the liquid is vaporizing rapidly (boiling). None of the changes described above involves the breaking of chemical bonds or the production of a new substance, so the changes are physical in nature, rather than chemical.

16. Electrolysis is the passage of an electrical current through a substance or solution to force a chemical reaction to occur. Electrolysis causes chemical changes to take place that would ordinarily not take place on their own. When an electrical current is passed through water, the current causes the water molecules to break down into their constituent elements (hydrogen gas and oxygen gas).

17. a. physical; this represents only a change in state
b. chemical; when the alcohol in the brandy burns, it is converted into other substances (carbon dioxide and water)
c. chemical; the acid reacts with the glass and converts it into other substances
d. physical; this represents only a change in freezing point
e. chemical; since the antacid produces a new substance (carbon dioxide), this can only be a chemical change
f. chemical; the metal case of the battery is converted into other substances, which allows the battery to leak
g. chemical; the cellulose which make up the cotton fibers is converted into other substances, leaving a hole
h. physical; milk contains protein, and when the vinegar is added, the acidity of the vinegar causes a change in the protein's shape, making it insoluble in water.
i. physical; when you stop pulling on the rubber, it goes back to its original shape
j. physical; this represents just a change in state
k. physical; the acetone dissolves the lacquers in the nail polish, forming a solution (which is not a chemical process)

18. a. chemical; the scorch represents the oxidation of the material
b. physical; the gas in the tires decreases in volume with temperature
c. chemical; tarnish on silver is caused by reaction of the silver with sulfur or oxygen
d. chemical; the ethyl alcohol in the wine is oxidized to acetic acid

e. chemical; the oven cleaner contains sodium hydroxide which converts greases in the oven into soaps

f. chemical; an ordinary flashlight battery is constructed with a zinc casing which serves as one of the electrodes. As the battery discharges, the zinc is oxidized.

g. chemical; the acids attack the calcium phosphate matrix of the teeth

h. chemical; the charring represents the breakdown of the sugar

i. chemical; iron in the blood catalyzes the decomposition of the hydrogen peroxide into oxygen gas (and water)

j. physical; this is just a change in state; carbon dioxide, only in the gaseous state, is still present after the sublimation

k. chemical; chlorine is an oxidizing agent and can change the chemical nature of the dyes in the fabrics

19. chemical

20. element

21. compounds

22. Compounds

23. When we say that a compound always has the same composition, we mean that the molecules of that compound always contain the same type and number of atoms of its constituent elements.

24. Typically, the properties of a compound and the elements that constitute it are very different. Consider the properties of liquid water and the hydrogen and oxygen gases from which the water was prepared. Consider the properties of sodium chloride (table salt) and the sodium metal and chlorine gas from which it might have been prepared.

25. the same

26. a variable

27. A solution is a homogeneous mixture in which the solute is uniformly dispersed on a molecular level in the solvent. Simple mixtures are typically nonhomogeneous and the dispersion of the components may not be on the molecular level.
28. solutions: window cleaner, shampoo, rubbing alcohol
   mixtures: salad dressing, jelly beans, the change in my pocket

29. a. mixture (fat, water, sugar, protein, etc.)
   b. mixture (wood pulp, cotton, coloring agents, etc.)
   c. pure substance
   d. mixture
   e. mixture (iron, and perhaps nickel, carbon, chromium, etc.)

30. a. mixture
   b. mixture
   c. mixture
   d. pure substance

31. a. homogeneous (if nothing has been mixed in with the ice cream)
   b. heterogeneous
   c. homogeneous
   d. heterogeneous
   e. heterogeneous

32. a. homogeneous
   b. heterogeneous
   c. heterogeneous
   d. homogeneous
   e. the paper itself is basically homogeneous in appearance

33. Consider a salt solution (sodium chloride in water). Since water boils at a much lower temperature than sodium chloride, the water can be boiled off from the solution, collected, and subsequently condensed back into the liquid state. This separates the two chemical substances.

34. Consider a mixture of salt (sodium chloride) and sand. Salt is soluble in water, sand is not. The mixture is added to water and stirred to dissolve the salt, and is then filtered. The salt solution passes through the filter, the sand remains on the filter. The water can then be evaporated from the salt.
35. If water is added to the sample, and the sample is then heated to boiling, this should dissolve the benzoic acid but not the charcoal. The hot sample could then be filtered, which would remove the charcoal. The solution which passed through the filter could then be cooled, which should cause some of the benzoic acid to crystallize, or the solution could be heated carefully to boil off the water leaving benzoic acid behind.

36. The solution is heated to vaporize (boil) the water. The water vapor is then cooled so that it condenses back to the liquid state, and the liquid is collected. After all the water is vaporized from the original sample, pure sodium chloride will remain. The process consists of physical changes.

37. energy

38. the calorie

39. In a sample of ice, the water molecules are held relatively rigidly in fixed positions in the crystal. As heat is applied, the water molecules begin to vibrate back and forth, but still basically remain in their original positions until the melting point is reached. When the melting point is reached, the forces which held the molecules in position in the crystal are overcome, and the molecules begin to move around much more freely, although they are still held within the bulk volume of the liquid. As the liquid continues to be heated, the molecules move more quickly and more freely until the boiling point is reached. At the boiling point, the forces holding the molecules in the liquid are overcome, and the individual molecules are moving so quickly and freely that they are able to escape from the volume of the liquid. As the molecules of the vapor continue to be heated, they move more and more quickly.

40. As the steam is cooled from 150 °C to 100 °C, the molecules of vapor gradually slow down as they lose kinetic energy. At 100 °C, the steam condenses into liquid water, and the temperature remains at 100 °C until all the steam has condensed. As the liquid water cools, the molecules in the liquid move more and more slowly as they lose kinetic energy. At 0 °C, the liquid water freezes.

41. lower

42. temperature

43. \[ 526 \, \text{J} \times \frac{25.0 \, \text{g}}{7.40 \, \text{g}} = 1.78 \times 10^3 \, \text{J} \]

44. \[ 526 \, \text{J} \times \frac{55 \, ^\circ \text{C}}{17 \, ^\circ \text{C}} = 1.7 \times 10^3 \, \text{J} \]
45.  a. \(7845 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 3.282 \times 10^4 \text{ J} = 32.82 \text{ kJ}\)

b. \(4.55 \times 10^4 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 1.90 \times 10^5 \text{ J} = 190. \text{ kJ}\)

c. \(62.142 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 2.600 \times 10^2 \text{ kJ} = 2.600 \times 10^5 \text{ J}\)

d. \(43,024 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 1.800 \times 10^5 \text{ J} = 180.0 \text{ kJ}\)

46. Since \(1.000 \text{ cal} = 4.184 \text{ J}\), then \(1.000 \text{ kcal} = 4.184 \text{ kJ}\)

a. \(462.4 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 110.5 \text{ kcal}\)

b. \(18.28 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 4.369 \text{ kcal}\)

c. \(1.014 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 0.2424 \text{ kcal}\)

d. \(190.5 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 45.53 \text{ kcal}\)

47.  a. \(55,322 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 55.322 \text{ kcal}\)

b. \(972 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 0.972 \text{ kcal}\)

c. \(442,800 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 442.8 \text{ kcal}\)

d. \(5.26 \times 10^4 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 52.6 \text{ kcal}\)

48.  a. \(12.30 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} = 12,300 \text{ cal} (1.230 \times 10^4 \text{ cal})\)

b. \(290.4 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} = 290,400 \text{ cal} (2.904 \times 10^5 \text{ cal})\)

c. \(940,000 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} = 940,000,000 \text{ cal} (9.4 \times 10^8 \text{ cal})\)

d. \(4201 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} = 4,201,000 \text{ cal} (4.201 \times 10^6 \text{ cal})\)
49. a. \[ 243,000 \ J \times \frac{1 \ kJ}{1000 \ J} = 243 \ kJ \]
b. \[ 4.184 \ J \times \frac{1 \ kJ}{1000 \ J} = 4.184 \times 10^{-3} \ kJ \]
c. \[ 0.251 \ J \times \frac{1 \ kJ}{1000 \ J} = 2.51 \times 10^{-4} \ kJ \]
d. \[ 450.3 \ J \times \frac{1 \ kJ}{1000 \ J} = 0.4503 \ kJ \]

50. a. \[ 45.62 \ kcal \times \frac{4.184 \ kJ}{1 \ kcal} = 190.9 \ kJ \]
b. \[ 72.94 \ kJ \times \frac{1 \ kcal}{4.184 \ kJ} = 17.43 \ kcal \]
c. \[ 2.751 \ kJ \times \frac{1 \ kcal}{4.184 \ kJ} \times \frac{1000 \ cal}{1 \ kcal} = 657.5 \ cal \]
d. \[ 5.721 \ kcal \times \frac{4.184 \ kJ}{1 \ kcal} \times \frac{1000 \ J}{1 \ kJ} = 2.394 \times 10^4 \ J \]

51. Temperature increase = \( 75.0 - 22.3 = 52.7^\circ C \)

\[ 145 \ g \times 4.184 \ J/g \ ^\circ C \times 52.7^\circ C \times \frac{1 \ cal}{4.184 \ J} = 7641.5 \ cal = 7.64 \times 10^3 \ cal \]

52. Heat = mass \times \text{specific heat capacity} \times \text{temperature change}

Specific heat capacity = \( \frac{\text{Heat}}{\text{mass} \times \text{temperature change}} \)

\[ 72.4 \ kJ = 72,400 \ J \]

Specific heat capacity = \( \frac{72,400 \ J}{(952 \ g)(10.7^\circ C)} = 7.11 \ J/g \ ^\circ C \)

53. Table 3.2 gives the specific heat capacity of gold as 0.13 J/g \ ^\circ C

Temperature increase = \( 155 - 120 = 35^\circ C \)

\[ 25.0 \ g \times 0.13 \ J/g \ ^\circ C \times 35^\circ C = 113.75 \ J = 1.1 \times 10^2 \ J \]

54. Table 3.2 gives the specific heat of silver as 0.24 J/g \ ^\circ C

Temperature increase = \( 15.2 - 12.0 = 3.2^\circ C \)

\[ 1.25 \ kJ = 1250 \ J \]

1250 J = (mass of silver) \times 0.24 J/g \ ^\circ C \times 3.2^\circ C

Mass of silver = \( 1627 \ g = 1.6 \times 10^3 \ g \) silver
55. Table 3.2 gives the specific heat capacity of iron as 0.45 J/g °C.

50. joules is the heat that is applied to the sample of iron, and must equal the product of the mass of iron, the specific heat capacity of the iron, and the temperature change undergone by the iron (which is what we want).

\[
\text{Heat} = \text{mass} \times \text{specific heat capacity} \times \text{temperature change}
\]

\[
50. \ J = 10. \ g \times 0.45 \ J/g \ ^\circ C \times \Delta T
\]

\[
\Delta T = \frac{50. \ J}{10. \ g \times 0.45 \ J/g \ ^\circ C} = 11 \ ^\circ C
\]

56. Table 3.2 gives the specific heat capacity of iron as 0.45 J/g °C.

\[
\text{Heat} = \text{mass} \times \text{specific heat capacity} \times \text{temperature change}
\]

Temperature change = 75.5 - 40.1 = 35.4 °C

\[
\text{Heat} = 852.2 \ g \times 0.45 \ J/g \ ^\circ C \times 35.4 \ ^\circ C = 1.4 \times 10^4 \ J
\]

57. \[
0.24 \ \frac{J}{g \ ^\circ C} \times \frac{1 \ \text{cal}}{4.184 \ J} = 0.057 \ \frac{\text{cal}}{g \ ^\circ C}
\]

58. \[
0.13 \ \frac{J}{g \ ^\circ C} \times \frac{1 \ \text{cal}}{4.184 \ J} = 0.031 \ \frac{\text{cal}}{g \ ^\circ C}
\]

59. It is not really necessary to calculate the temperature changes experienced by the metals. For metal samples of equal mass, the metal with the smallest specific heat capacity will experience the largest temperature change when a given amount of heat is applied. Remember that the specific heat capacity represents the ability of the substance to absorb heat energy.

60. Specific heat capacities are given in Table 3.2

for gold, \[ 25.0 \ g \times 0.13 \ J/g \ ^\circ C \times 20. \ ^\circ C = 65 \ J \]

for mercury, \[ 25.0 \ g \times 0.14 \ J/g \ ^\circ C \times 20. \ ^\circ C = 70 \ J \]

for carbon, \[ 25.0 \ g \times 0.71 \ J/g \ ^\circ C \times 20. \ ^\circ C = 360 \ J \]

61. Heat = mass \times \text{specific heat capacity} \times \text{temperature change}

The temperature change is 92 - 10. = 82 °C

\[
540. \ J = 22.5 \ g \times s \times 82 \ ^\circ C
\]

\[
s = \frac{540. \ J}{22.5 \ g \times 82 \ ^\circ C} = 0.29 \ J/g \ ^\circ C
\]
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62. \[ 1251 \text{ J} = 35.2 \text{ g} \times (\text{specific heat capacity}) \times 25.0 \, ^\circ\text{C} \]
   
   specific heat capacity = 1.42 \text{ J/g} \cdot \text{^\circ\text{C}}

63. mixture, compound

64. Since \( X \) is a pure substance, the fact that two different solids form when electrical current is passed indicates that \( X \) must be a compound.

65. Chalk must be a compound, since it loses mass when heated, and appears to change into a substance with different physical properties (the hard chalk turns into a crumbly substance).

66. Since vaporized water is still the \textit{same substance} as solid water, no chemical reaction has occurred. Sublimation is a physical change.

67. a. \[ 4.52 \text{ cal} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 4.52 \times 10^{-3} \text{ kcal} \]
   
b. \[ 5.27 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} \times \frac{1000 \text{ J}}{1 \text{ kJ}} = 2.20 \times 10^4 \text{ J} \]
   
c. \[ 852,000 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 3.56 \times 10^3 \text{ kJ} \]
   
d. \[ 352.4 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 1474 \text{ kJ} \]
   
e. \[ 5.72 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} = 1.37 \times 10^3 \text{ cal} \]
   
f. \[ 4.52 \times 10^3 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 4.52 \text{ kJ} \]

68. 2.5 kg of water = 2,500 g

Temperature change = 55.0 - 18.5 = 36.5 \text{ ^\circ\text{C}}

\[ 2500 \text{ g} \times 4.184 \text{ J/g} \cdot \text{^\circ\text{C}} \times 36.5 \, ^\circ\text{C} = 3.8 \times 10^6 \text{ J} \]

69. For a given mass of substance, the substance with the \textit{smallest} specific heat capacity (gold, 0.13 J/g \cdot ^\circ\text{C}) will undergo the \textit{largest} increase in temperature. Conversely, the substance with the \textit{largest} specific heat capacity (water, 4.184 J/g \cdot ^\circ\text{C}) will undergo the \textit{smallest} increase in temperature.

70. No calculation is necessary: aluminum will lose more heat because it has the \textit{higher} specific heat capacity.
71. Heat evolved by hydrogen: \(5.0 \, \text{g} \times 120. \, \text{J/g} = 600 \, \text{J} (6.0 \times 10^2 \, \text{J})\)

Heat evolved by methane: \(10. \, \text{g} \times 50. \, \text{J/g} = 500 \, \text{J} (5.0 \times 10^2 \, \text{J})\)

Total heat evolved by burning: \(600 + 500 = 1100 \, \text{J} (1.1 \times 10^3 \, \text{K})\)

For any substance, \(Q = m \times s \times \Delta T\). Let \(T_f\) represent the final temperature reached by the water. Then \(\Delta T = (T_f - 25 \, ^\circ\text{C})\)

\(1.1 \times 10^3 \, \text{J} = 500. \, \text{g} \times 4.184 \, \text{J/g} \, ^\circ\text{C} \times (T_f - 25 \, ^\circ\text{C})\)

\((T_f - 25 \, ^\circ\text{C}) = 0.53 \, ^\circ\text{C}\)

\(T_f = 25 + 0.53 = 25.5 \, ^\circ\text{C} = 26 \, ^\circ\text{C}\)

To the limits of measurement indicated in the problem, the temperature would not effectively increase. A relatively large amount of water was used to absorb the heat evolved by the combustion, and water has a very large heat capacity (this is why water is so often used as a coolant).

