Solution Manual for Principles of General Chemistry 3rd Edition Silberberg 0073402699 9780073402697

CHAPTER 2 THE COMPONENTS OF MATTER

END–OF–CHAPTER PROBLEMS

2.1 Plan: Refer to the definitions of an element and a compound.

Solution:

Unlike compounds, elements cannot be broken down by chemical changes into simpler materials. Compounds contain different types of atoms; there is only one type of atom in an element.

2.2 Plan: Refer to the definitions of a compound and a mixture.

Solution:

1) A compound has constant composition but a mixture has variable composition. 2) A compound has distinctly different properties than its component elements; the components in a mixture retain their individual properties.

2.3 Plan: Recall that a substance has a fixed composition.

Solution:

a) The fixed mass ratio means it has constant composition, thus, it is a **pure substance** (compound).
b) All the atoms are identical, thus, it is a **pure substance** (element).
c) The composition can vary, thus, this is an **impure substance** (a mixture).
d) The specific arrangement of different atoms means it has constant composition, thus, it is a **pure substance** (compound).

2.4 Plan: Remember that an element contains only one kind of atom while a compound contains at least two different elements (two kinds of atoms) in a fixed ratio. A mixture contains at least two different substances in a composition that can vary.

Solution:

a) The presence of more than one element (calcium and chlorine) makes this pure substance a **compound**.
b) There are only atoms from one element, sulfur, so this pure substance is an **element**.
c) This is a combination of two compounds and has a varying composition, so this is a **mixture**.
d) The presence of more than one type of atom means it cannot be an element. The specific, not variable, arrangement means it is a **compound**.

2.5 Plan: Recall that an element contains only one kind of atom; the atoms in an element may occur as molecules. A compound contains two kinds of atoms (different elements).

Solution:

a) This scene has 3 atoms of an element, 2 molecules of one compound (with one atom each of two different elements), and 2 molecules of a second compound (with 2 atoms of one element and one atom of a second element).

b) This scene has 2 atoms of one element, 2 molecules of a diatomic element, and 2 molecules of a compound (with one atom each of two different elements).

c) This scene has 2 molecules composed of 3 atoms of one element and 3 diatomic molecules of the same element.

2.6 Plan: Restate the three laws in your own words.

Solution:

a) The law of mass conservation applies to all substances — **elements, compounds, and mixtures**. Matter can neither be created nor destroyed, whether it is an element, compound, or mixture.

b) The law of definite composition applies to **compounds** only, because it refers to a constant, or definite, composition of elements within a compound.

c) The law of multiple proportions applies to **compounds** only, because it refers to the combination of elements to form compounds.

2.7 Plan: Review the three laws: law of mass conservation, law of definite composition, and law of multiple proportions.

Solution:

a) **Law of Definite Composition** — The compound potassium chloride, KCl, is composed of the same elements and same fraction by mass, regardless of its source (Chile or Poland).

b) **Law of Mass Conservation** — The mass of the substances inside the flashbulb did not change during the chemical reaction (formation of magnesium oxide from magnesium and oxygen).

c) **Law of Multiple Proportions** — Two elements, O and As, can combine to form two different compounds that have different proportions of As present.

2.8 Plan: The law of multiple proportions states that two elements can form two different compounds in which the proportions of the elements are different.

Solution:

Scene B illustrates the law of multiple proportions for compounds of chlorine and oxygen. The law of multiple proportions refers to the different compounds that two elements can form that have different proportions of the elements. Scene B shows that chlorine and oxygen can form both Cl_2O, dichlorine monoxide, and ClO_2, chlorine dioxide.

2.9 Plan: Review the definition of percent by mass.

Solution:

a) **No**, the mass percent of each element in a compound is fixed. The percentage of Na in the compound NaCl is 39.34% (22.99 amu/58.44 amu), whether the sample is 0.5000 g or 50.00 g.

b) **Yes**, the mass of each element in a compound depends on the mass of the compound. A 0.5000 g sample of NaCl contains 0.1967 g of Na (39.34% of 0.5000 g), whereas a 50.00 g sample of NaCl contains 19.67 g of Na (39.34% of 50.00 g).

2.10 Generally no, the composition of a compound is determined by the elements used, not their amounts. If too much of one element is used, the excess will remain as unreacted element when the reaction is over.

2.11 Plan: Review the mass laws: law of mass conservation, law of definite composition, and law of multiple proportions. For each experiment, compare the mass values before and after each reaction and examine the ratios of the mass of white compound to the mass of colorless gas.

Solution:

Experiment 1: mass before reaction = 1.00 g; mass after reaction = 0.64 g + 0.36 g = 1.00 g

Experiment 2: mass before reaction = 3.25 g; mass after reaction = 2.08 g + 1.17 g = 3.25 g

Both experiments demonstrate the **law of mass conservation** since the total mass before reaction equals the total mass after reaction.

Experiment 1: mass white compound/mass colorless gas = 0.64 g/0.36 g = 1.78

Experiment 2: mass white compound/mass colorless gas = 2.08 g/1.17 g = 1.78

Both Experiments 1 and 2 demonstrate the **law of definite composition** since the compound has the same composition by mass in each experiment.

2.12 Plan: Review the mass laws: law of mass conservation, law of definite composition, and law of multiple proportions. For each experiment, compare the mass values before and after each reaction and examine the ratios of the mass of reacted copper to the mass of reacted iodine.

Solution:

Experiment 1: mass before reaction = 1.27 g + 3.50 g = 4.77 g; mass after reaction = 3.81 g + 0.96 g = 4.77 g

Experiment 2: mass before reaction = 2.55 g + 3.50 g = 6.05 g; mass after reaction = 5.25 g + 0.80 g = 6.05 g

Both experiments demonstrate the **law of mass conversation** since the total mass before reaction equals the total mass after reaction.

Experiment 1: mass of reacted copper = 1.27 g; mass of reacted iodine = 3.50 g – 0.96 g = 2.54 g

Mass reacted copper/mass reacted iodine = 1.27 g/2.54 g = 0.50

Experiment 2: mass of reacted copper = 2.55 g – 0.80 g = 1.75 g; mass of reacted iodine = 3.50 g

Mass reacted copper/mass reacted iodine = 1.75 g/3.50 g = 0.50

Both Experiments 1 and 2 demonstrate the **law of definite composition** since the compound has the same composition by mass in each experiment.

2.13 Plan: Fluorite is a mineral containing only calcium and fluorine. The difference between the mass of fluorite and the mass of calcium gives the mass of fluorine. Mass fraction is calculated by dividing the mass of element by the mass of compound (fluorite) and mass percent is obtained by multiplying the mass fraction by 100.

Solution:

a) Mass (g) of fluorine = mass of fluorite – mass of calcium = 2.76 g – 1.42 g = **1.34 g fluorine**

b) Mass fraction of Ca = $\frac{\text{mass Ca}}{\text{mass fluorite}} = \frac{1.42 \text{ g Ca}}{2.76 \text{ g fluorite}}$ = 0.51449 = **0.514**

Mass fraction of F = $\frac{\text{mass F}}{\text{mass fluorite}} = \frac{1.34 \text{ g F}}{2.76 \text{ g fluorite}}$ = 0.48551 = **0.486**

c) Mass percent of Ca = 0.51449 x 100 = 51.449 = **51.4%** Mass percent of F = 0.48551 x 100 = 48.551 = **48.6%**

2.14 Plan: Galena is a mineral containing only lead and sulfur. The difference between the mass of galena and the mass of lead gives the mass of sulfur. Mass fraction is calculated by dividing the mass of element by the mass of compound (galena) and mass percent is obtained by multiplying the mass fraction by 100.

Solution:

a) Mass (g) of sulfur = mass of galena – mass of sulfur = 2.34 g – 2.03 g = **0.31 g sulfur**

b) Mass fraction of Pb = $\frac{\text{mass Pb}}{\text{mass galena}} = \frac{2.03 \text{ g Pb}}{2.34 \text{ g galena}}$ = 0.8675214 = **0.868**

$$\text{Mass fraction of S} = \frac{\text{mass S}}{\text{mass galena}} = \frac{0.31\text{ g S}}{2.34\text{ g galena}} = 0.1324786 = \mathbf{0.13}$$

c) Mass percent of Pb = (0.8675214)(100) = 86.752 = **86.8%** Mass percent of S = (0.1324786)(100) = 13.248 = **13%**

2.15 Plan: Since copper is a metal and sulfur is a nonmetal, the sample contains 88.39 g Cu and 44.61 g S. Calculate the mass fraction of each element in the sample by dividing the mass of element by the total mass of compound. Multiply the mass of the second sample of compound in grams by the mass fraction of each element to find the mass of each element in that sample.

Solution:

Mass (g) of compound = 88.39 g copper + 44.61 g sulfur = 133.00 g compound

$$\text{Mass fraction of copper} = \left(\frac{88.39\text{ g copper}}{133.00\text{ g compound}}\right) = 0.664586$$

$$\text{Mass (g) of copper} = (5264\text{ kg compound})\left(\frac{10^3\text{ g compound}}{1\text{ kg compound}}\right)\left(\frac{0.664586\text{ g copper}}{1\text{ g compound}}\right)$$

$$= 3.49838 \times 10^6 = \mathbf{3.498 \times 10^6\ g\ copper}$$

$$\text{Mass fraction of sulfur} = \left(\frac{44.61\text{ g sulfur}}{133.00\text{ g compound}}\right) = 0.335414$$

$$\text{Mass (g) of sulfur} = (5264\text{ kg compound})\left(\frac{10^3\text{ g compound}}{1\text{ kg compound}}\right)\left(\frac{0.335414\text{ g sulfur}}{1\text{ g compound}}\right)$$

$$= 1.76562 \times 10^6 = \mathbf{1.766 \times 10^6\ g\ sulfur}$$

2.16 Plan: Since cesium is a metal and iodine is a nonmetal, the sample contains 63.94 g Cs and 61.06 g I. Calculate the mass fraction of each element in the sample by dividing the mass of element by the total mass of compound. Multiply the mass of the second sample of compound by the mass fraction of each element to find the mass of each element in that sample.

Solution:

Mass of compound = 63.94 g cesium + 61.06 g iodine = 125.00 g compound

$$\text{Mass fraction of cesium} = \left(\frac{63.94\text{ g cesium}}{125.00\text{ g compound}}\right) = 0.51152$$

$$\text{Mass (g) of cesium} = (38.77\text{ g compound})\left(\frac{0.51152\text{ g cesium}}{1\text{ g compound}}\right) = 19.83163 = \mathbf{19.83\ g\ cesium}$$

$$\text{Mass fraction of iodine} = \left(\frac{61.06\text{ g iodine}}{\quad}\right) = 0.48848$$

$$\text{Mass (g) of iodine} = (38.77 \text{ g compound})\left(\frac{0.48848 \text{ g iodine}}{1 \text{ g compound}}\right) = 18.9384 = \mathbf{18.94 \text{ g iodine}}$$

125.00 g compound

2.17 Plan: The law of multiple proportions states that if two elements form two different compounds, the relative amounts of the elements in the two compounds form a whole-number ratio. To illustrate the law we must calculate the mass of one element to one gram of the other element for each compound and then compare this mass for the two compounds. The law states that the ratio of the two masses should be a small whole-number ratio such as 1:2, 3:2, 4:3, etc.

