32°F to account for the fact that 0°R is equivalent to 32°F.

(c) From (a), $? \text{°C} = \left(x\text{°R} \times \frac{5\text{°C}}{4\text{°R}}\right)$ Rearranging, we have $? \text{°R} = \left(x\text{°C} \times \frac{4\text{°R}}{5\text{°C}}\right)$

$\text{BP}_{\text{mercury}} (\text{°R}) = 356.6\text{°C} \times \frac{4\text{°R}}{5\text{°C}} = \mathbf{285.3°R}$

1-58. ***Refer to Section 1-12 and Examples 1-16 and 1-17.***

For Al: $? \text{°C} = \frac{1\text{°C}}{1 \text{K}} \times (933.6 \text{K} - 273.2 \text{K}) = \mathbf{660.4°C}$

$? \text{°F} = \left(660.4\text{°C} \times \frac{1.8\text{°F}}{1\text{°C}}\right) + 32\text{°F} = \mathbf{1221°F}$

For Ag: $? \text{°C} = \frac{1\text{°C}}{1 \text{K}} \times (1235.1 \text{K} - 273.2 \text{K}) = \mathbf{961.9°C}$

$? \text{°F} = \left(961.9\text{°C} \times \frac{1.8\text{°F}}{1\text{°C}}\right) + 32\text{°F} = \mathbf{1763°F}$

1-60. ***Refer to Section 1-12 and Examples 1-16 and 1-17.***

$? \,°C = \frac{1°C}{1.8°F} \times (102.0°F - 32.0°F) = \mathbf{38.9°C}$

$? \,K = 38.9\,°C + 273.2\,° = \mathbf{312.1\ K}$

1-62. ***Refer to Section 1-13, and Examples 1-18 and 1-19.***

amount of heat *gained* (J) = (mass of substance)(specific heat)(temp. change)
= 45.3 g x 0.895 J/g·°C x (62.5°C - 27.0°C)
= 1440 J (3 sig. figs.)

1-64. ***Refer to Section 1-13.***

(a) amount of heat *gained* (J) = (mass of substance)(specific heat)(temp. change)
= (69,700 g)(0.818 J/g·°C)(41.0°C - 25.0°C)
= $\mathbf{9.12 \times 10^5\ J}$

(b) Note that we will follow the convention of representing temperature (°C) as t and temperature (K) as T.

In any insulated system, the Law of Conservation of Energy states:
the amount of heat lost by Substance 1 = amount of heat gained by Substance 2

As will be discussed in later chapters, "heat lost" is a negative quantity and "heat gained" is a positive quantity. However, the "*amount* of heat lost" and the "*amount* of heat gained" quoted here call for *absolute* quantities without a sign associated with them. In other words, because we are using the words "lost" and "gained" the heat involved is positive and the differences in temperature are positive values as well in this exercise.

$$|\text{the amount of heat lost by Substance 1}| = |\text{amount of heat gained by Substance 2}|$$
$$|\text{(mass)(Sp. Ht.)(temp. change)}|_1 = |\text{(mass)(Sp. Ht.)(temp. change)}|_2$$

In this exercise,

$$|\text{(mass)(Sp. Ht.)(temp. change)}|_{\text{limestone}} = |\text{(mass)(Sp. Ht.)(temp. change)}|_{\text{air}}$$

Since any "change" is always defined as the final value minus the initial value, we have
(temp. change)$_{\text{limestone}}$ = (30.0°C - 41.0°C) and (temp. change)$_{\text{air}}$ = (t_{final} - 10.0°C)

for the limestone, $|30.0°C - 41.0°C| = |\text{negative value}| = (41.0°C - 30.0°C) = 11.0°C$
for the interior air, $|t_{\text{final}} - 10.0°C| = |\text{positive value}| = (t_{\text{final}} - 10.0°C)$

