GENERAL CHEMISTRY

Principles and Modern Applications

TENTH EDITION

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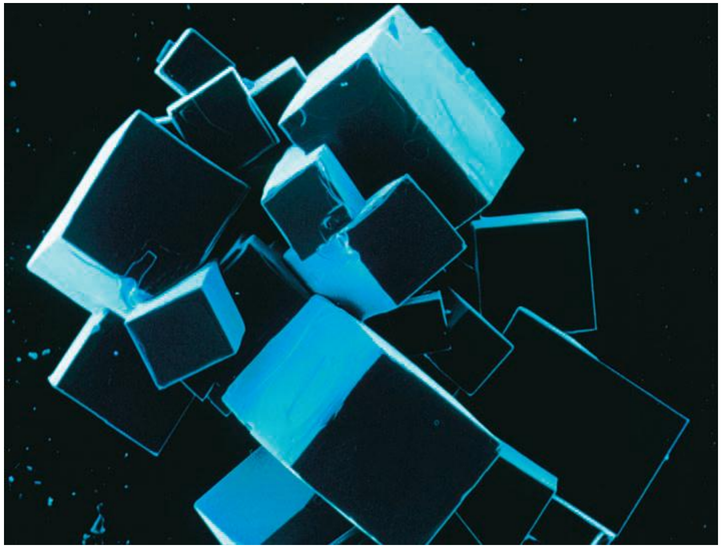
Chapter 3:

Chemical Compounds

CHAPTER :3 Chemical Compounds



Scanning electron microscope image of sodium chloride crystals.

CONTENTS

3-1 Types of Chemical Compounds and Their Formulas

- Two fundamental kinds of chemical bonds

  - *Covalent bonds, which involve a sharing of electrons* between atoms, give rise to molecular compounds.

  - *Ionic bonds, which involve a transfer of electrons* from one atom to another, give rise to ionic compounds.

Molecular Compounds

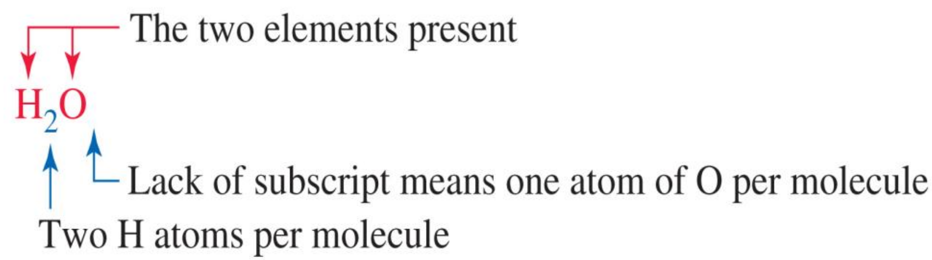
○ is made up of discrete units called molecules

○ a small number of *nonmetal atoms held together by covalent* bonds

○ are represented by **chemical formulas,** symbolic representations that, at minimum, indicate

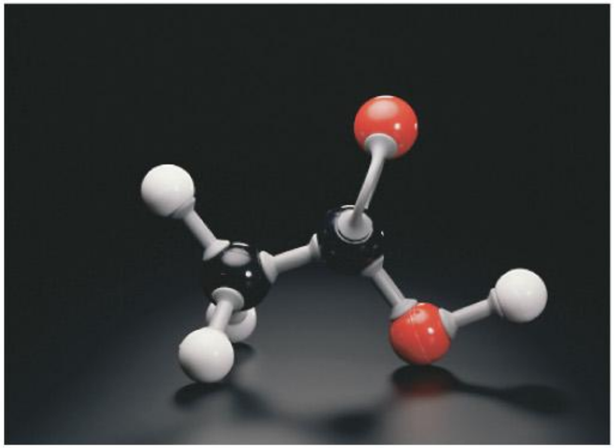
  ○the elements present

  ○the relative number of atoms of each element



CCl_4 : C and Cl (4) molecular compound

Molecular compounds

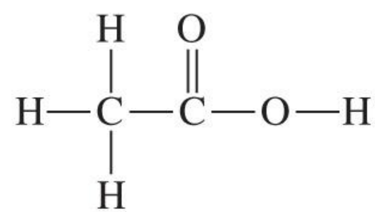


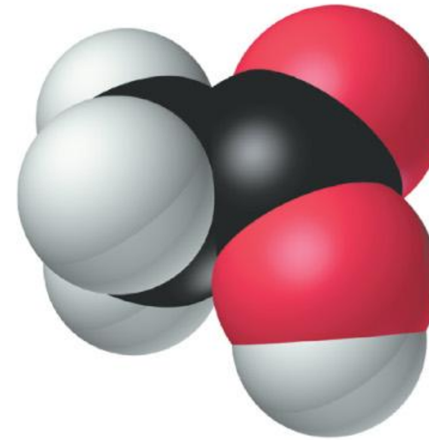
Molecular model:
("ball-and-stick")

Empirical formula: CH_2O

Molecular formula: $C_2H_4O_2$

Structural formula:





Molecular model:
("space-filling")

▲ Figure 3-1

Several representations of the compound acetic acid

Empirical formula:

- the simplest formula for a compound;
- shows the types of atoms present and their relative numbers.
- the subscripts in an empirical formula are reduced to their simplest whole-number ratio. P_2O_5 is an emp. form. for P_4O_{10}
  - Acetic acid, $C_2H_4O_2$
  - formaldehyde (used to make certain plastics and resins), CH_2O
  - and glucose $C_6H_{12}O_6$

  all have the same emprical formula of CH_2O

▲

Molecular formula:

- Is based on an actual molecule of a compound
- In some cases, the empirical and molecular formulas are identical, such as CH_2O for formaldehyde.
- In other cases, the molecular formula is a multiple of the empirical formula. E.g : Acetic acid, $C_2H_4O_2$, $(CH_2O)_2$

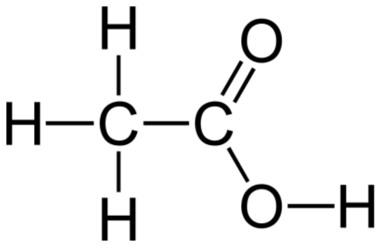
Empirical and molecular formulas tell us the combining ratio of the atoms in the compound, but they show nothing about how the atoms are attached to each other.

▲

Structural formula:

-shows the order in which atoms are bonded together in a molecule and by what types of bonds

4 H bonds – 3 (C-H), 1 (O-H)
2 O , 1 (C=O), 1 (C-O)



The covalent bonds in the structural formula are represented by lines or dashes (-) . One of the bonds is represented by a double dash (=) and is called a *double covalent bond,* stronger and tighter than a single bond

▲

Condensed Structural formula:

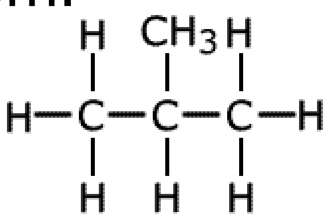
- written on a single line, is an alternative, less cumbersome way of showing how the atoms of a molecule are connected.

CH_3COOH or CH_3CO_2H

- Condensed structural formulas can also be used to show how a group of atoms is attached to another atom.

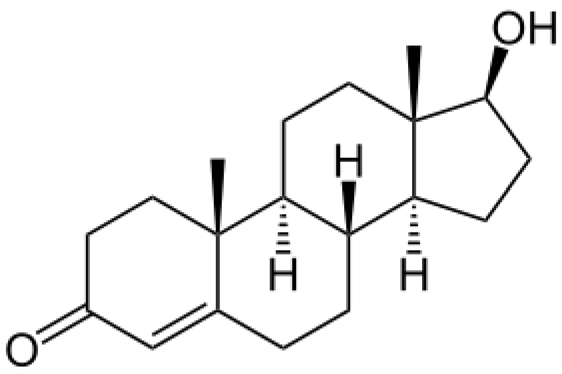
methylpropane, C_4H_{10} $CH_3CH(CH_3)CH_3$

CH(CH3)$_3$



line-angle formula (line structure)

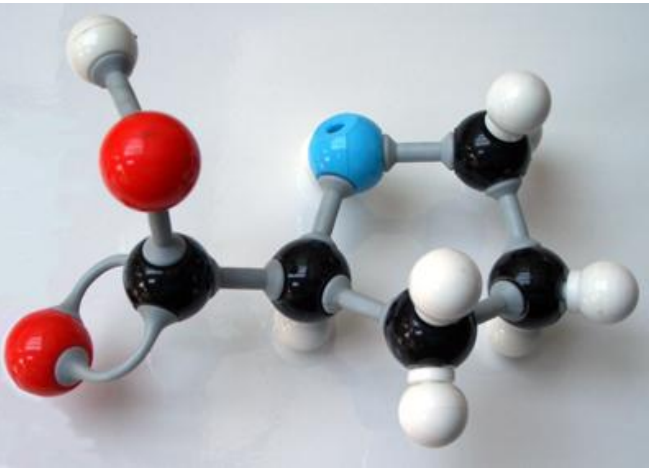
- used to represent organic compounds
- relatively quick and simple to draw
- lines represent chemical bonds.
- A carbon atom exists wherever a line ends or meets another line, and the number of H atoms needed to complete each carbon atoms four bonds are assumed to be present.
- The symbols of other atoms or groups of atoms and the bond lines joining them to C atoms are written explicitly.



Testosterone

Ball-and-stick models

- Molecules occupy space and have a three-dimensional shape, but empirical and molecular formulas do not convey any information about the spatial arrangements of atoms.



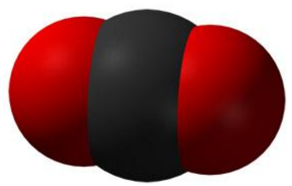
-Structural formulas can sometimes show this, but usually the only satisfactory way to represent the three-dimensional structure of molecules is with models.

- In a *ball-and-stick model,* atoms are represented by small balls, and the bonds between atoms by sticks.

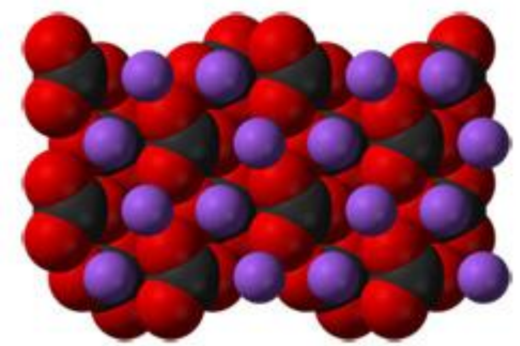
- Such models help us to visualize distances between the nuclei of atoms (bond lengths) and the geometrical shapes of molecules.

- Ball-and-stick models are easy to draw and interpret, but they can be somewhat misleading. Chemical bonds are forces that draw atoms in a molecule into direct contact. The atoms are not held apart as implied by a ball-and-stick model.

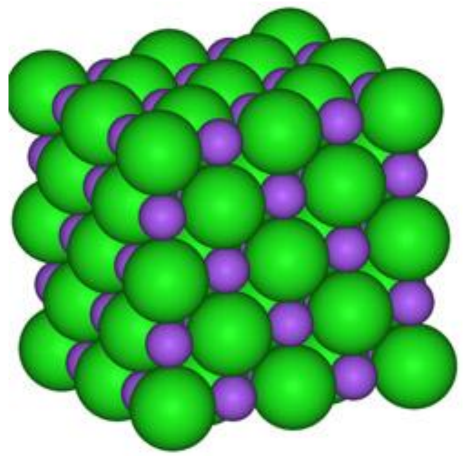
Space-filling model



- shows that the atoms in a molecule occupy space and that they are in actual contact with one another.

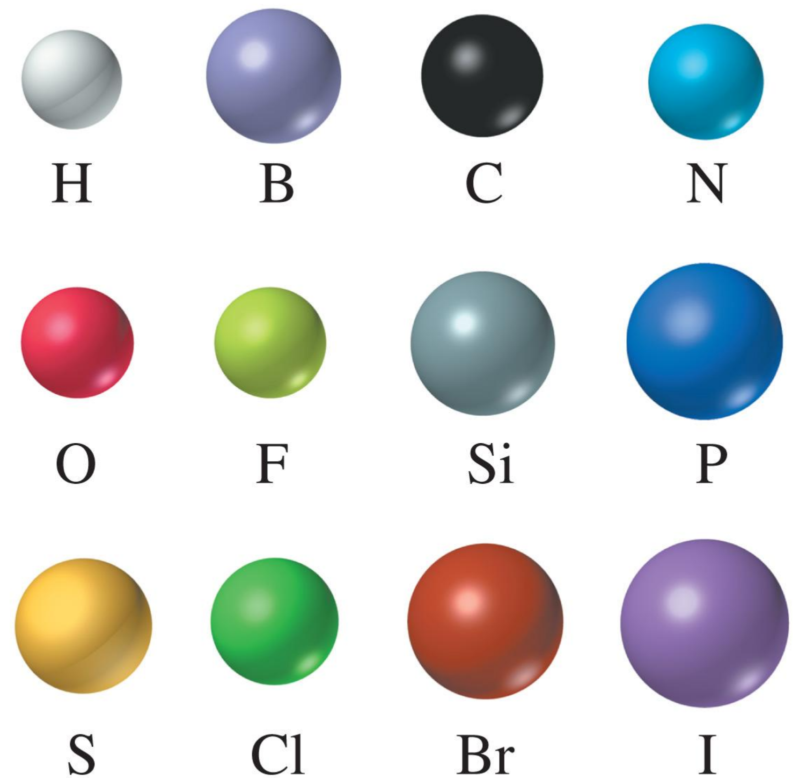


- certain computer programs generate images of space-filling models
- A space-filling model is a more accurate representation of the size and shape of a molecule because it is constructed to scale (that is, a nanometer size molecule is magnified to a millimeter or centimeter scale).