Solution:

Compound 1: $\frac{47.5 \text{ mass \% S}}{52.5 \text{ mass \% Cl}} = 0.90476 = 0.905$

Compound 2: $\frac{31.1 \text{ mass \% S}}{68.9 \text{ mass \% Cl}} = 0.451379 = 0.451$

Ratio: $\frac{0.905}{0.451} = 2.0067 = 2.00:1.00$

Thus, the ratio of the mass of sulfur per gram of chlorine in the two compounds is a small whole-number ratio of 2:1, which agrees with the law of multiple proportions.

2.18 Plan: The law of multiple proportions states that if two elements form two different compounds, the relative amounts of the elements in the two compounds form a whole-number ratio. To illustrate the law we must calculate the mass of one element to one gram of the other element for each compound and then compare this mass for the two compounds. The law states that the ratio of the two masses should be a small whole-number ratio such as 1:2, 3:2, 4:3, etc.

Solution:

Compound 1: $\frac{77.6 \text{ mass \% Xe}}{22.4 \text{ mass \% F}} = 3.4643 = 3.46$

Compound 2: $\frac{63.3 \text{ mass \% Xe}}{36.7 \text{ mass \% F}} = 1.7248 = 1.72$

Ratio: $\frac{3.46}{1.72} = 2.0116 = 2.01:1.00$

Thus, the ratio of the mass of xenon per gram of fluorine in the two compounds is a small whole-number ratio of 2:1, which agrees with the law of multiple proportions.

2.19 Plan: Calculate the mass percent of calcium in dolomite by dividing the mass of calcium by the mass of the sample and multiply by 100. Compare this mass percent to that in fluorite. The compound with the larger mass percent of calcium is the richer source of calcium. Solution:

$$\text{Mass percent calcium} = \frac{1.70 \text{ g calcium}}{7.81 \text{ g dolomite}} \times 100\% = 21.767 = \mathbf{21.8\% \text{ Ca}}$$

Fluorite (51.4%) is the richer source of calcium.

2.20 Plan: Determine the mass percent of sulfur in each sample by dividing the grams of sulfur in the sample by the total mass of the sample and multiplying by 100. The coal type with the smallest mass percent of sulfur has the smallest environmental impact.

Solution:

$$\text{Mass \% in Coal A} = \left(\frac{11.3 \text{ g sulfur}}{378 \text{ g sample}}\right)(100\%) = 2.9894 = 2.99\% \text{ S (by mass)}$$

$$\text{Mass \% in Coal B} = \left(\frac{19.0 \text{ g sulfur}}{495 \text{ g sample}}\right)(100\%) = 3.8384 = 3.84\% \text{ S (by mass)}$$

$$\text{Mass \% in Coal C} = \left(\frac{20.6 \text{ g sulfur}}{675 \text{ g sample}}\right)(100\%) = 3.0519 = 3.05\% \text{ S (by mass)}$$

Coal A has the smallest environmental impact.

2.21 Plan: This question is based on the law of definite composition. If the compound contains the same types of atoms, they should combine in the same way to give the same mass percentages of each of the elements.

Solution:

Potassium nitrate is a compound composed of three elements — potassium, nitrogen, and oxygen — in a specific ratio. If the ratio of these elements changed, then the compound would be changed to a different compound, for example, to potassium nitrite, with different physical and chemical properties. Dalton postulated that atoms of an element are identical, regardless of whether that element is found in India or Italy. Dalton also postulated that compounds result from the chemical combination of specific ratios of different elements. Thus, Dalton's theory explains why potassium nitrate, a compound comprised of three different elements in a specific ratio, has the same chemical composition regardless of where it is mined or how it is synthesized.

2.22 Plan: Review the discussion of the experiments in this chapter.

Solution:

Millikan determined the minimum *charge* on an oil drop and that the minimum charge was equal to the charge on one electron. Using Thomson's value for the *mass/charge ratio* of the electron and the determined value for the charge on one electron, Millikan calculated the mass of an electron (charge/(charge/mass)) to be $9.109\text{x}10^{-28}$ g.

2.23 Plan: The charges on the oil droplets should be whole-number multiples of a minimum charge. Determine that minimum charge by dividing the charges by small integers to find the common factor. Solution:

$-3.204\text{x}10^{-19}\ \text{C}/2 = -1.602\text{x}10^{-19}\ \text{C}$

$-4.806\text{x}10^{-19}\ \text{C}/3 = -1.602\text{x}10^{-19}\ \text{C}$

$-8.010\text{x}10^{-19}\ \text{C}/5 = -1.602\text{x}10^{-19}\ \text{C}$

$-1.442\text{x}10^{-18}\ \text{C}/4 = -1.602\text{x}10^{-19}\ \text{C}$

The value $\mathbf{-1.602\text{x}10^{-19}\ C}$ is the common factor and is the charge for the electron.

2.24 Rutherford and co-workers expected that the alpha particles would pass through the foil essentially unaffected, or perhaps slightly deflected or slowed down. The observed results (most passing through straight, a few deflected, a very few at large angles) were partially consistent with expectations, but the large-angle scattering could not be explained by Thomson's model. The change was that Rutherford envisioned a small (but massive) positively charged nucleus in the atom, capable of deflecting the alpha particles as observed.

2.25 Plan: Recall that the mass number is the sum of protons and neutrons while the atomic number is the number of protons.

Solution:

Mass number (protons plus neutrons) – atomic number (protons) = **number of neutrons (c)**.

2.26 Plan: The superscript is the mass number, the sum of the number of protons and neutrons. Consult the periodic table to get the atomic number (the number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons and electrons are equal.

Solution:

Isotope	Mass Number	# of Protons	# of Neutrons	# of Electrons
^{36}Ar	36	18	18	18
^{38}Ar	38	18	20	18
^{40}Ar	40	18	22	18

2.27 Plan: The superscript is the mass number, the sum of the number of protons and neutrons. Consult the periodic table to get the atomic number (the number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons and electrons are equal. Solution:

Isotope	Mass Number	# of Protons	# of Neutrons	# of Electrons
^{35}Cl	35	17	18	17
^{37}Cl	37	17	20	17

2.28 Plan: The superscript is the mass number (*A*), the sum of the number of protons and neutrons; the subscript is the atomic number (*Z*, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons.

Solution:

a)$^{16}_{8}O$and $^{17}_{8}O$ have the **same number of protons and electrons** (8), but different numbers of neutrons.

$^{16}_{8}O$and $^{17}_{8}O$ are isotopes of oxygen, and $^{16}_{8}O$ has 16 – 8 = 8 neutrons whereas $^{17}_{8}O$ has 17 – 8 = 9 neutrons.

Same *Z* value

b) $^{40}_{18}Ar$ and $_{19}{}^{41}K$ have the **same number of neutrons** (Ar: 40 – 18 = 22; K: 41 – 19 = 22) but different numbers of protons and electrons (Ar = 18 protons and 18 electrons; K = 19 protons and 19 electrons). **Same *N* value**

c) $_{27}{}^{60}Co$ and $^{60}_{28}Ni$ have different numbers of protons, neutrons, and electrons. Co: 27 protons, 27 electrons, and 60 – 27 = 33 neutrons; Ni: 28 protons, 28 electrons and 60 – 28 = 32 neutrons. However, both have a mass number of 60. **Same *A* value**

2.29 Plan: The superscript is the mass number (*A*), the sum of the number of protons and neutrons; the subscript is the atomic number (*Z*, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons.

Solution:

a))$^{3}_{1}H$ and $_{2}{}^{3}He$ have different numbers of protons, neutrons, and electrons. H: 1 proton, 1 electron, and 3 – 1 = 2 neutrons; He: 2 protons, 2 electrons, and 3 – 2 = 1 neutron. However, both have a mass number of 3.

Same *A* value

b) $^{14}_{6}C$ and $^{15}_{7}N$ have the **same number of neutrons** (C: 14 – 6 = 8; N: 15 – 7 = 8) but different numbers of protons and electrons (C = 6 protons and 6 electrons; N = 7 protons and 7 electrons). **Same *N* value**

c) $^{19}_{9}F$ and $^{18}_{9}F$ have the **same number of protons and electrons** (9), but different numbers of neutrons. $^{19}_{9}F$ and $^{18}_{9}F$ are isotopes of oxygen, and $^{19}_{9}F$ has 19 – 9 = 10 neutrons whereas $^{18}_{9}F$ has 18 – 9 = 9 neutrons.
Same *Z* value

2.30 Plan: Combine the particles in the nucleus (protons + neutrons) to give the mass number (superscript, *A*). The number of protons gives the atomic number (subscript, *Z*) and identifies the element.
Solution:

a) A = 18 + 20 = 38; Z = 18; $\mathbf{^{38}_{18}Ar}$

b) A = 25 + 30 = 55; Z = 25; $\mathbf{^{55}_{25}Mn}$

c) A = 47 + 62 = 109; Z = 47; $\mathbf{^{109}_{47}Ag}$

2.31 Plan: Combine the particles in the nucleus (protons + neutrons) to give the mass number (superscript, *A*). The number of protons gives the atomic number (subscript, *Z*) and identifies the element.
Solution:

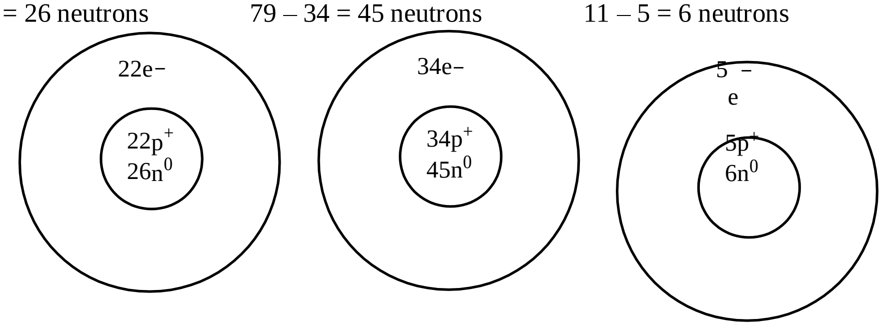
a) A = 6 + 7 = 13; Z = 6; $\mathbf{^{13}_{6}C}$

b) A = 40 + 50 = 90; Z = 40; $\mathbf{^{90}_{40}Zr}$

c) A = 28 + 33 = 61; Z = 28; $\mathbf{_{28}^{61}Ni}$

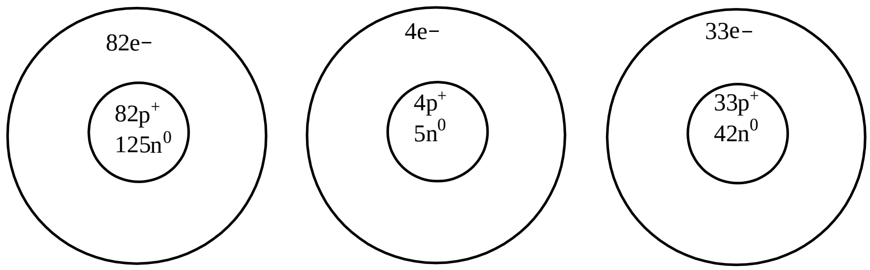
2.32 Plan: Determine the number of each type of particle. The superscript is the mass number (*A*) and the subscript is the atomic number (*Z*, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons. The protons and neutrons are in the nucleus of the atom.
Solution:

a) $^{48}_{22}Ti$	b) $^{79}_{34}Se$	c) $^{11}_{5}B$
22 protons	34 protons	5 protons
22 electrons	34 electrons	5 electrons
48 – 22 = 26 neutrons	79 – 34 = 45 neutrons	11 – 5 = 6 neutrons



2.33 Plan: Determine the number of each type of particle. The superscript is the mass number (*A*) and the subscript is the atomic number (*Z*, number of protons). The mass number – the number of protons = the number of neutrons. For atoms, the number of protons = the number of electrons. The protons and neutrons are in the nucleus of the atom.
Solution:

a) $^{207}_{82}Pb$	b) $^{9}_{4}Be$	c) $^{75}_{33}As$
82 protons	4 protons	33 protons
82 electrons	4 electrons	33 electrons
207 – 82 = 125 neutrons	9 – 4 = 5 neutrons	75 – 33 = 42 neutrons



2.34 Plan: To calculate the atomic mass of an element, take a weighted average based on the natural abundance of the isotopes: (isotopic mass of isotope 1 x fractional abundance) + (isotopic mass of isotope 2 x fractional abundance).