Before we start, we must first calculate the mass of air inside the house:

$$?\ \text{g air} = 2.83 \times 10^5 \text{ liters} \times \frac{1000 \text{ mL}}{1 \text{ L}} \times \frac{1.20 \times 10^{-5} \text{ g}}{1 \text{ mL}} = 3.40 \times 10^5 \text{ g}$$

$$69{,}700 \text{ g limestone} \times 0.818 \text{ J/g·°C} \times (41.0°C - 30.0°C) = 3.40 \times 10^5 \text{ g air} \times 1.004 \text{ J/g·°C} \times (t_{\text{final}} - 10.0°C)$$
$$6.27 \times 10^5 \text{ J} = (3.41 \times 10^5 \times t_{\text{final}}) \text{ J} - 3.41 \times 10^6 \text{ J}$$
$$4.04 \times 10^6 \text{ J} = (3.41 \times 10^5 \text{ J/°C}) \times t_{\text{final}}$$
$$t_{\text{final}} = \mathbf{11.8°C}$$

1-66. *Refer to Section 1-13 and Example 1-19.*

$$|\text{the amount of heat lost by Substance 1}| = |\text{amount of heat gained by Substance 2}|$$

$$|(\text{mass})(\text{Sp. Ht.})(\text{temp. change})|_{\text{metal}} = |(\text{mass})(\text{Sp. Ht.})(\text{temp. change})|_{\text{water}}$$

$$50.0\text{ g} \times (\text{Sp. Ht.}) \times (75.0°\text{C} - 18.3°\text{C}) = 100.\text{ g} \times 4.18\text{ J/g°C} \times (18.3°\text{C} - 15.0°\text{C})$$

(Sp. Ht.) x 2835 (remember: it has only 3 sig. figs.*) = 1379 (only 2 sig. figs.)

Solving, Sp. Ht. of the metal = **0.49 J/g°C** (2 significant figures set by the temperature change of the water)

* Note: it is better to carry all the numbers in your calculator and do your rounding to the correct number of significant figures at the end.

1-68. *Refer to Sections 1-9 and 1-10.*

(a) $? \text{ tons ore} = 5.79 \text{ tons hematite} \times \frac{100 \text{ tons ore}}{9.24 \text{ tons hematite}} = \mathbf{62.7 \text{ tons ore}}$

(b) $? \text{ kg ore} = 6.40 \text{ kg hematite} \times \frac{100 \text{ kg ore}}{9.24 \text{ kg hematite}} = \mathbf{69.3 \text{ kg ore}}$

1-70. *Refer to Appendix A, Section 1-9 and the conversion factors from Table 1-8.*

$? \text{ m} = 23.5 \text{ ft} \times \frac{12 \text{ in.}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{1 \text{ m}}{100 \text{ cm}} = \mathbf{7.16 \text{ m}}$

1-72. *Refer to Section 1-9 and Table 1-8.*

$? \text{ lethal dose} = 165 \text{ lb body wt} \times \frac{453.6 \text{ g body wt}}{1 \text{ lb body wt}} \times \frac{1 \text{ kg body wt}}{1000 \text{ g body wt}} \times \frac{1.5 \text{ mg drug}}{1 \text{ kg body wt}} = \mathbf{110 \text{ mg drug}}$ (2 sig. figs.)

1-74. *Refer to Sections 1-10 and 1-11.*

Plan: $\text{g ammonia} \overset{(1)}{\Rightarrow} \text{g solution} \overset{(2)}{\Rightarrow} \text{mL solution}$

Using 2 unit factors,
(1) Convert mass of ammonia to mass of solution using the definition of % by mass, then
(2) Convert mass of solution to volume (in mL) of solution using density

$? \text{ L solution} = 25.8 \text{ g ammonia} \times \frac{100 \text{ g soln}}{5 \text{ g ammonia}} \times \frac{1 \text{ mL soln}}{1.006 \text{ g soln}} = \mathbf{500 \text{ mL}}$ (1 significant figure due to 5% ammonia)

1-76. *Refer to Sections 1-3 and 1-11, Example 1-2, and Figure 1-7.*

(a) Box (i) represents the very ordered, dense solid state.

(b) Box (iii) represents the less ordered, slightly less dense liquid state.

(c) Box (ii) represents the disordered, much less dense gaseous state.

