Chapter 3
Chemical Compounds

The goal of this chapter is to introduce the different types of chemical compounds normally encountered by chemists—molecular compounds, organic compounds (which are usually molecular compounds), and ionic compounds. This chapter also explains how to name these compounds, and describes some of their physical and chemical properties. The chapter also revisits molecular weights and molar masses, and uses them to convert between percent composition by mass values and empirical and molecular formulas. Finally, the chapter introduces several biologically important elements.

Interpreting Molecular, Condensed, and Structural Formulas (3.1)

Chemists use a variety of formulas to describe chemical compounds. The simplest formula, called the molecular formula, is simply a list of the type and number of atoms present in the compound. The molecular formula of isooctane (the chemical used to determine a gasoline’s octane rating) shows that isooctane contains 8 carbon atoms and 18 hydrogen atoms, but does not show how these atoms are interconnected. While the condensed formula also contains subscripts showing the number of atoms present, it also shows how the atoms are connected together. The condensed formula of isooctane shows that one of the carbon atoms has three -CH3 groups attached to it, another carbon atom has two H atoms attached to it, and a third carbon atom has one H atom and two -CH3 groups attached to it. A structural formula shows the atom connections using lines between each of the atoms that are attached to each other.

```
  molecular formula: C8H18  
  structural formula:
     H     H     H     H     H     H     H     H     H     H
  condensed formula: (CH3)3CCH2CH(CH3)2
```

In each of these formulas, there is a trade-off between ease of writing and the amount of useful information provided. Molecular formulas are used when information about structural connections is not needed since they are easy to write; when more structural information is needed, the other two formulas are used.

Naming Binary Molecular Compounds (3.2 & 3.3)

Binary molecular compounds are compounds made of two types of atoms (binary) that form molecules. In binary molecular compounds, the two types of atoms are nonmetals (or sometimes metalloids). The general rule for naming binary compounds is to list the name of the first element and then list the name of the second element, replacing its ending with the ‘-ide’ suffix.

The problem with naming binary molecular compounds this way is that the same nonmetals can combine into different chemicals—for example, nitrogen and oxygen can form N2O, NO, N2O3, NO2, N2O4, and N2O5, but you can’t name them all “nitrogen oxide”! To alleviate this problem, there is an additional rule for naming binary molecular compounds: Before each element name, write a Greek prefix denoting how many atoms of each element are present in the molecule. The prefixes from 1 to 10 are: ‘mono’ (1), ‘di’ (2), ‘tri’ (3), ‘tetra’ (4), ‘penta’ (5), ‘hexa’ (6), ‘hepta’ (7), ‘octa’ (8), ‘nona’ (9), and ‘deca’ (10). You should probably memorize these prefixes because you will be using them frequently.

Example: Write the names for the following compounds: (a) P4S7, (b) SeF6, (c) Cl2O
Solution:
(a) The name for P₄S₇ is tetraphosphorus heptasulfide. The ‘tetra-’ tells you there are 4 P atoms, and the ‘hepta-’ tells you there are 7 S atoms (sulfide is named by dropping ‘-ur’ and replacing it with ‘-ide’).
(b) The name for SeF₆ is selenium hexafluoride. The prefix ‘mono-’ is almost always omitted, and if a prefix is missing it is assumed to be ‘mono-’. The ‘hexa-’ tells you there are 6 F atoms, and the ‘-ine’ ending on fluorine is changed to ‘-ide’ for fluoride.
(c) The name for Cl₂O is dichlorine monoxide. The ‘di-’ tells you there are 2 Cl atoms. When oxygen is the last element (replacing the ‘-ygen’ with ‘-ide’ to make oxide), ‘mono-’ is usually used. This is one of the few examples where ‘mono-’ is included. When oxygen is used and the prefix ends in ‘o’ or ‘a’, this last letter is usually dropped (i.e., monoxide instead of monooxide, pentoxide instead of pentoioxide, etc.).

There are some binary compounds that have common names that are used instead of the systematic names. The two most common examples are water (H₂O) and ammonia (NH₃).