72. For any substance, \(Q = m \times s \times \Delta T\). The quantity of heat gained by the water in this experiment must equal the total heat lost by the metals (i.e., the sum of the amount of heat lost by the iron and the amount of heat lost by the aluminum).

\[(m \times s \times \Delta T)_{\text{water}} = (m \times s \times \Delta T)_{\text{iron}} + (m \times s \times \Delta T)_{\text{aluminum}}\]

If \(T_f\) represents the final temperature reached by this system, then

\[m \times s \times (T_f - 22.5 \, ^\circ\text{C}) \]

\[= [m \times s \times (100 \, ^\circ\text{C} - T_f)]_{\text{iron}} + [m \times s \times (100 \, ^\circ\text{C} - T_f)]_{\text{aluminum}}\]

\[97.3 \, \text{g} \times 4.184 \, \text{J/g} \, ^\circ\text{C} \times (T_f - 22.5 \, ^\circ\text{C})\]

\[= [10.00 \, \text{g} \times 0.45 \, \text{J/g} \, ^\circ\text{C} \times (100 \, ^\circ\text{C} - T_f)]\]

\[+ [5.00 \, \text{g} \times 0.89 \, \text{J/g} \, ^\circ\text{C} \times (100 \, ^\circ\text{C} - T_f)]\]

\(407.1(T_f - 22.5) = 4.50(100 - T_f) + 4.45(100 - T_f) = 8.95(100 - T_f)\)

\(407.1T_f - 9160 = 895 - 8.95T_f\)

\(416.1T_f = 10,055\) which gives \(T_f = 24.2 \, ^\circ\text{C}\)

73. Let \(T_f\) represent the final temperature reached.

For the hot water, heat lost = \(50.0 \, \text{g} \times 4.184 \, \text{J/g} \, ^\circ\text{C} \times (100. - T_f)\)

For the cold water, heat gained = \(50.0 \, \text{g} \times 4.184 \, \text{J/g} \, ^\circ\text{C} \times (T_f - 25)\)

Heat lost by the hot water must equal heat gained by the cold water.
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\[ 50.0 \text{ g} \times 4.184 \text{ J/g °C} \times (100. - T_f) = 50.0 \text{ g} \times 4.184 \text{ J/g °C} \times (T_f - 25) \]
\[ 209.2 (100 - T_f) = 209.2 (T_f - 25) \]
\[ (100 - T_f) = (T_f - 25) \]
\[ 2T_f = 125 \text{ °C} \]
\[ T_f = 62.5 \text{ °C} = 63 \text{ °C} \]

74. \((m \times s \times \Delta T)_{\text{water}} = (m \times s \times \Delta T)_{\text{iron}}\)

Let \(T_f\) represent the final temperature reached by the system

\[ [75 \text{ g} \times 4.184 \text{ J/g °C} \times (T_f - 20.)] = [25.0 \text{ g} \times 0.45 \text{ J/g °C} \times (85 - T_f)] \]
\[ 314 (T_f - 20.) = 11.4 (85 - T_f) \]
\[ 314T_f - 6280 = 961 - 11.3T_f \]
\[ 325T_f = 7241 \]
\[ T_f = 22.3 \text{ °C} = 22 \text{ °C} \]

75. Liquids and gases both flow freely and take on the shape of their container. The molecules in liquids are relatively close together and interact with each other, whereas the molecules in gases are far apart from each other and do not interact with each other.

76. far apart

77. physical

78. chemical

79. physical

80. chemical

81. state

82. electrolysis

83. a. physical; milk contains protein, and when the vinegar is added, the acidity of the vinegar causes a change in the protein's shape, making it insoluble in water (see Chapter 21)

b. chemical; exposure to the oxygen of the air allows bacteria to grow which cause the chemical breakdown of components of the butter.
c. physical; salad dressing is a physical mixture of water soluble and insoluble components, which only combine temporarily when the dressing is shaken.

d. chemical; milk of magnesia is a base, which chemically reacts with and neutralizes the acid of the stomach.

e. chemical; steel consists mostly of iron, which chemically reacts with the oxygen of the atmosphere.

f. chemical; carbon monoxide combines chemically with the hemoglobin fraction of the blood, making it impossible for the hemoglobin to combine with oxygen.

g. chemical; cotton consists of the carbohydrate cellulose, which is broken down chemically by acids.

h. physical; sweat consists mostly of water, which consumes heat from the body in evaporating.

i. chemical; although the biochemical action of aspirin is not fully understood, the process is chemical in nature.

j. physical; oil molecules are not water soluble, and are repelled by the moisture in skin.

k. chemical; the fact that one substance is converted into two other substances demonstrates that this is a chemical process.

84. a. heterogeneous
   b. heterogeneous
   c. heterogeneous (unless you work hard to get all the lumps out!)
   d. although strictly heterogenous, it may appear homogeneous
   e. heterogeneous

85. a. heterogeneous
   b. homogeneous
   c. heterogeneous
   d. homogeneous (assuming there are no imperfections in the glass)
   e. heterogeneous

86. 9.0 J (It requires twice as much heat to warm twice as large a sample over the same temperature interval.)
87. Since it requires 103 J to heat the iron over a temperature change of 25° (from 25° to 50°C), it will require 206 J (twice as much heat) to heat the same sample of iron over a temperature change of 50° (from 25° to 75°).

88. a. \(44.21 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 185.0 \text{ J}\)

   b. \(162.4 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 679.5 \text{ J}\)

   c. \(3.721 \times 10^3 \text{ cal} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 1.557 \times 10^4 \text{ J}\)

   d. \(146.2 \text{ kcal} \times \frac{1000 \text{ cal}}{1 \text{ kcal}} \times \frac{4.184 \text{ J}}{1 \text{ cal}} = 6.117 \times 10^8 \text{ J}\)

89. a. \(52.18 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 12.47 \text{ kcal}\)

   b. \(4.298 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 1.027 \times 10^{-3} \text{ kcal}\)

   c. \(5.433 \times 10^3 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} \times \frac{1 \text{ kcal}}{1000 \text{ cal}} = 1.299 \text{ kcal}\)

   d. \(455.9 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 109.0 \text{ kcal}\)

90. a. \(5.442 \times 10^4 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 54.42 \text{ kJ}\)

   b. \(5.442 \times 10^4 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 1.301 \times 10^4 \text{ cal}\)

   c. \(352.6 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 1475 \text{ kJ}\)

   d. \(17.24 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 4.120 \text{ kcal}\)

91. Table 3.2 gives the specific heat capacity of gold as 0.13 J/g °C.

   Temperature increase = 75.0 - 20.0 = 55.0 °C

   Heat required = 25.0 g \times 0.13 \text{ J/g °C} \times 55.0 °C = 178.8 \text{ J} = 1.8 \times 10^2 \text{ J}

   (which is equivalent to 0.18 kJ, 0.043 kcal, 43 cal)
92. The specific heat capacity of water is 4.184 J/g °C.
Temperature increase = 39 - 25 = 14 °C.
75 g x 4.184 J/g °C x 14 °C = 4400 J (to 2 significant figures)

93. Table 3.2 gives the specific heat capacity of aluminum as 0.89 J/g °C.
Temperature increase = 85.2 - 22.1 = 63.1 °F.
63.1 °F x (100/180) = 35.1 °C. 37.5 lb = 1.70 x 10^4 g
Heat required = 1.70 x 10^4 g x 0.89 J/g °C x 35.1 °C = 5.3 x 10^5 J

94. For any substance, \( Q = m \times s \times \Delta T \). The basic calculation for each of the substances is the same (specific heat capacities are found in Table 3.2)
Heat required = 150. g x (specific heat capacity) x 11.2 °C

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat Capacity</th>
<th>Heat Required</th>
</tr>
</thead>
<tbody>
<tr>
<td>water (l)</td>
<td>4.184 J/g °C</td>
<td>7.03 x 10^3 J</td>
</tr>
<tr>
<td>water (s)</td>
<td>2.03 J/g °C</td>
<td>3.41 x 10^3 J</td>
</tr>
<tr>
<td>water (g)</td>
<td>2.0 J/g °C</td>
<td>3.4 x 10^3 J</td>
</tr>
<tr>
<td>aluminum</td>
<td>0.89 J/g °C</td>
<td>1.5 x 10^3 J</td>
</tr>
<tr>
<td>iron</td>
<td>0.45 J/g °C</td>
<td>7.6 x 10^2 J</td>
</tr>
<tr>
<td>mercury</td>
<td>0.14 J/g °C</td>
<td>2.4 x 10^2 J</td>
</tr>
<tr>
<td>carbon</td>
<td>0.71 J/g °C</td>
<td>1.2 x 10^3 J</td>
</tr>
<tr>
<td>silver</td>
<td>0.24 J/g °C</td>
<td>4.0 x 10^2 J</td>
</tr>
<tr>
<td>gold</td>
<td>0.13 J/g °C</td>
<td>2.2 x 10^2 J</td>
</tr>
</tbody>
</table>

95. Since, for any substance, \( Q = m \times s \times \Delta T \), we can solve this equation for the temperature change, \( \Delta T \)The results are tabulated:

<table>
<thead>
<tr>
<th>Substance</th>
<th>Specific Heat Capacity</th>
<th>Temperature Change</th>
</tr>
</thead>
<tbody>
<tr>
<td>water (l)</td>
<td>4.184 J/g °C</td>
<td>23.9 °C</td>
</tr>
<tr>
<td>water (s)</td>
<td>2.03 J/g °C</td>
<td>49.3 °C</td>
</tr>
<tr>
<td>water (g)</td>
<td>2.0 J/g °C</td>
<td>50. °C</td>
</tr>
<tr>
<td>aluminum</td>
<td>0.89 J/g °C</td>
<td>1.1 x 10^2 °C</td>
</tr>
<tr>
<td>iron</td>
<td>0.45 J/g °C</td>
<td>2.2 x 10^2 °C</td>
</tr>
<tr>
<td>mercury</td>
<td>0.14 J/g °C</td>
<td>7.1 x 10^2 °C</td>
</tr>
<tr>
<td>carbon</td>
<td>0.71 J/g °C</td>
<td>1.4 x 10^2 °C</td>
</tr>
<tr>
<td>silver</td>
<td>0.24 J/g °C</td>
<td>4.2 x 10^2 °C</td>
</tr>
<tr>
<td>gold</td>
<td>0.13 J/g °C</td>
<td>7.7 x 10^2 °C</td>
</tr>
</tbody>
</table>

96. Heat = mass x specific heat capacity x temperature change

The temperature change is 112.1 - 25.0 = 87.1 °C

1.351 kJ = 1351 J

\[ s = \frac{1351 J}{125 g \times 87.1 °C} = 0.124 J/g °C \]
Cumulative Review: Chapters 1, 2, and 3

1. Obviously, this answer depends on your own experiences in studying and learning about chemistry. We hope that by now you have at least gotten over any "fear" of chemistry you may have started out with. Perhaps you have begun to appreciate why one leading chemical manufacturer has as its corporate slogan "better living through chemistry".

2. By now, after having covered three chapters in this book, it is hoped that you have adopted an "active" approach to your study of chemistry. You may have discovered (perhaps through a disappointing grade on a quiz (though we hope not)), that you really have to get involved with chemistry. You can’t just sit and take notes, or just look over the solved examples in the textbook. You have to learn to solve problems. You have to learn how to interpret problems, and how to reduce them to the simple mathematical relationships you have studied. Whereas in some courses you might get by on just giving back on exams the facts or ideas presented in class, in chemistry you have to be able to extend and synthesize what has been discussed, and to apply the material to new situations. Don’t get discouraged if this is difficult at first: it’s difficult for everyone at first.

3. The steps of the scientific method, in brief, are: (1) make observations of the system and state the problem clearly; (2) formulate an explanation or hypothesis to try to explain your observations; and (3) perform one or more experiments to test the validity of your hypothesis. For the case of the liquid material, we first have to state the problem: we have a sample of clear liquid which may be either a pure compound or a mixture, and we want to determine which of these is correct. The liquid is completely homogeneous, so we can’t tell anything just by looking at it.

   If the unknown liquid is a mixture, we should be able to separate the components of the mixture by distillation (see Chapter 3). If the liquid is a mixture of two pure liquids, the liquids should boil at their characteristic temperatures during distillation. If the liquid is a solution of a solid material, then the liquid portion should boil off, leaving a solid residue behind. If the unknown liquid is a pure substance, rather than a mixture, it should boil away at a constant, uniform temperature. Let’s hypothesize that the liquid is a mixture.

   Now we perform the experiment: we begin heating the liquid in a distillation apparatus, monitoring the temperature with a thermometer. At 65°C, the liquid begins to boil, and a clear, sweet-smelling liquid begins to collect in the receiving container. The temperature remains at 65°C until approximately half the liquid has distilled, whereupon the temperature rises suddenly to 100°C. At this point, we change receiving flasks. The temperature remains at 100°C until the remainder of the liquid boils. We notice that this second fraction of liquid collected has no odor.

   Based on the observations that two separate fractions were collected, which had different boiling points, and that only one of the fractions had a noticeable odor, we can conclude that our hypothesis that the unknown liquid was a mixture is correct.
4. It is difficult sometimes for students (especially beginning students) to understand why certain subjects are required for a given college major. The faculty of your major department, however, have collectively many years of experience in the subject in which you have chosen to specialize. They really do know what courses will be helpful to you in the future. They may have had trouble with the same courses that now give you trouble, but they realize that all the work will be worth it in the end. Some courses you take, particularly in your major field itself, have obvious and immediate utility. Other courses, often times chemistry included, are provided to give you a general background knowledge, which may prove useful in understanding your own major or other subjects related to your major. In perhaps a burst of bravado, chemistry has been called "the central science" by one team of textbook authors. This moniker is very true however: in order to understand biology, physics, nutrition, farming, home economics, or whatever (it helps to have a general background in chemistry).

5. **Physical Quantity**  
**Basic SI Units**

<table>
<thead>
<tr>
<th>Physical Quantity</th>
<th>Basic SI Units</th>
</tr>
</thead>
<tbody>
<tr>
<td>mass</td>
<td>kilogram</td>
</tr>
<tr>
<td>distance</td>
<td>meter</td>
</tr>
<tr>
<td>time</td>
<td>second</td>
</tr>
<tr>
<td>temperature</td>
<td>kelvin</td>
</tr>
</tbody>
</table>

**Commonly Used Prefixes in the Metric System**

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Meaning</th>
<th>Power of Ten</th>
</tr>
</thead>
<tbody>
<tr>
<td>mega-</td>
<td>million</td>
<td>$10^6$</td>
</tr>
<tr>
<td>kilo-</td>
<td>thousand</td>
<td>$10^3$</td>
</tr>
<tr>
<td>deci-</td>
<td>tenth</td>
<td>$10^{-1}$</td>
</tr>
<tr>
<td>centi-</td>
<td>hundredth</td>
<td>$10^{-2}$</td>
</tr>
<tr>
<td>milli-</td>
<td>thousandth</td>
<td>$10^{-3}$</td>
</tr>
<tr>
<td>nano-</td>
<td>billionth</td>
<td>$10^{-9}$</td>
</tr>
</tbody>
</table>

The metric system is in use in most of the world because its system of units and multiples is simple to remember, and the system permits easy conversion between units. The various multiples and subdivisions of the basic units are based on factors of ten, which is also the basis for our number system. The United States uses a historical system in which there is no simple relationship between most units. Although several attempts have been made to gradually change the United States over to the metric system, no widespread support for the program has been achieved. Although 2-liter soda bottles were accepted without too much complaint (since they replaced a similar-sized 2-quart bottle, the thought of having an size 85 waist (in cm) may be repugnant to too many Americans! There obviously also would be a great cost to industry to retool all machinery and measurement devices in metric units.

