## EXAMPLE 2.4 — Formulas from Names for Type I Binary Compounds

Given the following systematic names, write the formula for each compound:

**a.** potassium iodide
**b.** calcium oxide
**c.** gallium bromide

### Solution

| Name | Formula | Comments |
|---|---|---|
| **a.** potassium iodide | KI | Contains $K^+$ and $I^-$. |
| **b.** calcium oxide | CaO | Contains $Ca^{2+}$ and $O^{2-}$. |
| **c.** gallium bromide | $GaBr_3$ | Contains $Ga^{3+}$ and $Br^-$. Must have $3Br^-$ to balance charge of $Ga^{3+}$. |

## EXAMPLE 2.5 — Naming Type II Binary Compounds

**1.** Give the systematic name for each of the following compounds:

**a.** CuCl **b.** HgO **c.** $Fe_2O_3$

**2.** Given the following systematic names, write the formula for each compound:

**a.** Manganese(IV) oxide
**b.** Lead(II) chloride

Type II binary ionic compounds contain a metal that can form more than one type of cation.

A compound must be electrically neutral.

### Solution

All of these compounds include a metal that can form more than one type of cation. Thus we must first determine the charge on each cation. This can be done by recognizing that a compound must be electrically neutral; that is, the positive and negative charges must exactly balance.

**1.**

| Formula | Name | Comments |
|---|---|---|
| **a.** CuCl | Copper(I) chloride | Because the anion is $Cl^-$, the cation must be $Cu^+$ (for charge balance), which requires a Roman numeral I. |
| **b.** HgO | Mercury(II) oxide | Because the anion is $O^{2-}$, the cation must be $Hg^{2+}$ [mercury(II)]. |
| **c.** $Fe_2O_3$ | Iron(III) oxide | The three $O^{2-}$ ions carry a total charge of 6−, so two $Fe^{3+}$ ions [iron(III)] are needed to give a 6+ charge. |

Mercury(II) oxide.

**2.**

| Name | Formula | Comments |
|---|---|---|
| **a.** Manganese(IV) oxide | $MnO_2$ | Two $O^{2-}$ ions (total charge 4−) are required by the $Mn^{4+}$ ion [manganese(IV)]. |
| **b.** Lead(II) chloride | $PbCl_2$ | Two $Cl^-$ ions are required by the $Pb^{2+}$ ion [lead(II)] for charge balance. |

## EXAMPLE 2.7 — Naming Compounds Containing Polyatomic Ions

**1.** Give the systematic name for each of the following compounds:

**a.** $Na_2SO_4$
**b.** $KH_2PO_4$
**c.** $Fe(NO_3)_3$
**d.** $Mn(OH)_2$
**e.** $Na_2SO_3$
**f.** $Na_2CO_3$

**2.** Given the following systematic names, write the formula for each compound:

**a.** Sodium hydrogen carbonate
**b.** Cesium perchlorate
**c.** Sodium hypochlorite
**d.** Sodium selenate
**e.** Potassium bromate

### Solution

**1.**

| Formula | Name | Comments |
|---|---|---|
| **a.** $Na_2SO_4$ | Sodium sulfate | |
| **b.** $KH_2PO_4$ | Potassium dihydrogen phosphate | |
| **c.** $Fe(NO_3)_3$ | Iron(III) nitrate | Transition metal—name must contain a Roman numeral. The $Fe^{3+}$ ion balances three $NO_3^-$ ions. |
| **d.** $Mn(OH)_2$ | Manganese(II) hydroxide | Transition metal—name must contain a Roman numeral. The $Mn^{2+}$ ion balances three $OH^-$ ions. |
| **e.** $Na_2SO_3$ | Sodium sulfite | |
| **f.** $Na_2CO_3$ | Sodium carbonate | |

## EXAMPLE 2.8 Naming Type III Binary Compounds

1. Name each of the following compounds:

   a. $PCl_5$ b. $PCl_3$ c. $SO_2$

2. From the following systematic names, write the formula for each compound:

   a. Sulfur hexafluoride
   b. Sulfur trioxide
   c. Carbon dioxide

### Solution

1.

| Formula | Name |
|---|---|
| a. $PCl_5$ | Phosphorus pentachloride |
| b. $PCl_3$ | Phosphorus trichloride |
| c. $SO_2$ | Sulfur dioxide |

2.

| Name | Formula |
|---|---|
| a. Sulfur hexafluoride | $SF_6$ |
| b. Sulfur trioxide | $SO_3$ |
| c. Carbon dioxide | $CO_2$ |

## EXAMPLE 3.2 Determining the Mass of a Sample of Atoms

Americium is an element that does not occur naturally. It can be made in very small amounts in a device known as a *particle accelerator*. Compute the mass in grams of a sample of americium containing six atoms.

### Solution

From the table inside the front cover of the text, we note that one americium atom has a mass of 243 amu. Thus the mass of six atoms is

$$6\ \cancel{\text{atoms}} \times 243\ \frac{\text{amu}}{\cancel{\text{atom}}} = 1.46 \times 10^{3}\ \text{amu}$$

Using the relationship

$$6.022 \times 10^{23}\ \text{amu} = 1\ \text{g}$$

we write the conversion factor for converting atomic mass units to grams:

$$\frac{1\ \text{g}}{6.022 \times 10^{23}\ \text{amu}}$$

The mass of six americium atoms in grams is

$$1.46 \times 10^{3}\ \cancel{\text{amu}} \times \frac{1\ \text{g}}{6.022 \times 10^{23}\ \cancel{\text{amu}}} = 2.42 \times 10^{-21}\ \text{g}$$

**Reality Check:** Since this sample contains only six atoms, the mass should be very small as the amount $2.42 \times 10^{-21}$ g indicates.

## EXAMPLE 3.4 Calculating Numbers of Atoms

A silicon chip used in an integrated circuit of a microcomputer has a mass of 5.68 mg. How many silicon (Si) atoms are present in the chip?

