1 The Foundations of Chemistry

1-1 (a) Biochemistry is the study of the chemistry of living things.
(b) Analytical chemistry studies the quantitative and qualitative composition analysis of substances.
(c) Geochemistry is the study of the properties and reactions of the substances that compose earth's crust.
(d) Nuclear chemistry is the study of the properties and reactions of atomic nuclei.
(e) Inorganic chemistry is the study of compounds of elements other than carbon; however, simple carbon compounds are also included, such as CO, CO_2, carbonates, and bicarbonates.

1-3 (a) Matter is anything that has mass and occupies space. An example of matter is your textbook.
(b) Kinetic energy is the energy of a moving object or the energy of an object due to its motion. A bowling ball has kinetic energy as it is rolling down the lane.
(c) Mass is a measure of the amount of matter in an object. The mass of a penny (a copper coin) is about 1 gram.
(d) An exothermic process is a process that releases heat energy. The combustion of gasoline is an exothermic process that is used in automobile engines.
(e) An intensive property is a property that is independent of the amount of material present. Density is an intensive property.

1-5 *Law of Conservation of Matter and Energy*: The combined amount of matter and energy available in the universe is fixed. This law recognizes that the energy released in a nuclear reaction comes from the conversion of matter into energy. The *Law of Conservation of Matter* and *Law of Conservation of Energy* refer to chemical (not nuclear) reactions and physical changes. In chemical reactions and physical changes, the quantity of mater has no detectable change and energy is neither created nor destroyed; energy is only converted from one form to another.

1-7 (a) Since energy can be converted from one type to another, a broad definition of exothermic is that the reaction releases energy. Since light is a form of energy, the production of light from a fluorescent light is a release of energy.
(b) In a similar manner, the production of light by a glow-in-the-dark object also releases light a form of energy.

1-9 (a) Exothermic. The gasoline gives off heat and light during combustion or burning.
(b) Exothermic. The ice cream is changing from a liquid to a solid. Heat must be lost for the particles to slow down and to freeze. This is the opposite of melting.
(c) Endothermic. The chocolate absorbs heat as it melts or changes from a solid to a liquid.
(d) Exothermic. As the temperature of the water drops, the heat energy is leaving the water and moving into the surroundings.
(e) Exothermic. Water vapor gives off heat as it condenses. The particles must cool to change from a gas to a liquid

(f) Exothermic. The match gives off heat as it burns. This heat can be used to light the wick of a candle.

1-11 (a) *Law of Conservation of Matter*: There is no detectable change in the quantity of matter during an ordinary chemical reaction or during a physical change. Examples—(i) when magnesium metal burns in oxygen, the mass of the product (magnesium oxide) is equal to the sum of the masses of the magnesium and oxygen that combine; (ii) when ice melts, its mass does not change.

(b) *Law of Conservation of Energy*: Energy cannot be created or destroyed in a chemical reaction or in a physical change; it can only be converted from one form to another. Example—in a hydroelectric plant, the mechanical (kinetic) energy of the falling water is converted into electrical energy; some of the energy is converted into heat.

(c) *Law of Conservation of Matter and Energy*: The combined amount of matter and energy available in the universe is fixed. Example—the energy released in a nuclear reaction comes from the conversion of matter into energy.

1-13 An incandescent light bulb converts electrical energy into light energy. A considerable portion of the electrical energy used is converted into heat energy. The Law of Conservation of Energy is observed since the sum of the heat energy and light energy produced is equal to the electrical energy consumed.

1-15 A *homogeneous mixture* has uniform composition and properties throughout. Among the examples given in this exercise, carbon dioxide (f) is the only pure substance. All samples of carbon dioxide would always contain the same ratio of carbon and oxygen. Examples (a), and (e) are homogeneous mixtures; examples (b), (c), (d), and (g) are heterogeneous mixtures. The heterogeneous mixtures have large particles that are suspended (mud, noodles, onion), floating (ice), or that are at the bottom of the container (chocolate chips, chunks of chicken); therefore, they are not homogeneous mixtures.

1-17 (a) *A gaseous element* is shown in box (i). The substance contains only one element because only blue spheres are shown, even though the element is diatomic. The substance is a gas because the particles have the maximum separation.

(b) *A gaseous compound* is shown in box (v). The substance is a compound because each particle contains two elements (two blue atoms and one red atom bonded together). The substance is a gas because the particles have the maximum separation.