6. Whenever a scientific measurement is made, we always employ the instrument or measuring device we are using to the limits of its precision. On a practical basis, this usually means that we estimate our
reading of the last significant figure of the measurement. An example of
the uncertainty in the last significant figure is given for measuring
the length of a pin in the text in Figure 2.5. Scientists appreciate the
limits of experimental techniques and instruments, and always assume
that the last digit in a number representing a measurement has been
estimated. Since the last significant figure in every measurement is
assumed to be estimated, it is never possible to exclude uncertainty
from measurements. The best we can do is to try to improve our
techniques and instruments so that we get more significant figures for
our measurements.

Scientists are careful about reporting their measurements to the
appropriate number of significant figures so as to indicate to their
colleagues the precision with which experiments were performed. That is,
the number of significant figures reported indicates how "carefully"
measurements were made. Suppose you were considering buying a new home,
and the real estate agent told you that a prospective new house was
"between 1000-2000 square feet of space" and had "five or six rooms,
more or less" and stood on "maybe an acre or two of land". Would you buy
the house or would you look for another real estate agent? When a
scientist says that a sample of material "has a mass of 3.126 grams" he
or she is narrowing down the limits as to the actual, true mass of the
sample: the mass is clearly slightly more than half way between 3.12 and
3.13 grams.

The rules for significant figures are covered in Section 2.5 of
the text. In brief, these rules for experimentally measured numbers are
as follows: (1) nonzero integers are always significant; (2) leading
zeroes are never significant, captive zeroes are always significant, and
trailing zeroes may be significant (if a decimal point is indicated);
(3) exact numbers (e.g., definitions) have an infinite number of
significant figures.

When we have to round off an answer to the correct number of
significant figures (as limited by whatever measurement was least
precise), we do this in a particular manner. If the digit to be removed
is equal to or greater than 5, the preceding digit is increased by 1. If
the digit to be removed is less than 5, the preceding digit is not
changed. If you are going to perform a series of calculations involving
a set of data, hold on to the digits in the intermediate calculations
until you arrive at the final answer, and then round off the final
answer to the appropriate number of significant figures.

When doing arithmetic with experimentally determined numbers, the
final answer is determined by the least precise measurement. In doing
multiplication or division calculations, the number of significant
figures in the result should be the same as the measurement with the
fewest significant figures. In performing addition or subtraction, the
number of significant figures in the result is limited by the
measurement with the fewest decimal places.

8. Dimensional analysis is a method of problem solving which pays
particular attention to the units of measurements and uses these units
as if they were algebraic symbols that multiply, divide, and cancel.
Consider the following example. A dozen of eggs costs $1.25. Suppose we
want to know how much one egg costs, and also how much three dozens of eggs will cost. To solve these problems, we need to make use of two equivalence statements:

1 dozen of eggs = 12 eggs

1 dozen of eggs = $1.25

The first of these equivalence statements is obvious: everyone knows that 12 eggs is "equivalent" to one dozen. The second statement also expresses an equivalence: if you give the grocer $1.25, he or she will give you a dozen of eggs. From these equivalence statements, we can construct the conversion factors we need to answer the two questions. For the first question (what does one egg cost) we can set up the calculation as follows

\[
\frac{\$1.25}{12 \text{ eggs}} = \frac{\$1.25}{12} = \$0.104 \approx \$0.10
\]

as the cost of one egg. Similarly, for the second question (the cost of 3 dozens of eggs), we can set up the conversion as follows

\[
3 \text{ dozens} \times \frac{\$1.25}{1 \text{ dozen}} = \$3.75
\]

as the cost of three dozens of eggs. See section 2.6 of the text for how we construct conversion factors from equivalence statements.

9. The Fahrenheit (°F) temperature scale is defined so that an ice/water bath has an equilibrium temperature of 32°F, and a boiling water bath a temperature of 212°F, under normal atmospheric pressure. There are 180 degree divisions between these two reference points. The Celsius (°C) temperature scale is defined so that an ice/water bath has a temperature of 0°C, whereas a boiling water bath has a temperature of 100°C, with 100 degree divisions between these reference points. Both the Fahrenheit and Celsius temperature scales are human inventions which choose convenient, stable, reproducible reference points as their definitions. The Kelvin (K) or Absolute temperature scale is based on a fundamental property of matter itself: the zero and lowest point (Absolute Zero) on the Kelvin temperature scale is the lowest possible temperature that can exist, and represents the temperature at which atoms and molecules are in their lowest possible energy states. All other temperatures on the Kelvin scale are positive relative to this. Since only this one reference point is used to define the Kelvin scale, the size of the Kelvin degree relative to this point was chosen to be the same size as the Celsius degree for convenience. The temperature of an ice/water bath is 273 K, and the temperature of a boiling water bath is 373 K.

10. Defining what scientists mean by "matter" often seems circular to students. Scientists say that matter is something that "has mass and occupies space", without ever really explaining what it means to "have mass" or to "occupy space"! The concept of matter is so basic and
fundamental, that it becomes difficult to give a good textbook definition other than to say that matter is the "stuff" of which everything is made. Matter can be classified and subdivided in many ways, depending on what we are trying to demonstrate.

On the most fundamental basis, all matter is composed of tiny particles (such as protons, electrons, neutrons, and the other subatomic particles). On one higher level, these tiny particles are combined in a systematic manner into units called atoms. Atoms in turn may be combined to constitute molecules. And finally, large groups of molecules may be placed together to form a bulk sample of substance that we can see.

Matter can also be classified as to the physical state a particular substance happens to take. Some substances are solids, some are liquids, and some are gases. Matter can also be classified as to whether it is a pure substance (one type of molecule) or a mixture (more than one type of molecule), and furthermore whether a mixture is homogeneous or heterogeneous.

11. The physical properties of a substance are the inherent characteristics of the substance, which result in no change in the composition of the substance when we measure or study these properties. Such properties include color, odor, physical state, density, solubility, melting point, boiling point, etc. The chemical properties of a given substance indicate how that substance reacts with other substances. For example, when we say that sodium is a grayish-white, soft, low-density metal, we are describing some of sodium's physical properties. When we say that sodium metal reacts with chlorine gas to form sodium chloride, we are describing a chemical property of sodium.

A physical change for a substance is a change in the substance that does not alter the identity or composition of the substance; physical changes typically represent changes in only the physical state (solid, liquid, vapor) of the substance. A chemical change for a substance results in the substance being converted into another substance or substances. For example, when we heat a piece of sodium metal in a sealed tube in a burner flame, the sodium melts and then vaporizes: the liquid and vapor are still sodium, however, and only physical changes have occurred. On the other hand, if we heat a piece of sodium in an open flame, the sodium reacts with oxygen in the air and is converted to a mixture of sodium oxides. The pure elemental substance sodium is converted into compounds and has undergone a chemical change.

12. Chemists tend to give a functional definition of what they mean by an "element": an element is a fundamental substance that cannot be broken down into any simpler substances by chemical methods. Compounds, on the other hand, can be broken down into simpler substances (the elements of which the compound is composed). For example, sulfur and oxygen are both elements (sulfur occurs as $S_8$ molecules and oxygen as $O_2$ molecules). When sulfur and oxygen are placed together and heated, the compound sulfur dioxide ($SO_2$) forms. When we analyze the sulfur dioxide produced, we notice that each and every molecule consists of one sulfur atom and two oxygen atoms, and on a mass basis, consists of 50% each of sulfur and oxygen. We describe this by saying that sulfur dioxide has a constant composition. The fact that a given compound has constant
composition is usually expressed in terms of the mass percentages of the elements present in the compound, but realize that the reason the mass percentages are constant is because of a constant number of atoms of each type present in the compound's molecules. If a scientist anywhere in the universe analyzed sulfur dioxide, he or she would find the same composition: if a scientist finds something that does not have the same composition, then the substance cannot be sulfur dioxide.

13. A mixture is a combination of two or more substances which may be varied in its composition. Most commonly in chemistry, a mixture is a combination of two or more pure substances (either elements or compounds). A solution is a particular type of mixture which appears completely homogeneous throughout. Although a solution is homogeneous in appearance, realize, however, that a solution is still a mixture of two or more pure substances: if it were possible to see the individual particles of a solution, we would notice that their were different types of molecules present.

In a sample of a pure substance, there is only one type of molecule present. There are two types of pure substances: elemental substances and compound substances. In an elemental substance, not only are all the molecules the same type, but all the atoms within those molecules are the same type. For example, the pure elemental substance oxygen consists of O₂ molecules (with no other type of molecule present). In addition, each O₂ molecule contains only one type of atom (O). In a sample of a compound substance, all the molecules are of the same type, but within each molecule are found atoms of different elements. For example, the pure compound substance water consists of H₂O molecules (with no other type of molecule present). Within each H₂O molecule, however, are found two different types of atoms (H and O). A compound is not a mixture, however, because the atoms of the different elements are chemically bonded (not just physically mixed), and the composition of the compound is not variable as to the relative amounts of each element present.

Two methods for separating mixtures are described in the text: filtration and distillation. Filtration can be used to separate a solid from a liquid. Distillation can be used to separate two liquids, or a dissolved solid from a liquid. There are many other separation methods beyond the scope of this text.

14. Scientists define energy as "the capacity to do work". As with trying to define "matter" earlier in this chapter, energy is such a fundamental concept that it is hard to define (what is "work"?). Although the SI unit of energy is the joule, until relatively recently, energies were more commonly given in terms of the calorie: one calorie is defined to be the amount of heat required to raise the temperature of one gram of water by one Celsius degree. The calorie is a "working" definition, and we can more easily appreciate this amount of energy. In terms of the SI unit, 1 calorie = 4.184 joule, so it takes 4.184 J to raise the temperature of one gram of water by one Celsius degree.

The specific heat capacity of a substance, in general, is the amount of energy required to raise the temperature of one gram of a substance by one Celsius degree. Therefore, the specific heat capacity
of water must be 1.000 cal/g°C or 4.184 J/g°C. To see how specific heat capacities may be used to calculate the energy change for a process, consider this example: How much energy is required to warm 25.0 g of water from 15.1°C to 35.2°C? The specific heat capacity of water is 1.000 cal/g°C: this is the quantity of energy required to raise the temperature of only one gram of water by only one Celsius degree. In this example, we are raising the temperature of 25.0 g of water, and we are raising the temperature by (35.2 - 15.1) = 20.1°C. So, using the specific heat capacity as a conversion factor, we can say

\[
\text{energy required} = \left( \frac{1.00 \text{ cal}}{\text{g} \cdot \text{°C}} \right) \times (25.0 \text{ g}) \times (20.1 \text{°C}) = 503 \text{ cal}
\]

Notice how the units of g and °C cancel, leaving the answer in energy units only. This sort of calculation of energy change can be done for any substance, using the substance's own specific heat capacity (see Table 3.2).

15. a. \(122.4 \times 10^5 = (1.224 \times 10^2) \times 10^5 = 1.224 \times 10^7\)
b. \(5.993 \times 10^{-4} = 0.0005993\)
c. \(0.0004321 \times 10^4 = (4.321 \times 10^{-4}) \times 10^4 = 4.321 \times 10^0 = 4.321\)
d. \(5.241 \times 10^2 = 524.1\)
e. \(0.000009814 = 9.814 \times 10^{-7}\)
f. \(14.2 \times 10^0 = 14.2\)

16. a. \(6.0 \text{ pt} \times \frac{1 \text{ qt}}{2 \text{ pt}} \times \frac{1 \text{ L}}{1.0567 \text{ qt}} = 2.8 \text{ L}\)
b. \(6.0 \text{ pt} \times \frac{1 \text{ qt}}{2 \text{ pt}} \times \frac{1 \text{ gal}}{4 \text{ qt}} = 0.75 \text{ gal}\)
c. \(5.91 \text{ yd} \times \frac{1 \text{ m}}{1.0936 \text{ yd}} = 5.40 \text{ m}\)
d. \(16.0 \text{ L} \times \frac{1 \text{ qt}}{0.94633 \text{ L}} \times \frac{32 \text{ fl. oz.}}{1 \text{ qt}} = 541 \text{ fl. oz.}\)
e. \(5.25 \text{ L} \times \frac{1 \text{ gal}}{3.7854 \text{ L}} = 1.39 \text{ gal}\)
f. \(62.5 \text{ mi} \times \frac{1 \text{ km}}{0.62137 \text{ mi}} = 101 \text{ km}\)
g. \(8.25 \text{ m} \times \frac{1.0936 \text{ yd}}{1 \text{ m}} \times \frac{36 \text{ in}}{1 \text{ yd}} = 325 \text{ in}\)
4.25 kg \times \frac{2.2046 \text{ lb}}{1 \text{ kg}} = 9.37 \text{ lb}

88.5 \text{ cm} \times \frac{10 \text{ mm}}{1 \text{ cm}} = 885 \text{ mm}

4.21 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 10.7 \text{ cm}

17. a. 10.20 + 4.1 + 26.001 + 2.4 = 42.701 = 42.7 (one decimal place)
   b. \left[1.091 - 0.991\right] + 1.2 = 1.3 (one decimal place)
   c. (4.06 + 5.1)(2.032 - 1.02) = (9.16)(1.012) = (9.2)(1.01) = 9.3
   d. (67.21)(1.003)(2.4) = 161.8 = 1.6 \times 10^2 \text{ (only 2 significant figures)}
   e. \left[\frac{(7.815 + 2.01)(4.5)}{1.9001}\right] = \left[\frac{(9.825)(4.5)}{1.9001}\right]
      = \left[\frac{(9.825)(4.5)}{1.9001}\right] = 23.27
      = 23 \text{ (only 2 significant figures)}
   f. (1.67 \times 10^{-9})(1.1 \times 10^{-4}) = 1.837 \times 10^{-13} = 1.8 \times 10^{-13}
   g. (4.02 \times 10^{-4})(2.91 \times 10^{3})/(9.102 \times 10^{-1}) = 1.29
   h. \left(1.04 \times 10^2 + 2.1 \times 10^3\right)/(4.51 \times 10^3)
      = \left(10.4 \times 10^1 + 2.1 \times 10^1\right)/(4.51 \times 10^3)
      = 2.77 \times 10^{-2}
   i. \left(1.51 \times 10^{-3}\right)^2/(1.074 \times 10^{-7}) = (2.2801 \times 10^{-6})/(1.074 \times 10^{-7})
      = 21.2 \text{ (only 3 significant figures)}
   j. \left(1.89 \times 10^2\right)/\left[(7.01 \times 10^{-3})(4.1433 \times 10^4)\right]
      = \left(1.89 \times 10^2\right)/\left[290.45\right]
      = 0.651 \text{ (only 3 significant figures)}

18. \(t_v = 1.80(t_c) + 32\)
    \(t_c = (t_v - 32)/1.80\)
    \(t_k = t_c + 273\)
   a. \(1.80(-50.1^\circ C) + 32 = -58.2^\circ F\)
   b. \((-30.7^\circ C - 32)/1.80 = -34.8^\circ C\)
   c. \(541 \text{ K} - 273 = 268^\circ C\)
Cumulative Review: Chapters 1, 2, and 3

d. \(221^\circ C + 273 = 494\) K

e. \(351\) K - 273 = 78°C

\[1.80(78°C) + 32 = 172.4 = 172^\circ F\]

f. \((72^\circ F - 32)/1.80 = 22.2^\circ C\)

\[22.2^\circ C + 273 = 295.2 = 295\) K

19. \(\text{density} = \frac{\text{mass}}{\text{volume}}\) \(\text{mass} = \text{volume} \times \text{density}\) \(\text{volume} = \frac{\text{mass}}{\text{density}}\)

a. \(\text{density} = 121.4\) g/42.4 cm\(^3\) = 2.86 g/cm\(^3\)

b. 0.721 lb = 327 g

\(\text{density} = 327\) g/241 cm\(^3\) = 1.36 g/cm\(^3\)

c. \(\text{mass} = 124.1\) mL \times 0.821 g/mL = 102 g

d. 4.51 L = 4,510 cm\(^3\)

\(\text{mass} = 4,510\) cm\(^3\) \times 1.15 g/cm\(^3\) = 5.19 \times 10^3 g

e. \(\text{volume} = 142.4\) g/0.915 g/mL = 156 mL

f. 4.2 lb = 1.9 \times 10^3 g

\(\text{volume} = 1.9 \times 10^3\) g/3.75 g/cm\(^3\) = 507 cm\(^3\) = 5.1 \times 10^2\) cm\(^3\)

20. a. \(459\) J \times \\(\frac{1\ \text{cal}}{4.184\ \text{J}}\) = 109.7 = 110. cal

b. \(7,031\) cal \times \\(\frac{4.184\ \text{J}}{1\ \text{cal}}\) \times \\(\frac{1\ \text{kJ}}{1000\ \text{J}}\) = 29.42 kJ

c. 55.31 kJ = 55,310 J = 5.531 \times 10^4\) J

d. 78.3 kcal \times \\(\frac{4.184\ \text{kJ}}{1\ \text{kcal}}\) = 327.6 = 328 kJ

e. 4,541 cal = 4.541 kcal

f. 84.1 kJ \times \\(\frac{1\ \text{kcal}}{4.184\ \text{kJ}}\) = 20.1 kcal
21. In general, heat required = mass × specific heat capacity × temp. change

(a) temperature change = 27 °C
mass = 1000 J/(4.184 J/g °C)(27 °C) = 8.9 g

(b) temperature change = 100. °C
mass = 1000 J/(0.45 J/g °C)(100. °C) = 22 g

(c) temperature change = 57 °C
mass = 1000 J/(0.71 J/g °C)(57 °C) = 25 g

(d) 56 °F = 13.3 °C  
75 °F = 23.9 °C
temperature change = 23.9 -13.3 = 10.6 °C
mass = 1000 J/(0.13 J/g °C)(10.6 °C) = 7.3 × 10² g

(e) temperature change = 385 K - 289 K = 96 K = 96 °C
mass = 1000 J/(0.24 J/g °C)(96 °C) = 43 g

(f) 85 °F = 29.4 °C
temperature change = 29.4 °C - -10 °C = 39.4 °C
mass = 1000 J(0.89 J/g °C)(39.4 °C) = 29 g
Chapter 4    Chemical Foundations: Elements, Atoms, and Ions

1. Although the number and nature of the elementary substances postulated by the ancient Greeks were incorrect, their idea that the matter we encounter in everyday life is composed of a few simpler substances is very similar to our modern concepts. Also, the idea that the simpler substances combine with each other in regular, fixed manners compares well with the modern theory of matter.