The rules described here are not used for binary compounds of C and H. These compounds are usually referred to as hydrocarbons (this name simply means ‘containing carbon and hydrogen’). The simplest hydrocarbons are called alkanes and have the general formula CₙH₂ₙ₊₂. The rule for naming alkanes is to use a prefix to show the number of carbon atoms and an ‘-ane’ ending. The prefixes for 1 to 4 C atoms are: ‘meth-’ (1), ‘eth-’ (2), ‘prop-’ (3), and ‘but-’ (4). The prefixes for 5 to 10 C atoms are the same as the Greek prefixes used in naming binary molecular compounds with the ‘a’ at the end of the prefixes dropped. So, CH₄ (1 C) is methane, C₃H₈ (3 C) is propane, C₅H₁₂ (5 C) is pentane, C₇H₁₆ (7 C) is heptane, etc.

Constitutional Isomers of Alkanes (3.4)
Isomers are two or more compounds that have the same molecular formula (same number and type of atoms) but have different structural formulas (the atoms within the molecules are bonded together in different ways). Isomers also have different physical and chemical properties. For alkanes, there are two basic forms of constitutional (structural) isomers. Straight-chain alkanes have the carbon atoms bonded together in a single chain, while branched alkanes do not. In straight-chain alkanes, you can draw a single path through all of the carbon atoms; you cannot draw a single path that goes through every carbon atom in a branched alkane. Also, branched alkanes contain a carbon atom bonded to three or four other carbon atoms (straight-chain alkanes have carbon atoms bonded to only one or two other carbon atoms).

Example: Write the structural formulas for the straight-chain isomer and a branched isomer of pentane.

Solution: The formula for pentane is C₅H₁₂ (CₙH₂ₙ₊₂, where \( n = 5 \)). The straight-chain alkane has the five C atoms in a single path. There are two branched alkane isomers for pentane. One of them has a C atom with three other C atoms attached to it, and the other has a C atom with four other C atoms attached to it.

Predicting the Charges of Monatomic Ions (3.5)
Metal atoms tend to form cations (positively-charged ions) by losing electrons, while nonmetals tend to form anions (negatively-charged ions) by gaining electrons. The charge of these monatomic ions can be determined from the number of protons and electrons in the ion: \( \text{Charge} = \#p^+ - \#e^- \). Metal cations are
named using the element name plus the word ‘ion’. So, Na⁺ is called the sodium ion and Ca²⁺ is called the calcium ion. For main group metals (metals in groups with an ‘A’ in their column number), the positive charge of the metal cation is equal to the metal’s column number. So, alkali metal ions (group 1A) have +1 charges (Li⁺, Na⁺, K⁺, etc.), alkaline earth metal ions (group 2A) have +2 charges (Mg²⁺, Ca²⁺, Sr²⁺, etc.), and so on. The metals at the bottom of groups 4A and 5A are an exception to this rule—they have positive charges equal to their column number minus two (Sn²⁺ and Pb²⁺ for 4A metals, and Bi³⁺ for the 5A metal).

Many of the transition elements can form more than one stable cation (e.g., copper can form Cu⁺ and Cu²⁺ ions). Since you cannot name both of these ions the “copper ion”, you must distinguish between these two ions by identifying the positive charge on the ion using Roman numerals and parentheses. So, Cu⁺ is called the copper(I) ion and Cu²⁺ is called the copper(II) ion.

For the nonmetals, the negative charge of the nonmetal anions equals eight minus the column number. So, the halide ions (group 7A) have –1 charges (F⁻, Cl⁻, Br⁻, etc.), group 6A ions have –2 charges (O²⁻, S²⁻, etc.), and so on. Nonmetal anions are named by replacing the ending of the elements name with ‘-ide ion’. So, Br⁻ is called the bromide ion and N₃⁻ is called the nitride ion.