-Different colors are used to distinguish the various types of atoms in ball-and-stick and space-filling models

- the colored spheres are of different sizes, which correspond to the size differences between the various atoms in the periodic table.



▲ Figure 3-3
Color scheme for use in molecular models

Structural formula ball and stick model space filling model

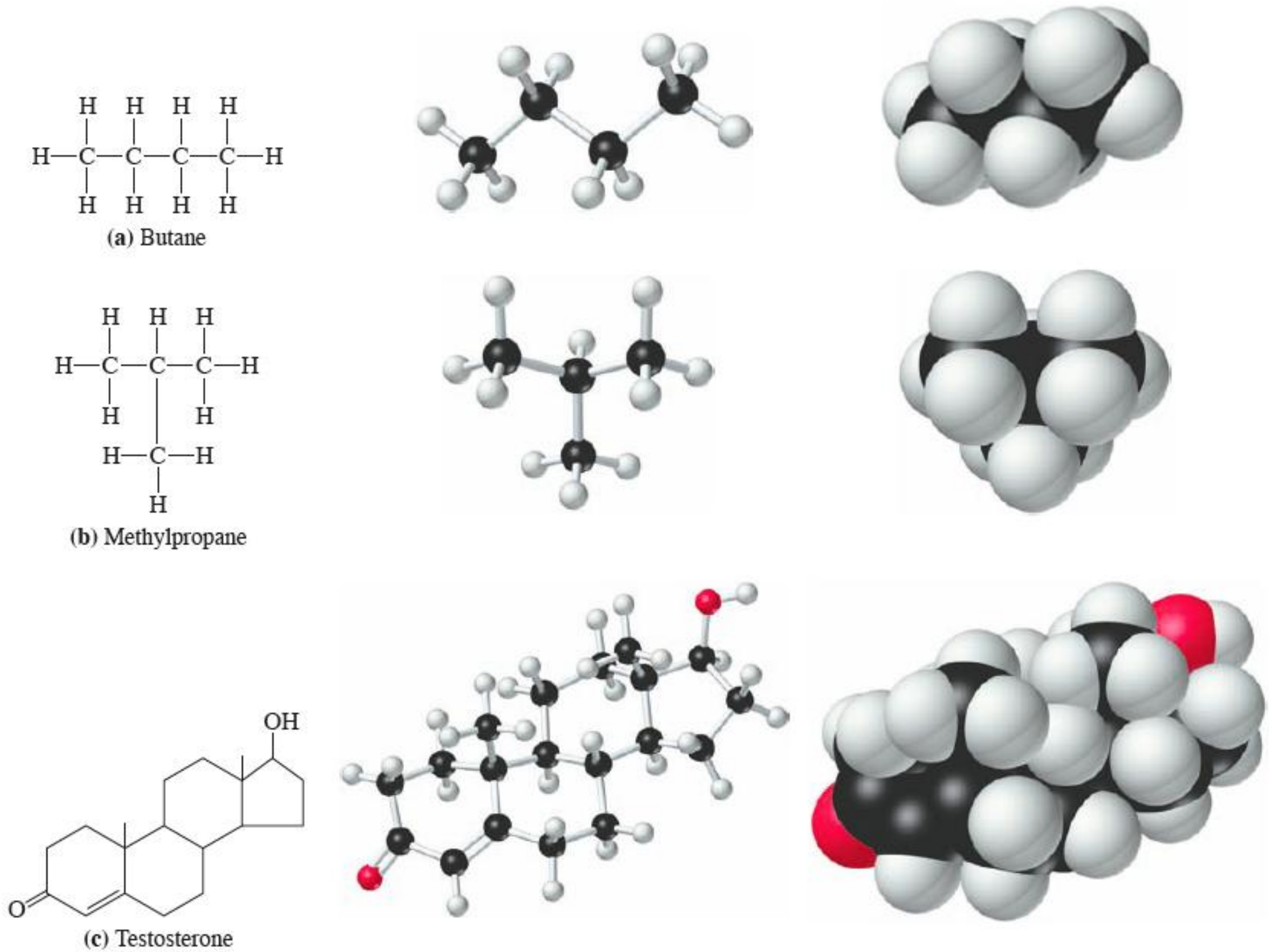


Figure 3-2 **Visualizations of (a) butane, (b) methylpropane, and (c) testosterone**

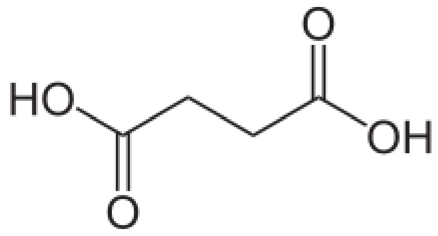
CONCEPT ASSESSMENT

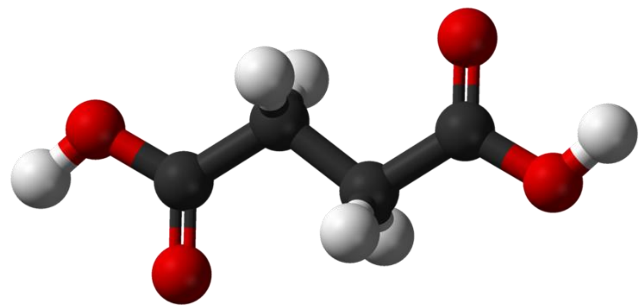
Represent the succinic acid molecule, HOOCCH2CH2COOH through molecular, empirical, structural, line-angle formulas and ball and stick model representation.

$C_4H_6O_4$

$C_2H_3O_2$

HOOC-$(CH_2)_2$-COOH



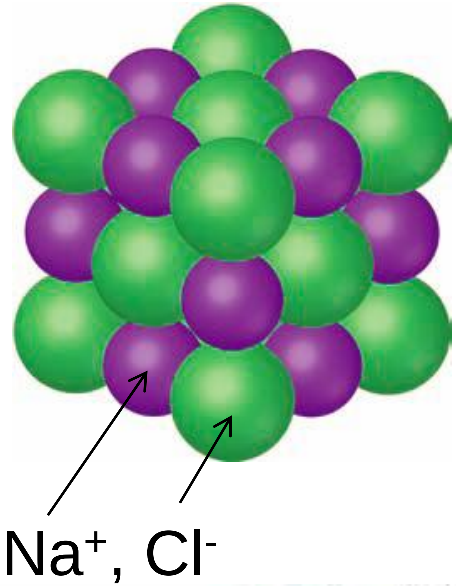


▲

Ionic Compounds

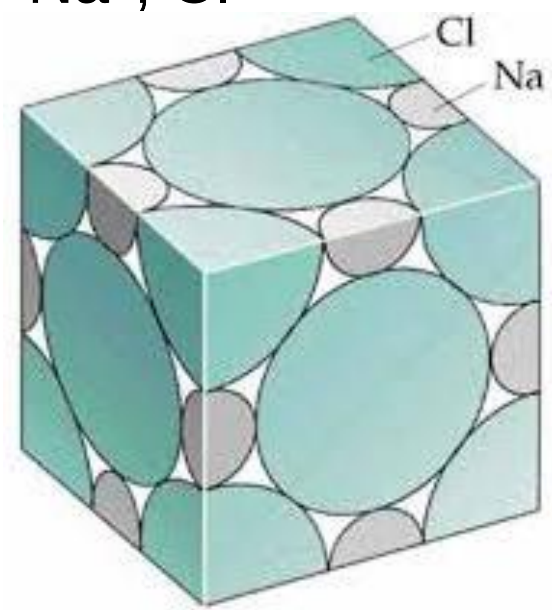
- Chemical combination of a metal and a nonmetal usually results in an ionic compound.
- An ionic compound is made up of positive and negative ions joined together by electrostatic forces of attraction
- Atoms of almost all elements can gain *or* lose electrons to form charged species called ions.
- Compounds composed of ions are known as **ionic compounds**.

  + Metals tend to lose electrons to form positively charged ions called **cations**.
  - Non-metals tend to gain electrons to form negatively charged ions called **anions**.



- An extended array of Na^+ and Cl^- ions. Each Na^+ ion in sodium chloride is surrounded by six Cl^- ions, and vice versa, and we cannot say that any one of these six ions belongs exclusively to a given ion. Yet, the ratio of to ions in sodium chloride is 1:1. we arbitrarily select a combination of one Na^+ ion and one Cl^- ion as a formula unit.

- *The formula unit of an ionic compound* is the smallest electrically neutral collection of ions. The simplest formula unit is NaCl



Monatomic ions, each consists of a single ionized atom.
Na^+, Cl^-, Mg^{2+}

Polyatomic ion, is made up of two or more atoms.
Nitrate ion, NO_3^- $Mg(NO_3^-)_2$

▲ FIGURE 3-4
Portion of an ionic crystal and a formula unit of NaCl

3-2 The Mole Concept and Chemical Compounds

Formula mass

the mass of a formula unit in atomic mass units (u)

Molecular mass

mass of a **molecule in atomic mass units.**

Weighted average mass

add up the weighted-average atomic masses

Exact Mass

add up the isotopic masses (see mass spectrometry)

The terms formula mass and molecular mass have essentially the same meaning, although when referring to ionic compounds, such as NaCl and $MgCl_2$ Formula mass is the proper term

molecular mass H_2O = 2(atomic mass H) + (atomic mass O)
= 2(1.00794 u) + 15.9994 u
= 18.0153 u

formula mass $MgCl_2$ = atomic mass Mg + 2(atomic mass Cl)
= 24.3050 u + 2(35.453 u)
= 95.211 u

Mole of a Compound

A mole: amount of substance having the same number of elementary entities as there are atoms in exactly 12 g of pure carbon-12.

A mole *of compound is an amount of compound* containing Avogadro's number of formula units or molecules.

The **molar mass is the mass of one mole of compound**

- one mole of molecules of a molecular compound and
- one mole of formula units of an ionic compound.

1 mol H_2O = 18.0153 g H_2O = 6.02214 * 10^{23} H_2O molecules

1 mol $MgCl_2$ = 95.211 g $MgCl_2$ = 6.02214 * 10^{23} $MgCl_2$ formula units

KEEP IN MIND

that although *molecular mass* and *molar mass sound similar* and are related, they are not the same. Molecular mass is the weighted-average mass of one molecule expressed in atomic mass units, u. *Molar mass is the mass* of Avogadro's number of molecules expressed in grams per mole, The two terms have the same numerical value but different units. g/mol.

Conversion factors

TABLE 3.1 Density, Molar Mass, and the Avogadro Constant as Conversion Factors	
Density, *d*	converts from volume to mass
Molar mass, *M*	converts from mass to amount (mol)
Avogadro constant, N_A	converts from amount (mol) to elementary entities

EXAMPLE 3-1 Relating Molar Mass, the Avogadro Constant, and Formula Units of an Ionic Compound

An analytical balance can detect a mass of 0.1 mg. How many ions are present in this minimally detectable quantity of $MgCl_2$?

Analyze

The central focus is the conversion of a measured quantity, 0.1 mg $MgCl_2$, to an amount in moles. After making the mass conversion, mg $\longrightarrow$ g, we can use the molar mass to convert from mass to amount in moles. Then, with the Avogadro constant as a conversion factor, we can convert from amount in moles to number of formula units. The final factor we need is based on the fact that there are *three* ions (*one* Mg^{2+} and *two* Cl^-) per formula unit (fu) of $MgCl_2$. It is often helpful to map out a conversion pathway that starts with the information given and proceeds through a series of conversion factors to the information sought. For this problem, we can begin with milligrams of $MgCl_2$ and make the following conversions:

$$\text{mg} \longrightarrow \text{g} \longrightarrow \text{mol} \longrightarrow \text{fu} \longrightarrow \text{number of ions}$$

Solve

The required conversions can be carried out in a stepwise fashion (as was done in Example 2-9), or they can be combined into a single line calculation. To avoid having to write down intermediate results and to avoid rounding errors, we'll use a single line calculation this time.

$$\begin{aligned} ?\ \text{ions} &= 0.1\ \text{mg MgCl}_2 \times \frac{1\ \text{g MgCl}_2}{1000\ \text{mg MgCl}_2} \times \frac{1\ \text{mol MgCl}_2}{95\ \text{g MgCl}_2} \\ &\times \frac{6.0 \times 10^{23}\ \text{fu MgCl}_2}{1\ \text{mol MgCl}_2} \times \frac{3\ \text{ions}}{1\ \text{fu MgCl}_2} \\ &= 2 \times 10^{18}\ \text{ions} \end{aligned}$$

Assess

The mass of the sample (0.1 mg) is given with one significant figure, and so the final answer is rounded to one significant figure. In the calculation above, the molar mass of $MgCl_2$ and the Avogadro constant are rounded off to two significant figures, that is, with one more significant figure than in the measured quantity.