2. The alchemists discovered several previously unknown elements (mercury, sulfur, antimony) and were the first to prepare several common acids.

3. Boyle’s most important contribution was his insistence that science should be firmly grounded in experiment. Boyle tried to limit the influence of any preconceptions about science, and only accepted as fact what could be demonstrated.

4. There are 112 elements presently known; of these 88 occur naturally and 24 are manmade. Table 4.1 lists the most common elements on the earth.

5. Oxygen is found in great abundance in the oceans (combined with hydrogen in water molecules) and in the earth itself (most rocks and minerals are oxygen compounds). Oxygen is found more commonly in compounds.

6. The four most abundant elements in living creatures are, respectively, oxygen, carbon, hydrogen, and nitrogen (see Table 4.2). In the nonliving world, the most abundant elements are, respectively, oxygen, silicon, aluminum, and iron (see Table 4.1).

7. B (boron)
   C (carbon)
   F (fluorine)
   H (hydrogen)
   I (iodine)
   K (potassium)
   N (nitrogen)
   O (oxygen)
   P (phosphorus)
   S (sulfur)
   U (uranium)
   V (vanadium)
   W (tungsten)
   Y (yttrium)

8. Sb (antimony)
   Cu (copper)
f. NaCl₂ is incorrect (NaCl is correct)
g. Cs₂Cl₂ is incorrect (CsCl is correct)

28. a. NH₄⁺, ammonium ion
   b. SO₃²⁻, sulfite ion
   c. NO₃⁻, nitrate ion
   d. SO₄²⁻, sulfate ion
   e. NO₂⁻, nitrite ion
   f. CN⁻, cyanide ion
   g. OH⁻, hydroxide ion
   h. ClO₄⁻, perchlorate ion
   i. ClO⁻, hypochlorite ion
   j. PO₄³⁻, phosphate ion

29. Na₂SO₃, CaSO₃, Al₂(SO₄)₃
   Na₂SO₄, CaSO₄, Al₂(SO₄)₃
   NaCN, Ca(CN)₂, Al(CN)₃
   NaClO₄, Ca(ClO₄)₂, Al(ClO₄)₃
   Na₃PO₄, Ca₃(PO₄)₂, AlPO₄

30. a. B₂O₃, diboron trioxide
    b. NO₂, nitrogen dioxide
    c. PCl₅, phosphorus trichloride
    d. N₂O₄, dinitrogen tetroxide
    e. P₂O₅, diphosphorus pentoxide
    f. ICl, iodine monochloride
    g. SF₆, sulfur hexafluoride
    h. N₂O₃, dinitrogen trioxide

31. a. K₂S
    c. HCl(aq)
    e. Al(NO₃)₃
    g. H₂S(aq)
    i. Mg(ClO₄)₂
    k. HNO₃(aq)
    m. CuBr₂
    o. AuCl₃
    b. NaH
    d. N₂O₄
    f. CaSO₄
    h. NH₄C₂H₃O₂
    j. P₂O₅
    l. Ag₂SO₄
    n. Ba₃(PO₄)₂
    p. MnCl₂
j. \text{Na}^+ \text{(11 protons, 10 electrons)}

k. \text{O}^{2-} \text{(8 protons, 10 electrons)}

l. \text{I}^- \text{(53 protons, 54 electrons)}

25. \text{KN} \text{ (potassium nitride); KBr (potassium bromide); KCl (potassium chloride); KH (potassium hydride); K}_2\text{O (potassium oxide); KI (potassium iodide)}

\text{Ca}_3\text{N}_2 \text{ (calcium nitride); CaBr}_2 \text{ (calcium bromide); CaCl}_2 \text{ (calcium chloride); CaH}_2 \text{ (calcium hydride); CaO (calcium oxide); CaI}_2 \text{ (calcium iodide)}

\text{AlN (aluminum nitride); AlBr}_3 \text{ (aluminum bromide); AlCl}_3 \text{ (aluminum chloride); AlH}_3 \text{ (aluminum hydride); Al}_2\text{O}_3 \text{ (aluminum oxide); AlI}_3 \text{ (aluminum iodide)}

\text{Ag}_2\text{N (silver nitride); AgBr (silver bromide); AgCl (silver chloride); AgH (silver hydride); Ag}_2\text{O (silver oxide); AgI (silver iodide)}

Realize that most of the hydrogen compounds of the nonmetallic elements which are given as ions in this question are, in fact, covalently bonded compounds (not ionic compounds): \text{H}_2\text{N (NH}_3\text{, ammonia); HBr (hydrogen bromide); HCl (hydrogen chloride); H}_2\text{ (elemental hydrogen); H}_2\text{O (water); HI (hydrogen iodide)}

\text{Na}_2\text{N (sodium nitride); NaBr (sodium bromide); NaCl (sodium chloride); NaH (sodium hydride); Na}_2\text{O (sodium oxide); NaI (sodium iodide).}

26. 

a. \text{FeCl}_3 \text{, iron(III) chloride, ferric chloride}

b. \text{Cu}_2\text{S, copper(I) sulfide, cuprous sulfide}

c. \text{CoBr}_2 \text{, cobalt(II) bromide, cobaltous bromide}

d. \text{Fe}_2\text{O}_3 \text{, iron(III) oxide, ferric oxide}

e. \text{AuI}_3 \text{, gold(III) iodide, auric iodide}

f. \text{Cr}_2\text{S}_3 \text{, chromium(III) sulfide, chromic sulfide}

g. \text{MnO}_2 \text{, manganese(IV) oxide, manganese dioxide (archaic)}

h. \text{CuO, copper(II) oxide, cupric oxide}

i. \text{NiS, nickel(II) sulfide, nickleous sulfide}

27. 

a. \text{NaS is incorrect (Na}_2\text{S is correct)}

b. \text{K}_2\text{S is correct}

c. \text{Rb}_3\text{N is correct}

d. \text{CaBr}_2 \text{ is incorrect (CaBr}_2 \text{ is correct)}

e. \text{AlI}_3 \text{ is correct}
23. **Atom** | **Simple Ion**
---|---
a. Mg | Mg$^{2+}$
b. F | F$^-$
c. Ag | Ag$^+$
d. Al | Al$^{3+}$
e. O | O$^2-$
f. Ba | Ba$^{2+}$
g. Na | Na$^+$
h. Br | Br$^-$
i. K | K$^+$
j. Ca | Ca$^{2+}$
k. S | S$^{2-}$
l. Li | Li$^+$
m. Cl | Cl$^-$

24. a. K$^+$ (19 protons, 18 electrons)
b. Ca$^{2+}$ (20 protons, 18 electrons)
c. Na$^{3-}$ (7 protons, 10 electrons)
d. Br$^-$ (35 protons, 36 electrons)
e. Al$^{3+}$ (13 protons, 10 electrons)
f. Ag$^+$ (47 protons, 46 electrons)
g. Cl$^-$ (17 protons, 18 electrons)
h. H$^+$ (1 proton, 0 electrons)
i. H$^-$ (1 proton, 2 electrons)
20. | Formula | Name     | Atomic Number |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>a. He</td>
<td>helium</td>
<td>2</td>
</tr>
<tr>
<td>b. B</td>
<td>boron</td>
<td>5</td>
</tr>
<tr>
<td>c. C</td>
<td>carbon</td>
<td>6</td>
</tr>
<tr>
<td>d. F</td>
<td>fluorine</td>
<td>9</td>
</tr>
<tr>
<td>e. S</td>
<td>sulfur</td>
<td>16</td>
</tr>
<tr>
<td>f. Ba</td>
<td>barium</td>
<td>56</td>
</tr>
<tr>
<td>g. Be</td>
<td>beryllium</td>
<td>4</td>
</tr>
<tr>
<td>h. O</td>
<td>oxygen</td>
<td>8</td>
</tr>
<tr>
<td>i. P</td>
<td>phosphorus</td>
<td>15</td>
</tr>
<tr>
<td>j. Si</td>
<td>silicon</td>
<td>14</td>
</tr>
</tbody>
</table>

21. | Atomic Number | Name     | Symbol |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>a. 19</td>
<td>potassium</td>
<td>K</td>
</tr>
<tr>
<td>b. 12</td>
<td>magnesium</td>
<td>Mg</td>
</tr>
<tr>
<td>c. 36</td>
<td>krypton</td>
<td>Kr</td>
</tr>
<tr>
<td>d. 92</td>
<td>uranium</td>
<td>U</td>
</tr>
<tr>
<td>e. 1</td>
<td>hydrogen</td>
<td>H</td>
</tr>
<tr>
<td>f. 6</td>
<td>carbon</td>
<td>C</td>
</tr>
<tr>
<td>g. 15</td>
<td>phosphorus</td>
<td>P</td>
</tr>
<tr>
<td>h. 20</td>
<td>calcium</td>
<td>Ca</td>
</tr>
<tr>
<td>i. 79</td>
<td>gold</td>
<td>Au</td>
</tr>
<tr>
<td>j. 82</td>
<td>lead</td>
<td>Pb</td>
</tr>
<tr>
<td>k. 29</td>
<td>copper</td>
<td>Cu</td>
</tr>
<tr>
<td>l. 35</td>
<td>bromine</td>
<td>Br</td>
</tr>
<tr>
<td>m. 2</td>
<td>helium</td>
<td>He</td>
</tr>
<tr>
<td>n. 8</td>
<td>oxygen</td>
<td>O</td>
</tr>
</tbody>
</table>

22. | protons | neutrons | electrons |
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>a. 4</td>
<td>He 2</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>b. 37</td>
<td>Cl 17</td>
<td>17</td>
<td>20</td>
</tr>
<tr>
<td>c. 79</td>
<td>Br 35</td>
<td>35</td>
<td>44</td>
</tr>
</tbody>
</table>
oxynions, the acid containing the anions will have the ending -ous if
the anion is the -ite anion and the ending -ic if the anion is the -ate
anion. For example, HNO₂ is nitrous acid and HNO₃ is nitric acid; H₂SO₄
is sulfurous acid and H₂SO₄ is sulfuric acid. The halogen elements
(Group 7) each form four oxynions, and consequently, four oxyacids. The
prefix hypo- is used for the oxyacid that contains fewer oxygen atoms
than the -ite anion, and the prefix per- is used for the oxyacid that
contains more oxygen atoms than the -ate anion. For example,

<table>
<thead>
<tr>
<th>acid</th>
<th>name</th>
<th>anion</th>
<th>anion name</th>
</tr>
</thead>
<tbody>
<tr>
<td>HBrO</td>
<td>hypobromous acid</td>
<td>BrO⁻</td>
<td>hypobromite</td>
</tr>
<tr>
<td>HBrO₂</td>
<td>bromous acid</td>
<td>BrO₂⁻</td>
<td>bromite</td>
</tr>
<tr>
<td>HBrO₃</td>
<td>bromic acid</td>
<td>BrO₃⁻</td>
<td>bromate</td>
</tr>
<tr>
<td>HBrO₄</td>
<td>perbromic acid</td>
<td>BrO₄⁻</td>
<td>perbromate</td>
</tr>
</tbody>
</table>

19. Name        | Symbol | Atomic Number |
---|--------|---------------|
 magnesium | Mg     | 12            |
 tin       | Sn     | 50            |
 lead      | Pb     | 82            |
 sodium    | Na     | 11            |
 hydrogen | H      | 1             |
 chlorine | Cl     | 17            |
 silver    | Ag     | 47            |
 potassium | K      | 19            |
 calcium   | Ca     | 20            |
 bromine   | Br     | 35            |
 neon      | Ne     | 10            |
 aluminum | Al     | 13            |
 gold      | Au     | 79            |
 mercury   | Hg     | 80            |
 iodine    | I      | 53            |
NO  nitrogen monoxide
NO2  nitrogen dioxide
N2O  dinitrogen monoxide
N2O4  dinitrogen tetroxide  (tetra or tetr means "four")

A polyatomic ion is an ion containing more than one atom. Some common polyatomic ions you should be familiar with are listed in Table 5.4. Parentheses are used in writing formulas containing polyatomic ions to indicate unambiguously how many of the polyatomic ion are present in the formula, to make certain that there is no mistake as to what is meant by the formula. For example, consider the substance calcium phosphate. The correct formula for this substance is Ca3(PO4)2, which indicates that three calcium ions are combined for every two phosphate ions (check the total number of positive and negative charges to see why this is so). If we did not write the parenthesis around the formula for the phosphate ion, that is, if we had written Ca3PO4, people reading this formula might think that there were 42 oxygen atoms present!

Several families of polyatomic anions contain an atom of a given element, combined with differing numbers of oxygen atoms. Such anions are called "oxyanions". For example, sulfur forms two common oxyanions, SO3^2- and SO4^2-. When there are two oxyanions in such a series (as for sulfur), the name of the anion with fewer oxygen atoms ends in -ite and the name of the anion with more oxygen atoms ends in -ate. Under this method, SO3^2- is named sulfite and SO4^2- is named sulfate. When there are more than two members of such a series, the prefixes hypo- and per- are used to indicate the members of the series with the fewest and largest number of oxygen atoms. For example, bromine forms four common oxyanions. The formulas and names of these oxyanions are listed below.

