## Example Exercise 9.1 Atomic Mass and Avogadro's Number

Refer to the atomic masses in the periodic table inside the front cover of this textbook. State the mass of Avogadro's number of atoms for each of the following elements:
(a) Copper
(c) sulfur
(b) Mercury
(d) helium

## Solution

The atomic mass of each element is listed below the symbol of the element in the periodic table: $\mathrm{Cu}=$ $63.55 \mathrm{amu}, \mathrm{Hg}=200.59 \mathrm{amu}, \mathrm{S}=32.07 \mathrm{amu}$, and $\mathrm{He}=4.00 \mathrm{amu}$. The mass of Avogadro's number of atoms is the atomic mass expressed in grams. Therefore, $6.02 \times 10^{23}$ atoms of
(a) $\mathrm{Cu}=63.55 \mathrm{~g}$
(c) $\mathrm{S}=32.07 \mathrm{~g}$
(b) $\mathrm{Hg}=200.59 \mathrm{~g}$
(d) $\mathrm{He}=4.00 \mathrm{~g}$

## Practice Exercise

Refer to the periodic table and state the mass for each of the following:
(a) 1 atom of Au
(b) $6.02 \times 1023$ atoms of Au

Answers: (a) 196.97 amu ; (b) 196.97 g

## Concept Exercise

What is the mass of an average platinum atom? What is the mass of Avogadro's number of Pt atoms?
Answer: See Appendix G.

## Example Exercise 9.2 Mole Calculations I

Calculate the number of sodium atoms in 0.120 mol Na .

## Strategy Plan

Step 1: What unit is asked for in the answer?
Step 2: What given value is related to the answer?
Step 1:

Step 2:


Since $1 \mathrm{~mol} \mathrm{Na}=6.02 \times 10^{23}$ atoms Na , the two unit factors are $1 \mathrm{~mol} \mathrm{Na} / 6.02 \times 1023$ atoms Na , and its reciprocal $6.02 \times 10^{23}$ atoms $\mathrm{Na} / 1 \mathrm{~mol} \mathrm{Na}$.

## Unit Analysis Map


unit in answer
$=$ ? atoms Na

## Example Exercise 9.2 Mole Calculations I

## Continued

## Solution

We apply the unit factor $6.02 \times 10^{23}$ atoms $\mathrm{Na} / 1 \mathrm{~mol} \mathrm{Na}$ to cancel moles $\mathrm{Na}(\mathrm{meH} \mathrm{Na})$, which appears in the denominator.

$$
0.120 \mathrm{mel} \mathrm{Na} \times \frac{6.02 \times 10^{23} \text { atoms } \mathrm{Na}}{1 \mathrm{mel} \mathrm{Na}}=7.22 \times 10^{22} \text { atoms } \mathrm{Na}
$$

The answer is rounded to three digits because the given value and unit factor each have three significant digits.

## Practice Exercise

Calculate the number of formula units in 0.0763 mol of sodium chloride, NaCl .

Answers: $4.59 \times 10^{22}$ formula units NaCl

## Concept Exercise

What is the number of molecules in 1.00 mol of any gas?
Answer: See Appendix G.

## Example Exercise 9.3 Mole Calculations I

Calculate the number of moles of potassium in $1.25 \times 10^{21}$ atoms K .

## Strategy Plan

Step 1: What unit is asked for in the answer?
Step 1:


Step 2: What given value is related to the answer? Step 2:

$$
1.25 \times 10^{21} \text { atoms K }
$$

Step 3: What unit factor(s) should we apply?
Since mol K $=6.02 \times 10^{23}$ atoms $K$, the two unit factors are $1 \mathrm{~mol} \mathrm{~K} / 6.02 \times 10^{23}$ atoms K , and its reciprocal $6.02 \times 10^{23}$ atoms $\mathrm{K} / 1 \mathrm{~mol} \mathrm{~K}$.