(c) *A homogeneous gaseous mixture* is shown in box (iv). A mixture is shown because there are two different types of particles (diatomic blue and a compound made of two blue and one red atom). The substance is a gas because the particles have the maximum separation.

(d) *A liquid solution* is shown in box (vi). A solution is a homogeneous liquid mixture. A mixture is shown because there are two different types of particles (a compound made of one red and two white atoms, with a second compound made of one red, one blue, and four white atoms). The substance is a liquid because the particles are much closer than in a gas, but the particles are not as close as a solid or in a regular repeating pattern as a solid.

(e) A *solid* is shown in box (ii). A solid is shown because the particles are shown very close together and are in a regular repeating pattern. A crystalline solid is depicted.

(f) A *pure liquid* is shown in box (iii). The substance is a liquid because the particles are all the same (maroon), are much closer than in a gas, but the particles are not as close as a solid or in a regular repeating pattern as a solid. The liquid happens to be diatomic. The liquid is pure because there is only one type of particle.

1-19 (a) Salt and water will form a homogeneous mixture, so to separate the salt from the water, you would need to evaporate or boil away the water to leave the salt behind.

(b) Iron filings and lead can be separated be using a magnet. Iron is attracted to a magnet, while lead is not.

(c) Elemental sulfur can be separated from sugar by using solubility properties. Sugar is soluble in water, while sulfur is not. Adding water to the mixture and pouring off the solution, sulfur will be left.

1-21 (a) *Chemical properties* are exhibited as matter undergoes changes in composition, whereas *physical properties* can be observed in the absence of any such change in composition.
Examples of chemical properties—(i) magnesium can combine with oxygen; (ii) gasoline is flammable.
Examples of physical properties—(i) water is a colorless liquid at room temperature; (ii) oxygen is a gas at room temperature and ordinary pressures; (iii) the melting point of bromine is –7.1˚C.

(b) *Intensive properties* are those properties that are independent of the amount of material examined, while *extensive properties* depend on the amount of material examined.
Examples of intensive properties—(i) magnesium can combine with oxygen; (ii) the melting point of bromine is –7.1˚C.
Examples of extensive properties—(i) the mass of a sample; (ii) the volume of a sample at specified conditions.

(c) *Chemical change*s occur when one or more substances react resulting in the formation of one or more new substances. *Physical changes* most often involve changes in physical state brought about by the absorption or release of energy
Example of chemical change—(i) alcohol reacting (burning) in oxygen to form carbon dioxide and water.
Examples of physical change—(i) ice melting to water with the absorption of heat; (ii) steam condensing to liquid water with the release of heat.

(d) *Mass* is a measure of the amount of matter in an object, while *weight* is a measure of gravitational attraction of the earth for an object.
An object having a mass of 454 g has a weight of one pound on Earth and the same object having a mass of 454 g would have zero weight in a zero gravitational field.

1-23 (a) Chemical process. Iron is combining with oxygen in the presence of water to form a new substance (rust).
(b) Physical process. Water as a solid (ice) is changing to liquid water. Melting does not change the composition.
(c) Chemical process. The wood is changed by the combustion or burning into ash, which is a new substance with none of the properties of the wood.
(d) Chemical process. The components of the potato are broken down into substances that can be absorbed by the digestive tract.
(e) Physical process. Dissolving sugar in water does not change the composition. If the water in the solution were allowed to evaporate, the sugar would be left behind.

1-25 (a) Kinetic energy (b) Potential energy
(c) Potential energy (d) Kinetic energy
(e) Kinetic energy (f) Potential energy

1-27 Both physical and chemical changes have taken place. The outer edge of the sugar cube melted (a physical change), then the sugar began to burn or oxidize (a chemical change). The heated portion has a different color and odor. The brown portion contains carbon left as the sugar decomposes.

1-29 (a) 6.50×10^2 (b) 6.30×10^{-2} (c) 8.60×10^3
(d) 8.600×10^3 (e) 1.6×10^4 (f) 1.0010×10^{-1}

1-31 (a) Exact (the result of counting)
(b) Exact (the result of counting)
(c) Exact (counted to the nearest penny)
(d) Not exact (obtained by measurement)
(e) Not exact (obtained by measurement)
(f) Exact (the result of counting)

1-33 Circumference = πd = (3.141593)(7.41 cm) = $\boxed{23.3 \text{ cm}}$

1-35 (a) 10^6 (b) 10^{-3} (c) 10^{-2} (d) 10^{-1} (e) 10^3 (f) 10^{-9}

1-37 5.31 cm = 5.31×10^{-2} m, 53.1 mm, 5.31×10^{-5} km, and 5.31×10^4 micrometers