### Solution

The strategy for doing this problem is to convert from milligrams of silicon to grams of silicon, then to moles of silicon, and finally to atoms of silicon:

Always check to see if your answer is sensible.

Paying careful attention to units and making sure the answer is reasonable can help you detect an inverted conversion factor or a number that was incorrectly entered in your calculator.

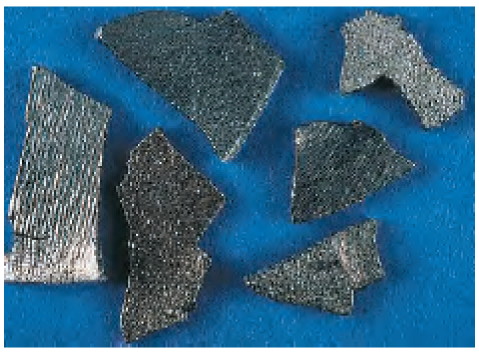
$$5.68\ \cancel{\text{mg Si}} \times \frac{1\ \text{g Si}}{1000\ \cancel{\text{mg Si}}} = 5.68 \times 10^{-3}\ \text{g Si}$$

$$5.68 \times 10^{-3}\ \cancel{\text{g Si}} \times \frac{1\ \text{mol Si}}{28.09\ \cancel{\text{g Si}}} = 2.02 \times 10^{-4}\ \text{mol Si}$$

$$2.02 \times 10^{-4}\ \cancel{\text{mol Si}} \times \frac{6.022 \times 10^{23}\ \text{atoms}}{1\ \cancel{\text{mol Si}}} = 1.22 \times 10^{20}\ \text{atoms}$$

## EXAMPLE 3.5 Calculating the Number of Moles and Mass

Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing $5.00 \times 10^{20}$ atoms and the mass of the sample.



Fragments of cobalt metal.

### Solution

Note that the sample of $5.00 \times 10^{20}$ atoms of cobalt is less than 1 mole ($6.022 \times 10^{23}$ atoms) of cobalt. What fraction of a mole it represents can be determined as follows:

$$5.00 \times 10^{20} \cancel{\text{atoms Co}} \times \frac{1 \text{ mol Co}}{6.022 \times 10^{23} \cancel{\text{atoms Co}}} = 8.30 \times 10^{-4} \text{ mol Co}$$

Since the mass of 1 mole of cobalt atoms is 58.93 g, the mass of $5.00 \times 10^{20}$ atoms can be determined as follows:

$$8.30 \times 10^{-4} \cancel{\text{mol Co}} \times \frac{58.93 \text{ g Co}}{1 \cancel{\text{mol Co}}} = 4.89 \times 10^{-2} \text{ g Co}$$

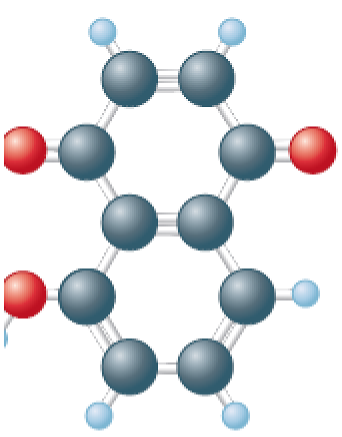
**Reality Check:** In this case the sample contains $5 \times 10^{20}$ atoms, which is approximately 1/1000 of a mole. Thus the sample should have a mass of about $(1/1000)(58.93) \cong 0.06$. Our answer of $\sim 0.05$ makes sense.

## EXAMPLE 3.6 Calculating Molar Mass I

Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants. The formula for juglone is $C_{10}H_6O_3$.

**a.** Calculate the molar mass of juglone.

**b.** A sample of $1.56 \times 10^{-2}$ g of pure juglone was extracted from black walnut husks. How many moles of juglone does this sample represent?



Juglone

### Solution

**a.** The molar mass is obtained by summing the masses of the component atoms. In 1 mole of juglone there are 10 moles of carbon atoms, 6 moles of hydrogen atoms, and 3 moles of oxygen atoms:

$$\begin{aligned} 10\text{ C:} &\quad 10 \times 12.01 \text{ g} = 120.1 \text{ g} \\ 6\text{ H:} &\quad 6 \times 1.008 \text{ g} = 6.048 \text{ g} \\ 3\text{ O:} &\quad 3 \times 16.00 \text{ g} = \underline{48.00 \text{ g}} \\ \text{Mass of 1 mol } C_{10}H_6O_3 &= 174.1 \text{ g} \end{aligned}$$

The mass of 1 mole of juglone is 174.1 g, which is the molar mass.

**b.** The mass of 1 mole of this compound is 174.1 g; thus $1.56 \times 10^{-2}$ g is much less than a mole. The exact fraction of a mole can be determined as follows:

$$1.56 \times 10^{-2} \cancel{\text{g juglone}} \times \frac{1 \text{ mol juglone}}{174.1 \cancel{\text{g juglone}}} = 8.96 \times 10^{-5} \text{ mol juglone}$$

## EXAMPLE 3.9 Calculating Mass Percent

Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula ($C_{10}H_{14}O$) and mass. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone.

