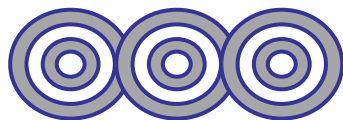




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Bruce E. Bursten; Catherine J. Murphy



# Chapter 3

## Stoichiometry

### Calculations with Chemical Formulas and Equations

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# Stoichiometry

Our focus will be on the use of chemical formulas to represent reactions and on the quantitative information we can obtain about the amounts of substances involved in reactions.

Stoichiometry is the area of study that examines the quantities of substances consumed and produced in chemical reactions.

The name derived from the Greek *stoicheion* “element” and *metron* “measure”.

# Law of Conservation of Mass

“We may lay it down as an incontestable axiom that, in all the operations of art and nature, nothing is created; an equal amount of matter exists both before and after the experiment. Upon this principle, the whole art of performing chemical experiments depends.”

-- Antoine Lavoisier, 1789.

Atoms are neither created nor destroyed during any chemical reaction.  
The changes that occur during any reaction merely rearrange the atoms.  
The same collection of atoms is present both before and after the reaction.  
-- Dalton's atomic theory.

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# **3.1**

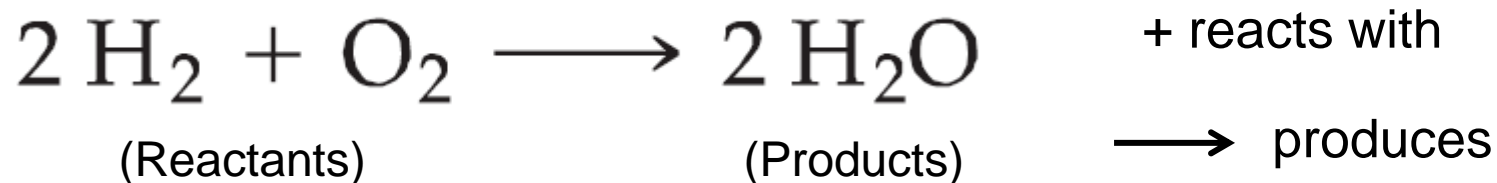
# **Chemical Equations**

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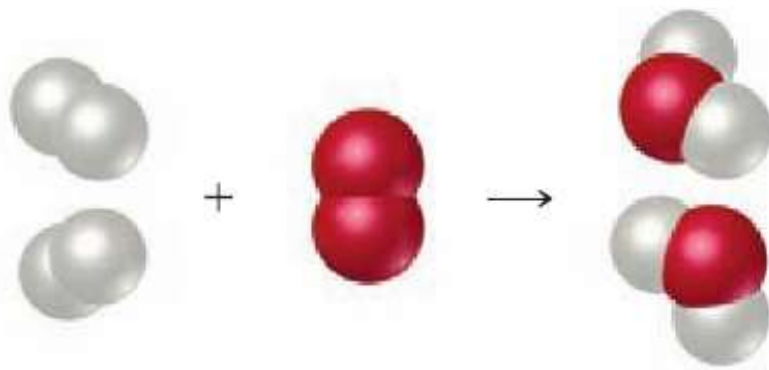
# Chemical Equations

Chemical equations are concise representations of chemical reactions.

When hydrogen gas  $\text{H}_2$  burns, it reacts with oxygen  $\text{O}_2$  in the air to form water  $\text{H}_2\text{O}$ .



The numbers in front of the formulas are coefficients (indicate the relative numbers of molecules of each kind involved in the reaction).


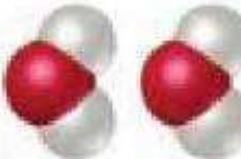



Molecular models

Balanced equation, a chemical equation have an equal number of atoms of each element on each side, because atoms are neither created nor destroyed in any reaction.

# Subscripts and Coefficients give different information

When balancing an equation, you should never change subscripts. In contrast, placing a suitable coefficient in front of a formula.