Solution:

Atomic mass of gallium = $(68.9256 \text{ amu})\left(\frac{60.11\%}{100\%}\right) + (70.9247 \text{ amu})\left(\frac{39.89\%}{100\%}\right) = 69.7230 =$ **69.72 amu**

Plan:

2.35 To calculate the atomic mass of an element, take a weighted average based on the natural abundance of the isotopes: (isotopic mass of isotope 1 x fractional abundance) + (isotopic mass of isotope 2 x fractional abundance) + (isotopic mass of isotope 3 x fractional abundance).

Solution:

$$\text{Atomic mass of Mg} = (23.9850\ \text{amu})\left(\frac{78.99\%}{100\%}\right) + (24.9858\ \text{amu})\left(\frac{10.00\%}{100\%}\right) + (25.9826\ \text{amu})\left(\frac{11.01\%}{100\%}\right)$$

$$= 24.3050 = \mathbf{24.31\ amu}$$

2.36 Plan: To find the percent abundance of each Cl isotope, let x equal the fractional abundance of ^{35}Cl and (1 – x) equal the fractional abundance of ^{37}Cl since the sum of the fractional abundances must equal 1. Remember that atomic mass = (isotopic mass of ^{35}Cl x fractional abundance) + (isotopic mass of ^{37}Cl x fractional abundance).

Solution:

Atomic mass = (isotopic mass of ^{35}Cl x fractional abundance) + (isotopic mass of ^{37}Cl x fractional abundance)

35.4527 amu = 34.9689 amu(x) + 36.9659 amu(1 – x)

35.4527 amu = 34.9689 amu(x) + 36.9659 amu – 36.9659 amu(x)

35.4527 amu = 36.9659 amu – 1.9970 amu(x)

1.9970 amu(x) = 1.5132 amu

x = 0.75774 and 1 – x = 1 – 0.75774 = 0.24226

% abundance **^{35}Cl = 75.774%** % abundance **^{37}Cl = 24.226%**

2.37 Plan: To find the percent abundance of each Cu isotope, let x equal the fractional abundance of ^{63}Cu and (1 – x) equal the fractional abundance of ^{65}Cu since the sum of the fractional abundances must equal 1. Remember that atomic mass = (isotopic mass of ^{63}Cu x fractional abundance) + (isotopic mass of ^{65}Cu x fractional abundance).

Solution:

Atomic mass = (isotopic mass of ^{63}Cu x fractional abundance) + (isotopic mass of ^{65}Cu x fractional abundance)

63.546 amu = 62.9396 amu(x) + 64.9278 amu(1 – x)

63.546 amu = 62.9396 amu(x) + 64.9278 amu – 64.9278 amu(x)

63.546 amu = 64.9278 amu – 1.9882 amu(x)

1.9882 amu(x) = 1.3818 amu

x = 0.69500 and 1 – x = 1 – 0.69500 = 0.30500

% abundance **^{63}Cu = 69.50%** % abundance **^{65}Cu = 30.50%**

2.38 Plan: Review the section in the chapter on the periodic table.

Solution:

a) In the modern periodic table, the elements are arranged in order of increasing atomic **number**.

b) Elements in a **column or group** (or family) have similar chemical properties, not those in the same period or row.

c) Elements can be classified as **metals**, metalloids, or nonmetals.

2.39 The **metalloids** lie along the "staircase" line, with properties intermediate between metals and nonmetals.

2.40 Plan: Review the section on the classification of elements as metals, nonmetals, or metalloids.

Solution:

To the left of the "staircase" are the metals, which are generally hard, shiny, malleable, ductile, good conductors of heat and electricity, and form positive ions by losing electrons. To the right of the "staircase" are the nonmetals, which are generally soft or gaseous, brittle, dull, poor conductors of heat and electricity, and form negative ions by gaining electrons.

2.41 Plan: Locate each element on the periodic table. The Z value is the atomic number of the element. Metals are to the left of the "staircase," nonmetals are to the right of the "staircase," and the metalloids are the elements that lie along the "staircase" line.

Solution:

a) Germanium	Ge	4A(14)	metalloid
b) Phosphorus	P	5A(15)	nonmetal
c) Helium	He	8A(18)	nonmetal
d) Lithium	Li	1A(1)	metal
e) Molybdenum	Mo	6B(6)	metal

2.42 Plan: Locate each element on the periodic table. The Z value is the atomic number of the element. Metals are to the left of the "staircase," nonmetals are to the right of the "staircase," and the metalloids are the elements that lie along the "staircase" line.

Solution:

a) Arsenic	As	5A(15)	metalloid
b) Calcium	Ca	2A(2)	metal
c) Bromine	Br	7A(17)	nonmetal
d) Potassium	K	1A(1)	metal
e) Aluminum	Al	3A(13)	metal

2.43 Plan: Review the section in the chapter on the periodic table. Remember that alkaline earth metals are in Group 2A(2), the halogens are in Group 7A(17), and the metalloids are the elements that lie along the "staircase" line; periods are horizontal rows. Solution:

a) The symbol and atomic number of the heaviest alkaline earth metal are **Ra** and **88**.
b) The symbol and atomic number of the lightest metalloid in Group 4A(14) are **Si** and **14**.
c) The symbol and atomic mass of the coinage metal whose atoms have the fewest electrons are **Cu** and **63.55 amu**.
d) The symbol and atomic mass of the halogen in Period 4 are **Br** and **79.90 amu**.

2.44 Plan: Review the section in the chapter on the periodic table. Remember that the noble gases are in Group 8A(18), the alkali metals are in Group 1A(1), and the transition elements are the groups of elements located between Groups 2A(s) and 3A(13); periods are horizontal rows and metals are located to the left of the "staircase" line.

Solution:

a) The symbol and atomic number of the heaviest nonradioactive noble gas are **Xe** and **54**, respectively.
b) The symbol and group number of the Period 5 transition element whose atoms have the fewest protons are **Y** and **3B(3)**.
c) The symbol and atomic number of the only metallic chalcogen are **Po** and **84**.
d) The symbol and number of protons of the Period 4 alkali metal atom are **K** and **19**.

2.45 Plan: Review the section of the chapter on the formation of ionic compounds.

Solution:

Reactive metals and nometals will form **ionic** bonds, in which one or more electrons are transferred from the metal atom to the nonmetal atom to form a cation and an anion, respectively. The oppositely charged ions attract, forming the ionic bond.

2.46 Plan: Review the section of the chapter on the formation of covalent compounds. Solution:
Two nonmetals will form **covalent** bonds, in which the atoms share two or more electrons.

Plan:

2.47 Plan: Assign charges to each of the ions. Since the sizes are similar, there are no differences due to the sizes.

Solution:

Coulomb's law states the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges*. The *product of charges* in MgO (+2 x –2 = –4) is greater than the *product of charges* in LiF (+1 x –1 = –1). Thus, **MgO** has stronger ionic bonding.

2.48 A metal and a nonmetal will form an ionic compound. Locate these elements on the periodic table and predict their charges.

Solution:

Magnesium chloride ($MgCl_2$) is an ionic compound formed from a metal (magnesium) and a nonmetal (chlorine). Magnesium atoms transfer electrons to chlorine atoms. Each magnesium atom loses two electrons to form a Mg^{2+} ion and the same number of electrons (10) as the noble gas neon. Each chlorine atom gains one electron to form a Cl^- ion and the same number of electrons (18) as the noble gas argon. The Mg^{2+} and Cl^- ions attract each other to form an ionic compound with the ratio of one Mg^{2+} ion to two Cl^- ions. The total number of electrons lost by the magnesium atoms equals the total number of electrons gained by the chlorine atoms.

2.49 Plan: Recall that ionic bonds occur between metals and nonmetals, whereas covalent bonds occur between nonmetals. Solution:

KNO_3 shows **both** ionic and covalent bonding, covalent bonding between the N and O in NO_3^- and ionic bonding between the NO_3^- and the K^+.

2.50 Plan: Locate these elements on the periodic table and predict what ions they will form. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8.

Solution:

Potassium (K) is in Group **1**A(1) and forms the $\mathbf{K^+}$ ion. Iodine (I) is in Group **7**A(17) and forms the $\mathbf{I^-}$ ion (7 – 8 = –1).

2.51 Plan: Locate these elements on the periodic table and predict what ions they will form. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8.

Solution:

Barium in Group **2**A(2) forms a +2 ion: $\mathbf{Ba^{2+}}$. Selenium in Group **6**A(16) forms a –2 ion: $\mathbf{Se^{2-}}$ (6 – 8 = –2).

2.52 Plan: Use the number of protons (atomic number) to identify the element. Add the number of protons and neutrons together to get the mass number. Locate the element on the periodic table and assign its group and period number.

Solution:

a) Oxygen (atomic number = 8)	mass number = 8p + 9n = 17	Group 6A(16)	Period 2
b) Fluorine (atomic number = 9)	mass number = 9p + 10n = 19	Group 7A(17)	Period 2
c) Calcium (atomic number = 20)	mass number = 20p + 20n = 40	Group 2A(2)	Period 4

2.53 Plan: Use the number of protons (atomic number) to identify the element. Add the number of protons and neutrons together to get the mass number. Locate the element on the periodic table and assign its group and period number.

Solution:

a) Bromine (atomic number = 35)	mass number = 35p + 44n = 79	Group 7A(17)	Period 4
b) Nitrogen (atomic number = 7)	mass number = 7p + 8n = 15	Group 5A(15)	Period 2
c) Rubidium (atomic number = 37)	mass number = 37p + 48n = 85	Group 1A(1)	Period 5

2.54 Plan: Determine the charges of the ions based on their position on the periodic table. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8.