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>BrO^-</td>
<td>hypobromite (fewest number of oxygens)</td>
</tr>
<tr>
<td>BrO2^-</td>
<td>bromite</td>
</tr>
<tr>
<td>BrO3^-</td>
<td>bromate</td>
</tr>
<tr>
<td>BrO4^-</td>
<td>perbromate (largest number of oxygens)</td>
</tr>
</tbody>
</table>

Acids, in general, are substances which produce protons (H+ ions) when dissolved in water. For acids which do not contain oxygen, the prefix hydro- and the suffix -ic are used with the root name of the element present in the acid (for example: HCl, hydrochloric acid; H2S, hydrosulfuric acid; HF, hydrofluoric acid). The nomenclature of acids whose anions contain oxygen is more complicated. A series of prefixes and suffixes is used with the name of the non-oxygen atom in the anion of the acid: these prefixes and suffixes indicate the relative (not actual) number of oxygen atoms present in the anion. Most of the elements that form oxyanions form two such anions: for example, sulfur forms sulfite ion (SO3^2-) and sulfate ion (SO4^2-), and nitrogen forms nitrite ion (NO2^-) and nitrate ion (NO3^-). For an element that forms two
14. When naming ionic compounds, we name the positive ion (cation) first. For simple binary Type I ionic compounds, the ending -ide is added to the root name of the element which is the negative ion (anion). For example, for the Type I ionic compound formed between potassium and sulfur, K₂S, the name would be potassium sulfide: potassium is the cation, sulfur is the anion (with the suffix -ide added). Type II compounds are named by either of two systems, the "ous-ic" system (which is falling out of use), and the "Roman numeral" system which is preferred by most chemists. Type II compounds involve elements which form more than one stable ion, and so it is necessary to specify which ion is present in a given compound. For example, iron forms two types of stable ion: Fe²⁺ and Fe³⁺. Iron can react with oxygen to form either of two stable oxides, FeO or Fe₂O₃, depending on which cation is involved. Under the Roman numeral naming system, FeO would be named iron(II) oxide to show that it contains Fe²⁺ ions; Fe₂O₃ would be named iron(III) oxide to indicate that it contains Fe³⁺ ions. The Roman numeral used in a name corresponds to the charge of the specific ion present in the compound. Under the less-favored "ous-ic" system, for an element that forms two stable ions, the ending -ous is used to indicate the lower-charged ion, whereas the ending -ic is used to indicate the higher-charged ion. FeO and Fe₂O₃ would thus be named ferrous oxide and ferric oxide, respectively. The "ous-ic" system has fallen out of favor since it does not indicate the actual charge on the ion, but only that it is the lower or higher charged of the two. This can lead to confusion; for example Fe₂O₃ is called ferric oxide in this system, but Cu²⁺ is called cupric ion (since there is also a Cu⁺ stable ion).

15. Type III binary compounds represent compounds involving only nonmetallic elements. In writing the name for such compounds, the element listed first in the formula is named first (using the full name of the element), and then the second element in the formula is named as though it were an anion (with the -ide ending). This is similar, thus far, to the method used for naming ionic compounds (Type I). Since there often may be more than one compound possible involving the same two nonmetallic elements, the naming system for Type III compounds goes one step further than the system for ionic compounds, by explicitly stating (by means of a numerical prefix) the number of atoms of each of the nonmetallic elements present in the molecules of the compound. For example, carbon and oxygen (both nonmetals) form two common compounds, CO and CO₂. To indicate clearly which compound is being discussed, the names of these compounds indicate explicitly the number of oxygen atoms present by using a numerical prefix.

CO  carbon monoxide  (mon or mono is the prefix meaning "one")

CO₂  carbon dioxide  (di is the prefix meaning "two")

The prefix mono is not normally used for the first element named in a compound if there is only one atom of the element present, but numerical prefixes are used for the first element if there is more than one atom of that element present. For example, nitrogen and oxygen form many binary compounds. Study closely how the examples following are named:
nucleus). These ions obviously contain fewer electrons than the atoms from which they are formed, however. A negative ion forms when an atom or molecule gains one or more electrons from an outside source (another atom or molecule). For example, chlorine atoms and oxygen atoms form ions as indicated below:

\[ \text{Cl}^{+}(\text{atom}) + e^- \rightarrow \text{Cl}^{-} (\text{ion}) \]
\[ \text{O}^{+}(\text{atom}) + 2e^- \rightarrow \text{O}^{2-} (\text{ion}) \]

Since the periodic table is arranged in terms of the electronic structure of the elements, in particular with the elements in the same vertical column having similar electronic structures, the mere location of an element in the periodic table can be an indication of what simple ions the element forms. For example, the Group 1 elements all form 1+ ions (Li\(^{+}\), Na\(^{+}\), K\(^{+}\), Rb\(^{+}\), Cs\(^{+}\)), while the Group 7 elements all form 1- ions (F\(^{-}\), Cl\(^{-}\), Br\(^{-}\), I\(^{-}\)). You will learn more about how the charge of an ion is related to an atom's electronic structure in a later chapter. For now, concentrate in learning the material shown in Figure 4.19.

12. Ionic compounds typically are hard, crystalline solids with high melting and boiling points. Ionic substances like sodium chloride, when dissolved in water or when melted, conduct electrical currents: chemists have taken this evidence to mean that ionic substances consist of positively and negatively charged particles (ions). Although an ionic substance is made up of positively and negatively charged particles, there is no net electrical charge on a sample of such a substance because the total number of positive charges is balanced by an equal number of negative charges. An ionic compound could not possibly exist of just cations or just anions: there must be a balance of charge or the compound would be very unstable (like charges repel each other).

13. The principle we use when writing the formula of an ionic compound is sometimes called the "principle of electroneutrality". This is just a long word that means that a chemical compound must have an overall net electrical charge of zero. For ionic compounds, this means that the total number of positive charges on the positive ions present must equal the total number of negative charges on the negative ions present. For example, with sodium chloride, if we realize that an individual sodium ion has a 1+ charge, and that an individual chloride ion has a 1- charge, then if we combine one of each of these ions, the compound will have an overall net charge of zero: (1+) + (1-) = 0. On the other hand for magnesium iodide, when we realize that an individual magnesium ion has a 2+ charge, then clearly one iodide ion with its 1- charge will not lead to a compound with an overall charge of zero: we would need two iodide ions, each with its 1- charge, to balance the 2+ charge of the magnesium ion: (2+) + 2(1-) = 0. If we considered magnesium oxide, however, we would need only one oxide ion, with its 2- charge, to balance with one magnesium with its 2+ charge [(2+) + (2-) = 0], and so the formula of magnesium oxide is just MgO.
(reflecting 6 protons plus 7 neutrons in the nucleus). The various isotopes of an element have identical chemical properties since the chemical properties of an atom are a function of the electrons in the atom (not the nucleus). The physical properties of the isotopes of an element (and compounds containing those isotopes) may differ because of the difference in mass of the isotopes.

The periodic table arranges the elements in order of increasing atomic number (from \( Z = 1 \) to \( Z = 112 \)). The table is further arranged by placing elements with similar electronic structure (and hence similar chemical properties) into the same vertical column (group), beginning with each period a new principal energy shell. Based on this arrangement by electronic structure, the metallic elements tend to be towards the left-hand side of the chart, while the non-metallic elements are found towards the right-hand, upper side. Since metallic nature increases going downward within any vertical column (as the outermost shell gets farther from the nucleus), there are also some metallic elements among the lower members of groups at the right-hand side of the table (many periodic tables indicate the dividing line between metallic and nonmetallic elements with a colored "stairstep". Some of the groups have been given common names; these are listed below:

<table>
<thead>
<tr>
<th>Group</th>
<th>Family Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Alkali Metals</td>
</tr>
<tr>
<td>2</td>
<td>Alkaline Earth Elements</td>
</tr>
<tr>
<td>6</td>
<td>Chalcogens (not used very commonly)</td>
</tr>
<tr>
<td>7</td>
<td>Halogens</td>
</tr>
<tr>
<td>8</td>
<td>Noble Gases</td>
</tr>
</tbody>
</table>

10. Most elements are too reactive to be found in nature in other than the combined form. Aside from the noble metals gold, silver, and platinum, the only other elements commonly found in nature in the uncombined state are some of the gaseous elements (such as \( O_2 \), \( N_2 \), He, Ar, etc.), and the solid nonmetals carbon and sulfur.

11. Ions are electrically charged particles formed from atoms or molecules which have gained or lost one or more electrons. Isolated atoms typically do not form ions on their own, but are induced to gain or lose electrons by some other species (which loses or gains the electrons). Positively charged ions are called cations, while negative ions are termed anions. A positive ion forms when an atom or molecule loses one or more of its electrons (negative charges). For example, sodium atoms and magnesium atoms form ions as indicated below

\[
Na(\text{atom}) \rightarrow Na^+(\text{ion}) + e^- \\
Mg(\text{atom}) \rightarrow Mg^{2+}(\text{ion}) + 2e^-
\]

The resulting ions contain the same number of protons and neutrons in their nuclei as do the atoms from which they are formed, since the only change that has taken place involves the electrons (which are not in the
The expression nuclear atom indicates that we view the atom as having a dense center of positive charge (called the nucleus) around which the electrons move through primarily empty space. Rutherford's experiment involved shooting a beam of particles at a thin sheet of metal foil. According to the then current "plum pudding" model of the atom, most of these positively charged particles should have passed right through the foil. However, Rutherford detected that a significant number of particles effectively bounced off something and were deflected backwards to the source of particles, and that other particles were deflected from the foil at large angles. Rutherford realized that his observations could be explained if the atoms of the metal foil had a small, dense, positively charged nucleus, with a significant amount of empty space between nuclei. The empty space between nuclei would allow most of the particles to pass through the atom. However, if a particle hit a nucleus head-on, it would be deflected backwards at the source. If a positively-charged particle passed near a positively charged nucleus (but did not hit the nucleus head-on), then the particle would be deflected by the repulsive forces between the positive charges. Rutherford's experiment conclusively disproved the "plum pudding" model for the atom, which envisioned the atom as a uniform sphere of positive charge, with enough negatively charged electrons scattered through the atom to balance out the positive charge.

The three fundamental particles from which atoms are composed are electrons, protons, and neutrons. The properties of these particles are summarized below:

<table>
<thead>
<tr>
<th>Particle</th>
<th>Relative Mass</th>
<th>Relative Charge</th>
<th>Location</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton</td>
<td>1836</td>
<td>1+</td>
<td>nucleus</td>
</tr>
<tr>
<td>neutron</td>
<td>1839</td>
<td>none</td>
<td>nucleus</td>
</tr>
<tr>
<td>electron</td>
<td>1</td>
<td>1-</td>
<td>outside nucleus</td>
</tr>
</tbody>
</table>

It is the number and arrangement of the electrons in an atom which is responsible for the chemical behavior of the atom. The electrons are found in nearly the entire region of space occupied by an atom, from just outside the nucleus all the way out to the outermost edge of the atom. When two atoms approach each other in space prior to a reaction taking place, it is the electrons which "see" the electrons of the other atom. The nucleus is so small, compared to the overall size of the atom, that the particles are too far away from the outside of the atom to interact with other atoms.

Isotopes represent atoms of the same element which have different atomic masses. Isotopes are a result of the fact that atoms of a given element may have different numbers of neutrons in their nuclei. Isotopes have the same atomic number (number of protons in the nucleus) but have different mass numbers (total number of protons and neutrons in the nucleus). The different isotopes of an atom are indicated by symbolism of the form $^A_ZX$ in which $Z$ represents the atomic number, and $A$ the mass number, of element $X$. For example, $^{13}_6C$ represents a nuclide of carbon with atomic number 6 (6 protons in the nucleus) and mass number 13.
Cumulative Review: Chapters 4 and 5

1. An element is a pure substance which cannot be broken down into simpler substances by chemical means. There are presently 112 elements recognized, of which 88 occur in nature (the remaining 24 have been synthesized by nuclear processes). The most abundant elements (by mass) on the earth are oxygen (49.2%), silicon (25.7%), and aluminum (7.50%), with less than 5% of each of the other elements present. A table of elemental abundances is given in the text as Table 4.1.

2. How many elements could you name? While you certainly don’t have to memorize all the elements, you should at least be able to give the symbol or name for the most common elements (listed in Table 4.3).

3. The symbols for some elements may refer to an archaic name for the element, or to the element’s name in a modern language other than English. Here are some examples:

<table>
<thead>
<tr>
<th>Element</th>
<th>English Name</th>
<th>Derivation of Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Na</td>
<td>sodium</td>
<td>Latin: natrium</td>
</tr>
<tr>
<td>K</td>
<td>potassium</td>
<td>Latin: kalium</td>
</tr>
<tr>
<td>Fe</td>
<td>iron</td>
<td>Latin: ferrum</td>
</tr>
<tr>
<td>W</td>
<td>tungsten</td>
<td>German: wolfram</td>
</tr>
</tbody>
</table>

Dalton’s atomic theory as presented in this text consists of five main postulates. Realize that although Dalton’s theory was exceptional scientific thinking for its time, some of the postulates have been modified as our scientific instruments and calculational methods have become increasingly more sophisticated. The main postulates of Dalton’s theory are as follows: (1) Elements are made up of tiny particles called atoms; (2) all atoms of a given element are identical; (3) although all atoms of a given element are identical, these atoms are different from the atoms of all other elements; (4) atoms of one element can combine with atoms of another element to form a compound, and such a compound will always have the same relative numbers and types of atoms for its composition; (5) atoms are merely rearranged into new groupings during an ordinary chemical reaction, and no atom is ever destroyed and no new atom is ever created during such a reaction.

A compound is a distinct, pure substance that is composed of two or more elements held together by chemical bonds. In addition, a given compound always contains exactly the same relative masses of its constituent elements. This latter statement is termed the law of constant composition. The law of constant composition is a result of the fact that a given compound is made up of molecules containing a particular type and number of each constituent atom. For example, water’s composition by mass (88.8% oxygen, 11.2% hydrogen) is a result of the fact that each and every water molecule contains one oxygen atom (relative mass 16.0) and two hydrogen atoms (relative mass 1.008 each). The law of constant composition is important to our study of chemistry because it means that we can always assume that any sample of a given pure substance, from whatever source, will be identical to any other sample.
e. sodium chlorite  
f. cobalt(III) sulfate.

90. a. CaCl₂  
b. Ag₂O  
c. Al₂S₃  
d. BeBr₂  
e. H₂S  
f. KH  
g. MgI₂  
h. CsF

91. a. SO₂  
b. N₂O  
c. XeF₄  
d. P₄O₁₀  
e. PCl₅  
f. SF₆  
g. NO₂

92. a. NaH₂PO₄  
b. LiClO₄  
c. Cu(HCO₃)₂  
d. KC₂H₃O₂  
e. BaO₂  
f. Cs₂SO₃

93. a. AgClO₄  
b. Co(OH)₃  
c. NaClO or NaOCl  
d. K₂Cr₂O₇  
e. NH₄NO₂  
f. Fe(OH)₃  
g. NH₄HCO₃  
h. KBrO₄
c. arsenic triiodide

d. dinitrogen tetraoxide (tetroxide)

e. dichlorine monoxide

f. sulfur hexafluoride

84. a. iron(III) acetate, ferric acetate

b. bromine monofluoride

c. potassium peroxide

d. silicon tetrabromide

e. copper(II) permanganate, cupric permanganate

f. calcium chromate

85. nitrate (the ending -ate always implies the larger number of oxygen atoms)

86. a. $\text{CO}_3^{2-}$

b. $\text{HCO}_3^-$

c. $\text{C}_2\text{H}_3\text{O}_2^-$

d. $\text{CN}^-$

87. a. $\text{Cr}^{2+}$

b. $\text{CrO}_4^{2-}$

c. $\text{Cr}^{3+}$

d. $\text{Cr}_2\text{O}_7^{2-}$

88. a. carbonate

b. chlorate

c. sulfate

d. phosphate

e. perchlorate

f. permanganate

89. a. lithium dihydrogen phosphate

b. copper(II) cyanide

c. lead(II) nitrate

d. sodium hydrogen phosphate
80. a. Incorrect. Si is the element silicon, not silver.
   b. Incorrect. Co is the symbol for cobalt, not copper.
   c. Incorrect. Hydrogen exists as the hydride ion in this compound.
   d. Correct
   e. Incorrect. P is just "phosphorus" not "phosphoric".