## Unit Analysis Map

given value
$1.25 \times 10^{21}$ atoms K

unit in answer
$=\quad ? \mathrm{~mol} \mathrm{~K}$

## Example Exercise 9.3 Mole Calculations I

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol} \mathrm{~K} / 6.02 \times 10^{23}$ atoms K to cancel atoms K (atems K , which appears in the denominator.

$$
1.25 \times 10^{21} \text { atems } \mathrm{K} \times \frac{1 \mathrm{~mol} \mathrm{~K}}{6.02 \times 10^{23} \text { atems K }}=2.08 \times 10^{-3} \mathrm{~mol} \mathrm{~K}
$$

The answer is rounded to three digits because the given value and unit factor each have three significant digits.

## Practice Exercise

Calculate the number of moles of potassium iodide in $5.34 \times 10^{25}$ formula units of KI.
Answer: 88.7 mol KI

## Concept Exercise

What is the number of molecules in 1.00 mol of iodine crystals?

Answer: See Appendix G.

## Example Exercise 9.4 Molar Mass Calculations

Calculate the molar mass for each of the following substances:
(a) silver metal, Ag
(b) ammonia gas, $\mathrm{NH}_{3}$
(c) magnesium nitrate, $\mathrm{Mg}\left(\mathrm{NO}_{3}\right)_{2}$

## Solution

We begin by finding the atomic mass of each element in the periodic table. The molar mass equals the sum of the atomic masses expressed in $\mathrm{g} / \mathrm{mol}$.
(a) The atomic mass of Ag is 107.87 amu , and the molar mass of silver equals $107.87 \mathrm{~g} / \mathrm{mol}$.
(b) The sum of the atomic masses for $\mathrm{NH}_{3}$ is $14.01 \mathrm{amu}+3(1.01) \mathrm{amu}=17.04 \mathrm{amu}$. The molar mass of ammonia equals $17.04 \mathrm{~g} / \mathrm{mol}$.
(c) The sum of the atomic masses for $\mathrm{Mg}\left(\mathrm{NO}_{3}\right)_{2}$ is
$24.31 \mathrm{amu}+2(14.01+16.00+16.00+16.00) \mathrm{amu}=148.33 \mathrm{amu}$.
The molar mass of magnesium nitrate equals $148.33 \mathrm{~g} / \mathrm{mol}$.

## Practice Exercise

Calculate the molar mass for each of the following substances:
$\begin{array}{lll}\text { (a) manganese metal, } \mathrm{Mn} & \text { (b) sulfur hexafluoride, } \mathrm{SF}_{6} & \text { (c) strontium acetate, } \mathrm{Sr}\left(\mathrm{C}_{2} \mathrm{H}_{3} \mathrm{O}_{2}\right)_{2}\end{array}$
Answers: (a) $54.94 \mathrm{~g} / \mathrm{mol}$; (b) $146.07 \mathrm{~g} / \mathrm{mol}$; (c) $205.72 \mathrm{~g} / \mathrm{mol}$

## Concept Exercise

The molecular mass of water is 18.02 amu . What is the mass of Avogadro's number of water molecules?
Answer: See Appendix G.

## Example Exercise 9.5 Mole Calculations II

What is the mass in grams of $2.01 \times 10^{22}$ atoms of sulfur?

## Strategy Plan

Step 1: What unit is asked for in the answer?
Step 2: What given value is related to the answer?


By definition, $1 \mathrm{~mol} \mathrm{~S}=6.02 \times 10^{23}$ atoms S , and the molar mass from the periodic table is $32.07 \mathrm{~g} \mathrm{~S}=1 \mathrm{~m}$ the two pairs of unit factors are shown in Step 3.
$\frac{1 \mathrm{~mol} \mathrm{~S}}{32.07 \mathrm{~g} \mathrm{~S}} \quad$ or $\quad \frac{32.07 \mathrm{~g} \mathrm{~S}}{1 \mathrm{~mol} \mathrm{~S}}$