(d) The physical states rank from least dense to most dense: gaseous state << liquid state < solid state

1-78. *Refer to Sections 1-4 and 1-5.*

Physical properties: zinc metal is a gray and shiny solid
zinc metal piece can be cut with scissors
copper chloride solution is blue in color
the new product is brown and granular

Physical changes: the zinc pieces reduced in size when cut with scissors
the zinc pieces reduced in size during the reaction
the solution became colorless and became warmer

Chemical changes: some of the zinc disappeared. It must have reacted, because zinc metal is not soluble in water
a new brown granular product formed
the reaction is exothermic and heat was released, making the flask warm to the touch

1-80. *Refer to Sections 1-4 and 1-5, and Exercise 1-79.*

Water is more dense than ice at 0°C because a cube of ice (less dense) will float in a glass of water (more dense). The first drawing shows liquid water molecules that are disorganized and slightly closer together, whereas the second drawing depicts the water molecules in a very rigid, ordered structure. When a sample has more mass per unit volume, it is more dense, so liquid water is more dense than solid water because its molecules are closer together.

1-82. *Refer to your life story.*

Chemical vocabulary and understanding can come from many experiences, besides the classroom. Perhaps you visited a science museum, or had a chemistry "magic show" come to your school. You may have been given a chemistry set as a present. There are many science-related shows on television and the internet has many, many links to science pages. Use your own life experiences to answer this question.

1-84. *Refer to Appendix A, Table 1-8 for conversion factors, and Example 1-4.*

Each cesium atom has a diameter = 2 x 2.65 Å = 5.30 Å

$$? \text{ Cs atoms} = 1.00 \text{ inch} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{1 \text{ m}}{100 \text{ cm}} \times \frac{1 \text{ Å}}{10^{-10} \text{ m}} \times \frac{1 \text{ atom}}{5.30 \text{ Å}} = \mathbf{4.79 \times 10^7 \text{ atoms}}$$

1-86. *Refer to Section 1-5 and your common sense.*

As a student writes out an End-of-Chapter Exercise, the direct chemical changes that occur include

(1) reactions (including irreversible adsorption) of the ink in the pen with the paper,
(2) the body's biochemical reactions,
(3) the creation of new neural pathways in the student's brain due to the new information she/he is learning.

More indirect chemical changes include the burning of coal or natural gas to provide the power for electricity, heat and light. If the student is doing a problem outside on a beautiful day, chemical changes might involve photosynthesis occurring in the plants around her/him providing oxygen for the student to breathe and the fusion reactions in the sun which provide heat and light, etc.

The complete answer is limited only by the student's imagination and understanding of the meaning of chemical changes. So, definitely yes, the answer involves knowledge not covered in Chapter 1.

1-88. ***Refer to Section 1-12 and Example 1-16.***

? °C of iron = $\frac{1°C}{1.8°F}$ x (65°F - 32°F) = **18°C**

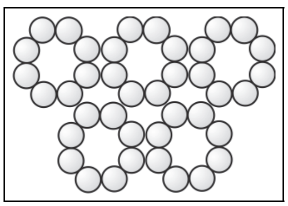
Therefore, the **water sample** at 65°C has a higher temperature than the iron sample at only 18°C.

1-90. ***Refer to Section 1-2.***

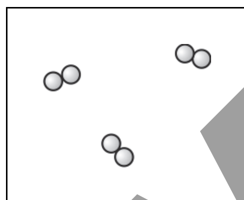
From left to right: NO, NO_2, N_2O, N_2O_3, N_2O_4 and N_2O_5.

1-92. ***Refer to Section 1-2, Figures 1-3 and 1-4, and Example 1-1.***

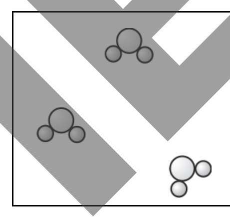
At room temperature, sulfur (rhombic) is a solid with formula, S_8, oxygen is a diatomic gas, O_2 and sulfur dioxide is a gas, SO_2.



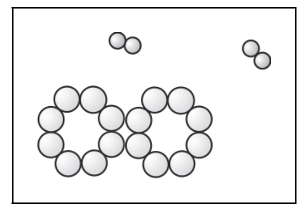
Sulfur, $S_8(s)$



Oxygen, $O_2(g)$



Sulfur dioxide, $SO_2(g)$



Mixture of S_8 and O_2

One similarity between S_8 and O_2 is that they are both elements composed of molecules. However, S_8 is a solid, with the molecular units arranged close together in a systematic way and O_2 is a gas, with its diatomic molecules relatively far apart.