Chemical symbol	Meaning		Composition
$\text{H}_2\text{O}$	One molecule of water:		Two H atoms and one O atom
$2 \text{H}_2\text{O}$	Two molecules of water:		Four H atoms and two O atoms
$\text{H}_2\text{O}_2$	One molecule of hydrogen peroxide:		Two H atoms and two O atoms

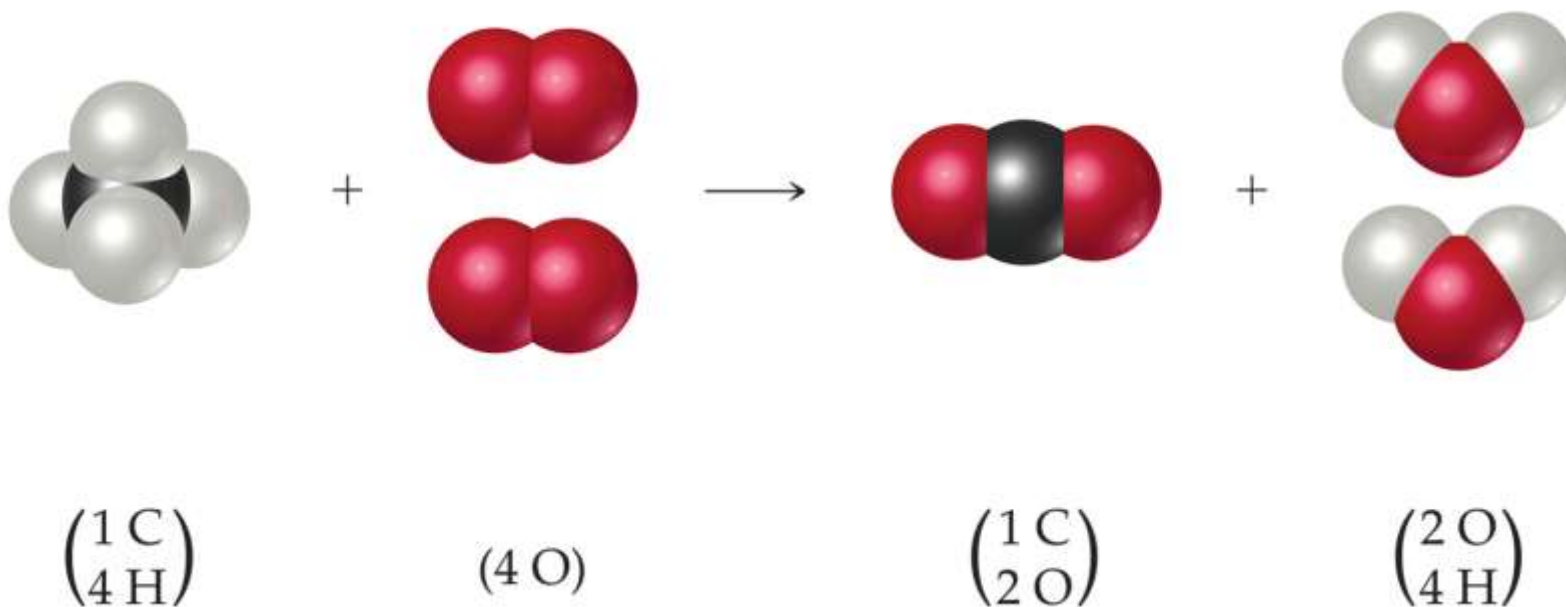
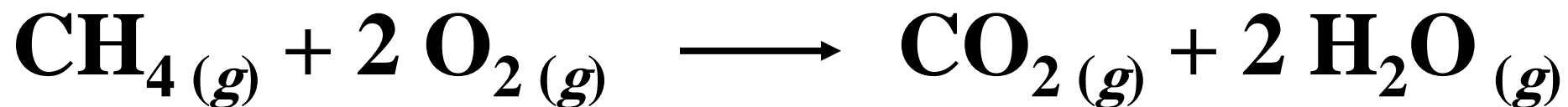
**Subscripts** tell the number of atoms of each element in a molecule.

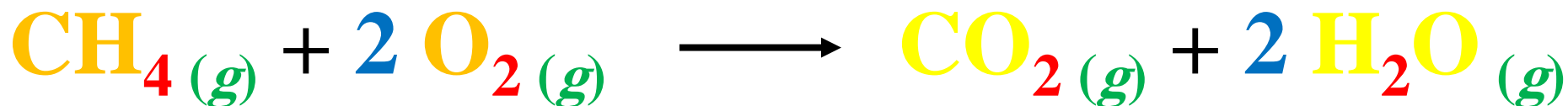
**Coefficients** tell the number of molecules.