81. a. Since the bromide ion must have a 1- charge, the iron ion must be in the 2+ state: the name is iron(II) bromide.
   b. Since sulfide ion always has a 2- charge, the cobalt ion must be in the 2+ state: the name is cobalt(II) sulfide.
   c. Since sulfide ion always has a 2- charge, and since there are three sulfide ions present, each cobalt ion must be in the 3+ state: the name is cobalt(III) sulfide.
   d. Since oxide ion always has a 2- charge, the tin ion must be in the 4+ state: the name is tin(IV) oxide.
   e. Since chloride ion always has a 1- charge, each mercury ion must be in the 1+ state: the name is mercury(I) chloride.
   f. Since chloride ion always has a 1- charge, the mercury ion must be in the 2+ state: the name is mercury(II) chloride.

82. a. Since bromide ions always have a 1- charge, the cobalt ion must have a 3+ charge: the name is cobaltic bromide.
   b. Since iodide ions always have a 1- charge, the lead ion must have a 4+ charge: the name is plumbic iodide.
   c. Since oxide ions always have a 2- charge, and since there are three oxide ions, each iron ion must have a 3+ charge: the name is ferric oxide.
   d. Since sulfide ions always have a 2- charge, the iron ion must have a 2+ charge: the name is ferrous sulfide.
   e. Since chloride ions always have a 1- charge, the tin ion must have a 4+ charge: the name is stannic chloride.
   f. Since oxide ions always have a 2- charge, the tin ion must have a 2+ charge: the name is stannous oxide.

83. a. Xenon hexafluoride
   b. Oxygen difluoride
These are some problems from the book "Nomenclature". They involve chemical reactions and balancing equations.

f. \( \text{P}(15e^-) + 3e^- \rightarrow \text{P}^-(18e^-) \)

77. a. none likely (element 36, Kr, is a noble gas)
b. \( \text{Ga}^{3+} \) (element 31, Ga, is in Group 3)
c. \( \text{Te}^{2-} \) (element 52, Te, is in Group 6)
d. \( \text{Tl}^{3+} \) (element 81, Tl, is in Group 3)
e. \( \text{Br}^- \) (element 35, Br, is in Group 7)
f. \( \text{Fr}^+ \) (element 87, Fr, is in Group 1)

78. a. Two 1+ ions are needed to balance a 2- ion, so the formula must have two Na+ ions for each S2- ion: \( \text{Na}_2\text{S} \).
b. One 1+ ion exactly balances a 1- ion, so the formula should have an equal number of K+ and Cl- ions: \( \text{KCl} \).
c. One 2+ ion exactly balances a 2- ion, so the formula must have an equal number of Ba2+ and O2- ions: \( \text{BaO} \).
d. One 2+ ion exactly balances a 2- ion, so the formula must have an equal number of Mg2+ and Se2- ions: \( \text{MgSe} \).
e. One 2+ ion requires two 1- ions to balance charge, so the formula must have twice as many Br- ions as Cu2+ ions: \( \text{CuBr}_2 \).
f. One 3+ ion requires three 1- ions to balance charge, so the formula must have three times as many I- ions as Al3+ ions: \( \text{AlI}_3 \).
g. Two 3+ ions give a total of 6+, whereas three 2- ions will give a total of 6-. The formula then should contain two Al3+ ions and three O2- ions: \( \text{Al}_2\text{O}_3 \).
h. Three 2+ ions are required to balance two 3- ions, so the formula must contain three Ca2+ ions for every two N3- ions: \( \text{Ca}_3\text{N}_2 \).

79. a. beryllium oxide
b. magnesium iodide
c. sodium sulfide
d. aluminum oxide
e. hydrogen chloride (gaseous); hydrochloric acid (aqueous)
f. lithium fluoride
g. silver(I) sulfide; usually called silver sulfide
h. calcium hydride
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<th>Formula</th>
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<td>Ni(HSO&lt;sub&gt;4&lt;/sub&gt;)&lt;sub&gt;2&lt;/sub&gt;</td>
<td>Ni(H&lt;sub&gt;2&lt;/sub&gt;PO&lt;sub&gt;4&lt;/sub&gt;)&lt;sub&gt;2&lt;/sub&gt;</td>
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<td>Ag&lt;sub&gt;2&lt;/sub&gt;SO&lt;sub&gt;4&lt;/sub&gt;</td>
<td>AgHSO&lt;sub&gt;4&lt;/sub&gt;</td>
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<td>BaCl&lt;sub&gt;2&lt;/sub&gt;</td>
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</table>

67. unreactive
68. helium
69. two

70. iodine (solid), bromine (liquid), fluorine and chlorine (gases)

71. 2-
72. 1-
73. 3+
74. 1-

75. [1] e
    [2] a
    [3] a
    [4] g
    [5] g
    [6] f
    [7] g
    [8] a
    [9] e
    [10] j

76. a. Al(13e<sup>-</sup>) → Al<sup>3+</sup>(10e<sup>-</sup>) + 3e<sup>-</sup>
    b. S(16e<sup>-</sup>) + 2e<sup>-</sup> → S<sup>-</sup>(18e<sup>-</sup>)
    c. Cu(29e<sup>-</sup>) → Cu<sup>+</sup>(28e<sup>-</sup>) + e<sup>-</sup>
    d. F(9e<sup>-</sup>) + e<sup>-</sup> → F<sup>-</sup>(10e<sup>-</sup>)
    e. Zn(30e<sup>-</sup>) → Zn<sup>2+</sup>(28e<sup>-</sup>) + 2e<sup>-</sup>
Ni^{2+}: NiCO_{3} \quad \text{nickel(II) carbonate}
Ni(BrO_{3})_{2} \quad \text{nickel(II) bromate}
Ni(C_{2}H_{3}O_{2})_{2} \quad \text{nickel(II) acetate}
Ni(OH)_{2} \quad \text{nickel(II) hydroxide}
Ni(HCO_{3})_{2} \quad \text{nickel(II) bicarbonate}
Ni_{3}(PO_{4})_{2} \quad \text{nickel(II) phosphate}
NiSO_{4} \quad \text{nickel(II) sulfate}
Ni(ClO_{4})_{2} \quad \text{nickel(II) perchlorate}
NiCl_{2} \quad \text{nickel(II) chloride}

Hg^{2+}: HgCO_{3} \quad \text{mercury(I) carbonate}
Hg(BrO_{3})_{2} \quad \text{mercury(I) bromate}
Hg(C_{2}H_{3}O_{2})_{2} \quad \text{mercury(I) acetate}
Hg(OH)_{2} \quad \text{mercury(I) hydroxide}
Hg(HCO_{3})_{2} \quad \text{mercury(I) bicarbonate}
(Hg_{2})_{3}(PO_{4})_{2} \quad \text{mercury(I) phosphate}
HgSO_{3} \quad \text{mercury(I) sulfite}
Hg(ClO_{4})_{2} \quad \text{mercury(I) perchlorate}
HgSO_{4} \quad \text{mercury(I) sulfate}
HgO \quad \text{mercury(I) oxide}
HgCl_{2} \quad \text{mercury(I) chloride}

Hg^{2+}: HgCO_{3} \quad \text{mercury(II) carbonate}
Hg(BrO_{3})_{2} \quad \text{mercury(II) bromate}
Hg(C_{2}H_{3}O_{2})_{2} \quad \text{mercury(II) acetate}
Hg(OH)_{2} \quad \text{mercury(II) hydroxide}
Hg(HCO_{3})_{2} \quad \text{mercury(II) bicarbonate}
Hg_{2}(PO_{4})_{2} \quad \text{mercury(II) phosphate}
HgSO_{3} \quad \text{mercury(II) sulfite}
Hg(ClO_{4})_{2} \quad \text{mercury(II) perchlorate}
HgSO_{4} \quad \text{mercury(II) sulfate}
HgO \quad \text{mercury(II) oxide}
HgCl_{2} \quad \text{mercury(II) chloride}

Ca(NO_{3})_{2} \quad \text{CaSO}_{4} \quad \text{Ca(HSO}_{4})_{2} \quad \text{Ca(H}_{2}PO_{4})_{2} \quad \text{CaO} \quad \text{CaCl}_{2}
Sr(NO_{3})_{2} \quad \text{SrSO}_{4} \quad \text{Sr(HSO}_{4})_{2} \quad \text{Sr(H}_{2}PO_{4})_{2} \quad \text{SrO} \quad \text{SrCl}_{2}
(NH_{4})NO_{3} \quad \text{NH}_{4}(SO_{4})_{2} \quad \text{NH}_{4}HSO_{4} \quad \text{NH}_{4}H_{2}PO_{4} \quad (NH_{4})_{2}O \quad \text{NH}_{4}Cl
Al(NO_{3})_{3} \quad \text{Al}_{2}(SO_{4})_{3} \quad \text{Al(HSO}_{4})_{3} \quad \text{Al(H}_{2}PO_{4})_{3} \quad \text{Al}_{2}O_{3} \quad \text{AlCl}_{3}
Fe(NO_{3})_{3} \quad \text{Fe}_{2}(SO_{4})_{3} \quad \text{Fe(HSO}_{4})_{3} \quad \text{Fe(H}_{2}PO_{4})_{3} \quad \text{Fe}_{2}O_{3} \quad \text{FeCl}_{3}
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<td>Fe³⁺:</td>
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<td>FeCl₃</td>
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</tbody>
</table>
63. Answers are given, respectively, for the $M^{2+}$, $M^{2+}$, and $M^{3+}$ ions:
   a. $M_2\text{CrO}_4$, $M\text{CrO}_4$, $M_2\text{(CrO}_4)_3$
   b. $M_2\text{Cr}_2\text{O}_7$, $M\text{Cr}_2\text{O}_7$, $M_2\text{(Cr}_2\text{O}_7)_3$
   c. $M_2\text{S}$, $M\text{S}$, $M_2\text{S}_3$
   d. $M\text{Br}_2$, $M\text{Br}_2$, $M\text{Br}_3$
   e. $M\text{HCO}_3$, $M\text{(HCO}_3)_2$, $M\text{(HCO}_3)_3$
   f. $M_2\text{HPO}_4$, $M\text{HPO}_4$, $M_2\text{(HPO}_4)_3$

64. $M^+$ compounds: $MD$, $M_2E$, $M_3F$
   $M^{2+}$ compounds: $MD_2$, $M_2E_2$, $M_3F_2$
   $M^{3+}$ compounds: $MD_3$, $M_2E_3$, $MF$

65. $Fe^{2+}$:
   - $Fe\text{CO}_3$: iron(II) carbonate; ferrous carbonate
   - $Fe\text{(BrO}_3)_2$: iron(II) bromate; ferrous bromate
   - $Fe\text{(C}_2\text{H}_3\text{O}_2)_2$: iron(II) acetate; ferrous acetate
   - $Fe\text{(OH)}_2$: iron(II) hydroxide; ferrous hydroxide
   - $Fe\text{(HCO}_3)_2$: iron(II) bicarbonate; ferrous bicarbonate
   - $Fe_3\text{(PO}_4)_2$: iron(II) phosphate; ferrous phosphate
   - $Fe\text{SO}_3$: iron(II) sulfite; ferrous sulfite
   - $Fe\text{(ClO}_4)_2$: iron(II) perchlorate; ferrous perchlorate
   - $Fe\text{SO}_4$: iron(II) sulfate; ferrous sulfate
   - $FeO$: iron(II) oxide; ferrous oxide
   - $Fe\text{Cl}_2$: iron(II) chloride; ferrous chloride

$Al^{3+}$:
   - $Al_2\text{(CO}_3)_3$: aluminum carbonate
   - $Al\text{(BrO}_3)_3$: aluminum bromate
   - $Al\text{(C}_2\text{H}_3\text{O}_2)_3$: aluminum acetate
   - $Al\text{(OH)}_3$: aluminum hydroxide
   - $Al\text{(HCO}_3)_3$: aluminum bicarbonate
   - $Al\text{PO}_4$: aluminum phosphate
   - $Al_2\text{(SO}_3)_3$: aluminum sulfite
   - $Al\text{(ClO}_4)_3$: aluminum perchlorate
   - $Al_2\text{(SO}_4)_3$: aluminum sulfate
   - $Al_2\text{O}_3$: aluminum oxide
   - $Al\text{Cl}_3$: aluminum chloride

$Na^+$:
   - $Na_2\text{CO}_3$: sodium carbonate
   - $Na\text{BrO}_3$: sodium bromate
   - $Na\text{C}_2\text{H}_3\text{O}_2$: sodium acetate
   - $Na\text{OH}$: sodium hydroxide
e. ammonia
f. silver(I) sulfate (usually called silver sulfate)
g. beryllium hydroxide

60. a. chloric acid
b. cobalt(III) chloride; cobaltic chloride
c. diboron trioxide
d. water
e. acetic acid
f. iron(III) nitrate; ferric nitrate
g. copper(II) sulfate; cupric sulfate

60. a. ammonium carbonate
b. ammonium hydrogen carbonate, ammonium bicarbonate
c. calcium phosphate
d. sulfurous acid
e. manganese(IV) oxide
f. iodic acid
g. potassium hydride

61. a. $\text{K}_2\text{O}$
b. $\text{MgO}$
c. $\text{FeO}$
d. $\text{Fe}_2\text{O}_3$
e. $\text{ZnO}$
f. $\text{PbO}$
g. $\text{Al}_2\text{O}_3$

62. a. $\text{M(C}_2\text{H}_3\text{O}_2)_4$
b. $\text{M(MnO}_4)_4$
c. $\text{MO}_7$
d. $\text{M(HPO}_4)_2$
e. $\text{M(OH)}_4$
f. $\text{M(NO}_2)_4$
51. Ionic compounds have a strong resistance to melting (compared to covalent compounds like sucrose) because of the strong electrical attraction of the positive ions for the negative ions in the crystal of ionic compound (each positive ion is surrounded by several negative ions and vice versa).

52. A moist paste of NaCl would contain Na\(^+\) and Cl\(^-\) ions in solution, and would serve as a conductor of electrical impulses.

53. A binary compound is a compound containing two and only two elements. A polyatomic anion is several atoms bonded together which, as a whole, carries a negative electrical charge. An oxyanion is a negative ion containing a particular element and one or more oxygen atoms.

54. 
   \[
   \begin{align*}
   H &\to H^+ (\text{hydrogen ion: a cation}) + e^- \\
   H + e^- &\to H^- (\text{hydride ion: an anion})
   \end{align*}
   \]

55. hydrogen

56. missing oxyanions: IO\(_3^-\); ClO\(_2^-\)

   missing oxyacids: HClO\(_4\); HClO; HBrO\(_2\)

57. a. calcium acetate
   b. phosphorus trichloride
   c. copper(II) permanganate, cupric permanganate
   d. iron(III) carbonate, ferric carbonate
   e. lithium hydrogen carbonate, lithium bicarbonate
   f. chromium(III) sulfide, chromic sulfide
   g. calcium cyanide

58. a. gold(III) bromide, auric bromide
   b. cobalt(III) cyanide, cobaltic cyanide
   c. magnesium hydrogen phosphate
   d. diboron hexahydride (diborane is its common name)
Chapter 5  Nomenclature  77

c. \( \text{HC}_2\text{H}_3\text{O}_2 \)
d. \( \text{HBr} \)
e. \( \text{HClO}_2 \)
f. \( \text{H}_2\text{Se} \)
g. \( \text{H}_2\text{SO}_3 \)
h. \( \text{HBrO}_2 \)

48. a. \( \text{HCN} \)
b. \( \text{HNO}_3 \)
c. \( \text{H}_2\text{SO}_4 \)
d. \( \text{H}_3\text{PO}_4 \)
e. \( \text{HClO} \) or \( \text{HOCl} \)
f. \( \text{HBr} \)
g. \( \text{HBrO}_2 \)
h. \( \text{HF} \)

49. a. \( \text{LiCl} \)
b. \( \text{Cu}_2\text{CO}_3 \)
c. \( \text{HBr} \)
d. \( \text{Ca(NO}_3)_2 \)
e. \( \text{NaClO}_4 \)
f. \( \text{Al(OH)}_3 \)
g. \( \text{Ba(HCO}_3)_2 \)
h. \( \text{FeSO}_4 \)
i. \( \text{B}_2\text{Cl}_6 \)
j. \( \text{PBr}_5 \)
k. \( \text{K}_2\text{SO}_3 \)
l. \( \text{Ba(C}_2\text{H}_3\text{O}_2)_2 \)

50. a. \( \text{Mg(HSO}_4)_2 \)
b. \( \text{CsClO}_4 \)
c. \( \text{FeO} \)
d. \( \text{H}_2\text{Te} \)
e. \( \text{Sr(NO}_3)_2 \)
f. \( \text{Sn(C}_2\text{H}_3\text{O}_2)_4 \)
43. a. $\text{PI}_3$
   b. $\text{SiCl}_4$
   c. $\text{N}_2\text{O}_5$
   d. $\text{IBr}$
   e. $\text{B}_2\text{O}_3$
   f. $\text{N}_2\text{O}_4$
   g. $\text{CO}$

44. a. $\text{CO}_2$
   b. $\text{SO}_2$
   c. $\text{N}_2\text{Cl}_4$
   d. $\text{Cl}_4$
   e. $\text{PF}_5$
   f. $\text{P}_2\text{O}_5$

45. a. $\text{NH}_4\text{NO}_3$
   b. $\text{Mg}(\text{C}_2\text{H}_3\text{O}_2)_2$
   c. $\text{CaO}_2$
   d. $\text{KHSO}_4$
   e. $\text{FeSO}_4$
   f. $\text{KHCO}_3$
   g. $\text{CoSO}_4$
   h. $\text{LiClO}_4$

46. a. $\text{Ca}_3(\text{PO}_4)_2$
   b. $\text{NH}_4\text{NO}_3$
   c. $\text{Al}(\text{HSO}_4)_3$
   d. $\text{BaSO}_4$
   e. $\text{Fe}_2(\text{NO}_3)_3$
   f. $\text{CuOH}$

47. a. $\text{H}_2\text{S}$
   b. $\text{HBrO}_4$
37. An acid is a substance which produces hydrogen ions, $\text{H}^+$, when dissolved in water.