**Common Polyatomic Ions (3.5)**

Polyatomic ions are simply ions (charged objects) that have more than one atom in them (polyatomic). There are no systematic rules for naming these ions, so you will simply have to memorize the names and formulas (including charges) of the more common polyatomic ions. Table 3.7 in your textbook (p. 88) has a list of common polyatomic ions; the most commonly used ions are: the acetate ion (CH₃COO⁻), the ammonium ion (NH₄⁺), the carbonate ion (CO₃²⁻), the cyanide ion (CN⁻), the hydroxide ion (OH⁻), the nitrate ion (NO₃⁻), the phosphate ion (PO₄³⁻), and the sulfate ion (SO₄²⁻).

**Properties of Ionic and Molecular Compounds (3.5 & 3.7)**

Molecular compounds, which are made of two or more nonmetal atoms, consist of individual molecules that are weakly attracted to each other. Because of this weak attraction, these compounds tend to be gases or liquids at room temperature (low melting and boiling points), and are not particularly hard solids (easily deformed). Because these compounds are neutral and have no positive or negative ions, they are poor electrical conductors (the flow of electricity is simply the flow of charged particles, like ions or electrons). Examples include H₂(g), CO₂(g), H₂O(ℓ), C₆H₁₄(ℓ), and C₁₂H₂₂O₁₁(s).

Ionic compounds, which are usually made from a combination of metal and nonmetal atoms, consist of cations and anions that are strongly attracted to each other through ionic bonding (the attraction of positive and negative ions). Unlike molecular compounds, ionic compounds do not have individual molecules in them. Instead, all of the ions in an ionic solid can be viewed as a single, incredibly large “molecule”. Because ionic compounds have strong attractions holding the ions together, they tend to be solids at room temperature (high melting and boiling points) and are very hard solids. Because ionic compounds have charged particles in them, they can conduct electricity if the ions are allowed to move. That is why ionic compounds are terrible electrical conductors in their solid form (the ions can’t move), but are good electrical conductors when the ions are allowed to move (e.g., in its liquid or melted state, and when dissolved in water). Examples include NaCl(s), Fe₂O₃(s), and NH₄CH₃COO(s).

**Naming Ionic Compounds (3.5 & 3.6)**

Ionic compounds are named by listing the name of the cation first and the name of the anion last (omitting the word ‘ion’). When naming an ionic compound, you need to determine the cation and anion present in the compound. As an example, what is the name of BaCO₃? You may not recognize Ba, but CO₃²⁻ is the carbonate ion (a polyatomic ion you should memorize). Since the overall charge of any compound must be zero, the Ba ion must have a +2 charge to balance the –2 charge on the carbonate ion. Ba²⁺ is the barium ion (column 2A), and its charge matches the rule (the positive charge of main group metals is equal to its column number). Since the compound is made of barium ions and carbonate ions, its name is barium
carbonate (cation name first, anion name last). Similarly, an ionic compound made from the aluminum ion and the iodide ion would be called aluminum iodide. What is the chemical formula for aluminum iodide?

When writing the formula of an ionic compound, the rule is that the overall charge of the compound must be zero (neutral). Since aluminum (Al) appears in column 3A, the charge for the aluminum cation must be +3 (Al^{3+}). Since iodine (the name of the element, I) appears in column 7A, the charge for the iodide ion must be −1 (I^−; if the number is missing before a ‘+’ or ‘−’ sign, it is assumed to be 1). In order for aluminum iodide to be neutral, the total positive charge must be equal to the total negative charge. So, there must be 3 I^- ions for every 1 Al^{3+} ion, or AlI_3 (cations are always listed first).