How many atoms of Mg, O, and H are represented by  $3 \text{Mg}(\text{OH})_2$ ?

3 atoms Mg, 6 atoms O, 6 atoms H

# Anatomy of a Chemical Equation





**Reactants** appear on the left side of the equation.

**Products** appear on the right side of the equation.

The **states** of the reactants and products are written in parentheses to the right of each compound; (*g*) gas, (*l*) liquid, (*s*) solid, (*aq*) aqueous solution.

**Subscripts** present within a formula and tell the number of atoms of each element in a molecule.

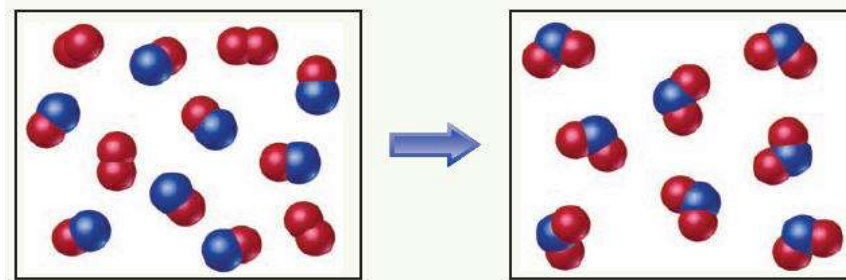
**Coefficients** are inserted in front of a formula to balance the equation.

Sometimes the conditions (such as temperature or pressure) under which the reaction proceeds appear above or below the reaction arrow.  $\Delta$  refer to temperature.



### Sample Exercise 3.1 Interpreting and Balancing Chemical Equations

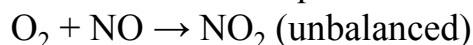
The following diagram represents a chemical reaction in which the red spheres are oxygen atoms and the blue spheres are nitrogen atoms. (a) Write the chemical formulas for the reactants and products. (b) Write a balanced equation for the reaction. (c) Is the diagram consistent with the law of conservation of mass?



#### Solution

(a) The left box, which represents the reactants, contains two kinds of molecules, those composed of two oxygen atoms ( $O_2$ ) and those composed of one nitrogen atom and one oxygen atom (NO). The right box, which represents the products, contains only molecules composed of one nitrogen atom and two oxygen atoms ( $NO_2$ ).

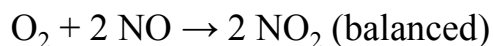
(b) The unbalanced chemical equation is



This equation has three O atoms on the left side of the arrow and two O atoms on the right side. We can increase the number of O atoms by placing a coefficient 2 on the product side:



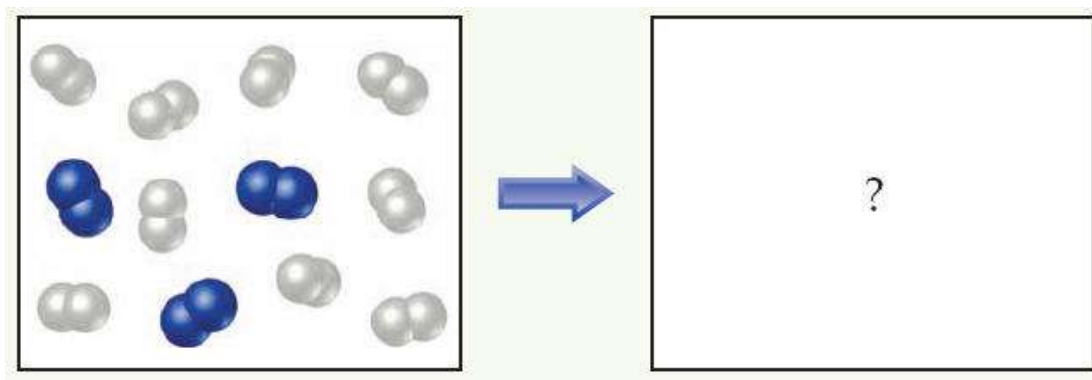
Now there are two N atoms and four O atoms on the right. Placing the coefficient 2 in front of NO balances both the number of N atoms and O atoms:



(c) The left box (reactants) contains four  $O_2$  molecules and eight NO molecules. Thus, the molecular ratio is one  $O_2$  for each two NO as required by the balanced equation. The right box (products) contains eight  $NO_2$  molecules. The number of  $NO_2$  molecules on the right equals the number of NO molecules on the left as the balanced equation requires. Counting the atoms, we find eight N atoms in the eight NO molecules in the box on the left. There are also  $4 \times 2 = 8$  O atoms in the  $O_2$  molecules and eight O atoms in the NO molecules, giving a total of 16 O atoms. In the box on the right, we find eight N atoms and  $8 \times 2 = 16$  O atoms in the eight  $NO_2$  molecules. Because there are equal numbers of both N and O atoms in the two boxes, the drawing is consistent with the law of conservation of mass.

## Practice Exercise

In the following diagram, the white spheres represent hydrogen atoms, and the blue spheres represent nitrogen atoms. To be consistent with the law of conservation of mass, how many  $\text{NH}_3$  molecules should be shown in the right box?



**Answer:** Six  $\text{NH}_3$  molecules.

### **SAMPLE EXERCISE 3.2** | Balancing Chemical Equations

Balance this equation:

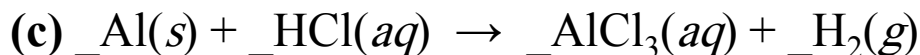
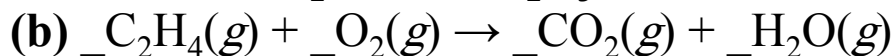
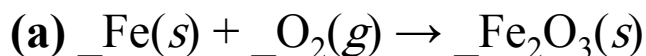


#### **SOLUTION**



### **Practice Exercise**

Balance the following equations by providing the missing coefficients:



**Answers:** (a) 4, 3, 2; (b) 1, 3, 2, 2; (c) 2, 6, 2, 3

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## **3.2**

# **Some Simple Patterns of Chemical Reactivity**

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# Reaction Types

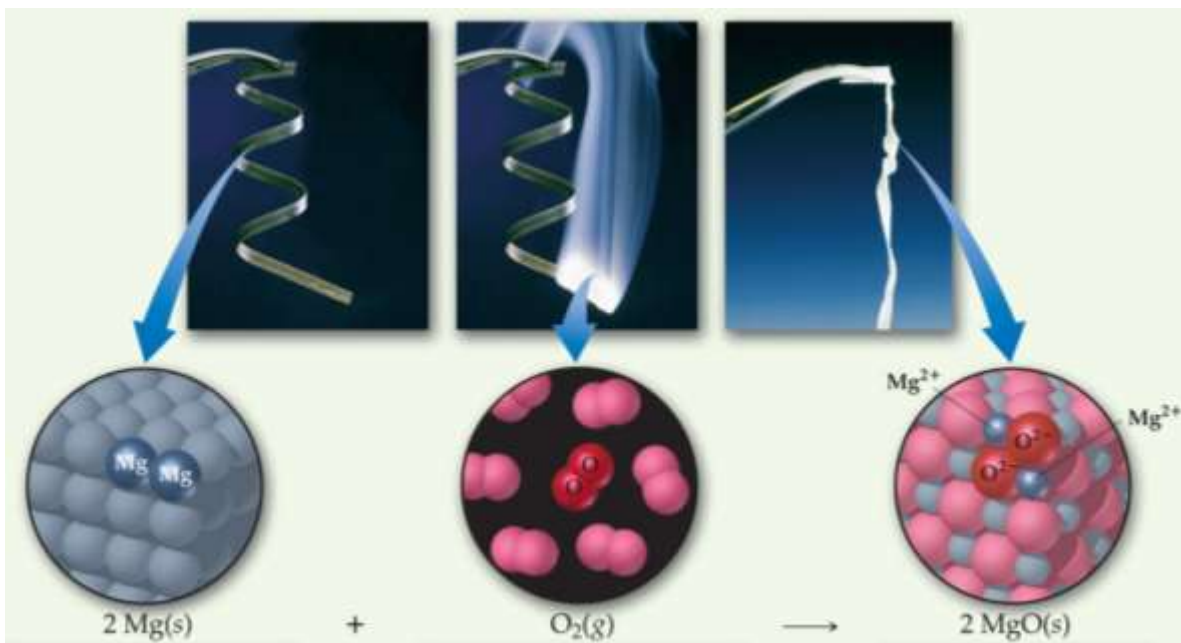
(some simple patterns of chemical reactivity)