38. oxygen (commonly referred to as oxyacids)

39. a. hydrochloric acid
   b. sulfuric acid
   c. nitric acid
   d. hydroiodic acid
   e. nitrous acid
   f. chloric acid
   g. hydrobromic acid
   h. hydrofluoric acid
   i. acetic acid

40. a. hypochlorous acid
    b. sulfurous acid
    c. bromic acid
    d. hypoiodous acid
    e. perbromic acid
    f. hydrosulfuric acid
    g. hydroselenic acid
    h. phosphorous acid

41. a. $\text{Li}_2\text{O}$
    b. $\text{AlI}_3$
    c. $\text{Ag}_2\text{O}$
    d. $\text{K}_2\text{N}$
    e. $\text{Ca}_3\text{P}_2$
    f. $\text{MgF}_2$
    g. $\text{Na}_2\text{S}$
    h. $\text{BaH}_2$

42. a. $\text{PbO}_2$
    b. $\text{SnBr}_2$
    c. CuS
    d. CuI
32. a. MgCl₂
   b. Ca(ClO)₂
   c. KClO₃
   d. Ba(ClO₄)₂

33. a. permanganate
   b. peroxide
   c. chromate
   d. dichromate
   e. nitrate
   f. sulfite

34. a. ammonium
   b. dihydrogen phosphate
   c. sulfate
   d. hydrogen sulfite (also called bisulfite)
   e. perchlorate
   f. iodate

35. a. iron(III) nitrate, ferric nitrate
   b. cobalt(II) phosphate, cobaltous phosphate
   c. chromium(III) cyanide, chromic cyanide
   d. aluminum sulfate
   e. chromium(II) acetate, chromous acetate
   f. ammonium sulfite

36. a. ammonium sulfate
   b. potassium perchlorate
   c. iron(III) sulfate, ferric sulfate
   d. calcium phosphate
   e. calcium hydroxide
   f. potassium carbonate
e. manganese(II) fluoride - ionic  
f. zinc oxide - ionic

23. polyatomic

24. An oxyanion is a polyatomic anion containing oxygen combined with another element. The following oxyanions of bromine illustrate the nomenclature

\[
\begin{align*}
\text{BrO}^- & \quad \text{hypobromite} \\
\text{BrO}_2^- & \quad \text{bromite} \\
\text{BrO}_3^- & \quad \text{bromate} \\
\text{BrO}_4^- & \quad \text{perbromate} \\
\end{align*}
\]

25. one fewer oxygen atom

26. perchlorate, \( \text{ClO}_4^- \)

27. \( \begin{align*}
\text{ClO}_4^- & \quad \text{perchlorate} \\
\text{ClO}^- & \quad \text{hypochlorite} \\
\text{ClO}_3^- & \quad \text{chlorate} \\
\text{ClO}_2^- & \quad \text{chlorite} \\
\end{align*} \)

28. hypobromite

\[
\begin{align*}
\text{IO}_3^- & \quad \text{periodate} \\
\text{IO}^- & \quad \text{or} \quad \text{IO}^{} \\
\end{align*}
\]

29. \( \begin{align*}
a. & \quad \text{p}^3^- \\
b. & \quad \text{PO}_4^{3^-} \\
c. & \quad \text{PO}_3^{3^-} \\
d. & \quad \text{HPO}_4^{2^-} \\
\end{align*} \)

30. \( \begin{align*}
a. & \quad \text{NO}_3^- \\
b. & \quad \text{NO}_2^- \\
c. & \quad \text{NH}_4^+ \\
d. & \quad \text{CN}^- \\
\end{align*} \)

31. \( \begin{align*}
a. & \quad \text{Cl}^- \\
b. & \quad \text{ClO}^- \\
\end{align*} \)
e. nitrogen triiodide
f. diboron trioxide

18. a. germanium tetrahydride
    b. dinitrogen tetrabromide
    c. diphosphorus pentasulfide
d. selenium dioxide
e. ammonia (nitrogen trihydride)
f. silicon dioxide

19. a. tin(IV) bromide, stannic bromide - ionic
    b. aluminum hydride - ionic
c. iron(II) oxide, ferrous oxide - ionic
d. copper(II) iodide, cupric iodide - ionic
e. oxygen difluoride - nonionic
    f. xenon hexafluoride - nonionic

20. a. diboron hexahydride - nonionic (common name: diborane)
b. calcium nitride - ionic
c. carbon tetrabromide - nonionic
d. silver sulfide - ionic
e. copper(II) chloride, cupric chloride - ionic
    f. chlorine monofluoride - nonionic

21. a. magnesium sulfide - ionic
    b. aluminum chloride - ionic
c. phosphorus trihydride (the common name phosphine is always used)
d. chlorine monobromide - nonionic
e. lithium oxide - ionic
    f. tetraphosphorus decoxide - nonionic

22. a. radium chloride - ionic
    b. selenium dichloride - nonionic
c. phosphorus trichloride - nonionic
d. sodium phosphide - ionic
f. Since each bromide ion has a 1- charge, the tin ion must have a 4+ charge: tin(IV) bromide.

a. Since each chloride ion has a 1- charge, the cobalt ion must have a 2+ charge: cobaltous chloride.

b. Since each bromide ion has a 1- charge, the chromium ion must have a 3+ charge: the name is chromic bromide.

c. Since the oxide ion has a 2- charge, the lead ion must have a 2+ charge: the name is plumbous oxide.

d. Since each oxide ion has a 2- charge, the tin ion must have a 4+ charge: the name is stannic oxide.

e. Since each oxide ion has a 2- charge, the iron ion must have a 3+ charge: the name is ferric oxide.

f. Since each chloride ion has a 1- charge, the iron ion must have a 3+ charge: the name is ferric chloride.

16. a. Since oxide ions have a 2- charge the lead ion must have a 4+ charge: the name is plumbic oxide.

b. Since bromide ions have a 1- charge, the tin ion must have a 2+ charge: the name is stannous bromide.

c. Since the sulfide ion has a 2- charge, the two copper ions must each have a 1+ charge: the name is cuprous sulfide.

d. Since the iodide ion has a 1- charge, the copper ion must have a 1+ charge: the name is cuprous iodide.

e. Since iodide ions have a 1- charge, each mercury must have a 1+ charge: the name is mercurous iodide.

f. Since fluoride ions have a 1- charge, the chromium ion must have a 3+ charge: the name is chromic fluoride.

17. Remember that for this type of compound of nonmetals, numerical prefixes are used to indicate how many of each type of atom is present. However, if only one atom of the first element mentioned in the compound is present in a molecule, the prefix mono- is not needed.

- a. iodine pentafluoride
- b. arsenic trichloride
- c. selenium monoxide
- d. xenon tetrafluoride
11. a. incorrect; BaH₂ is barium hydride  
b. incorrect; Na₂O is sodium oxide  
c. correct  
d. incorrect; SiO₂ is silicon dioxide  
e. correct  

12. a. incorrect; silver chloride is AgCl  
b. correct  
c. incorrect; sodium oxide would be Na₂O  
d. incorrect; barium chloride is BaCl₂  
e. incorrect; strontium oxide is SrO  

13. a. Since the bromide ion has a 1- charge, the tin ion must have a 2+ charge: the name is tin(II) bromide.  
b. Since the iodide ion has a 1- charge, the tin ion must have a 4+ charge: the name is tin(IV) iodide.  
c. Since the oxide ion has a 2- charge, the chromium ion must have a 2+ charge: the name is chromium(II) oxide.  
d. Since the oxide ion has a 2- charge, each chromium ion must have a 3+ charge: the name is chromium(III) oxide.  
e. Since the iodide ion has a 1- charge, each mercury ion must have a 1+ charge: the name is mercury(I) iodide.  
f. Since the iodide ion has a 1- charge, the mercury ion must have a 2+ charge: the name is mercury(II) iodide.  

14. a. Since each iodide ion has a 1- charge, the iron ion must have a 3+ charge: the name is iron(III) iodide.  
b. Since each chloride ion has a 1- charge, the manganese must have a 2+ charge: the name is manganese(II) chloride.  
c. Since the oxide ion has a 2- charge, the mercury ion must have a 2+ charge: mercury(II) oxide.  
d. Since the oxide ion has a 2- charge, the copper atoms must each have a 1+ charge: the name is copper(I) oxide.  
e. Since the oxide ion has a 2- charge, the copper ion must have a 2+ charge: copper(II) oxide.
Chapter 5  Nomenclature

1. A binary compound is one that contains only two elements. Examples are sodium chloride, water, and carbon dioxide.

2. compounds that contain a metal and a nonmetal; compounds containing two nonmetals

3. cation, anion

4. cation

5. positive ion (cation)

6. The substance "sodium chloride" consists of an extended lattice array of sodium ions, Na\(^+\), and chloride ions, Cl\(^-\). Each sodium ion is surrounded by several chloride ions, and each chloride ion is surrounded by several sodium ions. We write the formula as NaCl to indicate the relative number of each ion in the substance.

7. -ous

8. Roman numeral

9. a. sodium iodide
   b. calcium fluoride
   c. aluminum sulfide
   d. calcium bromide
   e. strontium oxide
   f. silver chloride [silver(I) chloride]
   g. cesium iodide
   h. lithium oxide

10. a. potassium chloride
    b. barium oxide
    c. rubidium sulfide
    d. sodium phosphide
    e. aluminum fluoride
    f. magnesium nitride
    g. calcium iodide
    h. radium chloride
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d.  44  K  
   19

e.  41  Ca  
   20

f.  35  K  
   19

107.  a.  22 protons, 19 neutrons, 22 electrons

b.  30 protons, 34 neutrons, 30 electrons

c.  32 protons, 44 neutrons, 32 electrons

d.  36 protons, 50 neutrons, 36 electrons

e.  33 protons, 42 neutrons, 33 electrons

f.  19 protons, 22 neutrons, 19 electrons

108.

<table>
<thead>
<tr>
<th>symbol</th>
<th>number of protons</th>
<th>number of neutrons</th>
<th>mass number</th>
</tr>
</thead>
<tbody>
<tr>
<td>a.  41 Ca</td>
<td>20</td>
<td>21</td>
<td>41</td>
</tr>
<tr>
<td>b.  55 Mn</td>
<td>25</td>
<td>30</td>
<td>55</td>
</tr>
<tr>
<td>c.  109 Ag</td>
<td>47</td>
<td>62</td>
<td>109</td>
</tr>
<tr>
<td>d.  45 Sc</td>
<td>21</td>
<td>24</td>
<td>45</td>
</tr>
</tbody>
</table>

109.  a.  C; Z = 6; nonmetal

b.  Se; Z = 34; nonmetal

c.  Rn; Z = 86; nonmetal; noble gases

d.  Be; Z = 4; metal; alkaline earth elements
b. zirconium
c. rubidium
d. radon
e. uranium
f. manganese
g. nickel
h. bromine

103. a. tellurium
    b. palladium
c. zinc
d. silicon
e. cesium
f. bismuth
g. fluorine
h. titanium

104. a. \( \text{CO}_2 \)
b. \( \text{AlCl}_3 \)
c. \( \text{HClO}_4 \)
d. \( \text{SCl}_6 \)

105. a. nitrogen, \( \text{N} \)
b. neon, \( \text{Ne} \)
c. sodium, \( \text{Na} \)
d. nickel, \( \text{Ni} \)
e. titanium, \( \text{Ti} \)
f. argon, \( \text{Ar} \)
g. krypton, \( \text{Kr} \)
h. xenon, \( \text{Xe} \)

106. a. \( \begin{array}{c} 13 \\ 6 \end{array} \text{C} \)
b. \( \begin{array}{c} 13 \\ 6 \end{array} \text{C} \)
c. \( \begin{array}{c} 13 \\ 6 \end{array} \text{C} \)
96. The chief use of gold in ancient times was as ornamentation, whether in statuary or in jewelry. Gold possesses an especially beautiful luster, and since it is relatively soft and malleable, it could be worked finely by artisans; among the metals, gold is particularly inert to attack by most substances in the environment.

97. Boyle defined a substance as an element if it could not be broken down into simpler substances by chemical means.

98. a. I
b. Si
c. W
d. Fe
e. Cu
f. Co

99. a. Ca
b. K
c. Cs
d. Pb
e. Pt
f. Au

100. a. Br
b. Bi
c. Hg
d. V
e. F
f. Ca

101. a. Ag
b. Al
c. Cd
d. Sb
e. Sn
f. As

102. a. osmium
total number of protons and neutrons happens to be the same for two atoms then the atoms will have the same mass number.

90. Most of the mass of an atom is concentrated in the nucleus: the protons and neutrons which constitute the nucleus have similar masses, and these particles are nearly two thousand times heavier than electrons. The chemical properties of an atom depend on the number and location of the electrons it possesses. Electrons are found in the outer regions of the atom, and are the particles most likely to be involved in interactions between atoms.

91. Yes. For example, carbon and oxygen form carbon monoxide (CO) and carbon dioxide (CO₂). The existence of more than one compound between the same elements does not in any way contradict Dalton's theory. For example, the relative mass of carbon in different samples of CO is always the same, and the relative mass of carbon in different samples of CO₂ is also always the same. Dalton did not say, however, that two different compounds would have to have the same relative masses of the elements present. In fact, Dalton said that two different compounds of the same elements would have different relative masses of the elements.