Example: Write the name or formula for the following ionic compounds.
(a) Ca_3(PO_4)_2, (b) CrN, (c) lead(II) nitrate, (d) strontium hydride

Solution:
(a) The ions present are the calcium ion (Ca^{2+}) and the phosphate ion (PO_4^{3−}), so this is calcium phosphate. Remember that the Greek prefixes are not used in ionic compounds, only molecular compounds.
(b) The ions present are the chromium(III) ion (Cr^{3+}) and the nitride ion (the ‘-ide’ ending tells you this is a monatomic anion of nitrogen; since N is in column 5A, its negative charge is 8 − 5 = 3 or N^{3−}), so this is chromium(III) nitride. Since chromium can have multiple charges (+2 or +3), the Roman numeral in parentheses tells you that the charge of the chromium ion in this compound is +3 (Cr^{3+}).
(c) The formula for the lead(II) ion is Pb^{2+} (the Roman numeral tells you its charge is +2), and the formula for nitrate is NO_3^{−}. The neutral formula is Pb(NO_3)_2 (1 Pb^{2+} for every 2 NO_3^{−}).
(d) The formula for strontium ion is Sr^{2+} (Sr is in column 2A, so its charge is +2), and the formula for hydride is H^{-} (the ‘-ide’ ending tells you this is a monatomic anion of hydrogen; since H can appear in column 7A, its negative charge is 8 − 7 = 1). The neutral formula is SrH_2 (1 Sr^{2+} for every 2 H^{-}).

Electrolytes and Nonelectrolytes (3.7)

Electrolytes are compounds whose aqueous solutions conduct electricity. In order for the solutions to conduct electricity, they must contain charged particles (ions). So, ionic compounds dissolved in water are expected to be electrolytes because they undergo dissociation into ions, but molecular compounds would not because they do not produce ions. The fact that aqueous solutions of ionic compounds conduct electricity shows that the ions in these solution have the freedom to move independently—if they were stuck together, then it would still be ‘neutral’ and the solution would not conduct electricity. Compounds whose aqueous solutions do not conduct electricity are called nonelectrolytes. Compounds that do not dissolve in water, or that do not produce ions when dissolved in water, are expected to be nonelectrolytes.

The Mole Concept (3.8)

The concept of the mole was introduced in the previous chapter of the textbook (section 2.7). The mole is simply the number of atoms present in exactly 12 g of carbon-12. This definition was made so that the masses listed in the periodic table for every element are both the atomic weights (in units of amu/atom) and the molar mass (in units of g/mol). The numerical value of a mole, called Avogadro’s number, is 6.022 × 10^{23} ‘things’/mol (the ‘thing’ can be anything—atoms, molecules, ions, etc.).

Calculating Molar Masses (3.8)

Calculating the molar mass of a compound is done by adding up the individual molar masses of the elements present in the compound. For example, the molar mass of sucrose (C_{12}H_{22}O_{11}) is calculated by adding up the molar mass of 12 carbon atoms, 22 hydrogen atoms, and 11 oxygen atoms.

Example: What is the molar mass of tetraethyl lead, Pb(C_2H_5)_4, the additive that made gasoline ‘leaded’?
**Solution:** The formula is C₈H₂₀Pb, and the molar mass would be the mass of 1 mol of C₈H₂₀Pb.

\[
1 \text{ mol } \text{C}_{8}\text{H}_{20}\text{Pb} = 8 \text{ mol C} \cdot \frac{12.01 \text{ g C}}{\text{mol C}} + 20 \text{ mol H} \cdot \frac{1.008 \text{ g H}}{\text{mol H}} + 1 \text{ mol Pb} \cdot \frac{207.2 \text{ g Pb}}{\text{mol Pb}} = 323.4 \text{ g C}_{8}\text{H}_{20}\text{Pb}
\]

So, the molar mass of tetraethyl lead (used as a conversion factor) is 323.4 g C₈H₂₀Pb / mol C₈H₂₀Pb.

**Converting Between Mass and Mole Values (3.8)**

Molar masses of compounds, just like molar masses of elements (introduced in section 2.8 of the textbook), are mathematical equalities. Using the example for tetraethyl lead (above), you can see that the mathematical equality is 1 mol C₈H₂₀Pb = 323.4 g C₈H₂₀Pb, and a conversion factor can be made from this molar mass value with either the mass value on top and the mole value on bottom.