PRACTICE EXAMPLE A: How many grams of $MgCl_2$ would you need to obtain 5.0×10^{23} Cl^- ions?

PRACTICE EXAMPLE B: How many nitrate ions, NO_3^-, and how many oxygen atoms are present in 1.00 μg of magnesium nitrate, $Mg(NO_3)_2$?

da Inc.

EXAMPLE 3-2 Combining Several Factors in a Calculation Involving Molar Mass

The volatile liquid ethyl mercaptan, C_2H_6S, is one of the most odoriferous substances known. It is sometimes added to natural gas to make gas leaks detectable. How many C_2H_6S molecules are contained in a 1.0 μL sample? The density of liquid ethyl mercaptan is 0.84 g/mL.

Analyze

The central focus is again the conversion of a measured quantity to an amount in moles. Because the density is given in g/mL, it will be helpful to convert the measured volume to milliliters. Then, density can be used as a conversion factor to obtain the mass in grams, and the molar mass can then be used to convert mass to amount in moles. Finally, the Avogadro constant can be used to convert the amount in moles to the number of molecules. In summary, the conversion pathway is $\mu L \rightarrow L \rightarrow g \rightarrow mol \rightarrow$ molecules.

Solve

As always, the required conversions can be combined into a single line calculation. However, it is instructive to break the calculation into three steps: (1) a conversion from volume to mass, (2) a conversion from mass to amount in moles, and (3) a conversion from amount in moles to molecules. These three steps emphasize, respectively, the roles played by density, molar mass, and the Avogadro constant in the conversion pathway. (See Table 3.1.)

Convert from volume to mass.

$$? \text{ g } C_2H_6S = 1.0\ \mu L \times \frac{1 \times 10^{-6}\ L}{1\ \mu L} \times \frac{1000\ mL}{1\ L} \times \frac{0.84\ g\ C_2H_6S}{1\ mL} = 8.4 \times 10^{-4}\ g\ C_2H_6S$$

Convert from mass to amount in moles.

$$? \text{ mol } C_2H_6S = 8.4 \times 10^{-4}\ g\ C_2H_6S \times \frac{1\ mol\ C_2H_6S}{62.1\ g\ C_2H_6S} = 1.4 \times 10^{-5}\ mol\ C_2H_6S$$

Convert from moles to molecules.

$$? \text{ molecules } C_2H_6S = 1.4 \times 10^{-5}\ mol\ C_2H_6S \times \frac{6.02 \times 10^{23}\ molecules\ C_2H_6S}{1\ mol\ C_2H_6S} = 8.1 \times 10^{18}\ molecules\ C_2H_6S$$

Assess

Remember to store intermediate results in your calculator *without* rounding off. Round off at the end. The answer is rounded to two significant figures because the volume and density are given with two significant figures. Rounding errors are avoided if the required conversions are combined into a single line calculation.

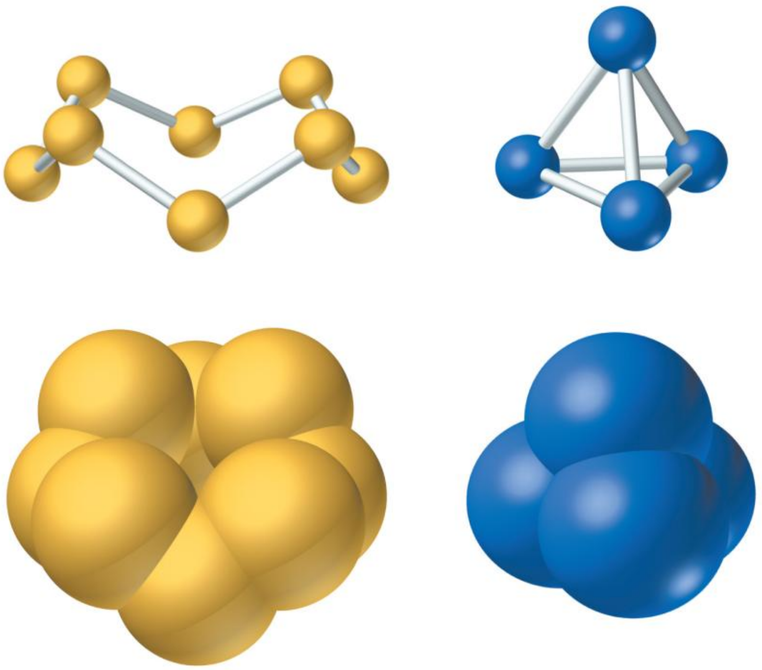
$$? \text{ molecules } C_2H_6S = 1.0\ \mu L \times \frac{1 \times 10^{-6}\ L}{1\ \mu L} \times \frac{1000\ mL}{1\ L} \times \frac{0.84\ g\ C_2H_6S}{1\ mL} \times \frac{1\ mol\ C_2H_6S}{62.1\ g\ C_2H_6S} \times \frac{6.02 \times 10^{23}\ molecules\ C_2H_6S}{1\ mol\ C_2H_6S} = 8.1 \times 10^{18}\ molecules\ C_2H_6S$$

PRACTICE EXAMPLE A: Gold has a density of 19.32 g/cm^3. A piece of gold foil is 2.50 cm on each side and 0.100 mm thick. How many atoms of gold are in this piece of gold foil?

PRACTICE EXAMPLE B: If the 1.0 μL sample of liquid ethyl mercaptan from Example 3-2 is allowed to evaporate and distribute itself throughout a chemistry lecture room with dimensions 62 ft $\times$ 35 ft $\times$ 14 ft, will the odor of the vapor be detectable in the room? The limit of detectability is $9 \times 10^{-4}\ \mu mol/m^3$.

Mole of an Element

But the atoms of some elements are joined together to form molecules.
H2, O2, N2, F2, Cl2,I2, P4, S8

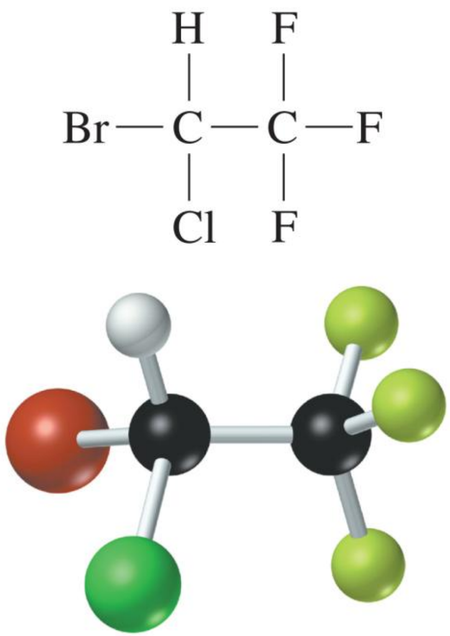


The existence of an element in more than one molecular form, referred to as allotropy.
Thus, oxygen exists in two allotropic forms, the predominantly abundant diatomic oxygen, and the much less abundant allotrope ozone, The molar mass of ordinary dioxygen is 31.9988 g and that of ozone is 47.9982 g O_3/mol O_3

Molecular forms of elemental sulfur and phosphorus. In a sample of solid sulfur, there are eight sulfur atoms in a sulfur molecule. In solid white phosphorus, there are four phosphorus atoms per molecule.

▲ Figure 3-5
Molecular forms of elemental sulfur and phosphorus

3-3 Composition of Chemical Compounds



Halothane $C_2HBrClF_3$

Mole ratio $n_C/n_{halothane}$

Mass ratio $m_C/m_{halothane}$

$$M(C_2HBrClF_3) = 2M_C + M_H + M_{Br} + M_{Cl} + 3M_F$$

$$= (2 \times 12.01) + 1.01 + 79.90 + 35.45 + (3 \times 19.00)$$

$$= 197.38 \text{ g/mol}$$

EXAMPLE 3-3 Using Relationships Derived from a Chemical Formula

How many moles of F atoms are in a 75.0 mL sample of halothane ($d = 1.871$ g/mL)?

Analyze

The conversion pathway for this problem is given above. First, convert the volume of the sample to mass; this requires density as a conversion factor. Next, convert the mass of halothane to its amount in moles; this requires the inverse of the molar mass as a conversion factor. The final conversion factor is based on the formula of halothane.

Solve

$$? \text{ mol F} = 75.0 \text{ mL } C_2HBrClF_3 \times \frac{1.871 \text{ g } C_2HBrClF_3}{1 \text{ mL } C_2HBrClF_3} \times \frac{1 \text{ mol } C_2HBrClF_3}{197.4 \text{ g } C_2HBrClF_3} \times \frac{3 \text{ mol F}}{1 \text{ mol } C_2HBrClF_3} = 2.13 \text{ mol F}$$

Assess

If we had been asked for the number of moles of C instead, the final conversion factor in the calculation above would have been (2 mol C/1 mol $C_2HBrClF_3$).

PRACTICE EXAMPLE A: How many grams of Br are contained in 25.00 mL of halothane ($d = 1.871$ g/mL)?

PRACTICE EXAMPLE B: How many milliliters of halothane would contain 1.00×10^{24} Br atoms?

Calculating Percent Composition from a Chemical Formula

1. Determine the molar mass of the compound. This is the *denominator in* equation (3.1).
2. Determine the contribution of the given element to the molar mass. This product of the formula subscript and the molar mass of the element appears in the *numerator of equation (3.1).*
3. Formulate the ratio of the mass of the given element to the mass of the compound as a whole. This is the ratio of the numerator from step 2 to the denominator from step 1.
4. Multiply this ratio by 100% to obtain the mass percent of the element

The mass composition of a compound is the collection of mass percentages of the individual elements in the compound

EXAMPLE 3-4 Calculating the Mass Percent Composition of a Compound

What is the mass percent composition of halothane, $C_2HBrClF_3$?

Analyze

Apply the four-step method described above. First, determine the molar mass of $C_2HBrClF_3$. Then formulate the mass ratios and convert them to mass percents. If we use molar masses that are rounded to two decimal places, the calculated mass percents will be accurate to two decimal places.

Solve

The molar mass of $C_2HBrClF_3$ is 197.38 g/mol. The mass percents are

$$\% \text{ C} = \frac{\left(2 \text{ mol C} \times \dfrac{12.01 \text{ g C}}{1 \text{ mol C}}\right)}{197.38 \text{ g } C_2HBrClF_3} \times 100\% = 12.17\% \text{ C}$$

$$\% \text{ H} = \frac{\left(1 \text{ mol H} \times \dfrac{1.01 \text{ g H}}{1 \text{ mol H}}\right)}{197.38 \text{ g } C_2HBrClF_3} \times 100\% = 0.51\% \text{ H}$$

$$\% \text{ Br} = \frac{\left(1 \text{ mol Br} \times \dfrac{79.90 \text{ g Br}}{1 \text{ mol Br}}\right)}{197.38 \text{ } C_2HBrClF_3} \times 100\% = 40.48\% \text{ Br}$$

$$\% \text{ Cl} = \frac{\left(1 \text{ mol Cl} \times \dfrac{35.45 \text{ g Cl}}{1 \text{ mol Cl}}\right)}{197.38 \text{ g } C_2HBrClF_3} \times 100\% = 17.96\% \text{ Cl}$$

$$\% \text{ F} = \left(\frac{3 \text{ mol F} \times \dfrac{19.00 \text{ g F}}{1 \text{ mol F}}}{197.38 \text{ g } C_2HBrClF_3}\right) \times 100\% = 28.88\% \text{ F}$$

Thus, the percent composition of halothane is 12.17% C, 0.51% H, 40.48% Br, 17.96% Cl, and 28.88% F.

Assess

The mass ratios appearing above are based on a sample that contains exactly *one mole* of halothane. Another approach is to calculate the mass of each element present in a sample that contains exactly *100* g of halothane. For example, in a *100 g* sample of halothane,

$$? \text{ g C} = 100 \text{ g } C_2HBrClF_3 \times \frac{1 \text{ mol } C_2HBrClF_3}{197.38 \text{ g } C_2HBrClF_3} \times \frac{2 \text{ mol C}}{1 \text{ mol } C_2HBrClF_3} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 12.17 \text{ g C}$$

and so halothane is 12.17% C. The mass of carbon in a 100 g sample is numerically equal to the mass percent of carbon. If you compare the calculation of g C with that for % C given earlier, you will see that both calculations involve exactly the same factors but in a slightly different order.