## Unit Analysis Map

```
    given value
2.01 }\times1\mp@subsup{0}{}{22}\mathrm{ atoms S
```



## Example Exercise 9.5 Mole Calculations II

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol} \mathrm{~S} / 6.02 \times 10^{23}$ atoms S to cancel atoms $£_{\text {(atems } S)}$, and $32.07 \mathrm{~g} \mathrm{~S} / 1 \mathrm{~mol} \mathrm{~S}$ to cancel mol S (mots):

$$
2.01 \times 10^{22} \text { atoms } \mathrm{S} \times \frac{1 \mathrm{mots}}{6.02 \times 10^{23} \text { atems S }} \times \frac{32.07 \mathrm{~g} \mathrm{~S}}{1 \mathrm{motS}}=1.07 \mathrm{~g} \mathrm{~S}
$$

## Practice Exercise

What is the mass of $7.75 \times 10^{22}$ formula units of lead(II) sulfide, PbS ?
Answer: 30.8 g PbS

## Concept Exercise

Sulfur occurs naturally as $S_{8}$ molecules. What is the mass of Avogadro's number of $S_{8}$ molecules?
Answer: See Appendix G.

## Example Exercise 9.6 Mole Calculations II

How many $\mathrm{O}_{2}$ molecules are present in 0.470 g of oxygen gas?

## Strategy Plan

Step 1: What unit is asked for in the answer?

Step 2: What given value is related to the answer?

Step 3: What unit factor(s) should we apply?


Step 3:

$\frac{1 \mathrm{~mol} \mathrm{O}_{2}}{6.02 \times 10^{23} \text { molecules } \mathrm{O}_{2}} \quad$ or $\quad \frac{6.02 \times 10^{23} \text { molecules } \mathrm{O}_{2}}{1 \mathrm{~mol} \mathrm{O}_{2}}$

## Unit Analysis Map



## Example Exercise 9.6 Mole Calculations II

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol} \mathrm{O}_{2} / 32.00 \mathrm{~g} \mathrm{O}_{2}$ to cancel $\mathrm{g} \mathrm{O}_{2}\left(\mathrm{~g} \mathrm{O}_{2}\right)$, and $6.02 \times 10^{23}$ molecules $\mathrm{O}_{2} / 1 \mathrm{~mol} \mathrm{O}_{2}$ to cancel moles $\mathrm{O}_{2}\left(\mathrm{meto}_{2}\right)$. Thus,

$$
0.470 \mathrm{~g}_{2} \times \frac{1 \mathrm{~mol}_{2}}{32.00 \mathrm{gO}_{2}^{2}} \times \frac{6.02 \times 10^{23} \mathrm{molecules} \mathrm{O}_{2}}{1 \mathrm{~mol}_{2}}=8.84 \times 10^{21} \text { molecules } \mathrm{O}_{2}
$$

## Practice Exercise

How many formula units of lithium fluoride are found in 0.175 g LiF ?
Answer: $4.06 \times 10^{21}$ formula units LiF

## Concept Exercise

Ozone occurs naturally as $\mathrm{O}_{3}$ molecules. What is the mass of Avogadro's number of ozone molecules?
Answer: See Appendix G.

## Example Exercise 9.7 Mass of a Molecule

Calculate the mass in grams for a single molecule of carbon dioxide, $\mathrm{CO}_{2}$ (given that $\mathrm{MM}=44.01 \mathrm{~g} / \mathrm{mol}$ ).

## Strategy Plan

Step 1: What unit is asked for in the answer?
Step 1: $\quad \mathrm{g} \mathrm{CO}_{2} /$ molecule $\mathrm{CO}_{2}$

Step 2: What given value is related to the answer?
Step 2:

$$
44.01 \mathrm{~g} \mathrm{CO}_{2} / 1 \mathrm{~mol} \mathrm{CO}_{2}
$$

Step 3: What unit factor(s) should we apply?
Step 3:


By definition, 1 mol $\mathrm{CO}_{2}=6.02 \times 10^{23}$ molecules.
Thus, the unit factors are shown in Step 3.