**Example:** You have a sample of 32.5 g of bromine pentafluoride. How many moles of bromine pentafluoride do you have? What is the mass (in g) of the same amount of disulfur tetranitride?

**Solution:** Before you go any further, you must determine the formulas of these two compounds. The formula for bromine pentafluoride is BrF₅, and the formula for disulfur tetranitride is S₂N₄. To answer the first question, you need to determine the molar mass of BrF₅ and use it to convert 32.5 g BrF₅ to mol BrF₅.

\[
1 \text{ mol BrF}_5 = 1 \text{ mol Br} \cdot \frac{79.90 \text{ g Br}}{\text{mol Br}} + 5 \text{ mol F} \cdot \frac{19.00 \text{ g F}}{\text{mol F}} = 174.90 \text{ g BrF}_5
\]

\[
32.5 \text{ g BrF}_5 \cdot \frac{1 \text{ mol BrF}_5}{174.90 \text{ g BrF}_5} = 0.186 \text{ mol BrF}_5
\]

For the second question, determine the molar mass of S₂N₄ and use it to convert mol S₂N₄ to g S₂N₄.

\[
1 \text{ mol S}_2\text{N}_4 = 2 \text{ mol S} \cdot \frac{32.07 \text{ g S}}{\text{mol S}} + 4 \text{ mol N} \cdot \frac{14.01 \text{ g N}}{\text{mol N}} = 120.18 \text{ g S}_2\text{N}_4
\]

\[
0.186 \text{ mol S}_2\text{N}_4 \cdot \frac{120.18 \text{ g S}_2\text{N}_4}{\text{mol S}_2\text{N}_4} = 22.3 \text{ g S}_2\text{N}_4
\]

**Hydrated Ionic Compounds (3.8)**

Some ionic compounds have water molecules trapped within the solid. These solids are referred to as ionic hydrates, and the water molecules trapped in the solid are called waters of hydration. Common washing soda has two sodium ions (Na⁺), one carbonate ion (CO₃²⁻), and ten water molecules in its formula unit. Its formula is Na₂CO₃⋅10H₂O, but could also be written Na₂CO₃(H₂O)₁₀. Hydrated ionic compounds are named the same way as the anhydrous compound (the compound without the waters of hydration), plus a Greek prefix indicating the number of waters of hydration and the word ‘hydrate’. So, Na₂CO₃⋅10H₂O is named sodium carbonate decahydrate.

**Determining Percent Composition by Mass (3.9)**

Percent composition by mass is a list of the percentages (by mass) of the elements present in a compound. It is also referred to as an elemental analysis, and can be used to identify the molecular formula of an unknown chemical compound. To determine whether a compound has the molecular formula you think it has, you can calculate the percent composition by mass data from the proposed formula and compare these values to the actual (measured) values. The formula for percent composition by mass of element X is:

\[
\% X = \frac{m_X}{m_{\text{total}}} \times 100\%
\]
Example: You have collected a sample of a suspicious chemical and sent it out to a commercial laboratory for elemental analysis. The percent composition by mass data returned to you by the laboratory are 40.0% C, 6.7% H, and 53.3% O. Is this compound more likely to be glucose (C₆H₁₂O₆), a common sugar used by the human body, or formaldehyde (CH₂O), a known carcinogen?

Solution: You need to determine the percent composition by mass for glucose and for formaldehyde. For each sample, assume that you have exactly 1 mol of the chemical, and determine the percent composition values for each element based on those masses.