### Solution

**Where are we going?**

- To find the mass percent of each element in carvone

*What do we know?*

✓ Molecular formula, $C_{10}H_{14}O$

*What information do we need to find the mass percent?*

✓ Mass of each element (we'll use 1 mol carvone)

✓ Molar mass of carvone

**How do we get there?**

*What is the mass of each element in 1 mol $C_{10}H_{14}O$?*

$$\text{Mass of C in 1 mol} = 10 \cancel{\text{mol}} \times 12.01 \frac{\text{g}}{\cancel{\text{mol}}} = 120.1 \text{ g}$$

$$\text{Mass of H in 1 mol} = 14 \cancel{\text{mol}} \times 1.008 \frac{\text{g}}{\cancel{\text{mol}}} = 14.11 \text{ g}$$

$$\text{Mass of O in 1 mol} = 1 \cancel{\text{mol}} \times 16.00 \frac{\text{g}}{\cancel{\text{mol}}} = 16.00 \text{ g}$$

*What is the molar mass of $C_{10}H_{14}O$?*

$$\begin{array}{ccccccc} 120.1 \text{ g} & + & 14.11 \text{ g} & + & 16.00 \text{ g} & = & 150.2 \text{ g} \\ C_{10} & + & H_{14} & + & O & = & C_{10}H_{14}O \end{array}$$

*What is the mass percent of each element?*

We find the fraction of the total mass contributed by each element and convert it to a percentage:

- $\text{Mass percent of C} = \dfrac{120.1 \text{ g C}}{150.2 \text{ g } C_{10}H_{14}O} \times 100\% = 79.96\%$
- $\text{Mass percent of H} = \dfrac{14.11 \text{ g H}}{150.2 \text{ g } C_{10}H_{14}O} \times 100\% = 9.394\%$
- $\text{Mass percent of O} = \dfrac{16.00 \text{ g O}}{150.2 \text{ g } C_{10}H_{14}O} \times 100\% = 10.65\%$

## EXAMPLE 3.10 Determining Empirical and Molecular Formulas I

Determine the empirical and molecular formulas for a compound that gives the following percentages on analysis (in mass percents):

71.65% Cl    24.27% C    4.07% H

The molar mass is known to be 98.96 g/mol.

**Solution**

**Where are we going?**

- To find the empirical and molecular formulas for the given compound

*What do we know?*

✓ Percent of each element

✓ Molar mass of the compound is 98.96 g/mol

*What information do we need to find the empirical formula?*

✓ Mass of each element in 100.00 g of compound

✓ Moles of each element

**How do we get there?**

*What is the mass of each element in 100.00 g of compound?*

Cl 71.65 g    C 24.27 g    H 4.07 g

*What are the moles of each element in 100.00 g of compound?*

$$71.65 \text{ g} \cancel{\text{Cl}} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g} \cancel{\text{Cl}}} = 2.021 \text{ mol Cl}$$

$$24.27 \text{ g} \cancel{\text{C}} \times \frac{1 \text{ mol C}}{12.01 \text{ g} \cancel{\text{C}}} = 2.021 \text{ mol C}$$

$$4.07 \text{ g} \cancel{\text{H}} \times \frac{1 \text{ mol H}}{1.008 \text{ g} \cancel{\text{H}}} = 4.04 \text{ mol H}$$

*What is the empirical formula for the compound?*

Dividing each mole value by 2.021 (the smallest number of moles present), we find the empirical formula $ClCH_2$.

*What is the molecular formula for the compound?*

Compare the empirical formula mass to the molar mass.

Empirical formula mass = 49.48 g/mol (Confirm this!)

Molar mass is given = 98.96 g/mol

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96 \text{ g/mol}}{49.48 \text{ g/mol}} = 2$$

$$\text{Molecular formula} = (ClCH_2)_2 = Cl_2C_2H_4$$

- This substance is composed of molecules with the formula $Cl_2C_2H_4$.

## EXAMPLE 3.14 Balancing a Chemical Equation II

At 1000°C, ammonia gas, $NH_3(g)$, reacts with oxygen gas to form gaseous nitric oxide, $NO(g)$, and water vapor. This reaction is the first step in the commercial production of nitric acid by the Ostwald process. Balance the equation for this reaction.

**Solution**

**1, 2 ▶** The unbalanced equation for the reaction is

$$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(g)$$

**3 ▶** Because all the molecules in this equation are of about equal complexity, where we start in balancing it is rather arbitrary. Let's begin by balancing the hydrogen. A coefficient of 2 for $NH_3$ and a coefficient of 3 for $H_2O$ give six atoms of hydrogen on both sides:

$$2NH_3(g) + O_2(g) \rightarrow NO(g) + 3H_2O(g)$$

The nitrogen can be balanced with a coefficient of 2 for NO:

$$2NH_3(g) + O_2(g) \rightarrow 2NO(g) + 3H_2O(g)$$

Finally, note that there are two atoms of oxygen on the left and five on the right. The oxygen can be balanced with a coefficient of $\frac{5}{2}$ for $O_2$:

$$2NH_3(g) + \frac{5}{2}O_2(g) \rightarrow 2NO(g) + 3H_2O(g)$$

However, the usual custom is to have whole-number coefficients. We simply multiply the entire equation by 2.

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

## EXAMPLE 3.15 Chemical Stoichiometry I

Solid lithium hydroxide is used in space vehicles to remove exhaled carbon dioxide from the living environment by forming solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can be absorbed by 1.00 kg of lithium hydroxide?

### Solution

#### Where are we going?

- To find the mass of $CO_2$ absorbed by 1.00 kg LiOH

*What do we know?*

✓ Chemical reaction

$$LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

✓ 1.00 kg LiOH

*What information do we need to find the mass of $CO_2$?*

✓ Balanced equation for the reaction

#### How do we get there?

1 ▶ *What is the balanced equation?*

$$2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(l)$$

2 ▶ *What are the moles of LiOH?*

To find the moles of LiOH, we need to know the molar mass.

*What is the molar mass for LiOH?*

$$6.941 + 16.00 + 1.008 = 23.95 \text{ g/mol}$$

Now we use the molar mass to find the moles of LiOH:

$$1.00 \cancel{\text{kg LiOH}} \times \frac{1000 \cancel{\text{g LiOH}}}{1 \cancel{\text{kg LiOH}}} \times \frac{1 \text{ mol LiOH}}{23.95 \cancel{\text{g LiOH}}} = 41.8 \text{ mol LiOH}$$

3 ▶ *What is the mole ratio between $CO_2$ and LiOH in the balanced equation?*

$$\frac{1 \text{ mol } CO_2}{2 \text{ mol LiOH}}$$

4 ▶ *What are the moles of $CO_2$?*

$$41.8 \cancel{\text{mol LiOH}} \times \frac{1 \text{ mol } CO_2}{2 \cancel{\text{mol LiOH}}} = 20.9 \text{ mol } CO_2$$

5 ▶ *What is the mass of $CO_2$ formed from 1.00 kg LiOH?*

$$20.9 \cancel{\text{mol } CO_2} \times \frac{44.0 \text{ g } CO_2}{1 \cancel{\text{mol } CO_2}} = 9.20 \times 10^2 \text{ g } CO_2$$

- Thus 920. g of $CO_2(g)$ will be absorbed by 1.00 kg of LiOH($s$).