## Unit Analysis Map

given value
$44.01 \mathrm{~g} \mathrm{CO}_{2}$
$1 \mathrm{~mol} \mathrm{CO}_{2}$


## Example Exercise 9.7 Mass of a Molecule

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol} \mathrm{CO} 2 / 6.02 \times 10^{23}$ molecules $\mathrm{CO}_{2}$ to cancel $\mathrm{mol} \mathrm{CO}_{2}$ ( $\mathrm{melCO} \mathrm{CO}_{2}$ ).

$$
\frac{44.01 \mathrm{~g}}{1 \mathrm{molCO}_{2}} \times \frac{1 \mathrm{molCO}_{2}}{6.02 \times 10^{23} \text { molecules }}=7.31 \times 10^{-23} \mathrm{~g} / \mathrm{molecule}
$$

## Practice Exercise

Calculate the mass in grams for a single molecule of carbon monoxide, CO.
Answer: $4.65 \times 10^{-23} \mathrm{~g} /$ molecule

## Concept Exercise

What is the mass of a carbon monoxide molecule expressed in amu?
Answer: See Appendix G.

## Example Exercise 9.8 Density of a Gas at STP

Calculate the density of ammonia gas, $\mathrm{NH}_{3}$, at STP. ( $\mathrm{MM}=17.04 \mathrm{~g} / \mathrm{mol}$ )

## Strategy Plan

Step 1: What unit is asked for in the answer?
Step 1:
$\mathrm{g} / \mathrm{mol}$

Step 2: What given value is related to the answer? Step 2:

$$
1.96 \mathrm{~g} / 1.00 \mathrm{~L}
$$

Step 3: What unit factor(s) should we apply?
Step 3:


From the molar volume of a gas at STP, $1 \mathrm{~mol}=22.4 \mathrm{~L}$; thus, the unit factors are $1 \mathrm{~mol} / 22.4 \mathrm{~L}$, and $22.4 \mathrm{~L} / 1 \mathrm{~mol}$.

## Unit Analysis Map


unit in answer
$=\frac{\mathrm{g}}{\mathrm{L}}$

## Example Exercise 9.8 Density of a Gas at STP

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol} / 22.4 \mathrm{~L}$ to cancel moles (mot), which appears in the denominator.

$$
\frac{17.04 \mathrm{~g}}{1 \mathrm{~mol}} \times \frac{1 \mathrm{~mol}}{22.4 \mathrm{~L}}=0.761 \mathrm{~g} / \mathrm{L}
$$

If we filled a balloon with ammonia gas ( $0.761 \mathrm{~g} / \mathrm{L}$ ), the balloon would float in air ( $1.29 \mathrm{~g} / \mathrm{L}$ ). Similarly, a less dense liquid floats on a more dense liquid; for example, gasoline ( $0.7 \mathrm{~g} / \mathrm{mL}$ ) floats on water $(1.0 \mathrm{~g} / \mathrm{mL})$.

## Practice Exercise

Calculate the density of hydrogen sulfide gas, $\mathrm{H}_{2} \mathrm{~S}$, at STP.
Answer: 1.52 g/L

## Concept Exercise

Why does the balloon filled with helium gas rise, while the balloon filled with nitrogen gas does not?

Answer: See Appendix G.


## Example Exercise 9.9 Molar Mass of a Gas

A fire extinguisher releases 1.96 g of an unknown gas that occupies 1.00 L at STP. What is the molar mass ( $\mathrm{g} / \mathrm{mol}$ ) of the unknown gas?

## Strategy Plan

Step 1: What unit is asked for in the answer?
Step 1:


Step 2: What given value is related to the answer? Step 2:


Step 3: What unit factor(s) should we apply?
Step 3:


From the molar volume of a gas at STP, $1 \mathrm{~mol}=22.4 \mathrm{~L}$; thus, the unit factors are $1 \mathrm{~mol} / 22.4 \mathrm{~L}$, and $22.4 \mathrm{~L} / 1 \mathrm{~mol}$.

## Unit Analysis Map

```
given value
    1.96 g
    1.00 L
```



$$
\begin{aligned}
& \text { unit in answer } \\
& \qquad=\frac{\mathrm{g}}{\mathrm{~mol}}
\end{aligned}
$$

## Example Exercise 9.9 Molar Mass of a Gas

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol} / 22.4 \mathrm{~L}$ to cancel liters $()$, which appears in the denominator.

$$
\frac{1.96 \mathrm{~g}}{1.00 \mathrm{~K}} \times \frac{22.4 \mathrm{~K}}{1 \mathrm{~mol}}=43.9 \mathrm{~g} / \mathrm{mol}
$$

Because the unknown gas is from a fire extinguisher, we suspect that it is carbon dioxide. By summing the molar mass for $\mathrm{CO}_{2}(44.01 \mathrm{~g} / \mathrm{mol})$, we help confirm that the unknown gas is carbon dioxide.