Next, determine the ratio of the charges to get the ratio of the ions.

Solution:

Lithium [Group 1A(1)] forms the Li^+ ion; oxygen [Group 6A(16)] forms the O^{2-} ion (6 – 8 = –2). The ionic compound that forms from the combination of these two ions must be electrically neutral, so two Li^+ ions combine with one O^{2-} ion to form the compound Li_2O. There are twice as many Li^+ ions as O^{2-} ions in a sample of Li_2O.

$$\text{Number of } O^{2-} \text{ ions} = \left(8.4\times10^{21}\ Li^+ \text{ ions}\right)\left(\frac{1\ O^{2-}\text{ ion}}{2\ Li^+\text{ ions}}\right) = \mathbf{4.2\times10^{21}\ O^{2-}\ ions}$$

Plan:

2.55 Determine the charges of the ions based on their position on the periodic table. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8. Next, determine the ratio of the charges to get the ratio of the ions.

Solution:

Ca [Group 2A(2)] forms Ca^{2+} and I [Group 7A(17)] forms I^- ions (7 – 8 = –1). The ionic compound that forms from the combination of these two ions must be electrically neutral, so one Ca^{2+} ion combines with two I^- ions to form the compound CaI_2. There are twice as many I^- ions as Ca^{2+} ions in a sample of CaI_2.

Number of I^- ions = $(7.4\text{x}10^{21}\ Ca^{2+}\text{ ions})\left(\frac{2\ I^-\text{ ions}}{1\ Ca^{2+}\text{ ion}}\right) = 1.48\text{x}10^{22}$ = **$1.5\text{x}10^{22}$ I^- ions**

2.56 Plan: The key is the size of the two alkali metal ions. The charges on the sodium and potassium ions are the same as both are in Group 1A(1), so there will be no difference due to the charge. The chloride ions are the same in size and charge, so there will be no difference due to the chloride ion.

Solution:

Coulomb's law states that the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges*. The *product of the charges* is the same in both compounds because both sodium and potassium ions have a +1 charge. Attraction increases as distance decreases, so the ion with the smaller radius, Na^+, will form a stronger ionic interaction (**NaCl**).

2.57 Plan: The key is the charge of the two metal ions. The sizes of the lithium and magnesium ions are about the same (magnesium is slightly smaller), so there will be little difference due to ion size. The oxide ions are the same in size and charge, so there will be no difference due to the oxide ion.

Solution:

Coulomb's law states the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges*. The *product of charges* in MgO (+2 x –2 = –4) is greater than the *product of charges* in Li_2O (+1 x –2 = –2). Thus, **MgO** has stronger ionic bonding.

2.58 Plan: Review the definitions of molecular and structural formulas.

Solution:

Both the structural and molecular formulas show the actual numbers of each type of atom in the molecule; in addition, the structural formula shows the arrangement of the atoms (i.e., how the atoms are connected to each other).

2.59 Plan: Review the concepts of atoms and molecules. Solution:

The mixture is similar to the sample of hydrogen peroxide in that both contain 20 billion oxygen atoms and 20 billion hydrogen atoms since both O_2 and H_2O_2 contain 2 oxygen atoms per molecule and both H_2 and H_2O_2 contain 2 hydrogen atoms per molecule. They differ in that they contain different types of molecules: H_2O_2 molecules in the hydrogen peroxide sample and H_2 and O_2 molecules in the mixture. In addition, the mixture contains 20 billion molecules (10 billion H_2 molecules + 10 billion O_2 molecules) while the hydrogen peroxide sample contains 10 billion molecules.

2.60 Plan: Write the symbol of each element present in the compound; the given number of each type of atom is represented with a subscript.

Solution:

a) Hydrazine has two nitrogen atoms and four hydrogen atoms: **N_2H_4**.

b) Glucose has six carbon atoms, twelve hydrogen atoms, and six oxygen atoms: **$C_6H_{12}O_6$**.

2.61 Plan: Write the symbol of each element present in the compound; the given number of each type of atom is represented with a subscript.

Solution:

a) Ethylene glycol has two carbon atoms, six hydrogen atoms, and two oxygen atoms: $\mathbf{C_2H_6O_2}$.
b) Peroxodisulfuric acid has two hydrogen atoms, two sulfur atoms, and eight oxygen atoms: $\mathbf{H_2S_2O_8}$.

2.62 Locate each of the individual elements on the periodic table, and assign charges to each of the ions. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix. Solution:
a) Sodium is a metal that forms a +1 (Group **1**A) ion and nitrogen is a nonmetal that forms a –3 ion (Group **5**A, 5 – 8 = –3).

+3 –3

+1 –3 +1

Na N Na$_3$N The compound is **Na_3N, sodium nitride**.

b) Oxygen is a nonmetal that forms a –2 ion (Group **6**A, 6 – 8 = –2) and strontium is a metal that forms a +2 ion (Group **2**A). +2 –2

Sr O The compound is **SrO, strontium oxide**.

c) Aluminum is a metal that forms a +3 ion (Group **3**A) and chlorine is a nonmetal that forms a –1 ion (Group **7**A, 7– 8 = –1). +3 –3

+3 –1 +3 –1

Al Cl AlCl$_3$ The compound is **$AlCl_3$, aluminum chloride**.

2.63 Plan: Locate each of the individual elements on the periodic table, and assign charges to each of the ions. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix.
Solution:
a) Cesium is a metal that forms a +1 (Group **1**A) ion and bromine is a nonmetal that forms a –1 ion (Group **7**A, 7 – 8 = –1).

+1 –1

Cs Br The compound is **CsBr, cesium bromide**.

b) Sulfur is a nonmetal that forms a –2 ion (Group **6**A, 6 – 8 = –2) and barium is a metal that forms a +2 ion (Group **2**A). +2 –2

Ba S The compound is **BaS, barium sulfide**.

c) Fluorine is a nonmetal that forms a –1 ion (Group **7**A, 7 – 8 = –1) and calcium is a metal that forms a +2 ion (Group **2**A). –2

+2 –1 +2 –1

Ca F CaF$_2$ The compound is **CaF_2, calcium fluoride**.

2.64 Plan: Based on the atomic numbers (the subscripts) locate the elements on the periodic table. Once the atomic numbers are located, identify the element and based on its position, assign a charge. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix.
Solution:
a) $_{12}$L is the element Mg (Z = 12). Magnesium [Group **2**A(2)] forms the Mg^{2+} ion. $_9$M is the element F (Z = 9). Fluorine [Group **7**A(17)] forms the F^- ion (7 – 8 = –1). The compound formed by the combination of these two elements is **MgF_2, magnesium fluoride**.
b) $_{30}$L is the element Zn (Z = 30). Zinc forms the Zn^{2+} ion (see Table 2.3). $_{16}$M is the element S (Z = 16). Sulfur [Group **6**A(16)] will form the S^{2-} ion (6 – 8 = –2). The compound formed by the combination of these two elements is **ZnS, zinc sulfide**.
c) $_{17}$L is the element Cl (Z = 17). Chlorine [Group **7**A(17)] forms the Cl^- ion (7 – 8 = –1). $_{38}$M is the element Sr (Z = 38). Strontium [Group **2**A(2)] forms the Sr^{2+} ion. The compound formed by the combination of these two elements is **$SrCl_2$, strontium chloride**.

Plan:

2.65 Plan: Based on the atomic numbers (the subscripts) locate the elements on the periodic table. Once the atomic numbers are located, identify the element and based on its position, assign a charge. For A group cations (metals), ion charge = group number; for anions (nonmetals), ion charge = group number minus 8. Find the smallest number of each ion that gives a neutral compound. To name ionic compounds with metals that form only one ion, name the metal, followed by the nonmetal name with an -ide suffix.

Solution:

a) $_{37}Q$ is the element Rb (Z = 37). Rubidium [Group **1**A(1)] forms the Rb^+ ion. $_{35}R$ is the element Br (Z = 35). Bromine [Group **7**A(17)] forms the $Br^–$ ion (7 – 8 = –1). The compound formed by the combination of these two elements is **RbBr, rubidium bromide**.

b) $_8Q$ is the O (Z = 8). Oxygen [Group **6**A(16)] will form the $O^{2–}$ ion (6 – 8 = –2). $_{13}R$ is the element Al (Z = 13). Aluminum [Group **3**A(13)] forms the Al^{3+} ion. The compound formed by the combination of these two elements is **Al_2O_3, aluminum oxide**.

c) $_{20}Q$ is the element Ca (Z = 20). Calcium [Group **2**A(2)] forms the Ca^{2+} ion. $_{53}R$ is the element I (Z = 53). Iodine [Group **7**A(17)] forms the $I^–$ ion (7 – 8 = –1). The compound formed by the combination of these two elements is **CaI_2, calcium iodide**.

2.66 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds, name the metal, followed by the nonmetal name with an -ide suffix. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name.

Solution:

a) tin(IV) chloride = **$SnCl_4$** The (IV) indicates that the metal ion is Sn^{4+} which requires 4 $Cl^–$ ions for a neutral compound.
b) $FeBr_3$ = **iron(III) bromide** (common name is ferric bromide); the charge on the iron ion is +3 to match the –3 charge of 3 $Br^–$ ions. The +3 charge of the Fe is indicated by (III).
c) cuprous bromide = **CuBr** (cuprous is +1 copper ion, cupric is +2 copper ion).
d) Mn_2O_3 = **manganese(III) oxide** Use (III) to indicate the +3 ionic charge of Mn: $\overset{+6\ \ -6}{\overset{+3\ \ -2}{Mn_2O_3}}$

2.67 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. Hydrates, compounds with a specific number of water molecules associated with them, are named with a prefix before the word hydrate to indicate the number of water molecules.

Solution:

a) Na_2HPO_4 **= sodium hydrogen phosphate** Sodium [Group 1A(1)] forms the Na^+ ion; $HPO_4^{2–}$ is the hydrogen phosphate ion.

b) potassium carbonate dihydrate = **K_2CO_3•$2H_2O$** Potassium [Group 1A(1)] forms the K^+ ion; carbonate is the $CO_3^{2–}$ ion. Two K^+ ions are required to match the –2 charge of the carbonate ion. Dihydrate indicates two water molecules ("waters of hydration") that are written after a centered dot. c) $NaNO_2$ = **sodium nitrite** $NO_2^–$ is the nitrite polyatomic ion.

d) ammonium perchlorate = **NH_4ClO_4** Ammonium is the polyatomic ion NH_4^+ and perchlorate is the polyatomic ion $ClO_4^–$. One NH_4^+ is required for every one $ClO_4^–$ ion.