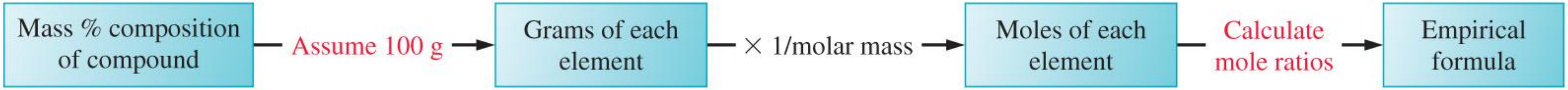
PRACTICE EXAMPLE A: Adenosine triphosphate (ATP) is the main energy-storage molecule in cells. Its chemical formula is $C_{10}H_{16}N_5P_3O_{13}$. What is its mass percent composition?

PRACTICE EXAMPLE B: *Without doing detailed calculations*, determine which two compounds from the following list have the same percent oxygen by mass: **(a)** CO; **(b)** CH_3COOH; **(c)** C_2O_3; **(d)** N_2O; **(e)** $C_6H_{12}O_6$; **(f)** $HOCH_2CH_2OH$.

Establishing Formulas from Experimentally Determined Percent Composition of Compounds

5 Step approach:

1. Choose an arbitrary sample size (100g).
2. Convert masses to amounts in moles.
3. Write a formula.
4. Convert formula to small whole numbers.
5. Multiply all subscripts by a small whole number to make the subscripts integral.



EXAMPLE 3-5 Determining the Empirical and Molecular Formulas of a Compound from Its Mass Percent Composition

Dibutyl succinate is an insect repellent used against household ants and roaches. Its composition is 62.58% C, 9.63% H, and 27.79% O. Its experimentally determined molecular mass is 230 u. What are the empirical and molecular formulas of dibutyl succinate?

Analyze

Use the five-step approach described above.

Solve

1. Determine the mass of each element in a 100.00 g sample.

   62.58 g C, 9.63 g H, 27.79 g O

2. Convert each of these masses to an amount in moles.

$$? \text{ mol C} = 62.58 \text{ g C} \times \frac{1 \text{ mol C}}{12.011 \text{ g C}} = 5.210 \text{ mol C}$$

$$? \text{ mol H} = 9.63 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.55 \text{ mol H}$$

$$? \text{ mol O} = 27.79 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 1.737 \text{ mol O}$$

3. Write a tentative formula based on these numbers of moles.

   $C_{5.21}H_{9.55}O_{1.74}$

4. Divide each of the subscripts of the tentative formula by the smallest subscript (1.74),

$$C_{\frac{5.21}{1.74}}H_{\frac{9.55}{1.74}}O_{\frac{1.74}{1.74}} = C_{2.99}H_{5.49}O$$

   and round off any subscripts that differ only slightly from whole numbers; that is, round 2.99 to 3.

   $C_3H_{5.49}O$

5. Multiply all subscripts by a small whole number to make them integral (here by the factor 2), and write the empirical formula.

$$C_{2\times3}H_{2\times5.49}O_{2\times1} = C_6H_{10.98}O_2$$

$$2 \times 5.49 = 10.98 \approx 11$$

   Empirical formula: $C_6H_{11}O_2$

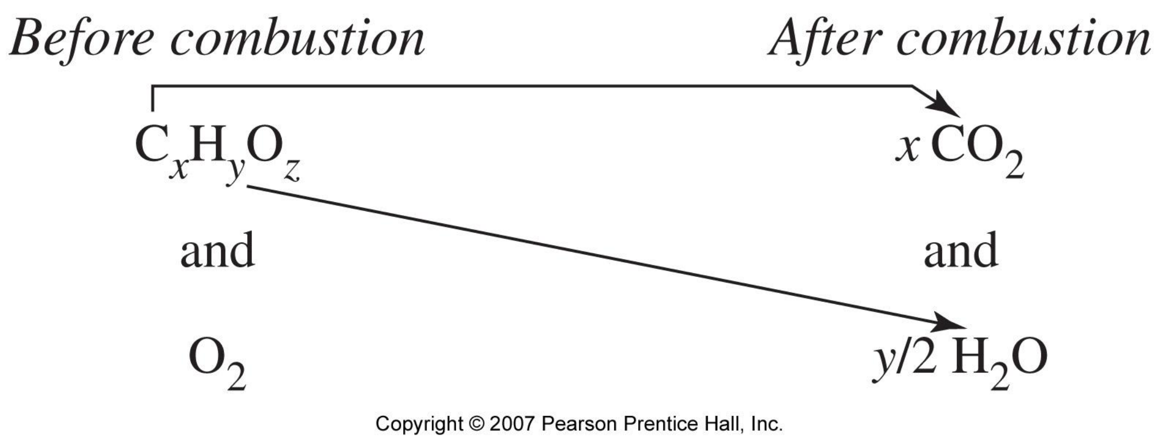
To establish the molecular formula, first determine the empirical formula mass.

$$[(6 \times 12.0) + (11 \times 1.0) + (2 \times 16.0)]\text{u} = 115 \text{ u}$$

Since the experimentally determined formula mass (230 u) is twice the empirical formula mass, the molecular formula is twice the empirical formula.

Molecular formula: $C_{12}H_{22}O_4$

Combustion Analysis:



After combustion,

- all the carbon atoms in the sample are found in the CO_2.
- all the H atoms are in the H_2O. Moreover, the only source of the carbon and hydrogen atoms was the sample being analyzed.
- Oxygen atoms in the CO_2 and H_2O could have come partly from the sample and partly from the oxygen gas consumed in the combustion. Thus, the quantity of oxygen in the sample has to be determined indirectly

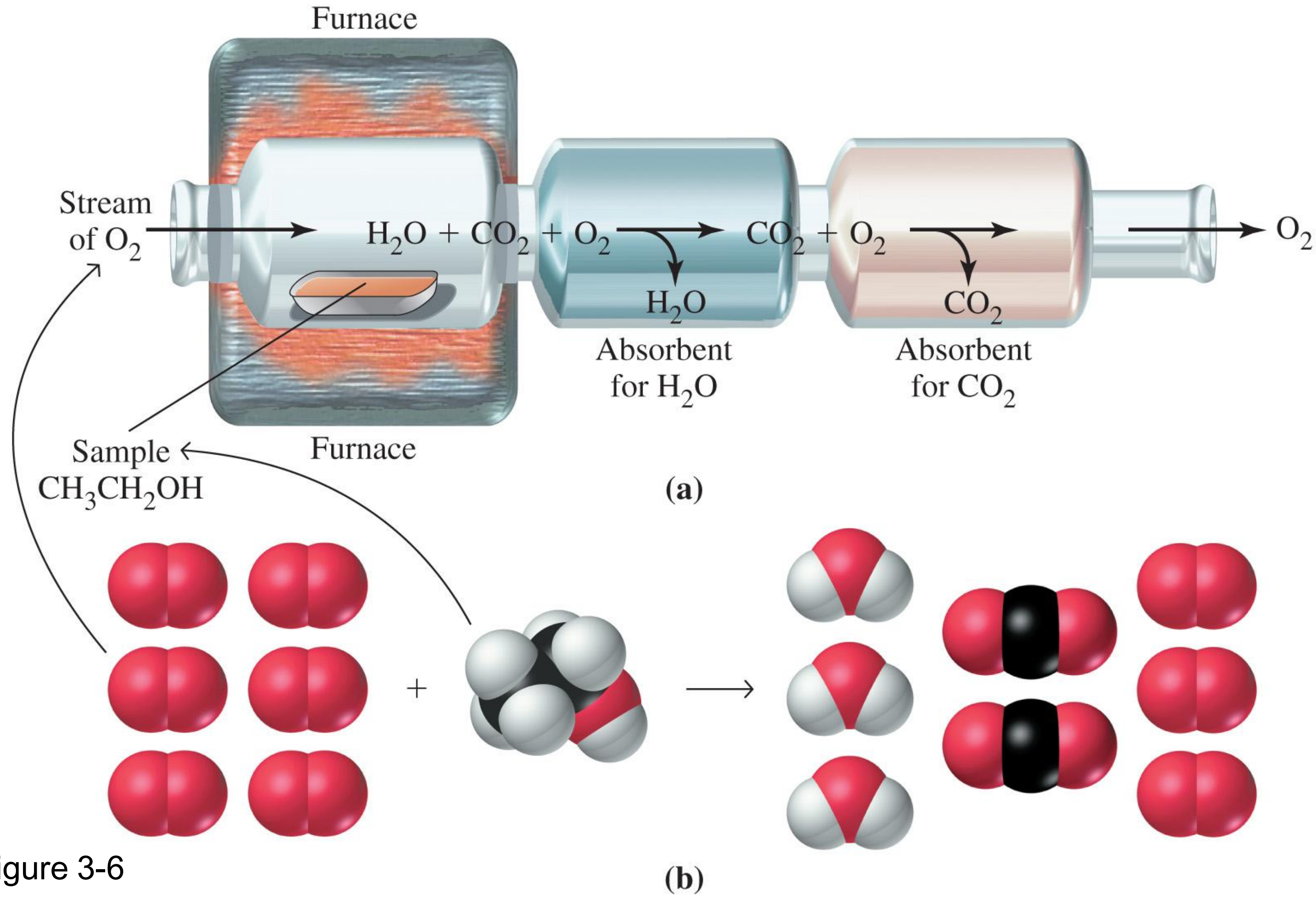


Figure 3-6

Apparatus for combustion analysis

▲ water vapor by magnesium perchlorate, and carbon dioxide gas by sodium hydroxide

EXAMPLE 3-6 Determining an Empirical Formula from Combustion Analysis Data

Vitamin C is essential for the prevention of scurvy. Combustion of a 0.2000 g sample of this carbon–hydrogen–oxygen compound yields 0.2998 g CO_2 and 0.0819 g H_2O. What are the percent composition and the empirical formula of vitamin C?

Analyze

After combustion, all the carbon atoms from the vitamin C sample are in CO_2 and all the hydrogen atoms are in H_2O. However, the oxygen atoms in CO_2 and H_2O come partly from the sample and partly from the oxygen gas consumed in the combustion. So, in the determination of the percent composition, we focus first on carbon and hydrogen and then on oxygen. To determine the empirical formula, we must calculate the amounts of C, H, and O in moles, and then calculate the mole ratios.

Solve

Percent Composition

First, determine the mass of carbon in 0.2988 g CO_2, by converting to mol C,

$$? \text{ mol C} = 0.2998 \text{ g } CO_2 \times \frac{1 \text{ mol } CO_2}{44.010 \text{ g } CO_2} \times \frac{1 \text{ mol C}}{1 \text{ mol } CO_2} = 0.006812 \text{ mol C}$$

and then to g C.

$$? \text{ g C} = 0.006812 \text{ mol C} \times \frac{12.011 \text{ g C}}{1 \text{ mol C}} = 0.08182 \text{ g C}$$

Proceed in a similar fashion for 0.0819 g H_2O to obtain

$$? \text{ mol H} = 0.0819 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} = 0.00909 \text{ mol H}$$

and

$$? \text{ g H} = 0.00909 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 0.00916 \text{ g H}$$

Obtain the mass of O in the 0.2000 g sample as the difference

$$? \text{ g O} = 0.2000 \text{ g sample} - 0.08182 \text{ g C} - 0.00916 \text{ g H} = 0.1090 \text{ g O}$$

Finally, multiply the mass fractions of the three elements by 100% to obtain mass percentages.

$$\% \text{ C} = \frac{0.08182 \text{ g C}}{0.2000 \text{ g sample}} \times 100\% = 40.91\% \text{ C}$$

$$\% \text{ H} = \frac{0.00916 \text{ g H}}{0.2000 \text{ g sample}} \times 100\% = 4.58\% \text{ H}$$

$$\% \text{ O} = \frac{0.1090 \text{ g O}}{0.2000 \text{ g sample}} \times 100\% = 54.50\% \text{ O}$$

Empirical Formula

At this point we can choose either of two alternatives. The first is to obtain the empirical formula from the mass percent composition, in the same manner illustrated in Example 3-5. The second is to note that we have already determined the number of moles of C and H in the 0.2000 g sample. The number of moles of O is

$$? \text{ mol O} = 0.1090 \text{ g O} \times \frac{1 \text{ mol O}}{15.999 \text{ g O}} = 0.006813 \text{ mol O}$$

From the numbers of moles of C, H, and O in the 0.2000 g sample, we obtain the tentative empirical formula

$$C_{0.006812}H_{0.00909}O_{0.006813}$$

Next, divide each subscript by the smallest (0.006812) to obtain

$$CH_{1.33}O$$

Finally, multiply all the subscripts by 3 to obtain

Empirical formula of vitamin C: $C_3H_4O_3$

Assess

The determination of the empirical formula does not require determining the mass percent composition as a preliminary calculation. The empirical formula can be based on a sample of any size, as long as the numbers of moles of the different atoms in that sample can be determined.