## Practice Exercise

Boron trifluoride gas is used in the manufacture of computer chips. Given that 1.505 g of the gas occupies 497 mL at STP, what is the molar mass of boron fluoride gas?

Answer: $67.8 \mathrm{~g} / \mathrm{mol}$

## Concept Exercise

What is the volume of one mole of any gas at STP?
Answer: See Appendix G.

## Example Exercise 9.10 Mole Calculations III

What is the mass of 3.36 L of ozone gas, O3, at STP?

## Strategy Plan

Step 1: What unit is asked for in the answer?
Step 2: What given value is related to the answer?


Step 3: What unit factor(s) should we apply?
From the molar volume of a gas at STP, $1 \mathrm{~mol} \mathrm{O}_{3}=22.4 \mathrm{~L} \mathrm{O}_{3}$. The molar mass from the periodic table is $48.00 \mathrm{~g} \mathrm{O}_{3}=1 \mathrm{~mol} \mathrm{O}_{3}$.
Thus, the two pairs of unit factors are shown in Step 3.

## Unit Analysis Map

given value
$3.36 \mathrm{~L} \mathrm{O}_{3}$

unit in answer
$=? \mathrm{~g} \mathrm{O}_{3}$

## Example Exercise 9.10 Mole Calculations III

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol} \mathrm{O}_{3} / 22.4 \mathrm{~L} \mathrm{O}_{3}$ to cancel $\mathrm{L} \mathrm{O}_{3}\left(\mathrm{LO}_{3}\right)$, and $48.00 \mathrm{~g} \mathrm{O}_{3} / 1 \mathrm{~mol} \mathrm{O} 3$ to cancel moles $\mathrm{O}_{3}\left(\mathrm{meHO}_{3}\right)$,

$$
3.36 \mathrm{LO}_{3} \times \frac{1 \mathrm{~mol}_{3}}{22.4 \mathrm{LO}_{3}} \times \frac{48.00 \mathrm{~g} \mathrm{O}_{3}}{1 \mathrm{~mol}_{3}}=7.20 \mathrm{~g} \mathrm{O}_{3}
$$

## Practice Exercise

What volume is occupied by 0.125 g of methane gas, $\mathrm{CH}_{4}$, at STP?
Answer: $0.174 \mathrm{~L} \mathrm{CH}_{4}(174 \mathrm{~mL})$

## Concept Exercise

How many molecules of ozone are in one mole of gas?
Answer: See Appendix G.

## Example Exercise 9.11 Mole Calculations III

How many molecules of hydrogen gas, $\mathrm{H}_{2}$, occupy 0.500 L at STP?

## Strategy Plan

Step 1: What unit is asked for in the answer?


Step 2: What given value is related to the answer?
$0.500 \mathrm{~L} \mathrm{H}_{2}$



From the molar volume of a gas at STP, $1 \mathrm{~mol} \mathrm{H}_{2}=22.4 \mathrm{~L} \mathrm{H}_{2}$ and by definition, $1 \mathrm{~mol} \mathrm{H}_{2}=6.02 \times 10^{23}$ molecules $\mathrm{H}_{2}$. Thus, the two pairs of unit factors are shown in Step 3.
Step 3: What unit factor(s) should we apply?