- **C**ombination Reactions
- **D**ecomposition Reactions
- **C**ombustion Reactions

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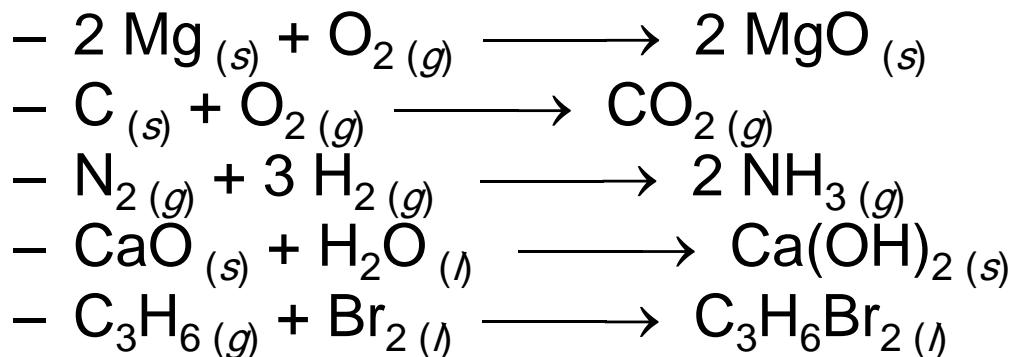
- Substitution reactions
- Addition reactions
- Elimination reactions
- Oxidation-reduction reactions .... etc.

# Combination Reactions

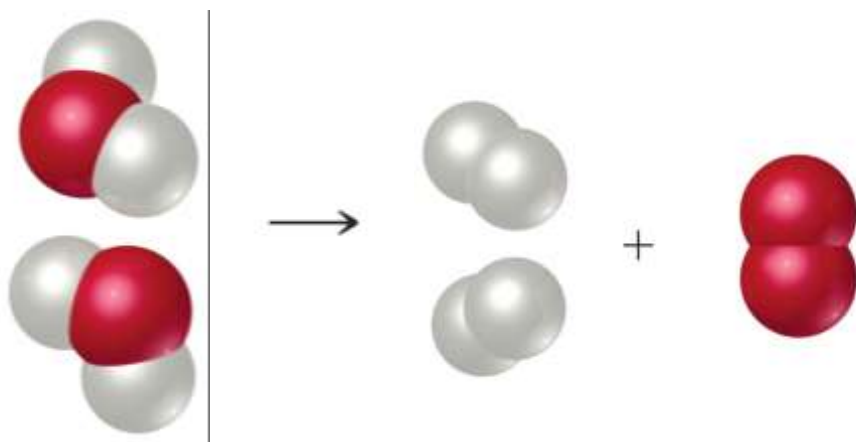


- In this type of reaction two or more substances react to form one product.
- A combination reaction between a metal and a nonmetal produce ionic solid.

- **Examples:**



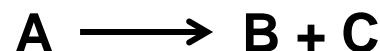
# Decomposition Reactions



- In this type one substance breaks down into two or more substances.