PRACTICE EXAMPLE A: Isobutyl propionate is the substance that flavors rum extract. Combustion of a 1.152 g sample of this carbon–hydrogen–oxygen compound yields 2.726 g CO_2 and 1.116 g H_2O. What is the empirical formula of isobutyl propionate?

PRACTICE EXAMPLE B: Combustion of a 1.505 g sample of thiophene, a carbon–hydrogen–sulfur compound, yields 3.149 g CO_2, 0.645 g H_2O, and 1.146 g SO_2 as the only products of the combustion. What is the empirical formula of thiophene?

3-4 Oxidation States: A Useful Tool in Describing Chemical Compounds

Metals tend to lose electrons.

$Na \longrightarrow Na^{+} + e^{-}$ **Reducing agents**

Non-metals tend to gain electrons.

$Cl + e^{-} \longrightarrow Cl^{-}$ **Oxidizing agents**

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 H 1.00794																	2 He 4.00260
3 Li 6.941	4 Be 9.01218											5 B 10.811	6 C 12.011	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.1797
11 Na 22.9898	12 Mg 24.3050											13 Al 26.9815	14 Si 28.0855	15 P 30.9738	16 S 32.066	17 Cl 35.4527	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.9559	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.9381	26 Fe 55.847	27 Co 58.9332	28 Ni 58.693	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.9216	34 Se 78.96	35 Br 79.904	36 Kr 83.80
37 Rb 85.4678	38 Sr 87.62	39 Y 88.9059	40 Zr 91.224	41 Nb 92.9064	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.906	46 Pd 106.42	47 Ag 107.868	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.757	52 Te 127.60	53 I 126.904	54 Xe 131.29
55 Cs 132.905	56 Ba 137.327	57 *La 138.906	72 Hf 178.49	73 Ta 180.948	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.967	80 Hg 200.59	81 Tl 204.383	82 Pb 207.2	83 Bi 208.980	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra 226.025	89 †Ac 227.028	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 (269)	111 (272)	112 (272)		114 (287)		116 (289)		118 (293)

*Lanthanide series	58 Ce 140.115	59 Pr 140.908	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.965	64 Gd 157.25	65 Tb 158.925	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.967
†Actinide series	90 Th 232.038	91 Pa 231.036	92 U 238.029	93 Np 237.048	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)

Oxidation State is used to keep track of the number of electrons that have been gained or lost by an element.

TABLE 3.2 Rules for Assigning Oxidation States

1. *The oxidation state (O.S.) of an individual atom in a free element (uncombined with other elements) is 0.*

[*Examples:* The O.S. of an isolated Cl atom is 0; the two Cl atoms in the molecule Cl_2 both have an O.S. of 0.]

2. *The total of the O.S. of all the atoms in*

(a) *neutral species, such as isolated atoms, molecules, and formula units, is 0;*

[*Examples:* The sum of the O.S. of all the atoms in CH_3OH and of all the ions in $MgCl_2$ is 0.]

(b) *an ion is equal to the charge on the ion.*

[*Examples:* The O.S. of Fe in Fe^{3+} is +3. The sum of the O.S. of all atoms in MnO_4^- is −1.]

3. *In their compounds, the group 1 metals have an O.S. of +1 and the group 2 metals have an O.S. of +2.*

[*Examples:* The O.S. of K is +1 in KCl and K_2CO_3; the O.S. of Mg is +2 in $MgBr_2$ and $Mg(NO_3)_2$.]

4. *In its compounds, the O.S. of fluorine is −1.*

[*Examples:* The O.S. of F is −1 in HF, ClF_3, and SF_6.]

5. *In its compounds, hydrogen usually has an O.S. of +1.*

[*Examples:* The O.S. of H is +1 in HI, H_2S, NH_3, and CH_4.]

6. *In its compounds, oxygen usually has an O.S. of −2.*

[*Examples:* The O.S. of O is −2 in H_2O, CO_2 and $KMnO_4$.]

7. *In binary (two-element) compounds with metals, group 17 elements have an O.S. of −1; group 16 elements, −2; and group 15 elements, −3.*

[*Examples:* The O.S. of Br is −1 in $MgBr_2$; the O.S. of S is −2 in Li_2S; and the O.S. of N is −3 in Li_3N.]

EXAMPLE 3-7 Assigning Oxidation States

What is the oxidation state of the underlined element in **(a)** $\underline{P}_4$; **(b)** $\underline{Al}_2O_3$; **(c)** $\underline{Mn}O_4^-$; **(d)** $Na\underline{H}$; **(e)** $H_2\underline{O}_2$; **(f)** $\underline{Fe}_3O_4$?

Solve

(a) P_4: This formula represents a molecule of elemental phosphorus. For an atom of a free element, the O.S. = 0 (rule 1). The O.S. of P in P_4 is 0.

(b) Al_2O_3: The total of the oxidation states of all the atoms in this formula unit is 0 (rule 2). The O.S. of oxygen is −2 (rule 6). The total for three O atoms is −6. The total for two Al atoms is +6. The O.S. of Al is +3.

(c) MnO_4^-: This is the formula for permanganate ion. The total of the oxidation states of all the atoms in the ion is −1 (rule 2). The total for the four O atoms is −8. The O.S. of Mn is +7.

(d) NaH: This is a formula unit of the ionic compound sodium hydride. Rule 3 states that the O.S. of Na is +1. Rule 5 indicates that H should also have an O.S. of +1. If both atoms had an O.S. of +1, the total for the formula unit would be +2. This violates rule 2. *Rules 2 and 3 take precedence over rule 5*. Na has an O.S. of +1; the total for the formula unit is 0; and the O.S. of H must be −1.

(e) H_2O_2: This is hydrogen peroxide. Rule 5, stating that H has an O.S. of +1, takes precedence over rule 6 (which says that oxygen has an O.S. of −2). The sum of the oxidation states of the two H atoms is +2 and that of the two O atoms must be −2. The O.S. of O must be −1.

(f) Fe_3O_4: The total of the oxidation states of four O atoms is −8. For three Fe atoms, the total must be +8. The O.S. per Fe atom is $\frac{8}{3}$ or $+2\frac{2}{3}$.

PRACTICE EXAMPLE A: What is the oxidation state of the underlined element in each of the following: $\underline{S}_8$; $\underline{Cr}_2O_7^{2-}$; $\underline{Cl}_2O$; $K\underline{O}_2$?

S_8 For an atom of a free element, the oxidation state is 0 (rule 1).

$Cr_2\ O_7$ The sum of all the oxidation numbers in the ion is -2 (rule 2). The O.S. of each oxygen is -2 (rule 6). Thus, the total for all seven oxygens is -14 . The total for both chromiums must be +12. Thus, each Cr has an O.S. = +6 .

Cl_2O The sum of all oxidation numbers in the compound is 0 (rule 2). The O.S. of oxygen is - (rule 6). The total for the two chlorines must be +2. Thus, each chlorine must have O.S. = +1.

KO_2 The sum for all the oxidation numbers in the compound is 0 (rule 2). The O.S. Of potassium is +1 (rule 3). The sum of the oxidation numbers of the two oxygens must be -1. Thus, each oxygen must have O.S. = - 1/2 .

PRACTICE EXAMPLE B: What is the oxidation state of the underlined element in each of the following: $\underline{S}_2O_3^{2-}$; $\underline{Hg}_2Cl_2$; $K\underline{Mn}O_4$; $H_2\underline{C}O$?

$S_2O_3^{-2}$ □ The sum of all the oxidation numbers in the ion is -2 (rule 2). The O.S. of oxygen is -2 (rule 6). Thus, the total for three oxygens must be -6 . The total for both sulfurs must be +4. Thus, each S has an O.S. = +2 .

Hg_2Cl_2 The O.S. of each Cl is -1 (rule 7). The sum of all O.S. is 0 (rule 2). Thus, the total for two Hg is +2 and each Hg has O.S. = +1.

$KMnO_4$ The O.S. of each O is -2 (rule 6). Thus, the total for 4 oxygens must be -8 . The K has O.S. = +1 (rule 3). The total of all O.S. is 0 (rule 2). Thus, the O.S. of Mn is +7.

H_2CO The O.S. of each H is +1 (rule 5), producing a total for both hydrogens of +2. The O.S. of O is -2 (rule 6). Thus, the O.S. of C is 0, because the total of all O.Ss. is 0 (rule 2).

3-5 Naming Compounds: Organic and Inorganic Compounds

-different compounds may have the same formula. In these instances, we will find it essential to distinguish among compounds by name.

- If all compounds were referred to by a common or trivial name, such as water (H_2O), ammonia (NH_3), or glucose ($C_6H_{12}O_6$) we would have to learn millions of unrelated names, an impossibility.

What we need is a systematic method of assigning names: Nomenclature

-We cannot give two substances the same name, yet we do want some similarities in the names of similar substances

Lead (IV) oxide PbO_2 Lead (II) oxide PbO

Figure 3-7 Two oxides of lead

-Compounds formed by carbon and hydrogen or carbon and hydrogen together with oxygen, nitrogen, and a few other elements are :

Organic compounds:

- has its own set of nomenclature rules.

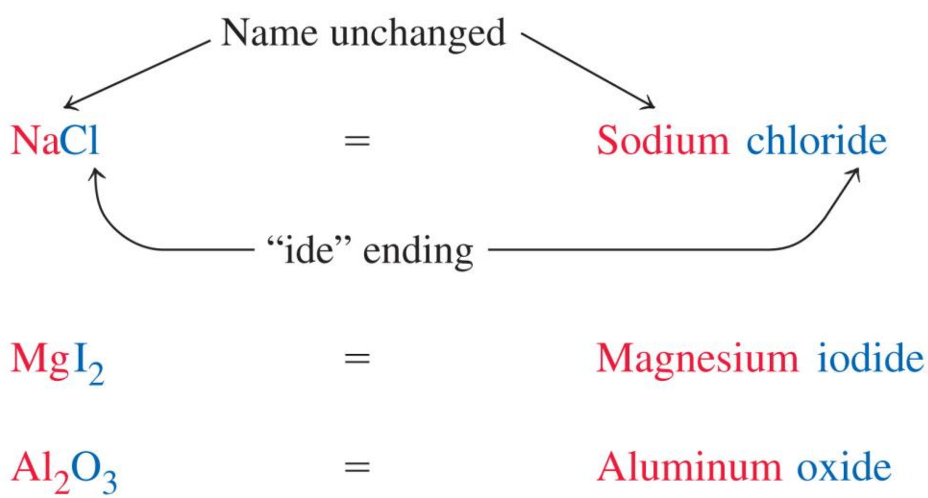
- Compounds that do not fit this description are **inorganic compounds**

3-6 Names and Formulas of Inorganic Compounds

Binary Compounds of Metals and Nonmetals

Binary compounds are those formed between two elements. If one of the elements is a metal and the other a nonmetal, the binary compound is usually made up of ions; that is, it is a binary *ionic compound. To name a binary compound* of a metal and a nonmetal,

- write the *unmodified* name of the metal
- then write the name of the nonmetal, modified to end in *–ide*



Ionic compounds, though made up of positive and negative ions, must be *electrically neutral*

TABLE 3.3 Some Simple Ions

Name	Symbol	Name	Symbol
Positive ions (cations)			
Lithium ion	Li^{+}	Chromium(II) ion	Cr^{2+}
Sodium ion	Na^{+}	Chromium(III) ion	Cr^{3+}
Potassium ion	K^{+}	Iron(II) ion	Fe^{2+}
Rubidium ion	Rb^{+}	Iron(III) ion	Fe^{3+}
Cesium ion	Cs^{+}	Cobalt(II) ion	Co^{2+}
Magnesium ion	Mg^{2+}	Cobalt(III) ion	Co^{3+}
Calcium ion	Ca^{2+}	Copper(I) ion	Cu^{+}
Strontium ion	Sr^{2+}	Copper(II) ion	Cu^{2+}
Barium ion	Ba^{2+}	Mercury(I) ion	Hg_2^{2+}
Aluminum ion	Al^{3+}	Mercury(II) ion	Hg^{2+}
Zinc ion	Zn^{2+}	Tin(II) ion	Sn^{2+}
Silver ion	Ag^{+}	Lead(II) ion	Pb^{2+}
Negative ions (anions)			
Hydride ion	H^{-}	Iodide ion	I^{-}
Fluoride ion	F^{-}	Oxide ion	O^{2-}
Chloride ion	Cl^{-}	Sulfide ion	S^{2-}
Bromide ion	Br^{-}	Nitride ion	N^{3-}

♠ some metals may form several ions, it is important to distinguish between them in naming their compounds.