\[ 1 \text{ mol } C₆H₁₂O₆ = 6 \text{ mol C} \left[ \frac{12.01 \text{ g C}}{\text{mol C}} \right] + 12 \text{ mol H} \left[ \frac{1.008 \text{ g H}}{\text{mol H}} \right] + 6 \text{ mol O} \left[ \frac{16.00 \text{ g O}}{\text{mol O}} \right] = 180.16 \text{ g C₆H₁₂O₆} \]

\[ \% \text{ C} = \frac{6 \text{ mol C} \times \frac{12.01 \text{ g C}}{\text{mol C}}}{180.16 \text{ g C₆H₁₂O₆}} \times 100\% = 40.00\% \text{ C} \]

\[ \% \text{ H} = \frac{12 \text{ mol H} \times \frac{1.008 \text{ g H}}{\text{mol H}}}{180.16 \text{ g C₆H₁₂O₆}} \times 100\% = 6.714\% \text{ H} \]

\[ \% \text{ O} = \frac{6 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}}}{180.16 \text{ g C₆H₁₂O₆}} \times 100\% = 53.29\% \text{ O} \]

\[ 1 \text{ mol CH₂O} = 1 \text{ mol C} \left[ \frac{12.01 \text{ g C}}{\text{mol C}} \right] + 2 \text{ mol H} \left[ \frac{1.008 \text{ g H}}{\text{mol H}} \right] + 1 \text{ mol O} \left[ \frac{16.00 \text{ g O}}{\text{mol O}} \right] = 30.03 \text{ g CH₂O} \]

\[ \% \text{ C} = \frac{1 \text{ mol C} \times \frac{12.01 \text{ g C}}{\text{mol C}}}{30.03 \text{ g CH₂O}} \times 100\% = 40.00\% \text{ C} \]

\[ \% \text{ H} = \frac{2 \text{ mol H} \times \frac{1.008 \text{ g H}}{\text{mol H}}}{30.03 \text{ g CH₂O}} \times 100\% = 6.714\% \text{ H} \]

\[ \% \text{ O} = \frac{1 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}}}{30.03 \text{ g CH₂O}} \times 100\% = 53.29\% \text{ O} \]

Even though glucose and formaldehyde have different chemical formulas (as well as different chemical and physical properties), they have the same percent composition by mass values. This is because they have the same relative ratios of atoms (C:H:O in a ratio of 1:2:1). Because all chemicals with the same atom ratios have the same percent composition by mass values, the compound’s molecular formula provides more information than the percent composition by mass values.

Determining Empirical and Molecular Formulas (3.10)

The percent composition by mass data measured in a chemistry laboratory consist of masses, but molecular formulas are written based on numbers of atoms (moles). In order to convert percent composition by mass data into a chemical formula (called the empirical formula), you need to use the molar masses of these elements to convert the mass data to mole data. The empirical formula represents the simplest ratio of atoms present in a compound. To determine the molecular formula (the actual chemical formula) you need to know the approximate molar mass of the compound.
Example: An ionic compound was sent for elemental analysis, and the laboratory reported that the chemical was 47.0% K, 14.5% C, and 38.5% O. What is the empirical formula of this compound? The molar mass of the compound was estimated to be about 170 g/mol. What is the molecular formula of this compound?

Solution: In order to convert the percent values into moles, assume that you have 100 g of the compound, and convert these masses to moles using each element’s molar mass. This assumption is made so that the percent values are now masses in g (in 100 g of sample, there is 47.0 g K, 14.5 g C, and 38.5 g O).

\[
\begin{align*}
47.0 \text{ g K} & \quad \frac{\text{mol K}}{39.10 \text{ g K}} = 1.20205 \text{ mol K (unrounded)} \\
14.5 \text{ g C} & \quad \frac{\text{mol C}}{12.01 \text{ g C}} = 1.20733 \text{ mol C (unrounded)} \\
38.5 \text{ g O} & \quad \frac{\text{mol O}}{16.00 \text{ g O}} = 2.40625 \text{ mol O (unrounded)}
\end{align*}
\]

The ratios of K:C:O are 1.20205:1.20733:2.40625, and you could write the formula \( \text{K}^{1.20205}\text{C}^{1.20733}\text{O}^{2.40625} \). However, molecular formulas are written with whole number subscripts because you can’t have fractions of K, C, or O atoms. To find the whole number ratios in the empirical formula, divide each subscript by the smallest subscript value.