## EXAMPLE 3.17 Stoichiometry: Limiting Reactant

Nitrogen gas can be prepared by passing gaseous ammonia over solid copper(II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. If a sample containing 18.1 g of $NH_3$ is reacted with 90.4 g of CuO, which is the limiting reactant? How many grams of $N_2$ will be formed?

### Solution

#### Where are we going?

- To find the limiting reactant
- To find the mass of $N_2$ produced

*What do we know?*

✓ The chemical reaction

$$NH_3(g) + CuO(s) \longrightarrow N_2(g) + Cu(s) + H_2O(g)$$

✓ 18.1 g $NH_3$

✓ 90.4 g CuO

*What information do we need?*

✓ Balanced equation for the reaction

✓ Moles of $NH_3$

✓ Moles of CuO

#### How do we get there?

**To find the limiting reactant**

*What is the balanced equation?*

$$2NH_3(g) + 3CuO(s) \longrightarrow N_2(g) + 3Cu(s) + 3H_2O(g)$$

*What are the moles of $NH_3$ and CuO?*

To find the moles, we need to know the molar masses.

$NH_3$ 17.03 g/mol
CuO 79.55 g/mol

$$18.1 \cancel{\text{g } NH_3} \times \frac{1 \text{ mol } NH_3}{17.03 \cancel{\text{g } NH_3}} = 1.06 \text{ mol } NH_3$$

$$90.4 \cancel{\text{g CuO}} \times \frac{1 \text{ mol CuO}}{79.55 \cancel{\text{g CuO}}} = 1.14 \text{ mol CuO}$$

*What is the mole ratio between $NH_3$ and CuO in the balanced equation?*

$$\frac{3 \text{ mol CuO}}{2 \text{ mol } NH_3}$$

*How many moles of CuO are required to react with 1.06 mol $NH_3$?*

$$1.06 \cancel{\text{mol } NH_3} \times \frac{3 \text{ mol CuO}}{2 \cancel{\text{mol } NH_3}} = 1.59 \text{ mol CuO}$$

- Thus 1.59 mol CuO is required to react with 1.06 mol $NH_3$. Since only 1.14 mol CuO is actually present, the amount of CuO is limiting; CuO will run out before $NH_3$ does. We can verify this conclusion by comparing the mole ratio of CuO and $NH_3$ required by the balanced equation:

$$\frac{\text{mol CuO}}{\text{mol } NH_3} (\text{required}) = \frac{3}{2} = 1.5$$

with the mole ratio actually present:

$$\frac{\text{mol CuO}}{\text{mol } NH_3} (\text{actual}) = \frac{1.14}{1.06} = 1.08$$

- Since the actual ratio is too small (smaller than 1.5), CuO is the limiting reactant.

**To find the mass of $N_2$ produced**

*What are the moles of $N_2$ formed?*

Because CuO is the limiting reactant, we must use the amount of CuO to calculate the amount of $N_2$ formed.

*What is the mole ratio between $N_2$ and CuO in the balanced equation?*

$$\frac{1 \text{ mol } N_2}{3 \text{ mol CuO}}$$

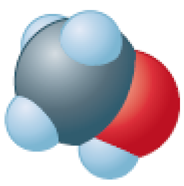
*What are the moles of $N_2$?*

$$1.14\ \cancel{\text{mol CuO}} \times \frac{1 \text{ mol } N_2}{3\ \cancel{\text{mol CuO}}} = 0.380 \text{ mol } N_2$$

*What mass of $N_2$ is produced?*
Using the molar mass of $N_2$ (28.02 g/mol), we can calculate the mass of $N_2$ produced:

- $0.380\ \cancel{\text{mol } N_2} \times \dfrac{28.02 \text{ g } N_2}{1\ \cancel{\text{mol } N_2}} = 10.6 \text{ g } N_2$

## EXAMPLE 3.18 Calculating Percent Yield



Methanol

Methanol ($CH_3OH$), also called *methyl alcohol*, is the simplest alcohol. It is used as a fuel in race cars and is a potential replacement for gasoline. Methanol can be manufactured by combining gaseous carbon monoxide and hydrogen. Suppose 68.5 kg CO(*g*) is reacted with 8.60 kg $H_2(g)$. Calculate the theoretical yield of methanol. If $3.57 \times 10^4$ g $CH_3OH$ is actually produced, what is the percent yield of methanol?

### Solution

**Where are we going?**

- To calculate the theoretical yield of methanol
- To calculate the percent yield of methanol

*What do we know?*

✓ The chemical reaction

$$H_2(g) + CO(g) \longrightarrow CH_3OH(l)$$

✓ 68.5 kg CO(*g*)
✓ 8.60 kg $H_2(g)$
✓ $3.57 \times 10^4$ g $CH_3OH$ is produced

*What information do we need?*

✓ Balanced equation for the reaction
✓ Moles of $H_2$
✓ Moles of CO
✓ Which reactant is limiting
✓ Amount of $CH_3OH$ produced

**How do we get there?**

**To find the limiting reactant**

*What is the balanced equation?*

$$2H_2(g) + CO(g) \longrightarrow CH_3OH(l)$$

*What are the moles of $H_2$ and CO?*
To find the moles, we need to know the molar masses.