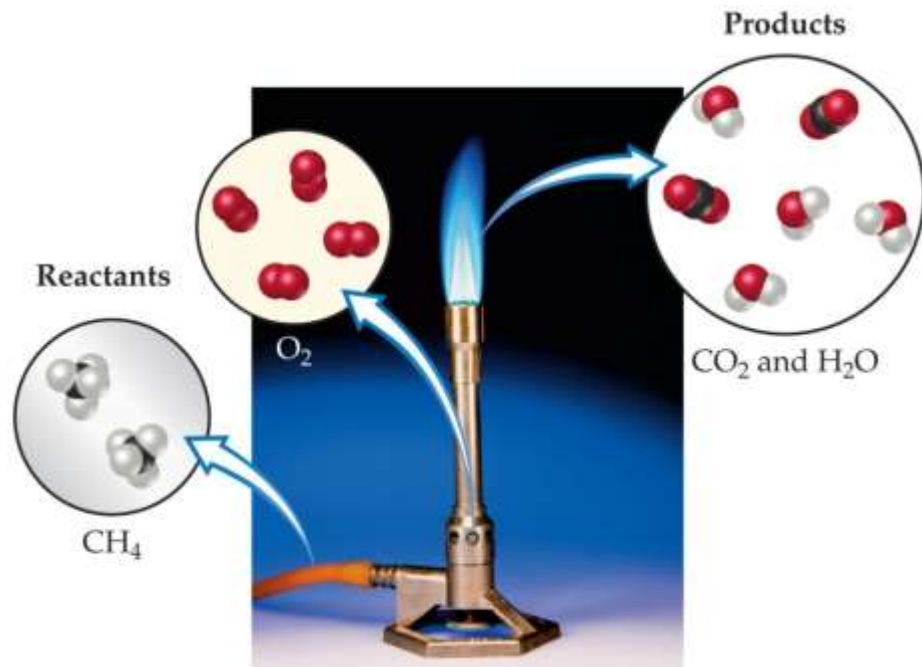
- Many compounds undergo decomposition reactions when heated.

- **Examples:**



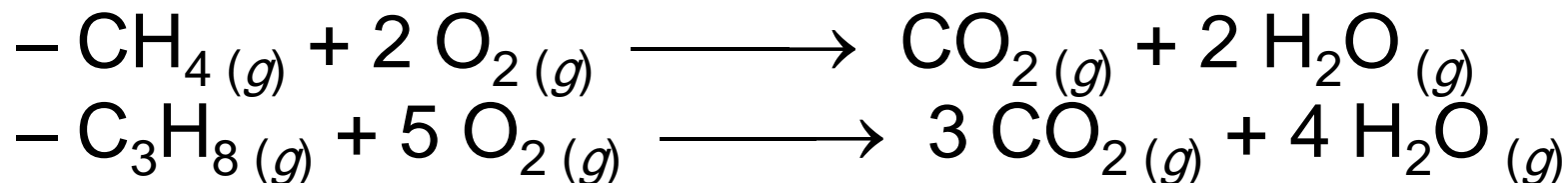
- $\text{CaCO}_3 (s) \longrightarrow \text{CaO} (s) + \text{CO}_2 (g)$
- $2 \text{KClO}_3 (s) \longrightarrow 2 \text{KCl} (s) + 3 \text{O}_2 (g)$
- $\text{Cu}(\text{OH})_2 (s) \longrightarrow \text{CuO} (s) + \text{H}_2\text{O} (l)$
- $\text{PbCO}_3 (s) \longrightarrow \text{PbO} (s) + \text{CO}_2 (g)$
- $2 \text{NaN}_3 (s) \longrightarrow 2 \text{Na} (s) + 3 \text{N}_2 (g)$

# Combustion Reactions



- These are generally rapid reactions that produce a flame.
- Most often involve hydrocarbons reacting with oxygen in the air.
- Hydrocarbon compounds contain only Carbons (**C**) and Hydrogen (**H**)).

- Examples:





When Na and S undergo a combination reaction, what is the chemical formula of the product?