\[
\begin{align*}
\text{K}_{1.20205} & \quad \text{C}_{1.20733} \quad \text{O}_{2.40625} \\
\frac{1}{1.20205} & \quad \frac{1}{1.20733} \quad \frac{1}{2.40625} \\
\text{K}^{1.0000} & \quad \text{C}^{1.0040} \quad \text{O}^{2.0020} = \text{KCO}_2
\end{align*}
\]

The empirical formula of this compound is KCO₂, but any whole-number multiple of this formula (K₂C₂O₄, K₃C₃O₆, etc.) could also be correct. To find the molecular formula (which is really a formula unit in this case, since this is an ionic compound and not a molecular compound), you need to compare the empirical mass to the molar mass of the compound, recognizing that the molar mass is some whole number times the empirical mass.

\[
\begin{align*}
M_{\text{molecular}} &= n \cdot M_{\text{empirical}}, \\
n &= \frac{M_{\text{molecular}}}{M_{\text{empirical}}} \\
1 \text{ mol KCO}_2 &= 1 \text{ mol K} \frac{39.10 \text{ g K}}{\text{mol K}} + 2 \text{ mol C} \frac{12.01 \text{ g C}}{\text{mol C}} + 2 \text{ mol O} \frac{16.00 \text{ g O}}{\text{mol O}} = 83.11 \text{ g KCO}_2 \\
n &= \frac{170 \text{ g/mol}}{83.11 \text{ g/mol}} = 2.05 = 2
\end{align*}
\]

This means that the molecular formula is exactly 2 times the empirical formula: \( \text{KCO}_2 \cdot 2 = \text{K}_2\text{C}_2\text{O}_4 \). The ionic compound is actually potassium oxalate, consisting of potassium (K⁺) ions and oxalate (C₂O₄²⁻) ions.

For ionic compounds, the empirical formula is usually reported and a molecular formula is not calculated because the formula unit of an ionic compound is the simplest ratio of ions (although there are exceptions).

**Biologically Important Elements (3.11)**

Although four elements (H, O, C, and N) make up the bulk of the human body’s mass, there are many other elements that are still required for human life to exist, just in smaller quantities. For example, phosphorus (P) is needed to make DNA, iron (Fe) is needed to make blood, calcium (Ca) is needed to make bones and teeth, chlorine (Cl) is needed to digest food, iodine (I) is needed to keep the thyroid gland functioning, and so on. The term **dietary mineral** is used to describe the elements (except for H, O, C, and N) needed to keep the human body functioning. Dietary minerals can be divided into two categories—**major minerals** (present in quantities greater than 0.01% of the body’s total mass) and **trace elements** (present in quantities less than 0.01% of the body’s total mass).
Chapter Review — Key Terms

The key terms that were introduced in this chapter are listed below, along with the section in which they were introduced. You should understand these terms and be able to apply them in appropriate situations.

alkane (3.3)  empirical formula (3.10)  molecular formula (3.1)
alkyl group (3.4)  formula unit (3.7)  molecular weight (3.8)
anion (3.5)  formula weight (3.8)  monatomic ion (3.5)
binary molecular compound (3.2)  functional group (3.1)  nonelectrolyte (3.7)
cation (3.5)  halide ion (3.6)  organic compound (3.1)
chemical bond (3.3)  hydrocarbon (3.3)  oxoanion (3.6)
condensed formula (3.1)  inorganic compound (3.1)  percent composition by mass (3.9)
constitutional isomer (3.4)  ionic bonding (3.7)  polyatomic ion (3.5)
Coulomb’s law (3.7)  ionic compound (3.5)  structural formula (3.1)
 crystal lattice (3.7)  ionic hydrate (3.8)  trace element (3.11)
dietary mineral (3.11)  isomer (3.4)  water of hydration (3.8)
dissociation (3.7)  major mineral (3.11)
electrolyte (3.7)  molecular compound (3.1)

Practice Test

After you have finished studying the chapter and the homework problems, the following questions can serve as a test to determine how well you have learned the chapter objectives.