| | |
|---|---|
| $H_2$ | 2.016 g/mol |
| CO | 28.02 g/mol |

$$68.5\ \cancel{\text{kg CO}} \times \frac{1000\ \cancel{\text{g CO}}}{1\ \cancel{\text{kg CO}}} \times \frac{1 \text{ mol CO}}{28.02\ \cancel{\text{g CO}}} = 2.44 \times 10^3 \text{ mol CO}$$

$$8.60\ \cancel{\text{kg } H_2} \times \frac{1000\ \cancel{\text{g } H_2}}{1\ \cancel{\text{kg } H_2}} \times \frac{1 \text{ mol } H_2}{2.016\ \cancel{\text{g } H_2}} = 4.27 \times 10^3 \text{ mol } H_2$$

*What is the mole ratio between $H_2$ and CO in the balanced equation?*

$$\frac{2 \text{ mol } H_2}{1 \text{ mol CO}}$$

*How does the actual mole ratio compare to the stoichiometric ratio?*
To determine which reactant is limiting, we compare the mole ratio of $H_2$ and CO required by the balanced equation

$$\frac{\text{mol } H_2}{\text{mol CO}} \text{ (required)} = \frac{2}{1} = 2$$

with the actual mole ratio

$$\frac{\text{mol } H_2}{\text{mol CO}} \text{ (actual)} = \frac{4.27 \times 10^3}{2.44 \times 10^3} = 1.75$$

- Since the actual mole ratio of $H_2$ to CO is smaller than the required ratio, *$H_2$ is limiting.*

**To calculate the theoretical yield of methanol**

*What are the moles of $CH_3OH$ formed?*
We therefore must use the amount of $H_2$ and the mole ratio between $H_2$ and $CH_3OH$ to determine the maximum amount of methanol that can be produced:

$$4.27 \times 10^3\ \cancel{\text{mol } H_2} \times \frac{1 \text{ mol } CH_3OH}{2\ \cancel{\text{mol } H_2}} = 2.14 \times 10^3 \text{ mol } CH_3OH$$

*What is the theoretical yield of $CH_3OH$ in grams?*

$$2.14 \times 10^3\ \cancel{\text{mol } CH_3OH} \times \frac{32.04 \text{ g } CH_3OH}{1\ \cancel{\text{mol } CH_3OH}} = 6.86 \times 10^4 \text{ g } CH_3OH$$

- Thus, from the amount of reactants given, the maximum amount of $CH_3OH$ that can be formed is $6.86 \times 10^4$ g. This is the *theoretical yield.*

*What is the percent yield of $CH_3OH$?*

- $\dfrac{\text{Actual yield (grams)}}{\text{Theoretical yield (grams)}} \times 100 = \dfrac{3.57 \times 10^4\ \cancel{\text{g } CH_3OH}}{6.86 \times 10^4\ \cancel{\text{g } CH_3OH}} \times 100\% = 52.0\%$

## EXAMPLE 4.2 Calculation of Molarity II

Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl in enough water to make 26.8 mL of solution.

### Solution

**Where are we going?**

- To find the molarity of HCl solution

*What do we know?*

✓ 1.56 g HCl
✓ 26.8 mL solution

*What information do we need to find molarity?*

✓ Moles solute

✓ Molarity $= \dfrac{\text{mol solute}}{\text{L solution}}$

### How do we get there?

*What are the moles of HCl (36.46 g/mol)?*

$$1.56 \cancel{\text{g HCl}} \times \frac{1 \text{ mol HCl}}{36.46 \cancel{\text{g HCl}}} = 4.28 \times 10^{-2} \text{ mol HCl}$$

*What is the volume of solution (in liters)?*

$$26.8 \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 2.68 \times 10^{-2} \text{ L}$$

*What is the molarity of the solution?*

- $$\text{Molarity} = \frac{4.28 \times 10^{-2} \text{ mol HCl}}{2.68 \times 10^{-2} \text{ L solution}} = 1.60\ M \text{ HCl}$$

## EXAMPLE 4.3 Concentration of Ions I

Give the concentration of each type of ion in the following solutions:

**a.** 0.50 $M$ $Co(NO_3)_2$
**b.** 1 $M$ $Fe(ClO_4)_3$

### Solution

### Where are we going?

- To find the molarity of each ion in the solution

*What do we know?*

✓ 0.50 $M$ $Co(NO_3)_2$
✓ 1 $M$ $Fe(ClO_4)_3$

*What information do we need to find the molarity of each ion?*

✓ Moles of each ion

### How do we get there?

**For $Co(NO_3)_2$**

*What is the balanced equation for dissolving the ions?*

$$Co(NO_3)_2(s) \xrightarrow{H_2O} Co^{2+}(aq) + 2NO_3^-(aq)$$

*What is the molarity for each ion?*

- $Co^{2+}$ $1 \times 0.50\ M = 0.50\ M\ Co^{2+}$
- $NO_3^-$ $2 \times 0.50\ M = 1.0\ M\ NO_3^-$

An aqueous solution of $Co(NO_3)_2$.

**For $Fe(ClO_4)_3$**

*What is the balanced equation for dissolving the ions?*

$$Fe(ClO_4)_3(s) \xrightarrow{H_2O} Fe^{3+}(aq) + 3ClO_4^-(aq)$$

*What is the molarity for each ion?*

- $Fe^{3+}$ $1 \times 1\ M = 1\ M\ Fe^{3+}$
- $ClO_4^-$ $3 \times 1\ M = 3\ M\ ClO_4^-$

## EXAMPLE 4.7 Concentration and Volume

What volume of 16 $M$ sulfuric acid must be used to prepare 1.5 L of a 0.10 $M$ $H_2SO_4$ solution?