## Unit Analysis Map

```
given value
    0.500 L H2
```



## Example Exercise 9.11 Mole Calculations III

## Continued

## Solution

We apply the unit factor $1 \mathrm{~mol}_{2} / 22.4 \mathrm{~L} \mathrm{H}_{2}$ to cancel $\mathrm{L}_{2}{ }_{\left(\mathrm{LH}_{2}\right)}$, and $6.02 \times 10^{23}$ molecules $\mathrm{H}_{2} / 1 \mathrm{~mol} \mathrm{H}_{2}$ to cancel moles $77 \mathrm{H}_{2}\left(\mathrm{molH}_{2}\right)$. Thus,

$$
0.500 \mathrm{LH}_{2} \times \frac{1 \mathrm{molH}_{2}}{22.4 \mathrm{LH}_{2}} \times \frac{6.02 \times 10^{23} \text { molecules } \mathrm{H}_{2}}{1 \mathrm{molH}_{2}}=1.34 \times 10^{22} \text { molecules } \mathrm{H}_{2}
$$

## Practice Exercise

What is the volume occupied by $3.33 \times 10^{21}$ atoms of helium gas, He , at STP?
Answer: 0.124 L He (124 mL)

## Concept Exercise

What is the mass of one mole of hydrogen gas?
Answer: See Appendix G.

## Example Exercise 9.12 Percent Composition of a Substance

TNT (trinitrotoluene) is a white crystalline substance that explodes at $240^{\circ} \mathrm{C}$. Calculate the percent composition of TNT, $\mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}$.

## Solution

Let's calculate the percent composition assuming there is 1 mol of TNT. For compounds with parentheses, it is necessary to count the number of atoms of each element carefully. That is, 1 mol of $\mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}$ contains 7 mol of C atoms, 5 mol of H atoms, 3 mol of $\mathrm{NO}_{2}$ (that is, 3 mol N atoms and 6 mol of O atoms).
We begin the calculation by finding the molar mass of $\mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}$ as follows:

$$
\begin{aligned}
7(12.01 \mathrm{~g} \mathrm{C})+5(1.01 \mathrm{~g} \mathrm{H})+3(14.01 \mathrm{~g} \mathrm{~N}+32.00 \mathrm{~g} \mathrm{O}) & =\mathrm{g} \mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3} \\
84.07 \mathrm{~g} \mathrm{C}+5.05 \mathrm{~g} \mathrm{H}+42.03 \mathrm{~g} \mathrm{~N}+96.00 \mathrm{~g} \mathrm{O} & =227.15 \mathrm{~g} \mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}
\end{aligned}
$$

Now, let's compare the mass of each element to the total molar mass of the compound, that is, 227.15 g .

$$
\begin{aligned}
& \frac{84.07 \mathrm{~g} \mathrm{C}}{227.15 \mathrm{~g} \mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}} \times 100 \%=37.01 \% \mathrm{C} \\
& \frac{5.05 \mathrm{~g} \mathrm{H}}{227.15 \mathrm{~g} \mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}} \times 100 \%=2.22 \% \mathrm{H} \\
& \frac{42.03 \mathrm{~g} \mathrm{~N}}{227.15 \mathrm{~g} \mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}} \times 100 \%=18.50 \% \mathrm{~N} \\
& \frac{96.00 \mathrm{~g} \mathrm{O}}{227.15 \mathrm{~g} \mathrm{C}_{7} \mathrm{H}_{5}\left(\mathrm{NO}_{2}\right)_{3}} \times 100 \%=42.26 \% \mathrm{O}
\end{aligned}
$$

The percent composition of TNT reveals that the explosive is mostly oxygen by mass. When we sum the individual percentages ( $37.01+2.22+18.50+42.26=99.99 \%$ ), we verify our calculation and, find that the total is approximately $100 \%$.

## Example Exercise 9.12 Percent Composition of a Substance

## Continued

## Practice Exercise

EDTA (ethylenediaminetetraacetic acid) is used as a food preservative and in the treatment of metallic lead poisoning. Calculate the percent composition of EDTA, $\mathrm{C}_{10} \mathrm{H}_{16} \mathrm{~N}_{2} \mathrm{O}_{8}$.

Answers: $41.09 \%$ C, $5.53 \% \mathrm{H}, 9.587 \%$ N, and $43.79 \%$ O

## Concept Exercise

If an analysis of sugar, $\mathrm{C}_{\mathrm{x}} \mathrm{H}_{\mathrm{y}} \mathrm{O}_{z}$, gave $40.0 \% \mathrm{C}$ and $6.7 \% \mathrm{H}$, what is the percent oxygen?
Answers: See Appendix G.