Fe^{+2} and Fe^{+3} *iron(II) ion, and iron(III) ion*

♠the Roman numeral immediately following the name of the metal indicates its oxidation state or simply the charge on the ion. Thus, is *iron(II) chloride, while is iron(III) chloride.*

♠ two different word endings to distinguish between *two* binary compounds containing the same two elements but in different proportions, such as

Cu_2O and CuO

cuprous oxide *cupric oxide*

$FeCl_2$ *ferrous chloride* $FeCl_3$ *is ferric chloride*

use the – *ous* ending for the lower oxidation state of the metal

– *ic for the higher oxidation* state.

♠Several inadequacies.

Do not help in naming the four oxides of vanadium: VO, and V_2O_3 , VO_2 and V_2O_5

EXAMPLE 3-8 Writing Formulas When Names of Compounds Are Given

Write formulas for the compounds barium oxide, calcium fluoride, and iron(III) sulfide.

Analyze

In each case, identify the cations and their charges, based on periodic table group numbers or on oxidation states appearing as Roman numerals in names: Ba^{2+}, Ca^{2+}, and Fe^{3+}. Then identify the anions and their charges: O^{2-}, F^{-}, and S^{2-}. Combine the cations and anions in the relative numbers required to produce electrically *neutral* formula units.

Solve

barium oxide: *one* Ba^{2+} and *one* O^{2-} = BaO

calcium fluoride: *one* Ca^{2+} and *two* F^{-} = CaF_2

iron(III) sulfide: *two* Fe^{3+} and *three* S^{2-} = Fe_2S_3

Assess

In the first case, an electrically neutral formula unit results from the combination of the charges 2+ and 2−; in the second case, 2+ and $2 \times (1-)$; and in the third case, $2 \times (3+)$ and $3 \times (2-)$.

PRACTICE EXAMPLE A: Write formulas for lithium oxide, tin(II) fluoride, and lithium nitride.

PRACTICE EXAMPLE B: Write formulas for the compounds aluminum sulfide, magnesium nitride, and vanadium(III) oxide.

EXAMPLE 3-9 Naming Compounds When Their Formulas Are Given

Write acceptable names for the compounds Na_2S, AlF_3, Cu_2O.

Analyze

This task is generally easier than that of Example 3-8 because all you need to do is name the ions present. However, you must recognize that copper forms two different ions and that the cation in Cu_2O is Cu^+, copper(I).

Solve

Na_2S: sodium sulfide
AlF_3: aluminum fluoride
Cu_2O: copper(I) oxide

Assess

Knowing when to use Roman numerals and when not to use them is a tricky aspect of naming ionic compounds. Because the metals of groups 1 and 2 have only one ionic form (one oxidation state), Roman numerals are not used in naming their compounds.

PRACTICE EXAMPLE A: Write acceptable names for CsI, CaF_2, FeO, $CrCl_3$.

PRACTICE EXAMPLE B: Write acceptable names for CaH_2, CuCl, Ag_2S, Hg_2Cl_2.

Binary Compounds of Two Non-Metals

- ♠ If the two elements in a binary compound are both nonmetals instead of a metal and a nonmetal, the compound is a molecular compounds usually write the positive OS element first. E.g. HCl hydrogen chloride

EXAMPLE 3-8 Write formulas for the compounds barium oxide, calcium fluoride, and iron(III) sulfide, lithium oxide, tin(II) fluoride, and lithium nitride.

- ♠ Some pairs of nonmetals form more than one compound

mono	1	penta	5
di	2	hexa	6
tri	3	hepta	7
tetra	4	octa	8

SO_2 = sulfur *dioxide*
SO_3 = sulfur *trioxide*
B_2Br_4 = *diboron tetrabromide*

TABLE 3.4 Naming Binary Molecular Compounds

Formula	Name[a]
BCl_3	Boron trichloride
CCl_4	Carbon tetrachloride
CO	Carbon monoxide
CO_2	Carbon dioxide
NO	Nitrogen monoxide
NO_2	Nitrogen dioxide
N_2O	Dinitrogen monoxide
N_2O_3	Dinitrogen trioxide
N_2O_4	Dinitrogen tetroxide
N_2O_5	Dinitrogen pentoxide
PCl_3	Phosphorus trichloride
PCl_5	Phosphorus pentachloride
SF_6	Sulfur hexafluoride

[a]When the prefix ends in *a* or *o* and the element name begins with *a* or *o*, the final vowel of the prefix is dropped for ease of pronunciation. For example, carbon *mon*oxide, not carbon *mono*oxide, and dinitrogen *tetr*oxide, not dinitrogen *tetra*oxide. However, PI_3 is phosphorus *tri*iodide, not phosphorus *tri*odide.

Binary Acids:

Acids produce H^+ when dissolved in water.

They are compounds that ionize in water.
A *binary acid is a two-element compound of* hydrogen and a nonmetal.

$H_2S(aq)$ = hydrosulfuric acid
$HI(aq)$ = hydroiodic acid
$HCl(aq)$ = hydrochloric acid
$HBr(aq)$ = hydrobromic acid
$HF(aq)$ = hydrofluoric acid

The symbol (*aq*) signifies *aqueous* solution.

Polyatomic Ions

In polyatomic ions, two or more atoms are joined together by covalent bonds. These ions are common, especially among the nonmetals.

TABLE 3.5 Some Common Polyatomic Ions

Name	Formula	Typical Compound
Cation		
Ammonium ion	NH_4^+	NH_4Cl
Anions		
Acetate ion	CH_3COO^-	$NaCH_3COO$
Carbonate ion	CO_3^{2-}	Na_2CO_3
Hydrogen carbonate ion[a] (or bicarbonate ion)	HCO_3^-	$NaHCO_3$
Hypochlorite ion	ClO^-	$NaClO$
Chlorite ion	ClO_2^-	$NaClO_2$
Chlorate ion	ClO_3^-	$NaClO_3$
Perchlorate ion	ClO_4^-	$NaClO_4$
Chromate ion	CrO_4^{2-}	Na_2CrO_4
Dichromate ion	$Cr_2O_7^{2-}$	$Na_2Cr_2O_7$
Cyanide ion	CN^-	$NaCN$
Hydroxide ion	OH^-	$NaOH$
Nitrite ion	NO_2^-	$NaNO_2$
Nitrate ion	NO_3^-	$NaNO_3$
Oxalate ion	$C_2O_4^{2-}$	$Na_2C_2O_4$

Oxalate ion	$C_2O_4^{2-}$	$Na_2C_2O_4$
Permanganate ion	MnO_4^-	$NaMnO_4$
Phosphate ion	PO_4^{3-}	Na_3PO_4
Hydrogen phosphate ion[a]	HPO_4^{2-}	Na_2HPO_4
Dihydrogen phosphate ion[a]	$H_2PO_4^-$	NaH_2PO_4
Sulfite ion	SO_3^{2-}	Na_2SO_3
Hydrogen sulfite ion[a] (or bisulfite ion)	HSO_3^-	$NaHSO_3$
Sulfate ion	SO_4^{2-}	Na_2SO_4
Hydrogen sulfate ion[a] (or bisulfate ion)	HSO_4^-	$NaHSO_4$
Thiosulfate ion	$S_2O_3^{2-}$	$Na_2S_2O_3$

[a]These anion names are sometimes written as a single word—for example, hydrogencarbonate, hydrogenphosphate, and so forth.

1. Polyatomic anions are more common than polyatomic cations. The most familiar polyatomic cation is the ammonium ion, NH_4^+.

2. Very few polyatomic anions carry the -*ide ending in their names. Of those* listed, only (hydroxide ion) and (cyanide ion) do. The common endings are -*ite and -ate, and some names carry prefixes, hypo- or per-.*

3. An element common to many polyatomic anions is *oxygen, usually in* combination with another nonmetal. Such anions are called oxoanions.

4. Certain nonmetals (such as Cl, N, P, and S) form a series of oxoanions containing different numbers of oxygen atoms. Their names are related to the oxidation state of the nonmetal atom to which the O atoms are bonded, ranging from *hypo- (lowest) to per- (highest) according to the following* scheme.

Increasing oxidation state of nonmetal $\rightarrow$

hypo___ite,   _______ite,   _________ate,   per___ate

Increasing number of oxygen atoms $\rightarrow$

5. All the common oxoanions of Cl, Br, and I carry a charge of (- 1).

6. Some series of oxoanions also contain various numbers of H atoms and are named accordingly. For example $HPO4^{2-}$, is the hydrogen phosphate ion and the H_2PO4^{-} dihydrogen phosphate ion.

7. The prefix thio- signifies that a sulfur atom has been substituted for an oxygen atom. (The sulfate ion has one S and four O atoms; thiosulfate ion has two S and three O atoms.)

OXOACIDS:

- The majority of acids are ternary compounds. They contain three different elements- hydrogen and two other nonmetals.
- If one of the nonmetals is oxygen, the acid is called an oxoacid. Think of oxoacids as combinations of hydrogen ions, H^+ and oxoanions.
- The scheme for naming oxoacids is similar to that outlined for oxoanions, except that the ending -ous is used instead of –ite and -ic instead of –ate.

TABLE 3.6 Nomenclature of Some Oxoacids and Their Salts

Oxidation State	Formula of Acid[a]	Name of Acid[b]	Formula of Salt[b]	Name of Salt
Cl: +1	$HClO$	*Hypo*chlor*ous* acid	$NaClO$	Sodium *hypo*chlor*ite*
Cl: +3	$HClO_2$	Chlor*ous* acid	$NaClO_2$	Sodium chlor*ite*
Cl: +5	$HClO_3$	Chlor*ic* acid	$NaClO_3$	Sodium chlor*ate*
Cl: +7	$HClO_4$	*Per*chlor*ic* acid	$NaClO_4$	Sodium *per*chlor*ate*
N: +3	HNO_2	Nitr*ous* acid	$NaNO_2$	Sodium nitr*ite*
N: +5	HNO_3	Nitr*ic* acid	$NaNO_3$	Sodium nitr*ate*
S: +4	H_2SO_3	Sulfur*ous* acid	Na_2SO_3	Sodium sulf*ite*
S: +6	H_2SO_4	Sulfur*ic* acid	Na_2SO_4	Sodium sulf*ate*

[a]In all these acids, H atoms are bonded to O atoms, not the central nonmetal atom. Often formulas are written to reflect this fact, for instance, HOCl instead of HClO and HOClO instead of $HClO_2$.

[b]In general, the *-ic* and *-ate* names are assigned to compounds in which the central nonmetal atom has an oxidation state equal to the periodic table group number minus 10. Halogen compounds are exceptional in that the *-ic* and *-ate* names are assigned to compounds in which the halogen has an oxidation state of +5 (even though the group number is 17).

EXAMPLE 3-10 Applying Various Rules for Naming Compounds

Name the compounds **(a)** $CuCl_2$; **(b)** ClO_2; **(c)** HIO_4; **(d)** $Ca(H_2PO_4)_2$.

Analyze

$CuCl_2$ and $Ca(H_2PO_4)_2$ are ionic compounds. To name these compounds, we must identify and name the ions. ClO_2 and HIO_4 are molecular compounds. ClO_2 is a binary compound of two nonmetals and HIO_4 is an oxoacid.

Solve

(a) In this compound, the oxidation state of Cu is +2. Because Cu can also exist in the oxidation state of +1, we must clearly distinguish between the two possible chlorides. $CuCl_2$ is copper(II) chloride.

(b) Both Cl and O are nonmetals. ClO_2 is a binary molecular compound called chlorine dioxide.

(c) The oxidation state of I is +7. By analogy to the chlorine-containing oxoacids in Table 3.6, we should name this compound periodic acid (pronounced "purr-eye-oh-dic" acid).

(d) The polyatomic anion $H_2PO_4^-$ is dihydrogen phosphate ion. Two of these ions are present for every Ca^{2+} ion in the compound calcium dihydrogen phosphate.

Assess

A fair bit of memorization is associated with naming compounds correctly. Mastery of this subject usually requires a lot of practice.

PRACTICE EXAMPLE A: Name the compounds SF_6, HNO_2, $Ca(HCO_3)_2$, $FeSO_4$.

PRACTICE EXAMPLE B: Name the compounds NH_4NO_3, PCl_3, HBrO, $AgClO_4$, $Fe_2(SO_4)_3$.

EXAMPLE 3-11 Applying Various Rules in Writing Formulas

Write the formula of the compound **(a)** tetranitrogen tetrasulfide; **(b)** ammonium chromate; **(c)** bromic acid; **(d)** calcium hypochlorite.