2.68 Plan: Review the rules for nomenclature covered in the chapter. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Compounds must be neutral. Solution:

a) Barium [Group **2**A(2)] forms Ba^{2+} and oxygen [Group 6A(16)] forms $O^{2–}$ (6 – 8 = –2) so the neutral compound forms from one Ba^{2+} ion and one $O^{2–}$ ion. Correct formula is **BaO**.

b) Iron(II) indicates Fe^{2+} and nitrate is $NO_3^–$ so the neutral compound forms from one iron(II) ion and two nitrate ions. Correct formula is **$Fe(NO_3)_2$**.

c) Mn is the symbol for manganese. Mg is the correct symbol for magnesium. Correct formula is **MgS**. Sulfide is the $S^{2–}$ ion and sulfite is the $SO_3^{2–}$ ion.

2.69 Plan: Review the rules for nomenclature covered in the chapter. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Compounds must be neutral.

Solution:

a) **copper(I) iodide** Cu is copper, not cobalt; since iodide is I^-, this must be copper(I).
b) **iron(III) hydrogen sulfate** HSO_4^- is hydrogen sulfate, and this must be iron(III) to be neutral.
c) **magnesium dichromate** Mg forms Mg^{2+} and $Cr_2O_7^{2-}$ is named dichromate ion.

2.70 Acids donate H^+ ion to the solution, so the acid is a combination of H^+ and a negatively charged ion. Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid. Oxoacids (H + an oxoanion) are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid.

Solution:

a) Hydrogen sulfate is HSO_4^-, so its source acid is **H_2SO_4**. Name of acid is **sulfuric acid** (-ate becomes -ic acid).
b) **HIO_3, iodic acid** IO_3^- is the iodate ion: -ate becomes -ic acid.
c) Cyanide is CN^-; its source acid is **HCN hydrocyanic acid** (binary acid).
d) **H_2S, hydrosulfuric acid** (binary acid).

2.71 Plan: Acids donate H^+ ion to the solution, so the acid is a combination of H^+ and a negatively charged ion. Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid. Oxoacids (H + an oxoanion) are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid.

Solution:

a) Perchlorate is ClO_4^-, so the source acid is **$HClO_4$**. Name of acid is **perchloric acid** (-ate becomes -ic acid).
b) **nitric acid, HNO_3** NO_3^- is the nitrate ion: -ate becomes -ic acid.
c) Bromite is BrO_2^-, so the source acid is **$HBrO_2$**. Name of acid is **bromous acid** (-ite becomes -ous acid). d) **hydrofluoric acid, HF** (binary acid)

2.72 Plan: This compound is composed of two nonmetals. The element with the lower group number is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound.

Solution: **disulfur tetrafluoride** **S_2F_4** Di- indicates two S atoms and tetra- indicates four F atoms.

2.73 Plan: This compound is composed of two nonmetals. When a compound contains oxygen and a halogen, the halogen is named first. Greek numerical prefixes are used to indicate the number of atoms of each element in the compound.

Solution:

dichlorine monoxide **Cl_2O** Di- indicates two Cl atoms and mono- indicates one O atom.

2.74 Plan: Break down each formula to the individual elements and count the number of atoms of each element by observing the subscripts. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.

Solution:

a) There are **12 atoms of oxygen** in $Al_2(SO_4)_3$. The molecular mass is:
Al = 2(26.98 amu) = 53.96 amu S = 3(32.07 amu) = 96.21 amu O = 12(16.00 amu) = 192.0 amu
342.2 amu

b) There are **9 atoms of hydrogen** in $(NH_4)_2HPO_4$. The molecular mass is:
N = 2(14.01 amu) = 28.02 amu
H = 9(1.008 amu) = 9.072 amu P = 1(30.97 amu) = 30.97 amu O = 4(16.00 amu) = 64.00 amu
132.06 amu

c) There are **8 atoms of oxygen** in $Cu_3(OH)_2(CO_3)_2$. The molecular mass is:

Plan:

Cu = 3(63.55 amu) = 190.6 amu
O = 8(16.00 amu) = 128.0 amu
H = 2(1.008 amu) = 2.016 amu C = 2(12.01 amu) = 24.02 amu
344.6 amu

2.75 Plan: Break down each formula to the individual elements and count the number of atoms of each element by observing the subscripts. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.
Solution:
a) There are **9 atoms of hydrogen** in $C_6H_5COONH_4$. The molecular mass is:
C = 7(12.01 amu) = 84.07 amu
H = 9(1.008 amu) = 9.072 amu
O = 2(16.00 amu) = 32.00 amu
N = 1(14.01 amu) = 14.01 amu
139.15 amu
b) There are **2 atoms of nitrogen** in $N_2H_6SO_4$. The molecular mass is:
N = 2(14.01 amu) = 28.02 amu
H = 6(1.008 amu) = 6.048 amu S = 1(32.07 amu) = 32.07 amu O = 4(16.00 amu) = 64.00 amu
130.14 amu
c) There are 1**2 atoms of oxygen** in $Pb_4SO_4(CO_3)_2(OH)_2$. The molecular mass is:
Pb = 4(207.2 amu) = 828.8 amu S = 1(32.07 amu) = 32.07 amu O = 12(16.00 amu) = 192.00 amu C = 2(12.01 amu) = 24.02 amu
H = 2(1.008 amu) = 2.016 amu
1078.9 amu

2.76 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.
Solution:
a) **$(NH_4)_2SO_4$** ammonium is NH_4^+ and sulfate is SO_4^{2-}
N = 2(14.01 amu) = 28.02 amu
H = 8(1.008 amu) = 8.064 amu S = 1(32.07 amu) = 32.07 amu O = 4(16.00 amu) = 64.00 amu
132.15 amu
b) **NaH_2PO_4** sodium is Na^+ and dihydrogen phosphate is $H_2PO_4^-$
Na = 1(22.99 amu) = 22.99 amu
H = 2(1.008 amu) = 2.016 amu P = 1(30.97 amu) = 30.97 amu O = 4(16.00 amu) = 64.00 amu
119.98 amu
c) **$KHCO_3$** potassium is K^+ and bicarbonate is HCO_3^-
K = 1(39.10 amu) = 39.10 amu
H = 1(1.008 amu) = 1.008 amu C = 1(12.01 amu) = 12.01 amu O = 3(16.00 amu) = 48.00 amu
100.12 amu

2.77 Plan: Review the rules for nomenclature covered in the chapter. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. The molecular (formula) mass is the sum of the atomic masses of all of the atoms.

Solution:

a) **$Na_2Cr_2O_7$** sodium is Na^+ and dichromate is $Cr_2O_7^{2-}$

Na = 2(22.99 amu) = 45.98 amu Cr = 2(52.00 amu) = 104.00 amu O = 7(16.00 amu) = 112.00 amu

261.98 amu

b) NH_4ClO_4 ammonium is NH_4^+ and perchlorate is ClO_4^-

N = 1(14.01 amu) = 14.01 amu
H = 4(1.008 amu) = 4.032 amu Cl = 1(35.45 amu) = 35.45 amu O = 4(16.00 amu) = 64.00 amu

117.49 amu

c) $Mg(NO_2)_2•3H_2O$ magnesium is Mg^{2+}, nitrite is NO_2^-, and trihydrate is $3H_2O$

Mg = 1(24.31 amu) = 24.31 amu N = 2(14.01 amu) = 28.02 amu
H = 6(1.008 amu) = 6.048 amu
O = 7(16.00 amu) = 112.00 amu

170.38 amu

2.78 Plan: Use the chemical symbols and count the atoms of each type to give a molecular formula. Use the nomenclature rules in the chapter to derive the name. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.

Solution:

a) Formula is SO_3. Name is **sulfur trioxide** (the prefix tri- indicates 3 oxygen atoms).

S = 1(32.07 amu) = 32.07 amu
O = 3(16.00 amu) = 48.00 amu

80.07 amu

b) Formula is C_3H_8. Since it contains only carbon and hydrogen it is a hydrocarbon and with three carbons its name is **propane**.

C = 3(12.01 amu) = 36.03 amu
H = 8(1.008 amu) = 8.064 amu

44.09 amu

2.79 Plan: Use the chemical symbols and count the atoms of each type to give a molecular formula. Use the nomenclature rules in the chapter to derive the name. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.

Solution:

a) Formula is N_2O. Name is **dinitrogen monoxide** (the prefix di- indicates 2 nitrogen atoms and mono- indicates 1 oxygen atom).

N = 2(14.01 amu) = 28.02 amu
O = 1(16.00 amu) = 16.00 amu

44.02 amu

b) Formula is C_2H_6. Since it contains only carbon and hydrogen it is a hydrocarbon and with three carbons its name is **ethane**.

C = 2(12.01 amu) = 24.02 amu
H = 6(1.008 amu) = 6.048 amu

30.07 amu

2.80 Plan: Review the law of mass conservation and law of definite composition. For each experiment, compare the mass values before and after each reaction and examine the ratios of the mass of reacted sodium to the mass of reacted chlorine.

Solution:

In each case, the mass of the starting materials (reactants) equals the mass of the ending materials (products), so the law of mass conservation is observed. Case 1: 39.34 g + 60.66 g = 100.00 g
Case 2: 39.34 g + 70.00 g = 100.00 g + 9.34 g
Case 3: 50.00 g + 50.00 g = 82.43 g + 17.57 g

Each reaction yields the product NaCl, not Na_2Cl or $NaCl_2$ or some other variation, so the law of definite composition is observed. In each case, the ratio of the mass of sodium to the mass of chlorine in the compound is the same.

Case 1: Mass Na/mass Cl_2 = 39.34 g/60.66 g = 0.6485

Case 2: Mass of reacted Cl_2 = initial mass – excess mass = 70.00 g – 9.34 g = 60.66 g Cl_2
Mass Na/mass Cl_2 = 39.34 g/60.66 g = 0.6485
Case 3: Mass of reacted Na = initial mass – excess mass = 50.00 g – 17.57 g = 32.43 g Na
Mass Na/mass Cl_2 = 32.43 g/50.00 g = 0.6486

2.81 Plan: Review the nomenclature rules in the chapter. For ionic compounds, name the metal, followed by the nonmetal name with an -ide suffix. For ionic compounds containing polyatomic ions, name the metal, followed by the name of the polyatomic ion. For metals, like many transition metals, that can form more than one ion each with a different charge, the ionic charge of the metal ion is indicated by a Roman numeral within parentheses immediately following the metal's name. Oxoacids (H + an oxoanion) are named by changing the suffix of the oxoanion: -ate becomes -ic acid and -ite becomes -ous acid. Greek numerical prefixes are used to indicate the number of atoms of each element in a compound composed of two nonmetals. Solution:

a) blue vitriol $CuSO_4$•$5H_2O$ **copper(II) sulfate pentahydrate**
SO_4^{2-} = sulfate; II is used to indicate the 2+ charge of Cu; penta- is used to indicate the 5 waters of hydration.

b) slaked lime $Ca(OH)_2$ **calcium hydroxide** The anion OH^- is hydroxide.

c) oil of vitriol H_2SO_4 **sulfuric acid**
SO_4^{2-} is the sulfate ion; since this is an acid, -ate becomes -ic acid.

d) washing soda Na_2CO_3 **sodium carbonate**
CO_3^{2-} is the carbonate ion.

e) muriatic acid HCl **hydrochloric acid**
Binary acids (H plus one other nonmetal) are named hydro- + nonmetal root + -ic acid.

f) Epsom salts $MgSO_4$•$7H_2O$ **magnesium sulfate heptahydrate**
SO_4^{2-} = sulfate; hepta- is used to indicate the 7 waters of hydration.

g) chalk $CaCO_3$ **calcium carbonate**
CO_3^{2-} is the carbonate ion.

h) dry ice CO_2 **carbon dioxide**
The prefix di- indicates 2 oxygen atoms; since there is only one carbon atom, no prefix is used.

i) baking soda $NaHCO_3$ **sodium hydrogen carbonate** HCO_3^- is the hydrogen carbonate ion.

j) lye NaOH **sodium hydroxide**
The anion OH^- is hydroxide.