1-39 $\underline{?}\ \$ = 14 \text{ gal} \times \frac{4 \text{ qt}}{1 \text{ gal}} \times \frac{1 \text{ L}}{1.056 \text{ qt}} \times \frac{\$0.861}{1L} = \boxed{\$45.66}$

1-41 $\underline{?} \text{ cm} = 8.25 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 20.955 \text{ cm}$ $\underline{?} \text{ cm} = 6.25 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 15.875 \text{ cm}$

21.0 cm x 15.9 cm = screen size

1-43 $\underline{?} \text{ g} = 10.25\text{g} + 5.5654\text{g} \times 105.4\text{g} = \boxed{121.2 \text{ g}}$

1-45 $\underline{?} \text{ g} = \frac{8.92 \text{ g}}{\text{cm}^3} \times 24.4 \text{ cm} \times 11.4 \text{ cm} \times 7.9 \text{ cm} = 19601\text{g} = \boxed{2.0 \times 10^4 \text{g}}$

1-47 $\underline{?} \text{ g} = 3.00 \text{ L} \times \frac{1000 \text{ cm}^3}{1 \text{ L}} \times \frac{1.0056 \text{ g}}{1 \text{ cm}^3} = \boxed{3.02 \times 10^3 g}$ (if three L has 3 sig. figs.)

1-49 (a) mass of water = 92.44 g – 78.91 g = 13.53 g water

volume of water $= 13.53 \text{ g} \times \frac{1 \text{ cm}^3}{1.0000 \text{ g}} = \boxed{13.53 \text{ cm}^3}$

(b) mass of unknown liquid = 88.42 g – 78.91 g = 9.51 g

density of unknown liquid $= \frac{\text{M}}{\text{V}} = \frac{9.51g}{13.53cm^3} = \boxed{0.703 \text{ g/cm}^3}$

1-51 $\underline{?} \text{ g} = 250.0 \text{ mL} \times \frac{1.049 \text{ g soln.}}{\text{mL}} \times \frac{40.0\ g \text{ acetic acid}}{100 \text{ g soln.}} = 104.9 \Rightarrow \boxed{105 \text{ g acetic acid}}$

1-53 (a) $\underline{?} \text{ K} = 245°\text{C} + 273.15° = \boxed{518 \text{ K}}$

(b) $\underline{?}\ °C = 25.2\ K - 273.15°C = -247.95°C = \boxed{-248.0°C}$

(c) $\underline{?}\ °F = \left(-42°C \times \frac{1.8°F}{1.0°C}\right) + 32°F = -43.6 = \boxed{-44°F}$

(d) First convert °F to °C, then °C to K.

$\underline{?}\ °C = \frac{1.0°C}{1.8°F} \times (110.0°F - 32°F) = 43.3°C$ with only 2 sig figs

$\underline{?}\ K = (43°C + 273.15°) = \boxed{316K}$

1-55 (a) $\underline{?}\ °F = \left(20°C \times \frac{1.8°F}{1.0°C}\right) + 32°F = 68°F$ so $\boxed{20°C}$ or 68°F is higher than 20°F

(b) $\underline{?}\ °F = \left(100°C \times \frac{1.8°F}{1.0°C}\right) + 32°F = 212°F$ so $\boxed{100°C}$ or 212°F is higher than 180°F

(c) $\underline{?}\ °C = \frac{1.0°C}{1.8°F} \times (100°F - 32°F) = 23.6°C$ so $\boxed{60°C}$ is higher than 100°F or 23.6°C

(d) $\underline{?}\ °F = \left(-12°C \times \frac{1.8°F}{1.0°C}\right) + 32°F = 10.4°F$ so $\boxed{20°F}$ is higher than –12°C or 10.4°F

1-57 He: $\underline{?}\ °C = 4.2\ K - 273.15°C = \boxed{-269.0°C}$

$\underline{?}\ °F = \left(\frac{1.8°F}{1.0°C} \times (-269.0°C)\right) + 32°F = \boxed{452.2°F}$

N_2: $\underline{?}\ °C = 77.4\ K - 273.15°C = \boxed{-195.8°C}$

$\underline{?}\ °F = \left(\frac{1.8°F}{1.0°C} \times (-195.8°C)\right) + 32°F = \boxed{-320.4°F}$