$\text{Na}_2\text{S}$

## PERIODIC TABLE OF THE ELEMENTS

Table of Selected Radioactive Isotopes

Table of Selected Radioactive Isotopes																	
1 1.00794 H Hydrogen	2 4.00260 He Helium	3 6.941 Li Lithium	4 9.01218 Be Beryllium	5 10.0129 B Boron	6 12.011 C Carbon	7 14.0031 N Nitrogen	8 15.9994 O Oxygen	9 18.9984 F Fluorine	10 20.1797 Ne Neon	11 22.98977 Na Sodium	12 24.304 Mg Magnesium	13 26.98154 Al Aluminum	14 28.0855 Si Silicon	15 30.97376 P Phosphorus	16 32.06 S Sulfur	17 35.453 Cl Chlorine	18 39.9624 Ar Argon
19 39.0983 K Potassium	20 40.078 Ca Calcium	21 44.9559 Sc Scandium	22 47.87 Ti Titanium	23 50.9415 V Vanadium	24 51.996 Cr Chromium	25 54.9380 Mn Manganese	26 55.935 Fe Iron	27 58.9332 Co Cobalt	28 58.9332 Ni Nickel	29 63.546 Cu Copper	30 65.38 Zn Zinc	31 69.723 Ga Gallium	32 72.61 Ge Germanium	33 74.9216 As Arsenic	34 75.92 Se Selenium	35 78.96 Br Bromine	36 83.80 Kr Krypton
37 85.4678 Rb Rubidium	38 87.62 Sr Strontium	39 88.9059 Y Yttrium	40 91.224 Zr Zirconium	41 90.9064 Nb Niobium	42 90.9064 Mo Molybdenum	43 92.90638 Tc Technetium	44 101.07 Ru Ruthenium	45 101.07 Rh Rhodium	46 101.07 Pd Palladium	47 106.905 Ag Silver	48 112.41 Cd Cadmium	49 114.904 In Indium	50 115.73 Sn Tin	51 121.757 Sb Antimony	52 127.60 Te Tellurium	53 126.905 I Iodine	54 131.29 Xe Xenon
55 132.9054 Cs Cesium	56 137.33 Ba Barium	57 138.9055 La Lanthanum	72 175.44 Hf Hafnium	73 180.9479 Ta Tantalum	74 183.84 W Tungsten	75 186.207 Re Rhenium	76 186.207 Os Osmium	77 190.23 Ir Iridium	78 195.08 Pt Platinum	79 196.967 Au Gold	80 200.59 Hg Mercury	81 204.38 Tl Thallium	82 207.2 Pb Lead	83 208.9804 Bi Bismuth	84 209 Po Polonium	85 210 At Astatine	86 222 Rn Radon
87 87 Fr Francium	88 88 Ra Radium	89 89 Ac Actinium	104 104 Rf Rutherfordium	105 105 Db Dubnium	106 106 Sg Seaborgium	107 107 Bh Bohrium	108 108 Hs Hassium	109 109 Mt Meitnerium	110 110 Uun Ununium	111 111 Uuh Ununium	112 112 Uub Ununium	113 113 Uut Ununium	114 114 Uuq Ununium	115 115 Uuq Ununium	116 116 Uuq Ununium	117 117 Uuq Ununium	118 118 Uuq Ununium

\* Estimated Values



58 140.12 Ce Cerium	59 140.9077 Pr Praseodymium	60 140.91 Nd Neodymium	61 (141) Pm Promethium	62 150.36 Sm Samarium	63 151.964 Eu Europium	64 157.25 Gd Gadolinium	65 158.925 Tb Terbium	66 158.93 Dy Dysprosium	67 162.5023 Ho Holmium	68 167.26 Er Erbium	69 168.9342 Tm Thulium	70 173.04 Yb Ytterbium	71 174.967 Lu Lutetium
90 223.021 Th Thorium	91 223.021 Pa Protactinium	92 226.025 U Uranium	93 237.0437 Np Neptunium	94 238.0289 Pu Plutonium	95 244.0642 Am Americium	96 247.0771 Cm Curium	97 247.0771 Bk Berkelium	98 247.0771 Cf Californium	99 247.0771 Es Einsteinium	100 247.0771 Fm Fermium	101 257.10 Md Mendelevium	102 258.10 No Nobelium	103 259.10 Lr Lawrencium

NOTES:  
(1) Black — solid.  
(2) Based upon carbon-12, (1) indicates most stable or least known isotope.  
(3) Shaded marked with diagonal refer to the ground state of 273 K and 1 atm and are given in units of g/L.

The A & B subgroup designations are those recommended by the International Union of Pure and Applied Chemistry.

Catalog Number WLS-18806

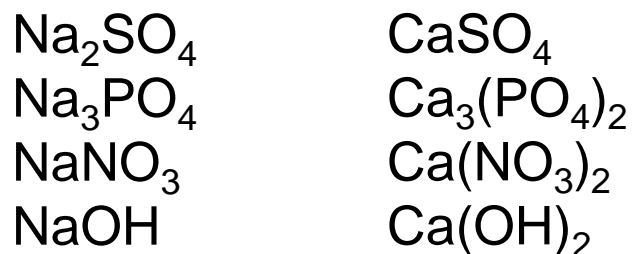
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## SELECTED POLYATOMIC IONS