2.82 Plan: Review the discussion on separations.
Solution:
Separating the components of a mixture requires physical methods only; that is, no chemical changes (no changes in composition) take place and the components maintain their chemical identities and properties throughout. Separating the components of a compound requires a chemical change (change in composition).

2.83 Plan: Review the definitions of homogeneous and heterogeneous.
Solution:
A homogeneous mixture is uniform in its macroscopic, observable properties; a heterogeneous mixture shows obvious differences in properties (density, color, state, etc.) from one part of the mixture to another.

2.84 A solution (such as salt or sugar dissolved in water) is a homogeneous mixture.

2.85 Plan: Review the definitions of homogeneous and heterogeneous. The key is that a homogeneous mixture has a uniform composition while a heterogeneous mixture does not. A mixture consists of two or more substances physically mixed together while a compound is a pure substance. Solution:

a) Distilled water is a **compound** that consists of H_2O molecules only.

b) Gasoline is a **homogeneous mixture** of hydrocarbon compounds of uniform composition that can be separated by physical means (distillation).

c) Beach sand is a **heterogeneous mixture** of different size particles of minerals and broken bits of shells.

d) Wine is a **homogeneous mixture** of water, alcohol, and other compounds that can be separated by physical means (distillation).
e) Air is a **homogeneous mixture** of different gases, mainly N_2, O_2, and Ar.

2.86 Plan: Review the definitions of homogeneous and heterogeneous. The key is that a homogeneous mixture has a uniform composition while a heterogeneous mixture does not. A mixture consists of two or more substances physically mixed together while a compound is a pure substance.
Solution:
a) Orange juice is a **heterogeneous mixture** of water, juice, and bits of orange pulp.
b) Vegetable soup is a **heterogeneous mixture** of water, broth, and vegetables.
c) Cement is a **heterogeneous mixture** of various substances.
d) Calcium sulfate is a **compound** of calcium, sulfur, and oxygen in a fixed proportion.
e) Tea is a **homogeneous mixture**.

2.87 Plan: Use the equation for the volume of a sphere in part a) to find the volume of the nucleus and the volume of the atom. Calculate the fraction of the atom volume that is occupied by the nucleus. For part b), calculate the total mass of the two electrons; subtract the electron mass from the mass of the atom to find the mass of the nucleus. Then calculate the fraction of the atom's mass contributed by the mass of the nucleus.
Solution:

a) Volume (m^3) of nucleus = $\frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(2.5\text{x}10^{-15}\text{ m}\right)^3 = 6.54498\text{x}10^{-44}\text{ m}^3$

Volume (m^3) of atom = $\frac{4}{3}\pi r^3 = \frac{4}{3}\pi\left(3.1\text{x}10^{-11}\text{ m}\right)^3 = 1.24788\text{x}10^{-31}\text{ m}^3$

Fraction of volume = $\frac{\text{volume of nucleus}}{\text{volume of atom}} = \frac{6.54498\text{x}10^{-44}\text{ m}^3}{1.24788\text{x}10^{-31}\text{ m}^3} = 5.2449\text{x}10^{-13} = \mathbf{5.2x10^{-13}}$

b) Mass of nucleus = mass of atom – mass of electrons
= $6.64648\text{x}10^{-24}$ g – $2(9.10939\text{x}10^{-28}$ g) = $6.64466\text{x}10^{-24}$ g

Fraction of mass = $\frac{\text{mass of nucleus}}{\text{mass of atom}} = \frac{\left(6.64466\text{x}10^{-24}\text{ g}\right)}{\left(6.64648\text{x}10^{-24}\text{ g}\right)}$ = 0.99972617 = **0.999726**

As expected, the volume of the nucleus relative to the volume of the atom is small while its relative mass is large.

2.88 Plan: Use the chemical symbols and count the atoms of each type to give a molecular formula. Use the nomenclature rules in the chapter to derive the name. These compounds are composed of two nonmetals. Greek numerical prefixes are used to indicate the number of atoms of each element in each compound. The molecular (formula) mass is the sum of the masses of each atom times its atomic mass.
Solution:
a) Formula is BrF_3. When a compound is composed of two elements from the same group, the element with the higher period number is named first. The prefix tri- indicates 3 fluorine atoms. A prefix is used with the first word in the name only when more than one atom of that element is present. The name is **bromine trifluoride**.

Br = 1(79.90 amu) = 79.90 amu F = 3(19.00 amu) = 57.00 amu
136.90 amu

b) The formula is SCl_2. The element with the lower group number is the first word in the name. The prefix di- indicates 2 chlorine atoms. A prefix is used with the first word in the name only when more than one atom of that element is present. The name is **sulfur dichloride**.

S = 1(32.07 amu) = 32.07 amu Cl = 2(35.45 amu) = 70.90 amu

102.97 amu

c) The formula is PCl_3. The element with the lower group number is the first word in the name. The prefix tri- indicates 3 chlorine atoms. A prefix is used with the first word in the name only when more than one atom of that element is present. The name is **phosphorus trichloride**.

P	=	1(30.97 amu)	=	30.97 amu
Cl	=	3(35.45 amu)	=	106.35 amu
				137.32 amu

d) The formula is N_2O_5. The element with the lower group number is the first word in the name. The prefix di- indicates 2 nitrogen atoms and the prefix penta- indicates 5 oxygen atoms. Only the second element is named with the suffix -ide. The name is **dinitrogen pentoxide**.

N	=	2(14.01 amu)	=	28.02 amu
O	=	5(16.00 amu)	=	80.00 amu
				108.02 amu

2.89 Plan: Determine the percent oxygen in each oxide by subtracting the percent nitrogen from 100%. Express the percentage in amu and divide by the atomic mass of the appropriate elements. Then divide each amount by the smaller number and convert to the simplest whole-number ratio. To find the mass of oxygen per 1.00 g of nitrogen, divide the mass percentage of oxygen by the mass percentage of nitrogen.

Solution:

a) I (100.00 – 46.69 N)% = 53.31% O

$$\left(\frac{46.69 \text{ amu N}}{14.01 \text{ amu N}}\right) = 3.3326 \text{ N} \qquad \left(\frac{53.31 \text{ amu O}}{16.00 \text{ amu O}}\right) = 3.3319 \text{ O}$$

$$\frac{3.3326 \text{ N}}{3.3319} = 1.0002 \text{ N} \qquad \frac{3.3319 \text{ O}}{3.3319} = 1.0000 \text{ O}$$

The simplest whole-number ratio is **1:1 N:O**.

II (100.00 – 36.85 N)% = 63.15% O

$$\left(\frac{36.85 \text{ amu N}}{14.01 \text{ amu N}}\right) = 2.6303 \text{ N} \qquad \left(\frac{63.15 \text{ amu O}}{16.00 \text{ amu O}}\right) = 3.9469 \text{ O}$$

$$\frac{2.6303\text{N}}{2.6303} = 1.0000 \text{ mol N} \qquad \frac{3.9469 \text{ O}}{2.6303} = 1.5001 \text{ O}$$

The simplest whole-number ratio is 1:1.5 N:O = **2:3 N:O**.

III (100.00 – 25.94 N)% = 74.06% O

$$\left(\frac{25.94 \text{ amu N}}{14.01\text{amu N}}\right) = 1.8515 \text{ N} \qquad \left(\frac{74.06 \text{ amuO}}{16.00\text{amuO}}\right) = 4.6288 \text{ O}$$

$$\frac{1.8515 \text{ N}}{1.8515} = 1.0000 \text{ N} \qquad \frac{4.6288 \text{ O}}{1.8515} = 2.5000 \text{ O}$$

The simplest whole-number ratio is 1:2.5 N:O = **2:5 N:O.**

b) I $\left(\frac{53.31 \text{ amu O}}{46.69 \text{ amu N}}\right) = 1.1418 =$ **1.14 g O**

II $\left(\dfrac{63.15 \text{ amu O}}{36.85 \text{ amu N}}\right) = 1.7137 = \mathbf{1.71\ g\ O}$

III $\left(\dfrac{74.06 \text{ amu O}}{25.94 \text{ amu N}}\right) = 2.8550 = \mathbf{2.86\ g\ O}$

2.90 Plan: Recall the definitions of solid, liquid, gas (from Chapter 1), element, compound, and homogeneous and heterogeneous mixtures.
Solution:

a) Gas is the phase of matter that fills its container. A mixture must contain at least two different substances. B, F, G, and I each contain only one gas. **D and E** each contain a mixture; E is a mixture of two different gases while D is a mixture of a gas and a liquid of a second substance.
b) An element is a substance that cannot be broken down into simpler substances. **A, C, G, and I** are elements.
c) The solid phase has a very high resistance to flow since it has a fixed shape. **A** shows a solid element.
d) A homogeneous mixture contains two or more substances and has only one phase. **E and H** are examples of this. E is a homogeneous mixture of two gases and H is a homogeneous mixture of two liquid substances.
e) A liquid conforms to the container shape and forms a surface. **C** shows one element in the liquid phase.
f) A diatomic particle is a molecule composed of two atoms. **B and G** contain diatomic molecules of gas.
g) A compound can be broken down into simpler substances. **B and F** show molecules of a compound in the gas phase.
h) The compound shown in **F** has molecules composed of two white atoms and one blue atom for a 2:1 atom ratio.
i) Mixtures can be separated into the individual components by physical means. **D, E, and H** are each a mixture of two different substances.
j) A heterogeneous mixture like **D** contains at least two different substances with a visible boundary between those substances.
k) Compounds obey the law of definite composition. **B and F** depict compounds.