1-59 If °F = 2x and °C = x, Then 2x = 1.8 x + 32   0.2x = 32   x = 160

to check to see what °F are if °C = 160

$\underline{?}\ °F = \left(\frac{1.8°F}{1.0°C} \times 160°C\right) + 32°F = \boxed{320°F}$

1-61 Temperature change = 32.0°C – 10.0°C = 22.0°C

$\underline{?}$ J = mass of substance x specific heat x temperature change

$= 78.2\ g \times \frac{4.184\ J}{g \bullet °C} \times 22.0°C = \boxed{7.20 \times 10^3 J}$

1-63 $\underline{?}$ J = mass of substance x specific heat x temperature change

$= 15.5\ g \times \frac{4.184 J}{g \bullet °C} \times (38.2°C - 90.0°C) = \boxed{-3.36 \times 10^3 J \text{ or } 3.36 \times 10^3 J \text{ must be removed}}$

1-65 $\underline{?}\text{ g } H_2O = 245\text{ mL} \times \frac{0.997\text{ g}}{1\text{ mL}} = 244\text{ g } H_2O$

$\underline{?}\text{ J} = 244\text{ g} \times \frac{4.184\text{J}}{\text{g}\bullet°\text{C}} \times (85.°\text{C} - 25.°\text{C}) = 6.1 \times 10^4\text{ J}$

$\underline{?}\text{ kJ/min} = \frac{6.1 \times 10^4\text{ J}}{2.00\text{ min}} \times \frac{1\text{ kJ}}{1000\text{ J}} = 30.5 = \boxed{31\text{ kJ/min}}$

1-67 (a) $\underline{?}$ g calcium carbonate = 75.45 g sample x $\frac{25.8\text{g calcium carbonate}}{100\text{g sample}} = \boxed{19.5\text{ g calcium carbonate}}$

(b) $\underline{?}$ g sample = 18.8 g calcium carbonate x $\frac{100\text{g sample}}{25.8\text{g calcium carbonate}} = \boxed{72.9\text{ g sample}}$

1-69 Radius of earth's orbit (m) = $1.5 \times 10^8\text{ km} \times \frac{10^3\text{ m}}{1\text{ km}} = 1.5 \times 10^{11}\text{ m}$

Radius of hydrogen atom (m) = $0.37\text{ Å} \times \frac{10^{-10}\text{ m}}{1\text{ Å}} = 3.7 \times 10^{-11}\text{ m}$

Ratio = $\frac{1.5 \times 10^{11}\text{m}}{3.7 \times 10^{-11}\text{m}} = \boxed{4.1 \times 10^{21}}$

1-71 $\frac{?\text{ km}}{h} = \frac{65\text{ mi}}{\text{h}} \times \frac{1.609\text{ km}}{\text{mi}} = \left(\frac{104.6\text{ km}}{\text{h}}\right) = \boxed{1.0 \times 10^2\text{ km/h}}$

1-73 (a) $\frac{1.00\text{ mi}}{4.90\text{ min}} \times \frac{60\text{ min}}{\text{h}} \times \frac{1.609\text{ km}}{\text{mi}} = \boxed{19.7\text{ km/h}}$

(b) $\frac{19.7\text{ km}}{\text{h}} \times \frac{1\text{ h}}{60\text{ min}} \times \frac{1\text{ min}}{60\text{ s}} \times \frac{1000\text{ m}}{1\text{ km}} \times \frac{100\text{ cm}}{1\text{ m}} = \boxed{547\text{ cm/s}}$

(c) $1500.\text{ m} \times \frac{1\text{ h}}{19.7\text{ km}} \times \frac{1\text{ km}}{1000\text{ m}} \times \frac{60\text{ min}}{1\text{ h}} \times \frac{60\text{ s}}{1\text{ min}} = 274\text{ s} \Rightarrow \boxed{4\text{ min }34\text{s}}$

1-75 If you wanted the pot or pan to heat up quickly, you would select material that has a small specific heat value. If you wanted the pot or pan to retain its temperature once it is hot, then you would select material that would have a higher specific heat. Most individuals desire some of both of these traits but feel that the first is the more important.

1-77 The density of newly-minted penny, g/cm^3 :
= 0.027)(8.72 g/cm^3) + (1.000 – 0.027)(7.14 g/cm^3)
= 0.24 g/cm^3 + 6.95 g/cm^3 = 7.18 g/cm^3

1-79 We know that water must be more dense, because ice floats in water.

1-81 The correct answer is (a). The particles would be the same size but closer together at the lower temperature.