### Solution

### Where are we going?

- To find the volume of $H_2SO_4$ required to prepare the solution

*What do we know?*

✓ 1.5 L of 0.10 $M$ $H_2SO_4$ is required
✓ We have 16 $M$ $H_2SO_4$

*What information do we need to find the volume of $H_2SO_4$?*

✓ Moles of $H_2SO_4$ in the required solution

### How do we get there?

*What are the moles of $H_2SO_4$ required?*

$$M \times V = \text{mol}$$

$$1.5 \cancel{\text{L solution}} \times \frac{0.10 \text{ mol } H_2SO_4}{\cancel{\text{L solution}}} = 0.15 \text{ mol } H_2SO_4$$

*What volume of 16 $M$ $H_2SO_4$ contains 0.15 mol $H_2SO_4$?*

$$V \times \frac{16 \text{ mol } H_2SO_4}{\text{L solution}} = 0.15 \text{ mol } H_2SO_4$$

## EXAMPLE 7.6 — Electron Subshells

For principal quantum level $n = 5$, determine the number of allowed subshells (different values of $\ell$), and give the designation of each.

### Solution

For $n = 5$, the allowed values of $\ell$ run from 0 to 4 ($n - 1 = 5 - 1$). Thus the subshells and their designations are

| $\ell = 0$ | $\ell = 1$ | $\ell = 2$ | $\ell = 3$ | $\ell = 4$ |
|---|---|---|---|---|
| 5*s* | 5*p* | 5*d* | 5*f* | 5*g* |

## EXAMPLE 7.7 — Electron Configurations

Give the electron configurations for sulfur (S), cadmium (Cd), hafnium (Hf), and radium (Ra) using the periodic table inside the front cover of this book.

### Solution

*Sulfur* is element 16 and resides in Period 3, where the 3*p* orbitals are being filled (see Fig. 7.30). Since sulfur is the fourth among the "3*p* elements," it must have four 3*p* electrons. Its configuration is

$$\text{S:}\quad 1s^22s^22p^63s^23p^4 \quad \text{or} \quad [\text{Ne}]3s^23p^4$$

*Cadmium* is element 48 and is located in Period 5 at the end of the 4*d* transition metals, as shown in Fig. 7.30. It is the tenth element in the series and thus has 10 electrons in the 4*d* orbitals, in addition to the 2 electrons in the 5*s* orbital. The configuration is

$$\text{Cd:}\quad 1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10} \quad \text{or} \quad [\text{Kr}]5s^24d^{10}$$

## EXAMPLE 7.8 — Trends in Ionization Energies

The first ionization energy for phosphorus is 1060 kJ/mol, and that for sulfur is 1005 kJ/mol. Why?

### Solution

Phosphorus and sulfur are neighboring elements in Period 3 of the periodic table and have the following valence electron configurations: Phosphorus is $3s^23p^3$, and sulfur is $3s^23p^4$.

Ordinarily, the first ionization energy increases as we go across a period, so we might expect sulfur to have a greater ionization energy than phosphorus. However, in this case the fourth *p* electron in sulfur must be placed in an already occupied orbital. The electron–electron repulsions that result cause this electron to be more easily removed than might be expected.

## EXAMPLE 7.10 — Trends in Radii

Predict the trend in radius for the following ions: $Be^{2+}$, $Mg^{2+}$, $Ca^{2+}$, and $Sr^{2+}$.

### Solution

All these ions are formed by removing two electrons from an atom of a Group 2A element. In going from beryllium to strontium, we are going down the group, so the sizes increase:

$$\underset{\substack{\uparrow \\ \text{Smallest radius}}}{Be^{2+}} < Mg^{2+} < Ca^{2+} < \underset{\substack{\uparrow \\ \text{Largest radius}}}{Sr^{2+}}$$

## EXAMPLE 8.1 — Relative Bond Polarities

Order the following bonds according to polarity: H—H, O—H, Cl—H, S—H, and F—H.

### Solution

The polarity of the bond increases as the difference in electronegativity increases. From the electronegativity values in Fig. 8.3, the following variation in bond polarity is expected (the electronegativity value appears in parentheses below each element):

$$\underset{(2.1)(2.1)}{\text{H—H}} < \underset{(2.5)(2.1)}{\text{S—H}} < \underset{(3.0)(2.1)}{\text{Cl—H}} < \underset{(3.5)(2.1)}{\text{O—H}} < \underset{(4.0)(2.1)}{\text{F—H}}$$

Electronegativity difference: 0 | 0.4 | 0.9 | 1.4 | 1.9

Covalent bond $\xrightarrow[\text{Polarity increases}]{}$ Polar covalent bond

## EXAMPLE 8.4 — Relative Ion Size II

Choose the largest ion in each of the following groups.

**a.** $Li^+$, $Na^+$, $K^+$, $Rb^+$, $Cs^+$
**b.** $Ba^{2+}$, $Cs^+$, $I^-$, $Te^{2-}$

### Solution

**a.** The ions are all from Group 1A elements. Since size increases down a group (the ion with the greatest number of electrons is largest), $Cs^+$ is the largest ion.
**b.** This is an isoelectronic series of ions, all of which have the xenon electron configuration. The ion with the smallest nuclear charge is largest:

$$Te^{2-} > I^- > Cs^+ > Ba^{2+}$$

## EXAMPLE 8.6 Writing Lewis Structures

Give the Lewis structure for each of the following.