## Example Exercise 9.13 Empirical Formula from Mass Composition

In a laboratory experiment, 0.500 g of scandium was heated and allowed to react with oxygen from the air. The resulting product oxide had a mass of 0.767 g . Now, let's find the empirical formula for scandium oxide, $\mathrm{Sc}_{?} \mathrm{O}_{\text {? }}$.

## Solution

The empirical formula is the simplest whole-number ratio of scandium and oxygen in the compound scandium oxide. This ratio is experimentally determined from the moles of each reactant. The moles of scandium are calculated as follows:

$$
0.500 \mathrm{gSc} \times \frac{1 \mathrm{~mol} \mathrm{Sc}}{44.96 \mathrm{gSc}}=0.0111 \mathrm{~mol} \mathrm{Sc}
$$

The moles of oxygen are calculated after first subtracting the mass of Sc from the product, $\mathrm{Sc}_{?} \mathrm{O}_{?}$ :

$$
0.767 \mathrm{~g} \mathrm{Sc}_{\mathrm{x}} \mathrm{O}_{\mathrm{y}}-0.500 \mathrm{~g} \mathrm{Sc}=0.267 \mathrm{~g} \mathrm{O}
$$

The moles of oxygen are calculated from the mass of oxygen that reacted:

$$
0.267 \mathrm{~g} O \times \frac{1 \mathrm{~mol} \mathrm{O}}{16.00 \mathrm{~g} \theta}=0.0167 \mathrm{~mol} \mathrm{O}
$$

The mole ratio in scandium oxide is $\mathrm{Sc}_{0.0111} \mathrm{O}_{0.0167}$. To simplify the ratio and obtain small whole numbers, we divide by the smaller number:

$$
\mathrm{Sc} \frac{0.0111}{0.0111} \mathrm{O} \frac{0.0167}{0.0111}=\mathrm{Sc}_{1.00} \mathrm{O}_{1.5 \mathrm{~d}}
$$

Since the calculated ratio is not close to whole numbers, we cannot round off the experimental ratio $\mathrm{Sc}_{1.00} \mathrm{O}_{1.50}$. However, we can double the ratio to obtain whole numbers, that is, $\mathrm{Sc}_{2.00} \mathrm{O}_{3.00}$. Thus, the empirical formula is $\mathrm{Sc}_{2} \mathrm{O}_{3}$.

## Example Exercise 9.13 Empirical Formula from Mass Composition

## Continued

## Practice Exercise

Iron can react with chlorine gas to give two different compounds, $\mathrm{FeCl}_{2}$ and $\mathrm{FeCl}_{3}$. If 0.558 g of metallic iron reacts with chlorine gas to yield 1.621 g of iron chloride, which iron compound is produced in the experiment?

Answer: $\mathrm{FeCl}_{3}$

## Concept Exercise

If 0.500 mol of yellow powder sulfur reacts with 0.500 mol of oxygen gas, what is the empirical formula of the sulfur oxide?

Answer: See Appendix G.

## Example Exercise 9.14 Empirical Formula from Percent Composition

Glycine is an amino acid found in protein. An analysis of glycine gave the following data: $32.0 \%$ carbon, $6.7 \%$ hydrogen, $18.7 \%$ nitrogen, and $42.6 \%$ oxygen. Calculate the empirical formula of the amino acid.

## Solution

If we assume a 100 g sample, then the percentage of each element equals its mass in 100 g of glycine, that is, $32.0 \mathrm{~g} \mathrm{C}, 6.7 \mathrm{~g} \mathrm{H}, 18.7 \mathrm{~g} \mathrm{~N}$, and 42.6 g O . We can determine the empirical formula as follows:

$$
\begin{aligned}
& 32.0 \mathrm{ge} \times \frac{1 \mathrm{~mol} \mathrm{C}}{12.01 \mathrm{ge}}=2.66 \mathrm{~mol} \mathrm{C} \\
& 6.7 \mathrm{gH} \times \frac{1 \mathrm{~mol} \mathrm{H}}{1.01 \mathrm{gH}}=6.6 \mathrm{~mol} \mathrm{H} \\
& 18.7 \mathrm{gH} \times \frac{1 \mathrm{~mol} \mathrm{~N}}{14.01 \mathrm{gN}}=1.33 \mathrm{~mol} \mathrm{~N} \\
& 42.6 \mathrm{~g} \theta \times \frac{1 \mathrm{~mol} \mathrm{O}}{16.00 \mathrm{~g} \theta}=2.66 \mathrm{~mol} \mathrm{O}
\end{aligned}
$$