1. Complete the following chart of molecular, condensed, and structural formulas.

<table>
<thead>
<tr>
<th>molecular formula</th>
<th>condensed formula</th>
<th>structural formula</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>NH₂CH₂CH₂NH₂</td>
<td></td>
</tr>
<tr>
<td></td>
<td>(CH₃)₂CHCH₂C≡N</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>O—O</td>
</tr>
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<td>O—O</td>
</tr>
<tr>
<td></td>
<td></td>
<td>O—O</td>
</tr>
</tbody>
</table>

2. Consider the following structural formulas.

I  II  III  IV

(a) Can all of these structures be isomers?
(b) Are any of these structures identical (i.e., the same compound)?
(c) Which of these structures are constitutional isomers of each other?
3. Determine the systematic names for the following molecular compounds.
   (a) water, $\text{H}_2\text{O}$
   (b) ammonia, $\text{NH}_3$
   (c) hydrazine, $\text{N}_2\text{H}_4$
   (d) nitric oxide, NO
   (e) nitrous oxide, $\text{N}_2\text{O}$
   (f) phosphine, $\text{PH}_3$

4. For each of the elements listed below, predict the charge of its monatomic ion.
   (a) carbon
   (b) fluorine
   (c) iron
   (d) lithium
   (e) selenium
   (f) strontium

5. Predict the formulas of the ionic compounds made from the following components.
   (a) bromine and calcium
   (b) carbon and potassium
   (c) chromium and the hydroxide ion
   (d) magnesium and oxygen
   (e) zinc and the cyanide ion
   (f) sulfur and the ammonium ion

6. Complete the following chart of chemical compounds using the compound’s molecular formulas.
   For the ‘compound type’ column, your choices are molecular, organic, ionic, molecular hydrate, and ionic hydrate. For the ‘ions present’ column, the correct answer may be none.

<table>
<thead>
<tr>
<th>molecular formula</th>
<th>compound type</th>
<th>ions present</th>
<th>chemical name</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{B}_2\text{O}_3$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{C}_3\text{H}_8$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{CoCl}_2\cdot6\text{H}_2\text{O}$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{FeCO}_3$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{IF}_7$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{K}_2\text{Se}$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>$\text{S}_2\text{N}_6$</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

7. Complete the following chart of chemical compounds using the compound’s names (use the same choices for the ‘compound type’ column and the ‘ions present’ column as in question 6).

<table>
<thead>
<tr>
<th>molecular formula</th>
<th>compound type</th>
<th>ions present</th>
<th>chemical name</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td>aluminum sulfate</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>butane</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>chlorine trifluoride</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>iron(III) nitrate nonahydrate</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>magnesium hydroxide</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>nonane</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>sodium nitride</td>
</tr>
</tbody>
</table>
8. A compound of titanium and chlorine has the formula TiCl₄. This compound has a melting point of –25°C and is a liquid at room temperature, but does not conduct electricity. 
(a) Based on the physical properties of this compound, is it a molecular or ionic compound?  
(b) Given the elements in this compound, is this surprising?  
(c) Write the name of this compound, based on its physical properties.

9. Calcium phosphate, Ca₃(PO₄)₂, is a major component in fertilizers. 
(a) What is the molar mass of calcium phosphate?  
(b) What is the mass of 1.25 mol of calcium phosphate?  
(c) How many moles of phosphorus atoms are present in 100.0 g of calcium phosphate?  
(d) What is the percent phosphorus in this compound?

10. Calculate the percent sulfur by mass for the following compounds.  
(a) Hydrogen sulfide, H₂S (the smell of rotten eggs)  
(b) Iron pyrite, FeS₂ (fool’s gold)  
(c) Sulfur dioxide, SO₂ (a major component of acid rain)  
(d) Sulfuric acid, H₂SO₄ (the burn from cutting onions)

11. The elemental analysis of a dark-green lustrous chemical was found to be 33.3% Fe, 28.6% C, and 38.1% O.  
(a) What is the empirical formula for this compound?  
(b) If its molar mass is approximately 500 g/mol, what is its molecular formula?