2.91 Plan: To find the mass percent divide the mass of each substance in mg by the amount of seawater in mg and multiply by 100. The percent of an ion is the mass of that ion divided by the total mass of ions.
Solution:

a) Mass (mg) of seawater = $(1 \text{ kg})\left(\dfrac{1000 \text{ g}}{1 \text{ kg}}\right)\left(\dfrac{1000 \text{ mg}}{1 \text{ g}}\right) = 1\text{x}10^6 \text{ mg}$

Mass % = $\left(\dfrac{\text{mass of substance}}{\text{mass of seawater}}\right)(100\%)$

Mass % Cl^- = $\left(\dfrac{18{,}980 \text{ mg } Cl^-}{1\text{x}10^6 \text{ mg seawater}}\right)(100\%) = \mathbf{1.898\%\ Cl^-}$

$$\text{Mass \% Na}^+ = \left(\frac{10{,}560 \text{ mg Na}^+}{1\text{x}10^6 \text{ mg seawater}}\right)(100\%) = \mathbf{1.056\%\ Na^+}$$

$$\text{Mass \% SO}_4^{2-} = \left(\frac{2650 \text{ mg SO}_4^{2-}}{1\text{x}10^6 \text{ mg seawater}}\right)(100\%) = \mathbf{0.265\%\ SO_4^{2-}}$$

$$\text{Mass \% Mg}^{2+} = \left(\frac{1270 \text{ mg Mg}^{2+}}{1\text{x}10^6 \text{ mg seawater}}\right)(100\%) = \mathbf{0.127\%\ Mg^{2+}}$$

$$\text{Mass \% Ca}^{2+} = \left(\frac{400 \text{ mg Ca}^{2+}}{1\text{x}10^6 \text{ mg seawater}}\right)(100\%) = \mathbf{0.04\%\ Ca^{2+}}$$

$$\text{Mass \% K}^+ = \left(\frac{380 \text{ mg K}^+}{1\text{x}10^6 \text{ mg seawater}}\right)(100\%) = \mathbf{0.038\%\ K^+}$$

$$\text{Mass \% HCO}_3^- = \left(\frac{140 \text{ mg HCO}_3^-}{1\text{x}10^6 \text{ mg seawater}}\right)(100\%) = \mathbf{0.014\%\ HCO_3^-}$$

The mass percents do not add to 100% since the majority of seawater is H_2O.

b) Total mass of ions in 1 kg of seawater

= 18,980 mg + 10,560 mg + 2650 mg + 1270 mg + 400 mg + 380 mg + 140 mg = 34,380 mg

$$\% \text{ Na}^+ = \left(\frac{10{,}560 \text{ mg Na}^+}{34{,}380 \text{ mg total ions}}\right)(100) = 30.71553 = \mathbf{30.72\%}$$

c) Alkaline earth metal ions are Mg^{2+} and Ca^{2+} (Group 2 ions). Total mass % = 0.127% Mg^{2+} + 0.04% Ca^{2+} = 0.167%

Alkali metal ions are Na^+ and K^+ (Group 1 ions). Total mass % = 1.056% Na^+ + 0.038% K^+ = 1.094%

$$\frac{\text{Mass \% of alkali metal ions}}{\text{Mass \% of alkaline earth metal ions}} = \frac{1.094\%}{0.167\%} = 6.6$$

Total mass percent for alkali metal ions is **6.6 times greater** than the total mass percent for alkaline earth metal ions. Sodium ions (alkali metal ions) are dominant in seawater.

d) Anions are Cl^-, SO_4^{2-}, and HCO_3^-.

Total mass % = 1.898% Cl^- + 0.265% SO_4^{2-} + 0.014% HCO_3^- = 2.177% anions

Cations are Na^+, Mg^{2+}, Ca^{2+}, and K^+.

Total mass % = 1.056% Na^+ + 0.127% Mg^{2+} + 0.04% Ca^{2+} + 0.038% K^+ = 1.2610 = 1.26% cations

The mass fraction of **anions** is larger than the mass fraction of cations. Is the solution neutral since the mass of anions exceeds the mass of cations? Yes, although the mass is larger, the number of positive charges equals the number of negative charges.

2.92 Plan: Review the mass laws in the chapter.

Solution:

The **law of mass conservation** is illustrated in this change. The first flask has six oxygen atoms and six nitrogen atoms. The same number of each type of atom is found in both of the subsequent flasks. The mass of the substances did not change. The **law of definite composition** is also illustrated. During both temperature changes, the same compound, N_2O, was formed with the same composition.

.

2.93 Plan: Use the density values to convert volume of each element to mass. Find the mass ratio of Ba to S in the compound and compare that to the mass ratio present.

Solution:

For barium sulfide the barium to sulfur mass ratio is (137.3 g Ba/32.07 g S) = 4.281 g Ba/g S

$$\text{Mass (g) of barium} = \left(2.50\ \text{cm}^3\ \text{Ba}\right)\left(\frac{3.51\ \text{g Ba}}{1\ \text{cm}^3\ \text{Ba}}\right) = 8.775\ \text{g Ba}$$

$$\text{Mass (g) of sulfur} = \left(1.75\ \text{cm}^3\ \text{S}\right)\left(\frac{2.07\ \text{g S}}{1\ \text{cm}^3\ \text{S}}\right) = 3.6225\ \text{g S}$$

$$\text{Barium to sulfur mass ratio} = \frac{8.775\ \text{g Ba}}{3.6225\ \text{g S}} = 2.4224 = 2.42\ \text{g Ba/g S}$$

No, the ratio is too low; there is insufficient barium.

2.94 Plan: First, count each type of atom present to produce a molecular formula. The molecular (formula) mass is the sum of the atomic masses of all of the atoms. Divide the mass of each element in the compound by the molecular mass and multiply by 100 to obtain the mass percent of each element.

Solution:

The molecular formula of succinic acid is $\mathbf{C_4H_6O_4}$.

C = 4(12.01 amu) = 48.04 amu

H = 6(1.008 amu) = 6.048 amu

O = 4 (16.00 amu) = 64.00 amu

118.09 amu

$$\%\ \text{C} = \left(\frac{48.04\ \text{amu C}}{118.088\ \text{amu}}\right)100\% = 40.6815 = \mathbf{40.68\%\ C}$$

$$\%\ \text{H} = \left(\frac{6.048\ \text{amu H}}{118.088\ \text{amu}}\right)100\% = 5.1216 = \mathbf{5.122\%\ H}$$

$$\%\ \text{O} = \left(\frac{64.00\ \text{amu O}}{118.088\ \text{amu}}\right)100\% = 54.1969 = \mathbf{54.20\%\ O}$$

Check: Total = (40.68 + 5.122 + 54.20)% = 100.00% The answer checks.

2.95 Plan: The toxic level of fluoride ion for a 70-kg person is 0.2 g. Convert this mass to mg and use the concentration of fluoride ion in drinking water to find the volume of water that contains the toxic amount. Convert the volume of the reservoir to liters and use the concentration of 1mg of fluoride ion per liter of water to find the mass of sodium fluoride required.

Solution:

A 70-kg person would have to consume 0.2 mg of F^- to reach the toxic level.

Mass (mg) of fluoride for a toxic level = $(0.2\ g\ F^-)\left(\frac{1\ mg\ F^-}{0.001\ g\ F^-}\right)$ = 200 mg F^-

Volume (L) of water = $(200\ mg)\left(\frac{1\ L\ water}{1\ mg\ F^-}\right)$ = 200 = **2x10^2 L water**

Volume (L) of reservoir = $8.50x10^7\ gal\left(\frac{4\ qt}{1\ gal}\right)\left(\frac{1\ L}{1.057\ qt}\right)$ = 3.26651x10^8 L

The molecular mass of NaF = 22.99 amu Na + 19.00 amu F = 41.99 amu. There are 19.00 mg of F^- in every 41.99 mg of NaF.

Mass (kg) of NaF = $(3.216651x10^8\ L)\left(\frac{1\ mg\ F^-}{1\ L\ H_2O}\right)\left(\frac{41.99\ mg\ NaF}{19.00\ mg\ F^-}\right)\left(\frac{10^{-3}\ g}{1\ mg}\right)\left(\frac{1\ kg\ NaF}{10^3\ g\ NaF}\right)$

= 710.88 = **711 kg NaF**

2.96 Plan: Z = the atomic number of the element. A is the mass number. To find the percent abundance of each Sb isotope, let x equal the fractional abundance of one isotope and (1 – x) equal the fractional abundance of the second isotope since the sum of the fractional abundances must equal 1. Remember that atomic mass = (isotopic mass of the first isotope x fractional abundance) + (isotopic mass of the second isotope x fractional abundance).

Solution:

a) Antimony is element 51so Z = 51. Isotope of mass 120.904 amu has a mass number of 121: $^{121}_{51}\mathbf{Sb}$

Isotope of mass 122.904 amu has a mass number of 123: $^{123}_{51}\mathbf{Sb}$

b) Let x = fractional abundance of antimony-121. This makes the fractional abundance of antimony-123 = 1 – x

x(120.904 amu) + (1 – x) (122.904 amu) = 121.8 amu

120.904 amu(x) + 122.904 amu – 122.904 amu(x) = 121.8 amu

2x = 1.104

x = 0.552 = **0.55 fraction of antimony-121**

1 – x = 1 – 0.552 = **0.45 fraction of antimony-123**

2.97 Plan: List all possible combinations of the isotopes. Determine the masses of each isotopic composition. The molecule consisting of the lower abundance isotopes (N-15 and O-18) is the least common, and the one containing only the more abundant isotopes (N-14 and O-16) will be the most common.

Solution:

a) b)

Formula Mass (amu)

$^{15}\mathbf{N_2}{}^{18}\mathbf{O}$	2(15 amu N) + 18 amu O = **48**	**least common**
$^{15}\mathbf{N_2}{}^{16}\mathbf{O}$	2(15 amu N) + 16 amu O = **46**	
$^{14}\mathbf{N_2}{}^{18}\mathbf{O}$	2(14 amu N) + 18 amu O = **46**	
$^{14}\mathbf{N_2}{}^{16}\mathbf{O}$	2(14 amu N) + 16 amu O = **44**	**most common**
$^{15}\mathbf{N}{}^{14}\mathbf{N}{}^{18}\mathbf{O}$	1(15 amu N) + 1(14 amu N) + 18 amu O = **47**	
$^{15}\mathbf{N}{}^{14}\mathbf{N}{}^{16}\mathbf{O}$	1(15 amu N) + 1(14 amu N) + 16 amu O = **45**	

2.98 Plan: Review the information about the periodic table in the chapter.

Solution:

a) Nonmetals are located in the upper-right portion of the periodic table: **Black, red, green, and purple**

b) Metals are located in the large left portion of the periodic table: **Brown and blue**

c) Some nonmetals, such as oxygen, chlorine, and argon, are gases: **Red, green, and purple**

d) Most metals, such as sodium and barium are solids; carbon is a solid: **Brown, blue, and black**

e) Nonmetals form covalent compounds; most noble gases do not form compounds:
Black and red or black and green or red and green

f) Nonmetals form covalent compounds; most noble gases do not form compounds:
Black and red or black and green or red and green

g) Metals react with nonmetals to form ionic compounds. For a compound with a formula of MX, the ionic charges of the metals and nonmetal must be equal in magnitude like Na^+ and Cl^- or Ba^{2+} and O^{2-}: **Brown and green or blue and red**

h) Metals react with nonmetals to form ionic compounds. For a compound with a formula of MX, the ionic charges of the metals and nonmetal must be equal in magnitude like Na^+ and Cl^- or Ba^{2+} and O^{2-}: **Brown and green or blue and red**

i) Metals react with nonmetals to form ionic compounds. For a compound with a formula of M_2X, the ionic charge of the nonmetal must be twice as large as that of the metal like Na^+ and O^{2-} or Ba^{2+} and C^{4-}: **Brown and red or blue and black**

j) Metals react with nonmetals to form ionic compounds. For a compound with a formula of MX_2, the ionic charge of the metal must be twice as large as that of the nonmetal like Ba^{2+} and Cl^-: **Blue and green**

k) Most Group 8A(18) elements are unreactive: **Purple**

l) Different compounds often exist between the same two nonmetal elements. Since oxygen exists as O^{2-} or $O_2{}^{2-}$, metals can sometimes form more than one compound with oxygen: **Black and red or red and green or black and green or brown and red or blue and red**

2.99 Plan: Convert the mass of compound in mg to kg and use the absolute mass of the atomic mass unit to find the number of amu of compound. Divide by the formula mass of the compound to obtain molecules of compound and then number of As atoms. To find mass percent of metal, divide the atomic mass of the metal by the total mass of metal-dimercaprol complex and multiply by 100.