**a.** HF  **d.** $CH_4$
**b.** $N_2$  **e.** $CF_4$
**c.** $NH_3$  **f.** $NO^+$

### Solution

In each case we apply the three steps for writing Lewis structures. Recall that lines are used to indicate shared electron pairs and that dots are used to indicate nonbonding pairs (lone pairs). We have the following tabulated results:

| | Total Valence Electrons | Draw Single Bonds | Calculate Number of Electrons Remaining | Use Remaining Electrons to Achieve Noble Gas Configurations | Check Number of Electrons |
|---|---|---|---|---|---|
| **a.** HF | $1 + 7 = 8$ | H—F | 6 | H—$\ddot{\underset{\cdot\cdot}{\mathrm{F}}}$: | H, 2<br>F, 8 |
| **b.** $N_2$ | $5 + 5 = 10$ | N—N | 8 | :N≡N: | N, 8 |
| **c.** $NH_3$ | $5 + 3(1) = 8$ | H—N(—H)—H | 2 | H—$\ddot{\mathrm{N}}$(—H)—H | H, 2<br>N, 8 |
| **d.** $CH_4$ | $4 + 4(1) = 8$ | $CH_4$: H—C(—H)(—H)—H | 0 | H—C(—H)(—H)—H | H, 2<br>C, 8 |
| **e.** $CF_4$ | $4 + 4(7) = 32$ | F—C(—F)(—F)—F | 24 | :$\ddot{\underset{\cdot\cdot}{\mathrm{F}}}$—C(—$\ddot{\underset{\cdot\cdot}{\mathrm{F}}}$:)(—$\ddot{\underset{\cdot\cdot}{\mathrm{F}}}$:)—$\ddot{\underset{\cdot\cdot}{\mathrm{F}}}$: | F, 8<br>C, 8 |
| **f.** $NO^+$ | $5 + 6 - 1 = 10$ | N—O | 8 | [:N≡O:]$^+$ | N, 8<br>O, 8 |

## EXAMPLE 8.7 Lewis Structures for Molecules That Violate the Octet Rule I

Write the Lewis structure for $PCl_5$.

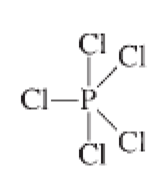
### Solution

We can follow the same stepwise procedure we used above for sulfur hexafluoride.

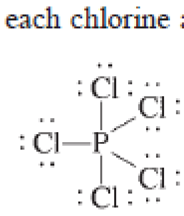
**1 ▶** Sum the valence electrons.

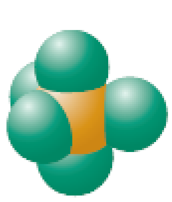
$$\underset{\mathrm{P}}{\underset{\uparrow}{5}} + 5(\underset{\mathrm{Cl}}{\underset{\uparrow}{7}}) = 40 \text{ electrons}$$

**2 ▶** Indicate single bonds between bound atoms.



**3 ▶** Distribute the remaining electrons. In this case, 30 electrons (40 − 10) remain. These are used to satisfy the octet rule for each chlorine atom. The final Lewis structure is





$PCl_5$

Note that phosphorus, which is a third-row element, has exceeded the octet rule by two electrons.

## EXAMPLE 22.2 Isomerism and Nomenclature

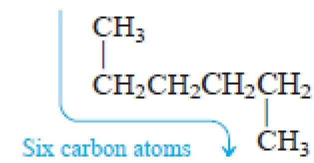
Draw the structural isomers for the alkane $C_6H_{14}$ and give the systematic name for each one.

### Solution

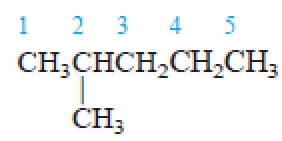
We will proceed systematically, starting with the longest chain and then rearranging the carbons to form the shorter, branched chains.

1. $CH_3CH_2CH_2CH_2CH_2CH_3$ Hexane

   Note that although a structure such as

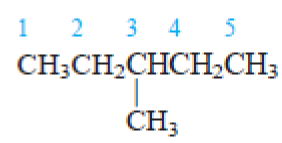



   may look different it is still hexane, since the longest carbon chain has six atoms.
2. We now take one carbon out of the chain and make it a methyl substituent.



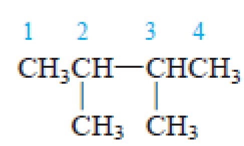

2-Methylpentane

   Since the longest chain consists of five carbons, this is a substituted pentane: 2-methylpentane. The 2 indicates the position of the methyl group on the chain. Note that if we numbered the chain from the right end, the methyl group would be on carbon 4. Because we want the smallest possible number, the numbering shown is correct.
3. The methyl substituent can also be on carbon 3 to give



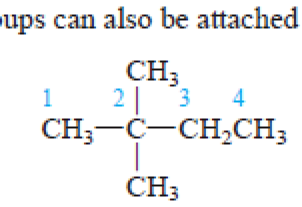

3-Methylpentane

   Note that we have now exhausted all possibilities for placing a single methyl group on pentane.
4. Next, we can take two carbons out of the original six-member chain:



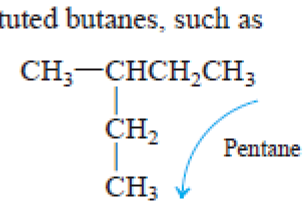

2,3-Dimethylbutane

   Since the longest chain now has four carbons, the root name is butane. Since there are two methyl groups, we use the prefix *di-*. The numbers denote that the two methyl groups are positioned on the second and third carbons in the butane chain. Note that when two or more numbers are used, they are separated by a comma.
5. The two methyl groups can also be attached to the same carbon atom as shown here:

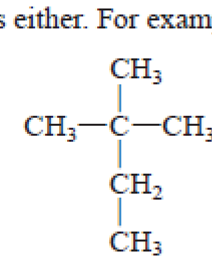



2,2-Dimethylbutane

   We might also try ethyl-substituted butanes, such as




   However, note that this is instead a pentane (3-methylpentane), since the longest chain has five carbon atoms. Thus it is not a new isomer. Trying to reduce the chain to three atoms provides no further isomers either. For example, the structure

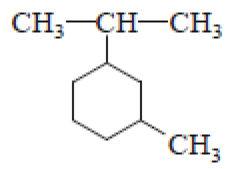
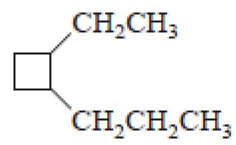


   is actually 2,2-dimethylbutane.

Thus there are only five distinct structural isomers of $C_6H_{14}$: hexane, 2-methylpentane, 3-methylpentane, 2,3-dimethylbutane, and 2,2-dimethylbutane.