Analyze

We must apply our knowledge of prefixes (such as *tetra-*) and endings (such as *-ic*), as well as the names of common polyatomic ions (such as *ammonium*, *chromate*, and *hypochlorite*).

Solve

(a) Molecules of this compound consist of *four* N atoms and *four* S atoms. The formula is N_4S_4.

(b) Two ammonium ions (NH_4^+) must be present for every chromate ion (CrO_4^{2-}). Place parentheses around NH_4^+, followed by the subscript 2. The formula is $(NH_4)_2CrO_4$. (This formula is read as "N-H-4, taken twice, C-R-O-4.")

(c) The *-ic* acid for the oxoacids of the halogens (group 17) has the halogen in the oxidation state of +5. Bromic acid is $HBrO_3$ (analogous to $HClO_3$ in Table 3.6).

(d) Here there are one Ca^{2+} and two ClO^- ions in a formula unit. This leads to the formula $Ca(ClO)_2$.

Assess

Notice that in writing formulas for compounds containing two or more polyatomic ions of the *same type*, as in **(b)** and **(d)**, we put parentheses around the formula of the ion (without the charge), followed by a subscript indicating the number of ions of that type. Proper use and placement of parentheses in writing formulas is important.

PRACTICE EXAMPLE A: Write formulas for the compounds **(a)** boron trifluoride, **(b)** potassium dichromate, **(c)** sulfuric acid, **(d)** calcium chloride.

PRACTICE EXAMPLE B: Write formulas for the compounds **(a)** aluminum nitrate, **(b)** tetraphosphorous decoxide, **(c)** chromium(III) hydroxide, **(d)** iodic acid.

Some Compounds of Greater Complexity:

Some complex substances: Hydrates.

In a hydrate, each formula unit of the compound has associated with it a certain number of water molecules. This does not mean that the compounds are wet, however. The water molecules are incorporated in the solid structure of the compound. The formula shown below signifies six H_2O molecules per formula unit of $CoCl_2$.

$CoCl_2 \cdot 6\ H_2O$ cobalt(II) chloride *hexahydrate.*

129.839 u + (6 * 18.0153 u) = 237.931 u.

$$\%H_2O = \frac{\left(6\text{ mol }H_2O \times \left(\frac{18.02\text{ g }H_2O}{1\text{ mol }H_2O}\right)\right)}{237.9\text{ g }CoCl_2\cdot 6\ H_2O} \times 100\% = 45.45\%\ H_2O$$

The water present in compounds as water of hydration can generally be removed, in part or totally, by heating.

- When the water is totally removed, the resulting compound is said to be anhydrous (without water).
- Anhydrous compounds can be used as water absorbers.
- $CoCl_2$ gains and loses water quite readily and indicates this through a color change. This fact can be used to make a simple moisture detector .

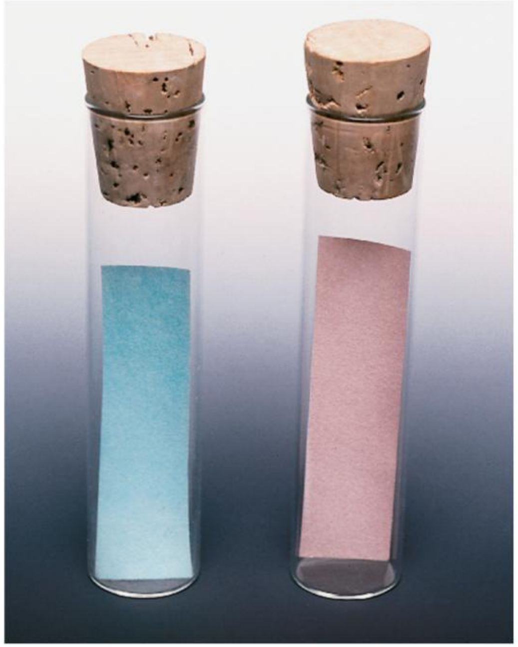


Figure 3-8 Effect of moisture on $CoCl_2$.
Blue anhydrous $CoCl_2$
Pink hexahydrate
$CoCl_2 \cdot 6\ H_2O$

3-7 Names and Formulas of Organic Compounds

Organic compounds abound in nature

- Fats, carbohydrates and proteins → foods.
- Propane, gasoline, kerosene, oil → fuels.
- Drugs and plastics → chemical industries

-Carbon atoms form chains and rings, act as the framework of molecules.

-All organic compounds contain carbon atoms; almost all contain hydrogen atoms; and many common ones also have oxygen, nitrogen, or sulfur atoms.

- These possibilities allow for an almost limitless number of different organic compounds. Organic compounds are mostly molecular; a few are ionic.

-There are millions of organic compounds, many comprising highly complex molecules. Their names are equally complicated.
- A systematic approach to naming these compounds is crucial, and the rules for naming inorganic compounds are of little use here.

Hydrocarbons:

-Compounds containing only carbon and hydrogen are called hydrocarbons.
-Hydrocarbons containing only single bonds are called alkanes.
-The simplest alkane is methane, followed by ethane, then propane

TABLE 3.7 Word Stem (or Prefix) Indicating the Number of Carbon Atoms in Simple Organic Molecules

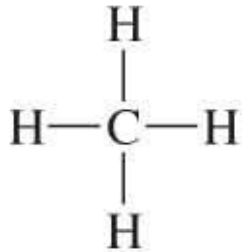
Stem (or prefix)	Number of C Atoms
Meth-	1
Eth-	2
Prop-	3
But-	4
Pent-	5
Hex-	6
Hept-	7
Oct-	8
Non-	9
Dec-	10

The names of the alkanes are composed of two parts:

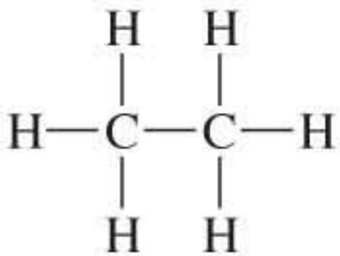
1- a word stem and the

2- ending (suffix) *–ane-* indicating that the molecule is an alkane

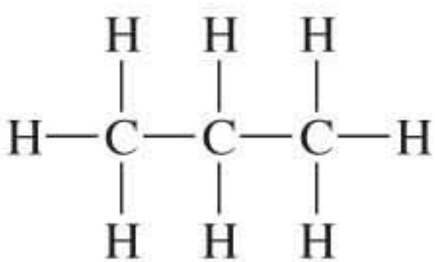
▲



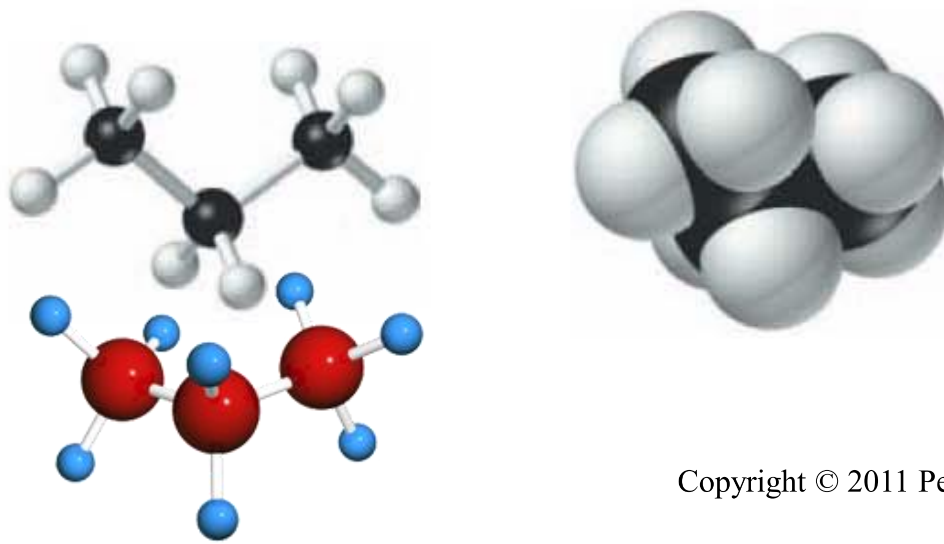
(a) Methane



(b) Ethane



(c) Propane



Alkenes:

-Hydrocarbon molecules with one or more double bonds between carbon atoms

- The simplest of the alkenes is ethene, its name consists of the stem *eth- and the ending -ene. Benzene,* C_6H_6 is a molecule with six carbon atoms arranged in a hexagonal ring

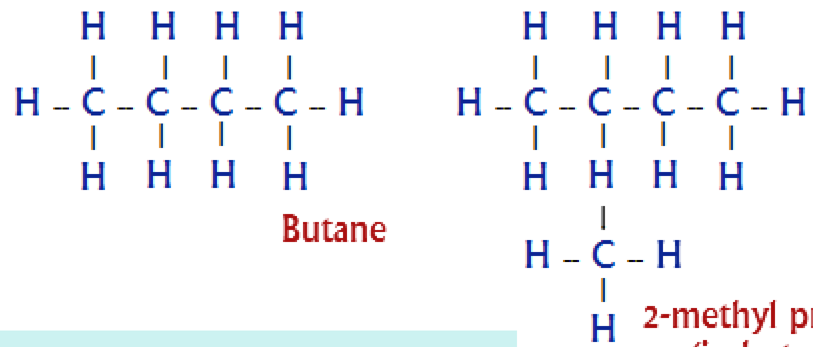
(d) Ethene (ethylene)

(e) Benzene

Isomers:

Butane and Methylpropane have the same molecular formula, C_4H_{10} but different structural formulas.

2 Isomers of C_4H_{10}



Butane

2-methyl propane (isobutane)

Isomers have the same molecular formula but have different arrangements of atoms in space

EXAMPLE 3-12 Recognizing Isomers

Are the following pairs of molecules isomers or not?

(a) $CH_3CH(CH_3)(CH_2)_3CH_3$ and $CH_3CH_2CH(CH_3)(CH_2)_3CH_3$

(b) $CH_3—CH(CH_2—CH_3)—CH_2—CH_3$ and $CH_3—CH(CH_3)—CH_2—CH_2—CH_3$

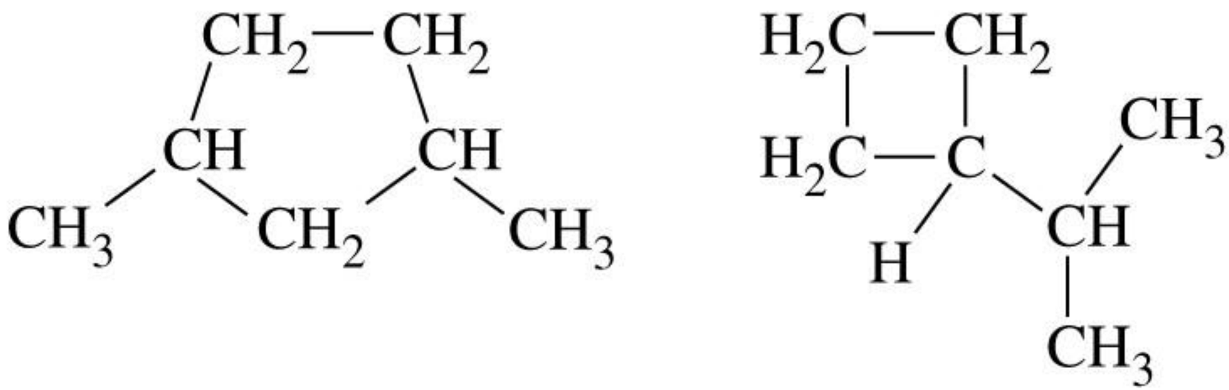
PRACTICE EXAMPLE A: Are the following pairs of molecules isomers?

(a) $CH_3C(CH_3)_2(CH_2)_3CH_3$ and $CH_3CH(CH_3)CH(CH_3)(CH_2)_3CH_3$

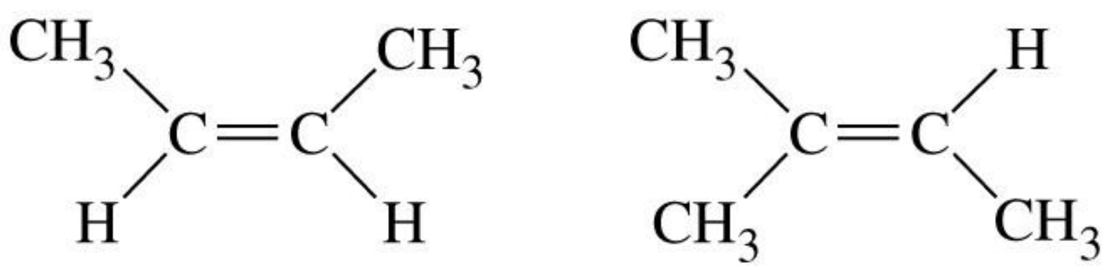
(b) $$\begin{array}{l} CH_3-\underset{\displaystyle CH_2-CH_2-CH_3}{\underset{|}{CH}}-CH_2-CH_3 \end{array} \quad \text{and} \quad CH_3-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_2-\underset{\displaystyle CH_3}{\underset{|}{CH}}-CH_3$$

PRACTICE EXAMPLE B: Are the pairs of molecules represented by the following structural formulas isomers?