The mole ratio of the elements in the amino acid is $\mathrm{C}_{2.66} \mathrm{H}_{6.6} \mathrm{~N}_{1.33} \mathrm{O}_{2.66}$. We can find a small whole-number ratio by dividing by the smallest number:

$$
\mathrm{C} \frac{2.66}{1.33} \mathrm{H} \frac{6.6}{1.33} \mathrm{~N} \frac{1.33}{1.33} \mathrm{O} \frac{2.66}{1.33}=\mathrm{C}_{2.00} \mathrm{H}_{5.0} \mathrm{~N}_{1.00} \mathrm{O}_{2.00}
$$

Simplifying, we find that the empirical formula for the amino acid glycine is $\mathrm{C}_{2} \mathrm{H}_{5} \mathrm{NO}_{2}$.

## Example Exercise 9.14 <br> Empirical Formula from Percent Composition

## Continued

## Practice Exercise

Calculate the empirical formula for caffeine given the following percent composition: $49.5 \% \mathrm{C}, 5.15 \% \mathrm{H}$, $28.9 \% \mathrm{~N}$, and $16.5 \%$ O.

Answer: $\mathrm{C}_{4} \mathrm{H}_{5} \mathrm{~N}_{2} \mathrm{O}$

## Concept Exercise

If 0.500 mol of yellow powder sulfur reacts with 0.750 mol of oxygen gas, what is the empirical formula of the sulfur oxide?

Answer: See Appendix G.

## Example Exercise 9.15 Molecular Formula from Empirical Formula

The empirical formula for fructose, or fruit sugar, is $\mathrm{CH}_{2} \mathrm{O}$. If the molar mass of fructose is $180 \mathrm{~g} / \mathrm{mol}$, find the actual molecular formula for the sugar.

## Solution

We can indicate the molecular formula of fructose as (CH2O)n. The molar mass of the empirical formula $\mathrm{CH}_{2} \mathrm{O}$ is $12 \mathrm{~g} \mathrm{C}+2(1 \mathrm{~g} \mathrm{H})+16 \mathrm{~g} \mathrm{O}=30 \mathrm{~g} / \mathrm{mol}$. Thus, the number of multiples of the empirical formula is

$$
\text { Fructose: } \begin{aligned}
\frac{\left(\mathrm{CH}_{2} \mathrm{O}\right)_{n}}{\mathrm{CH}_{2} \mathrm{O}} & =\frac{180 \mathrm{~g} / \mathrm{mol}}{30 \mathrm{~g} / \mathrm{mol}} \\
n & =6
\end{aligned}
$$

Thus, the molecular formula of fructose is $\left(\mathrm{CH}_{2} \mathrm{O}\right)_{6}$ or $\mathrm{C}_{6} \mathrm{H}_{12} \mathrm{O}_{6}$.

## Practice Exercise

Ethylene dibromide was used as a grain pesticide until it was banned. Calculate (a) the empirical formula and (b) the molecular formula for ethylene dibromide given its approximate molar mass of $190 \mathrm{~g} / \mathrm{mol}$ and its percent composition: $12.7 \%$ C, $2.1 \% \mathrm{H}$, and $85.1 \% \mathrm{Br}$.
Answers: (a) $\mathrm{CH}_{2} \mathrm{Br}$; (b) $\mathrm{C}_{2} \mathrm{H}_{4} \mathrm{Br}_{2}$

## Concept Exercise

If the molecular formula of hydrogen peroxide is $\mathrm{H}_{2} \mathrm{O}_{2}$, what is the empirical formula?
Answer: See Appendix G.