92. \( C_6H_{12}O_6 \)

93. FeO and Fe₂O₃

94. a. 29 protons; 34 neutrons; 29 electrons
   b. 35 protons; 45 neutrons; 35 electrons
   c. 12 protons; 12 neutrons; 12 electrons

95. \[
\begin{array}{ccc}
\text{mass number} & \text{symbol} & \text{number of neutrons} \\
24 & Al & 11 \\
25 & Al & 12 \\
26 & Al & 13 \\
28 & Al & 15 \\
29 & Al & 16 \\
30 & Al & 17 \\
\end{array}
\]

These are all considered to be aluminum atoms because they all have 13 protons in the nucleus.
87. Group 1

<table>
<thead>
<tr>
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<tbody>
<tr>
<td>hydrogen</td>
<td>H</td>
<td>1</td>
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<tr>
<td>lithium</td>
<td>Li</td>
<td>3</td>
</tr>
<tr>
<td>sodium</td>
<td>Na</td>
<td>11</td>
</tr>
<tr>
<td>potassium</td>
<td>K</td>
<td>19</td>
</tr>
<tr>
<td>beryllium</td>
<td>Be</td>
<td>4</td>
</tr>
<tr>
<td>magnesium</td>
<td>Mg</td>
<td>12</td>
</tr>
<tr>
<td>calcium</td>
<td>Ca</td>
<td>20</td>
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<tr>
<td>strontium</td>
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Group 2

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<td>sulfur</td>
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</tr>
<tr>
<td>selenium</td>
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<tr>
<td>tellurium</td>
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Group 6

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<tr>
<td>chlorine</td>
<td>Cl</td>
<td>17</td>
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<tr>
<td>bromine</td>
<td>Br</td>
<td>35</td>
</tr>
<tr>
<td>iodine</td>
<td>I</td>
<td>53</td>
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Group 7

88. Group 3

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<tr>
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<tbody>
<tr>
<td>boron</td>
<td>B</td>
<td>5</td>
</tr>
<tr>
<td>aluminum</td>
<td>Al</td>
<td>13</td>
</tr>
<tr>
<td>gallium</td>
<td>Ga</td>
<td>31</td>
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<tr>
<td>indium</td>
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Group 5

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<tr>
<td>nitrogen</td>
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<tr>
<td>phosphorus</td>
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<tr>
<td>arsenic</td>
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<tr>
<td>antimony</td>
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Group 8

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<tr>
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<tr>
<td>argon</td>
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<td>18</td>
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<tr>
<td>krypton</td>
<td>Kr</td>
<td>36</td>
</tr>
</tbody>
</table>

89. The atomic number represents the number of protons in the nucleus of an atom. The mass number represents the total number of protons and neutrons. No two different elements have the same atomic number. If the
Chapter 4  Chemical Foundations: Elements, Atoms, and Ions  

e. One 2- ion balances one 2+ ion: MgO  
f. The smallest common multiple of 2 and 3 is 6; two 3- ions are required to balance three 2+ ions: \( \text{Mg}_2\text{N}_2 \)  
g. Three 1+ ions are required to balance one 3- ion: \( \text{Na}_3\text{P} \)  
h. Two 1+ ions are required to balance one 2- ion: \( \text{Na}_2\text{S} \)  

84. a. One 2+ ion is exactly balanced by one 2- ion: \( \text{BaO} \)  
b. Three 1+ ions are needed to balance one 3- ion: \( \text{K}_3\text{P} \)  
c. Two 2+ ions are required to balance one 4- ion: \( \text{Ca}_2\text{C} \)  
d. The smallest common multiple of 3 and 2 is 6; two 3+ ions are balanced by three 2- ions: \( \text{Al}_2\text{S}_3 \)  
e. The smallest common multiple of 2 and 3 is 6; three 2+ ions are balanced by two 3- ions: \( \text{Sr}_3\text{P}_2 \)  
f. One 1+ ion is balanced by one 1- ion: \( \text{NaI} \)  
g. One 3+ ion is balanced by three 1- ions: \( \text{CoCl}_3 \)  
h. One 4+ ion is balanced by four 1- ions: \( \text{SnBr}_4 \)  

85. a. At; \( Z = 85 \)  
b. Xe; \( Z = 54 \)  
c. Ra; \( Z = 88 \)  
d. Sr; \( Z = 38 \)  
e. Pb; \( Z = 82 \)  
f. Se; \( Z = 34 \)  
g. Ar; \( Z = 18 \)  
h. Cs; \( Z = 55 \)  

86. a. 7; halogens  
b. 8; noble gases  
c. 2; alkaline earth elements  
d. 2; alkaline earth elements  
e. 4  
f. 6; (the members of group 6 are sometimes called the chalcogens)  
g. 8; noble gases  
h. 1; alkali metals
77. a. I⁻
   b. Sr²⁺
   c. Cs⁺
   d. Ra²⁺
   e. F⁻
   f. Al³⁺

78. a. Ra²⁺ (element 88, Ra, is in Group 2)
   b. Te²⁻ (element 52, Te, is in Group 6)
   c. I⁻ (element 53, I, is in Group 7)
   d. Fr⁺ (element 87, Fr, is in Group 1)
   e. At⁻ (element 85, At, is in Group 7)
   f. no ion is likely (element 86, Rn, is a noble gas)

79. A compound which has a high melting point (many hundreds of degrees) and which conducts an electrical current when melted or dissolved in water almost certainly consists of ions.

80. Sodium chloride is an ionic compound, consisting of Na⁺ and Cl⁻ ions. When NaCl is dissolved in water, these ions are set free, and can move independently to conduct the electrical current. Sugar crystals, although they may visually appear similar contain no ions. When sugar is dissolved in water, it dissolves as uncharged molecules. There are no electrically charged species present in a sugar solution to carry the electrical current.

81. In the solid state, although ions are present, they are rigidly held in fixed positions in the crystal of the substance. In order for ionic substances to be able to pass an electrical current, the ions must be able to move, which is possible when the solid is converted to the liquid state.

82. The total number of positive charges must equal the total number of negative charges so that there will be no net charge on the crystals of an ionic compound. A macroscopic sample of compound must ordinarily not have any net charge.

83. a. One 3⁻ ion is needed to balance one 3⁺ ion: FeP
   b. The smallest common multiple of 3 and 2 is 6; three 2⁻ ions are required to balance two 3⁺ ions: Fe₂S₃
   c. Three 1⁻ ions are required to balance one 3⁺ ion: FeCl₃
   d. Two 1⁻ ions are required to balance one 2⁺ ion: MgCl₂
71. metallic
72. nonmetallic
73. [1] b
    [2] d
    [3] b
    [4] h
    [5] f
    [7] a
    [8] c
    [9] g.
    [10] i
74. a. Co^{2+}: 27 protons, 25 electrons  CoO
    b. Co^{3+}: 27 protons, 24 electrons  Co_2O_3
    c. Cl^-: 17 protons, 18 electrons  CaCl_2
    d. K^+: 19 protons, 18 electrons  K_2O
    e. S^{2-}: 16 protons, 18 electrons  CaS
    f. Sr^{2+}: 38 protons, 36 electrons  SrO
    g. Al^{3+}: 13 protons, 10 electrons  Al_2O_3
    h. P^{3-}: 15 protons, 18 electrons  Ca_3P_2
75. a. Ca: 20 protons, 20 electrons  Ca^{2+}: 20 protons, 18 electrons
    b. P: 15 protons, 15 electrons  P^{3-}: 15 protons, 18 electrons
    c. Br: 35 protons, 35 electrons  Br^-: 35 protons, 36 electrons
    d. Fe: 26 protons, 26 electrons  Fe^{3+}: 26 protons, 23 electrons
    e. Al: 13 protons, 13 electrons  Al^{3+}: 13 protons, 10 electrons
    f. N: 7 protons, 7 electrons  N^{3-}: 7 protons, 10 electrons
76. a. 2e^{-}
    b. 2e^{-}
    c. 2e^{-}
    d. 3e^{-}
    e. 1e^{-}
    f. 3e^{-}
54. name | symbol | atomic number | group number | metal/nonmetal
---|---|---|---|---
rubidium | Rb | 37 | 1 | metal
germanium | Ge | 32 | 4 | metalloid
magnesium | Mg | 12 | 2 | metal
titanium | Ti | 22 | - | transition metal
iodine | I | 53 | 7 | nonmetal

55. compounds (and mixtures of compounds)

56. Most of the elements are too reactive to be found in the uncombined form in nature, and are found only in compounds.

57. argon

58. These elements are found uncombined in nature and do not readily react with other elements. For many years it was thought that these elements formed no compounds at all, although this has now been shown to be untrue.

59. diatomic

60. Diatomic gases: H₂, N₂, O₂, Cl₂, and F₂
Monatomic gases: He, Ne, Kr, Xe, Rn, and Ar

61. electricity

62. chlorine

63. liquids: bromine, mercury, gallium
gases: hydrogen, nitrogen, oxygen, fluorine, chlorine, and the noble gases (helium, neon, argon, krypton, xenon, radon)

64. diamond

65. zero

66. electrons

67. 1+

68. 2+

69. cations, anions

70. -ide
48. hydrogen, nitrogen, oxygen, fluorine, chlorine, plus all the group 8 elements (noble gases)

49. The only metal which ordinarily occurs as a liquid is mercury. The only nonmetallic element which occurs as a liquid at room temperature is bromine (elements such as oxygen and nitrogen are frequently obtainable as liquids, but these result from compression of the gases into cylinders at very low temperatures).

50. The metalloids are the elements found on either side of the "stairstep" region that is marked on most periodic tables. The metalloid elements show some properties of both metals and nonmetals.

51. a. Group 1; alkali metals
b. Group 2; alkaline earth elements
c. Group 8; noble gases
d. Group 7; halogens
e. Group 2; alkaline earth elements
f. Group 8; noble gases
g. Group 1; alkali metals

52. a. Group 7; halogens
b. Group 2; alkaline earth elements
c. Group 1; alkali metals
d. Group 1; alkali metals
e. Group 8; noble gases
f. Group 1; alkali metals
g. Group 8; noble gases

53. a. Sr; Z = 38; Group 2; metal
b. I; Z = 53; Group 7; nonmetal
c. Si; Z = 14; Group 4; metalloid
d. Cs; Z = 55; Group 1; metal
e. S; Z = 16; Group 6; nonmetal
### Chapter 4: Chemical Foundations: Elements, Atoms, and Ions

#### 41. Elements and Their Properties

<table>
<thead>
<tr>
<th>Element</th>
<th>Symbol</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Number of Neutrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium</td>
<td>Na</td>
<td>11</td>
<td>23</td>
<td>12</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N</td>
<td>7</td>
<td>15</td>
<td>8</td>
</tr>
<tr>
<td>Barium</td>
<td>Ba</td>
<td>56</td>
<td>136</td>
<td>80</td>
</tr>
<tr>
<td>Lithium</td>
<td>Li</td>
<td>3</td>
<td>9</td>
<td>6</td>
</tr>
<tr>
<td>Boron</td>
<td>B</td>
<td>5</td>
<td>11</td>
<td>6</td>
</tr>
</tbody>
</table>

#### 42. Elements and Their Properties (Continued)

<table>
<thead>
<tr>
<th>Element</th>
<th>Neutrons</th>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Nitrogen</td>
<td>6</td>
<td>7</td>
<td>13</td>
<td>N</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>7</td>
<td>7</td>
<td>14</td>
<td>N</td>
</tr>
<tr>
<td>Lead</td>
<td>124</td>
<td>82</td>
<td>206</td>
<td>Pb</td>
</tr>
<tr>
<td>Iron</td>
<td>31</td>
<td>26</td>
<td>57</td>
<td>Fe</td>
</tr>
<tr>
<td>Krypton</td>
<td>48</td>
<td>36</td>
<td>84</td>
<td>Kr</td>
</tr>
</tbody>
</table>

The elements are listed in the periodic table in order of increasing atomic number (number of protons in the nucleus). The periodic table originally was arranged on the basis of mass.

43. Elements with similar chemical properties are aligned vertically in families known as *groups*.

44. Metals are excellent conductors of heat and electricity, and are malleable, ductile, and generally shiny (lustrous) when a fresh surface is exposed.

45. Metallic elements are found towards the left and bottom of the periodic table; there are far more metallic elements than there are nonmetals.

46. Mercury is a liquid at room temperature.
38. a. $^1_{12}\text{B}$
   b. $^1_{15}\text{N}$
   c. $^1_{35}\text{Cl}$
   d. $^1_{235}\text{U}$
   e. $^1_{14}\text{C}$
   f. $^1_{31}\text{P}$

39. a. 94 protons, 150 neutrons, 94 electrons
   b. 95 protons, 146 neutrons, 95 electrons
   c. 89 protons, 138 neutrons, 89 electrons
   d. 55 protons, 78 neutrons, 55 electrons
   e. 77 protons, 116 neutrons, 77 electrons
   f. 25 protons, 31 neutrons, 25 electrons

40. a. 6 protons, 6 neutrons, 6 electrons
   b. 27 protons, 33 neutrons, 27 electrons
   c. 17 protons, 20 neutrons, 17 electrons
   d. 55 protons, 77 neutrons, 55 electrons
   e. 92 protons, 146 neutrons, 92 electrons
   f. 26 protons, 30 neutrons, 26 electrons
False; atoms that have the same number of protons, with different numbers of neutrons, represent isotopes.

False; the mass number represents the total number of protons and neutrons in the nucleus.

the same

mass

Dalton's original theory proposed that all atoms of a given element were identical. We now realize that different atoms of the same element must have a particular number of protons and electrons (the atomic number), but may have different numbers of neutrons (leading to different mass numbers).

Atoms of the same element (i.e., atoms with the same number of protons in the nucleus) may have different numbers of neutrons, and so will have different masses.

35. a. 24  
b. 13  
c. 34  
d. 18  
e. 55  
f. 7  
g. 26  
h. 17

36. a. 32  
b. 30  
c. 24  
d. 74  
e. 38  
f. 27  
g. 4  
h. 3

37. a. 17  
b. 37

0  
8  
Cl  
17
19. a. PCl₃  
b. B₂H₆  
c. CaCl₂  
d. CBr₄  
e. Fe₂O₃  
f. H₃PO₄  

20. a. C₆H₆  
b. N₂O₄  
c. CaCl₂  
d. FeBr₃  
e. NaNO₃  
f. Ca₃N₂  

21. a. False; Thomson obtained beams of identical particles whose nature did not depend on what gas was used to generate them.  
b. True  
c. False; the atom was envisioned as a sphere of positive charge in which negatively charged electrons were randomly distributed.  

22. a. False; Rutherford's bombardment experiments with metal foil suggested that the alpha particles were being deflected by coming near a dense, positively charged atomic nucleus.  
b. False; The proton and the electron have opposite charges, but the mass of the electron is much smaller than the mass of the proton.  
c. True  

23. neutrons  
24. protons  

25. The proton and the neutron have similar (but not identical) masses. Either of these particles has a mass approximately 2000 times greater than that of an electron. The combination of the protons and the neutrons make up the bulk of the mass of an atom, but the electrons make the greatest contribution to the chemical properties of the atom.  

26. neutron; electron  
27. 10⁻¹³ cm = 10⁻¹⁵ m  
28. electrons
13. a. potassium  
    b. germanium  
    c. phosphorus  
    d. carbon  
    e. nitrogen  
    f. sodium  
    g. neon  
    h. iodine

14. a. copper  
    b. cobalt  
    c. calcium  
    d. carbon  
    e. chromium  
    f. cesium  
    g. chlorine  
    h. cadmium

15. a. False; most materials occur as mixtures of compounds.  
    b. False; a given compound always contains the same relative number of atoms of its various elements.  
    c. False; molecules are made up of tiny particles called atoms.

16. According to Dalton, a given compound is always made up of the same number and type of atoms, and so the composition of the compound on a mass percentage basis will always be the same, no matter what the source of the compound is. For example, water molecules from the Atlantic Ocean and the Pacific Ocean all contain two hydrogen atoms bonded to one oxygen atom.

17. A compound is a distinct substance that is composed of two or more elements and always contains exactly the same relative masses of those elements.

18. According to Dalton, all atoms of the same element are identical; in particular, every atom of a given element has the same mass as every other atom of that element. If a given compound always contains the same relative numbers of atoms of each kind, and those atoms always have the same masses, then it follows that the compound made from those elements would always contain the same relative masses of its elements.
Au (gold)
Pb (lead)
Hg (mercury)
K (potassium)
Ag (silver)
Na (sodium)
Sn (tin)
W (tungsten)
Fe (iron)

9. a. Ne
   b. Ni
   c. K
   d. Si
   e. Ba
   f. Ag

10. a. Al
    b. Fe
    c. F
    d. Ca
    e. Au
    f. Hg

11. Symbol  Name
    Fe     iron
    Cl     chlorine
    S      sulfur
    U      uranium
    Ne     neon
    K      potassium

12. Ir     iridium
    Ta     tantalum
    Bi     bismuth
    Pu     plutonium
    Fr     francium
    At     astatine