1-83 (a) Let x = the reading on the Celsius thermometer = the reading on the Fahrenheit thermometer

$$x\,°F = \left(x°C \times \frac{1.8°F}{1.0°C}\right) + 32°F \text{ or, without units,}$$

$$x = \frac{1.8}{1.0}\,x + 32 \;;\quad 0.8x = -32 \;;\; x = \boxed{-40°C}\,,\, x = \boxed{-40°F}$$

(b) $$2x = \frac{1.8}{1.0}\,x + 32 \;;\; 2x = 1.8x + 32 \;;\; 0.2x = 32; \quad x = \boxed{160°C}\,,\; 2x = \boxed{320°F}$$

(c) $$-x = \frac{1.8}{1.0}\,x + 32 \;;\; 2.8x = -32; \; x = \boxed{-11.4°C}\,,\; -x = \boxed{+11.4°F}$$

1-85 The balloons filled with substances that are lighter than air will float. Assuming that the balloons are all the same volume, the He and Ne balloons should float, while the Ar and Kr balloons will sink.

1-87 Students know many chemical terms before they begin to read this textbook. A few of the terms that they are likely to know are: compound, distillation, and chemical reaction.

1-89 (a) A *gas* is shown in boxes (iii), (iv), (vii), and (ix). The particles are in the gas phase because the particles have the maximum separation and are in a random arrangement.

(b) A *liquid* is shown in boxes (v -the blue particles in the top right of the diagram) and (viii). These are liquids because the particles are much closer than in a gas, but the particles are not as close as a solid or in a regular repeating pattern as a solid.

(c) A *solid* is shown in boxes (i), (ii), (v- the brown particles in the bottom left of the diagram), and (vi). A solid is shown because the particles shown very close together and are in a regular repeating pattern.

(d) An *element* is shown in boxes (i), (iv), and (vi). The particles are all the same color (blue), even though the blue atoms in box (vi) are shown as diatomic particles.

(e) A *compound* is shown in box (iii). The compound depicted here is composed of one blue atom and one brown atom, since one blue is attached to one brown throughout. The arrangement of the compound shows that it happens to be in the gaseous state.

(f) A *mixture* is shown in boxes (ii), (vii), and (ix). Mixtures contain two or more different types of particles. Boxes (ii) and (vii) contain both blue and brown particles. Box (ix) contains diatomic blue particles and single brown particles. Boxes (v) and (viii) show two types of particles, but these particles are not yet mixed.

(g) A *pure substance* is shown in boxes (i), (iii), (iv), and (vi). A pure substance contains particles that are identical.

1-91 Chlorine is an element. The atom of chlorine could be represented by a single sphere or 1 Cl. The diatomic molecule of chlorine would be represented by two adjoining spheres, to depict 2 atoms of chlorine in the molecule. These two differ in that the molecule has 2 atoms.

It is impossible to draw an atom of methane. Methane is a compound. A single methane molecule is made up of 1 atom of carbon and 4 atoms of hydrogen. One could draw a carbon or hydrogen atom from the molecule, but there is no atom of methane; the molecule is the smallest unit.

Methane is a compound, while chlorine is an element.

1-93 The change from solid, black carbon and colorless oxygen gas to colorless carbon dioxide gas is a chemical change because a new substance is formed (carbon dioxide). The colors and states of matter of these substances are physical properties.

1-95 To say the length of an animal is 51 doesn't give any units, so the understanding is very limited. An animal of 51 meters is very different from an animal of 51 millimeters. One must use units that others understand. It is also a problem to give a measurement with a unit that has no understanding. To know the width of a room is 7.36 bleams is also no understanding, since there is no understanding of the unit 'bleams'.

1-97 $\underline{?}\ \text{K} = 5500°\text{C} + 273.15 = \boxed{5773\text{K}}$

$$\underline{?}\ °\text{F} = \left(5500°\text{C} \times \frac{1.8°\text{F}}{1.0°\text{C}}\right) + 32°\text{F} = \boxed{9932\ °\text{F or } 9.9 \times 10^{3}\ °\text{F in 2 significant digits}}$$

The astronomer is referring to degrees Fahrenheit, but is confused about the significant numbers if he/she is referring to 5500° C. The astronomer could be the larger number, which does give 10,000 degrees Fahrenheit.

$$\underline{?}\ °\text{F} = \left(6000°\text{C} \times \frac{1.8°\text{F}}{1.0°\text{C}}\right) + 32°\text{F} = \boxed{10{,}832\ °\text{F or } 1 \times 10^{4}\ °\text{F in 1 significant digit}}$$