The compound, SO_2, and the sample of S_8 mixed with O_2 both contain the elements, sulfur and oxygen, but SO_2 sample contains S and O in the definite ratio of 1:2 in each molecule and the individual gaseous SO_2 molecules are far apart. The mixture of S_8 and O_2 contains solid sulfur and molecular oxygen and the ratio of S to O can be variable. The mixture is heterogeneous, because $S_8(s)$ and $O_2(g)$ are present in different phases.

1-94. ***Refer to Section 1-11 and Appendix A.***

The calculation only involves multiplying and dividing. The number of significant figures in the answer is then set by the value with the least number of significant figures. Since density (=8.92 g/mL) has only 3 significant figures, the answer can only have 3 significant figures, which includes the first doubtful digit. The answer is V = 475 cm^3 and "5" is the first doubtful digit.

1-96. ***Refer to Section 1-9 and Appendix A.***

Many calculations in chemistry can be done in different ways. Consider the conversion of 3475 cm to miles.

(1) ? miles = 3475 cm x $\frac{1 \text{ in.}}{2.54 \text{ cm}}$ x $\frac{1 \text{ ft}}{12 \text{ in.}}$ x $\frac{1 \text{ mile}}{5280 \text{ ft}}$ = 0.021592649 miles or 0.02159 miles

Note: The following conversions are exact: 1 in. = 2.54 cm, 1 ft = 12 in., 1 mile = 5280 ft, so 2.54, 12, and 5280 have infinite numbers of significant figures. The number of significant figures in the answer is then set by the data: 4.

(2) ? miles = 3475 cm x $\frac{1\text{ m}}{100\text{ cm}}$ x $\frac{1\text{ km}}{1000\text{ m}}$ x $\frac{1\text{ mile}}{1.609\text{ km}}$ = 0.021597265 miles or 0.02160 miles

Note: Exact conversions: 1 m = 100 cm, 1 km = 1000 m. Inexact conversion: 1 mile = 1.609 km to 4 significant figures. The number of significant figures in the answer is set by the data (4 sig. figs.) but the answer has extra source of error since the conversion from kilometers to miles is only good to 4 sig. figs.

Method (1) uses all exact conversions and will give a more accurate answer than Method (2). If you really wanted to use Method (2), be sure that the inexact conversion contains more significant figures than your data. For example, if you used 1 mile = 1.6093 km, your answer would have been 0.021593239, and to 4 significant figures, both methods would have given essentially the same answer, differing only in the doubtful digit.

1-98. ***Refer to Sections 1-12 and 1-13, and the Key Terms for Chapter 1.***

Students often get the terms, heat, specific heat and temperature confused. Here are the formal definitions:

Heat: A form of energy that flows between two samples of matter because of their difference in temperature, measured in joules (J).

Specific heat: The amount of heat required to raise the temperature of one gram of a substance one degree Celsius. Its units are J/g·^{o}C.

Temperature: A measure of the intensity of heat, that is, the hotness or coldness of a sample or object. Temperature also refers to molecular motion. The warmer a substance is, the more its molecules are moving. Scientists usually work in ^{o}C or K.

If two samples of the same element are at different temperatures, their atoms have different kinetic energies and are moving at different average speeds. If the two samples touch, energy (heat) will transfer from the hotter to the colder element until their temperatures are the same and the average speed of their respective molecules are the same.

Different substances require different amounts of heat to change their temperatures. Specific heat is the constant that gives that information. It has units of J/g·^{o}C and is the amount of heat required (in joules) to heat up 1 gram of a substance by 1^{o}C.

As a final note, consider a 5.0 gram block of iron and a 15 gram block of iron, both at 25^{o}C. They are both at the same temperature, so if they came into contact, neither would change temperature. However, the 15 g iron block contains three times more heat than the 5.0 gram block. In other words, three times more heat is required to change the temperature of the 15 gram block of iron to 26^{o}C, as the 5.0 gram block of iron.