Solution:

a) As atoms = $(250.\text{mg dimercaprol})\left(\frac{10^{-3}\text{ g}}{1\text{ mg}}\right)\left(\frac{1\text{ kg}}{10^{3}\text{ g}}\right)\left(\frac{1\text{ amu}}{1.66054\text{x}10^{-27}\text{ kg}}\right)\left(\frac{1\text{ dimercaprol}}{124.23\text{ amu}}\right)\left(\frac{1\text{ As atom}}{1\text{ dimercaprol}}\right)$

= $1.2119\text{x}10^{21}$ As atoms = $\mathbf{1.21x10^{21}}$ **As atoms**

b) Mass % Hg = $\frac{200.6\text{ amu}}{(200.6 + 124.23)\text{ amu}}$ x 100% = 61.7554 = **61.76% Hg**

Mass % Tl = $\frac{204.4\text{ amu}}{(204.4 + 124.23)\text{ amu}}$ x 100% = 62.1976 = **62.20% Tl**

Mass % Cr = $\dfrac{52.00 \text{ amu}}{(52.00 + 124.23) \text{ amu}}$ x 100% = 29.5069 = **29.51% Cr**

2.100 Plan: Use Coulomb's Law which states that the energy of attraction in an ionic bond is directly proportional to the *product of charges* and inversely proportional to the *distance between charges.* The strongest ionic bonding occurs between ions with the largest ionic charges and the smallest radii.
Solution:
Of the cations, Mg^{2+} and Ba^{2+} have the largest ionic charges but Mg^{2+} is significantly smaller (72 pm vs 135 pm); pairing Mg^{2+} with the anion with the largest charge, O^{2-}, would give the strongest ionic bond: **Mg^{2+} and O^{2-}**. The weakest ionic bonding occurs between ions with the smallest ionic charges and the largest radii. Choose the largest +1 cation, Rb^+ (152 pm) and the largest –1 anion, I^- (220 pm): **Rb^+ and I^-**.

2.101 Plan: First, determine the fraction of each element in each mineral by dividing the total atomic mass of element by the total molecular mass of the mineral. The percent mass of each element in the rock can be found by multiplying the mass fraction of each element in each mineral by the mass fraction of that mineral in the rock and then multiplying by 100. Solution:

Molecular mass of Fe_2SiO_4 = 2(55.85 amu) + 28.09 amu + 4(16.00 amu) = 203.79 amu

Mass fraction of Fe = $\dfrac{2(55.85 \text{ amu})}{203.79 \text{ amu}}$ = 0.5481

Mass fraction of Si = $\dfrac{28.09 \text{ amu}}{203.79 \text{ amu}}$ = 0.1378

Mass fraction of O = $\dfrac{4(16.00 \text{ amu})}{203.79 \text{ amu}}$ = 0.3140

Molecular mass of Mg_2SiO_4 = 2(24.31 amu) + 28.09 amu + 4(16.00 amu) = 140.71 amu

Mass fraction of Mg = $\dfrac{2(24.31 \text{ amu})}{140.71 \text{ amu}}$ = 0.3455

Mass fraction of Si = $\dfrac{28.09 \text{ amu}}{140.71 \text{ amu}}$ = 0.1996

Mass fraction of O = $\dfrac{4(16.00 \text{ amu})}{140.71 \text{ amu}}$ = 0.4548

Molecular mass of SiO_2 = 28.09 amu + 2(16.00 amu) = 60.09 amu

Mass fraction of Si = $\dfrac{28.09 \text{ amu}}{60.09 \text{ amu}}$ = 0.4675

Mass fraction of O = $\dfrac{2(16.00 \text{ amu})}{60.09 \text{ amu}}$ 2(16.00 amu) ÷ 60.09 amu = 0.5325

Mass percent Fe = (0.050)(0.5481)(100%) = **2.7% Fe**
Mass percent Mg = (0.070)(0.3455)(100%) = **2.4% Mg**
Mass percent Si = [(0.050)(0.1378) + (0.070)(0.1996) + (0.880)(0.4675)](100%) = **43.2% Si**
Mass percent O = [(0.050)(0.3140) + (0.070)(0.4548) + (0.880)(0.5325)](100%) = **51.6% O**

2.102 Plan: To find the formula mass of potassium fluoride, add the atomic masses of potassium and fluorine. Fluorine has only one naturally occurring isotope, so the mass of this isotope equals the atomic mass of fluorine. The atomic mass of potassium is the weighted average of the two isotopic masses: (isotopic mass of isotope 1 x fractional abundance) + (isotopic mass of isotope 2 x fractional abundance).
Solution:

Average atomic mass of K =

(isotopic mass of ^{39}K x fractional abundance) + (isotopic mass of ^{41}K x fractional abundance)

$$\text{Average atomic mass of K} = (38.9637\ \text{amu})\left(\frac{93.258\%}{100\%}\right) + (40.9618\ \text{amu})\left(\frac{6.730\%}{100\%}\right) = 39.093\ \text{amu}$$

The formula for potassium fluoride is KF, so its molecular mass is (39.093 + 18.9984) = **58.091 amu**

2.103 Plan: One molecule of NO is released per atom of N in the medicine. Divide the total mass of NO released by the molecular mass of the medicine and multiply by 100 for mass percent. Solution:

NO = (14.01 + 16.00) amu = 30.01 amu

Nitroglycerin:

$C_3H_5N_3O_9$ = 3(12.01 amu C) + 5(1.008 amu H) + 3(14.01 amu N) + 9(16.00 amu O) = 227.10 amu

In $C_3H_5N_3O_9$ (molecular mass = 227.10 amu), there are 3 atoms of N; since 1 molecule of NO is released per atom of N, this medicine would release 3 molecules of NO. The molecular mass of NO = 30.01 amu.

$$\text{Mass percent of NO} = \frac{\text{total mass of NO}}{\text{mass of compound}}(100) = \frac{3(30.01\ \text{amu})}{227.10\ \text{amu}}(100) = 39.6433 = \mathbf{39.64\%}$$

Isoamyl nitrate:

$C_5H_{11}NO_3$ = 5(12.01 amu C) + 11(1.008 amu H) + 1(14.01 amu N) + 3(16.00 amu O) = 133.15 amu

In $(CH_3)_2CHCH_2CH_2ONO_2$ (molecular mass = 133.15 amu), there is one atom of N; since 1 molecule of NO is released per atom of N, this medicine would release 1 molecule of NO.

$$\text{Mass percent of NO} = \frac{\text{total mass of NO}}{\text{mass of compound}}(100) = \frac{1(30.01\ \text{amu})}{133.15\ \text{amu}}(100) = 22.5385 = \mathbf{22.54\%}$$

2.104 Plan: First, count each type of atom present to produce a molecular formula. Determine the mass fraction of each element. Mass fraction = $\frac{\text{total mass of the element}}{\text{molecular mass of TNT}}$. The mass of TNT multiplied by the mass fraction of each element gives the mass of that element.

Solution:

The molecular formula for TNT is $C_7H_5O_6N_3$. The molecular mass of TNT is:

C	=	7(12.01 amu)	=	84.07 amu
H	=	5(1.008 amu)	=	5.040 amu
O	=	6(16.00 amu)	=	96.00 amu
N	=	3(14.01 amu)	=	42.03 amu
				227.14 amu

The mass fraction of each element is:

$$C = \frac{84.07\ \text{amu}}{227.14\ \text{amu}} = 0.3701\ C \qquad H = \frac{5.040\ \text{amu}}{227.14\ \text{amu}} = 0.02219\ H$$

$$O = \frac{96.00\ \text{amu}}{227.14\ \text{amu}} = 0.4226\ O \qquad N = \frac{42.03\ \text{amu}}{227.14\ \text{amu}} = 0.1850\ N$$

Masses of each element in 1.00 lb of TNT = mass fraction of element x 1.00 lb.

Mass (lb) C = 0.3701 x 1.00 lb = **0.370 lb C**

Mass (lb) H = 0.02219 x 1.00 lb = **0.0222 lb H**

Mass (lb) O = 0.4226 x 1.00 lb = **0.423 lb O**

Mass (lb) N = 0.1850 x 1.00 lb = **0.185 lb N**

2.105 Plan: Determine the mass percent of platinum by dividing the mass of Pt in the compound by the molecular mass of the compound and multiplying by 100. For part b), divide the total amount of money available by the cost of Pt per gram to find the mass of Pt that can be purchased. Use the mass percent of Pt to convert from mass of Pt to mass of compound.

Solution:

a) The molecular formula for platinol is $Pt(NH_3)_2Cl_2$. Its molecular mass is:

Pt	=	1(195.1 amu)	=	195.1 amu
N	=	2 (14.01 amu)	=	28.02 amu
H	=	6(1.008 amu)	=	6.048 amu
Cl	=	2(35.45 amu)	=	70.90 amu
				300.1 amu

$$\text{Mass \% Pt} = \frac{\text{mass of Pt}}{\text{molecular mass of compound}}(100) = \frac{195.1\text{ amu}}{300.1\text{ amu}}(100) = 65.012 = \mathbf{65.01\%\ Pt}$$

$$\text{b) Mass (g) of Pt} = \left(\$1.00 \times 10^6\right)\left(\frac{1\text{ g Pt}}{\$32}\right) = 31{,}250\text{ g Pt}$$

$$\text{Mass (g) of platinol} = \left(31{,}250\text{ g Pt}\right)\left(\frac{100\text{ g platinol}}{65.01\text{ g Pt}}\right) = 4.8070\text{x}10^4 = \mathbf{4.8x10^4\ g\ platinol}$$

2.106 Plan: A change is physical when there has been a change in physical form but not a change in composition. In a chemical change, a substance is converted into a different substance.

Solution:

1) Initially, all the molecules are present in blue-blue or red-red pairs. After the change, there are no red-red pairs, and there are now red-blue pairs. Changing some of the pairs means there has been a **chemical change**.

2) There are two blue-blue pairs and four red-blue pairs both before and after the change, thus no chemical change occurred. The different types of molecules are separated into different boxes. This is a **physical change**.

3) The identity of the box contents has changed from pairs to individuals. This requires a **chemical change**.

4) The contents have changed from all pairs to all triplets. This is a change in the identity of the particles, thus, this is a **chemical change**.

5) There are four red-blue pairs both before and after, thus there has been no change in the identity of the individual units. There has been a **physical change**.