(a)



(b)



Functional Groups

-Functional groups are individual atoms or groupings of atoms that are attached to the carbon chains or rings of organic molecules and give the molecules their characteristic properties.

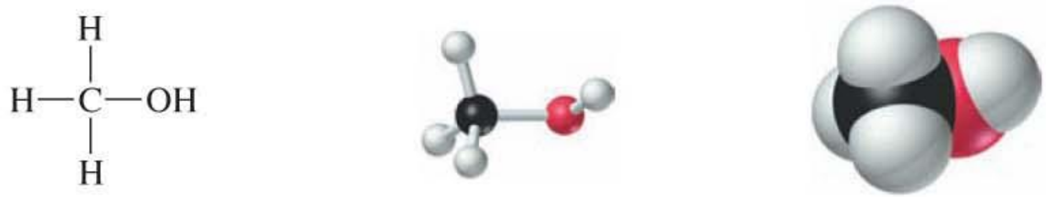
- Compounds with the same functional group generally have similar properties.

- The – OH group in alcohols is one of the many functional groups found in organic compounds

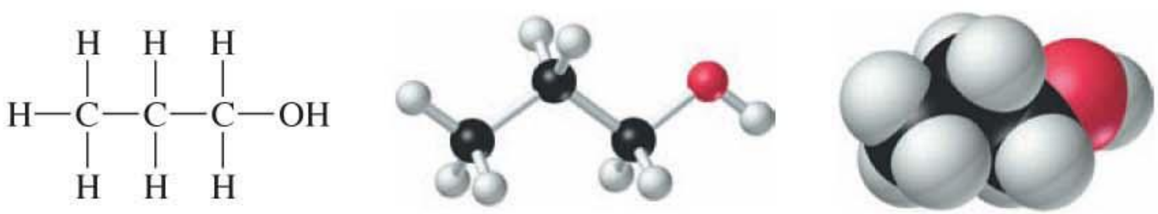
- The – OH group is called the hydroxyl group. The suffix -ol designates the presence of the - OH group in a class of organic molecules called alcohols.

Visualizations of some alcohols

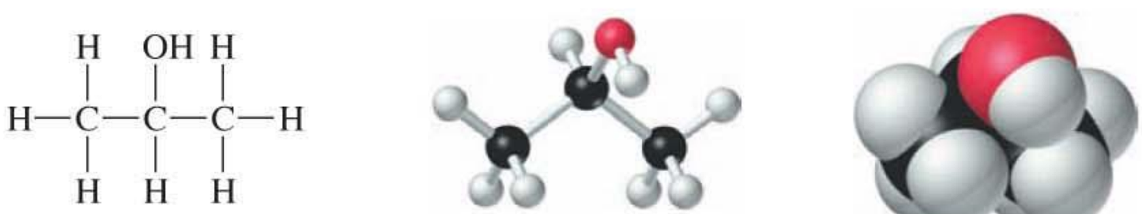
(a) Ethanol, or
Ethyl alcohol



(b) Methanol, or
Methyl alcohol



(c) 1-Propanol, or
Propyl alcohol

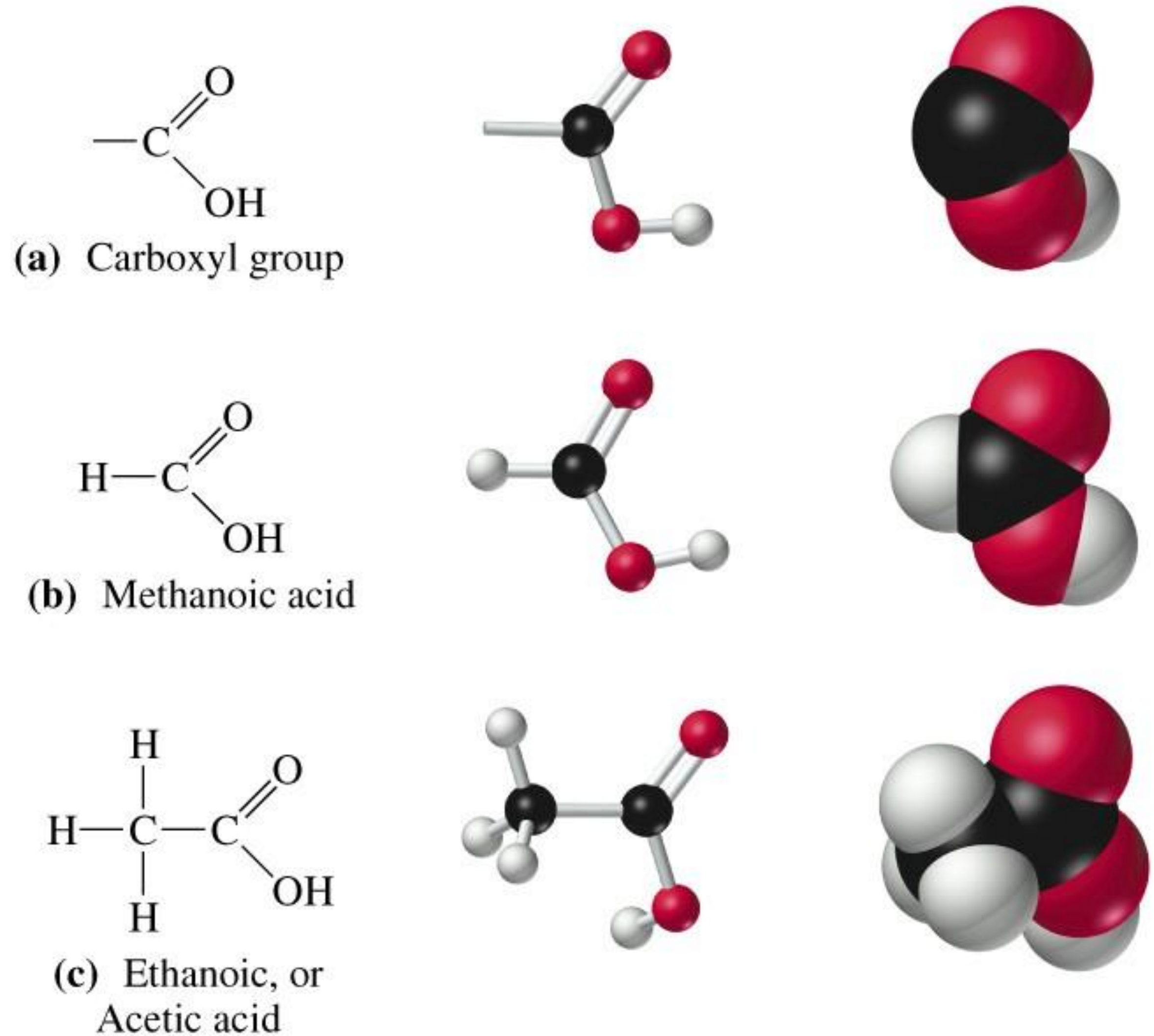


(d) 2-Propanol, or
Isopropyl alcohol

◀ FIGURE 3-10
Visualizations of some alcohols

Another important functional group is the *carboxyl group*, *-COOH*

- The hydrogen attached to one of the O atoms in a carboxyl group is *ionizable or acidic*.
- Compounds containing the carboxyl group are called **carboxylic acids.**



▲ Figure 3-11

The carboxyl group and visualizations of two carboxylic acids

EXAMPLE 3-13 Recognizing Types of Organic Compounds

What type of compound is each of the following?

(a) $CH_3CH_2CH_2CH_3$

(b) $CH_3CHClCH_2CH_3$

(c) $CH_3CH_2CO_2H$

(d) $CH_3CH_2CH(OH)CH_2CH_3$

a) alkane, butane

b) Chloro alkane, 2-chloro butane

c) Carboxylic acid, propanoic acid

d) Alcohol, 3-pentanol

▲

PRACTICE EXAMPLE A: What types of compounds correspond to each of the following formulas?

(a) $CH_3CH_2CH_3$ (b) $ClCH_2CH_2CH_3$

(c) $CH_3CH_2CH_2CO_2H$ (d) $CH_3CHCHCH_3$

PRACTICE EXAMPLE B: What types of compounds correspond to these formulas?

(a) $CH_3CH(OH)CH_3$ (b) $CH_3CH(OH)CH_2CO_2H$

(c) $CH_2ClCH_2CO_2H$ (d) $BrCHCHCH_3$

A)

a) alkane, propane

b) Chloro alkane, 1-chloro propane

c) Carboxylic acid, butanoic acid

d) Alkene, 2-butene

B)

a) alcohol, 2-propanol

b) Carboxylic acid, butanoic acid
Also contains alcohol

c) Carboxylic acid, propanoic acid,
also contains chloride,
chloro carbocylic acid

d) Alkene, contains Br. Bromoalkene.

▲

EXAMPLE 3-14 Naming Organic Compounds

Name these compounds.

(a) $CH_3CH_2CH_2CH_2CH_3$ (b) $CH_3CHFCH_2CH_3$

(c) $CH_3CH_2CO_2H$ (d) $CH_3CH_2CH(OH)CH_2CH_3$

a) alkane, pentane

b) Floro alkane, 2-floro butane

c) Carboxylic acid, propanoic acid

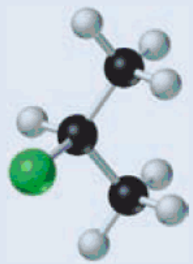
d) Alcohols, 3-pentanol

▲

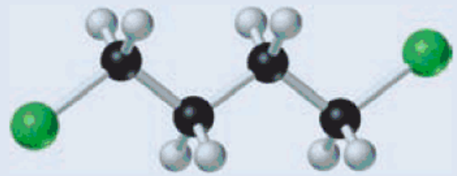
PRACTICE EXAMPLE A: Name the following compounds: **(a)** $CH_3CH(OH)CH_3$; **(b)** $ICH_2CH_2CH_3$; **(c)** $CH_3CH(CH_3)CH_2CO_2H$; **(d)** CH_3CHCH_2.

PRACTICE EXAMPLE B: Give plausible names for the molecules that correspond to the following ball-and-stick models.

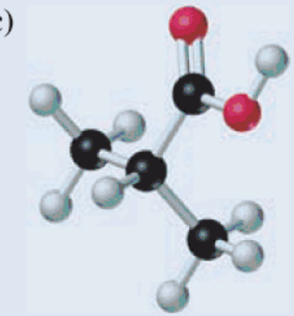
(a)



(b)



(c)



A)

a) 2-propanol

b) 1-iodopropane

c) 3-methyl butanoic acid

d) propene

B)

a) 2-chloro propane

b) 1,4-dichlorobutane

c) 2-methyl propanoic acid

EXAMPLE 3-15 Writing Structural Formulas from the Names of Organic Compounds

Write the condensed structural formula for the organic compounds: **(a)** butane, **(b)** butanoic acid, **(c)** 1-chloropentane, **(d)** 1-hexanol.

Solve

(a) The word stem *but-* indicates a structure with a four-carbon chain, and the suffix *-ane* indicates an alkane. No functional groups are indicated; hence, the condensed structural formula is $CH_3(CH_2)_2CH_3$.

(b) The *-oic* ending indicates that the end carbon atom of the four-carbon chain is part of a carboxylic acid group. The condensed structural formula is $CH_3(CH_2)_2CO_2H$.

(c) The prefix *chloro-* indicates the substitution of a chlorine atom for a H atom, and the 1- designates that it is on the first C atom of the carbon chain. The carbon chain is five C atoms long, as signified by the word stem *pent-*. The condensed structural formula is $CH_3(CH_2)_3CH_2Cl$.

(d) The suffix *-ol* indicates the presence of a hydroxyl group in place of a H atom, and the 1- designates that it is on the first C atom of the carbon chain. The word stem *hex-* signifies that the carbon chain is six C atoms long. The condensed structural formula is $CH_3(CH_2)_4CH_2OH$.

PRACTICE EXAMPLE A: Write the condensed structural formula for the organic compounds **(a)** pentane, **(b)** ethanoic acid, **(c)** 1-iodooctane (pronounced eye-oh-dough-octane), **(d)** 1-pentanol.

PRACTICE EXAMPLE B: Write the line-angle formula for the organic compounds **(a)** propene, **(b)** 1-heptanol, **(c)** chloroacetic acid, **(d)** hexanoic acid.

A)

(a) pentane: $CH_3(CH2)_3CH_3$

(b) ethanoic acid: CH_3CO_2H

(c) 1-iodooctane: $ICH_2(CH_2)_6CH_3$

(d) 1-pentanol: $CH_2(OH)(CH_2)_3CH_3$