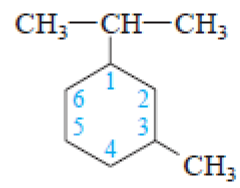
## EXAMPLE 22.4 Naming Cyclic Alkanes

Name each of the following cyclic alkanes.

a.  b. 

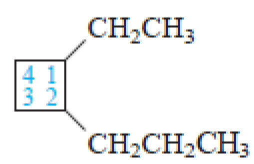
### Solution

a. The six-carbon cyclohexane ring is numbered as follows:



There is an isopropyl group at carbon 1 and a methyl group at carbon 3. The name is 1-isopropyl-3-methylcyclohexane, since the alkyl groups are named in alphabetical order.

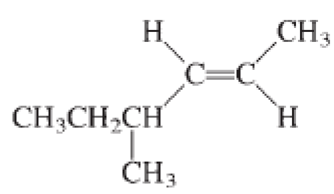
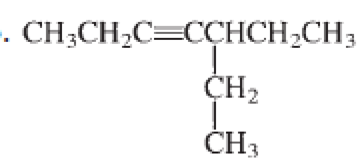
b. This is a cyclobutane ring, which is numbered as follows:



The name is 1-ethyl-2-propylcyclobutane.

## EXAMPLE 22.5 Naming Alkenes and Alkynes

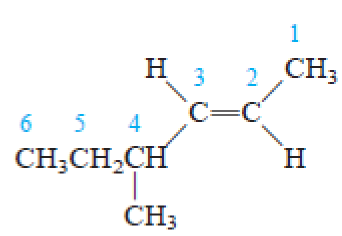
Name each of the following molecules.

a.  b. 


A worker using an oxyacetylene torch.

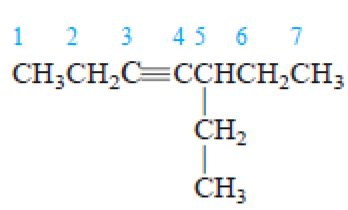
### Solution

a. The longest chain, which contains six carbon atoms, is numbered as follows:



Thus the hydrocarbon is a 2-hexene. Since the hydrogen atoms are located on opposite sides of the double bond, this molecule corresponds to the *trans* isomer. The name is *trans*-4-methyl-2-hexene.

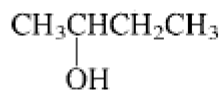
b. The longest chain, consisting of seven carbon atoms, is numbered as shown (giving the triple bond the lowest possible number):



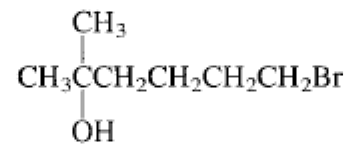
The hydrocarbon is a 3-heptyne. The full name is 5-ethyl-3-heptyne, where the position of the triple bond is indicated by the lower-numbered carbon atom involved in this bond.

## EXAMPLE 22.6 Naming and Classifying Alcohols

For each of the following alcohols, give the systematic name and specify whether the alcohol is primary, secondary, or tertiary.

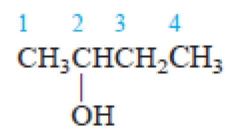
a. 

b. $ClCH_2CH_2CH_2OH$

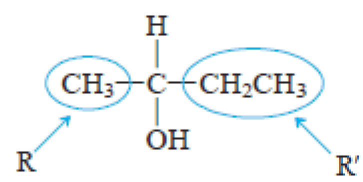
c. 

**Solution**

a. The chain is numbered as follows:



The compound is called 2-butanol, since the —OH group is located at the number 2 position of a four-carbon chain. Note that the carbon to which the —OH is attached also has $—CH_3$ and $—CH_2CH_3$ groups attached:

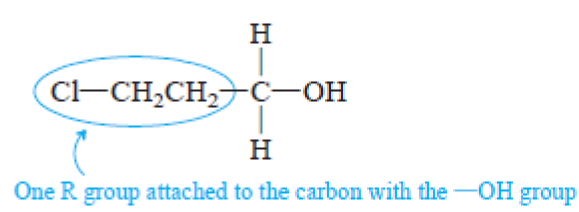


Therefore, this is a *secondary* alcohol.

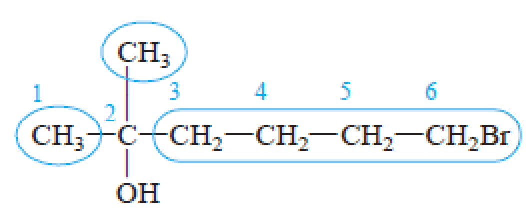
b. The chain is numbered as follows:

$$\overset{3}{Cl—CH_2}—\overset{2}{CH_2}—\overset{1}{CH_2}—OH$$

The name is 3-chloro-1-propanol. This is a *primary* alcohol:




c. The chain is numbered as follows:

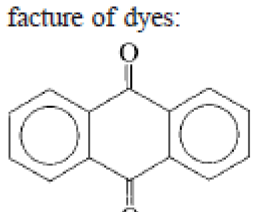


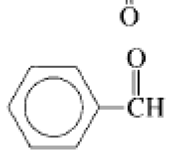
The name is 6-bromo-2-methyl-2-hexanol. This is a *tertiary* alcohol since the carbon where the —OH is attached also has three other R groups attached.

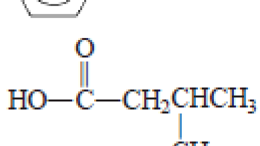
### Functional Groups

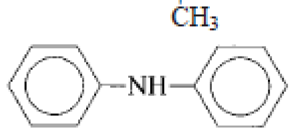
47. Identify each of the following compounds as a carboxylic acid, ester, ketone, aldehyde, or amine.

a. Anthraquinone, an important starting material in the manufacture of dyes:



b. 

c. 

d. 

isomers of methylcyclobutane **47. a.** ketone; **b.** aldehyde; **c.** carboxylic acid; **d.** amine