$\text{Hg}_2^{2+}$	dimercury (I)	$\text{CrO}_4^{2-}$	chromate
$\text{NH}_4^+$	ammonium	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
$\text{C}_2\text{H}_3\text{O}_2^-$	} acetate	$\text{MnO}_4^-$	permanganate
$\text{CH}_3\text{COO}^-$		$\text{MnO}_4^{2-}$	manganate
$\text{CN}^-$	cyanide	$\text{NO}_2^-$	nitrite
$\text{CO}_3^{2-}$	carbonate	$\text{NO}_3^-$	nitrate
$\text{HCO}_3^-$	hydrogen carbonate	$\text{OH}^-$	hydroxide
$\text{C}_2\text{O}_4^{2-}$	oxalate	$\text{PO}_4^{3-}$	phosphate
$\text{ClO}^-$	hypochlorite	$\text{SCN}^-$	thiocyanate
$\text{ClO}_2^-$	chlorite	$\text{SO}_3^{2-}$	sulfite
$\text{ClO}_3^-$	chlorate	$\text{SO}_4^{2-}$	sulfate
$\text{ClO}_4^-$	perchlorate	$\text{HSO}_4^-$	hydrogen sulfate
		$\text{S}_2\text{O}_3^{2-}$	thiosulfate

### Examples:



### **SAMPLE EXERCISE 3.3** | Writing Balanced Equations for Combination and Decomposition Reactions

Write balanced equations for the following reactions: (a) The combination reaction that occurs when lithium metal and fluorine gas react. (b) The decomposition reaction that occurs when solid barium carbonate is heated. (Two products form: a solid and a gas.)



### **PRACTICE EXERCISE**

Write balanced chemical equations for the following reactions: (a) Solid mercury(II) sulfide decomposes into its component elements when heated. (b) The surface of aluminum metal undergoes a combination reaction with oxygen in the air.

*Answers:* (a)  $\text{HgS}(s) \longrightarrow \text{Hg}(l) + \text{S}(s)$ ; (b)  $4 \text{Al}(s) + 3 \text{O}_2(g) \longrightarrow 2 \text{Al}_2\text{O}_3(s)$

### **SAMPLE EXERCISE 3.4** | Writing Balanced Equations for Combustion Reactions

Write the balanced equation for the reaction that occurs when methanol,  $\text{CH}_3\text{OH}(l)$ , is burned in air.



### **PRACTICE EXERCISE**

Write the balanced equation for the reaction that occurs when ethanol,  $\text{C}_2\text{H}_5\text{OH}(l)$ , is burned in air.

*Answer:*  $\text{C}_2\text{H}_5\text{OH}(l) + 3 \text{O}_2(g) \longrightarrow 2 \text{CO}_2(g) + 3 \text{H}_2\text{O}(g)$

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# **3.3**

## **Formula Weights**

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# Formula Weights (FW)

A formula weight is the sum of the atomic weights for the atoms in a chemical formula.

Formula weights are generally reported for ionic compounds.

For example, the formula weight of calcium chloride,  $\text{CaCl}_2$ , would be

$$\begin{array}{r} \text{Ca: } 1(40.1 \text{ amu}) \\ + \text{Cl: } 2(35.5 \text{ amu}) \\ \hline 111.1 \text{ amu} \end{array}$$

$$\text{FW of NaCl} = 23.0 \text{ amu} + 35.5 \text{ amu} = 58.5 \text{ amu}$$

$$\begin{aligned} \text{FW of H}_2\text{SO}_4 &= 2(\text{AW of H}) + (\text{AW of S}) + 4(\text{AW of O}) \\ &= 2(1.0 \text{ amu}) + 32.1 \text{ amu} + 4(16.0 \text{ amu}) \\ &= 98.1 \text{ amu} \end{aligned}$$

If the chemical formula is merely the chemical symbol of an element, such as Na, then the formula weight equals the atomic weight of the element.

# Molecular Weight (MW)

A molecular weight is the sum of the atomic weights of the atoms in a molecule.

For example, the molecular weight of the ethane molecule,  $C_2H_6$ , would be

$$\begin{array}{r} \text{C: } 2(12.0 \text{ amu}) \\ + \text{H: } 6(1.0 \text{ amu}) \\ \hline 30.0 \text{ amu} \end{array}$$

$$\text{MW of } C_6H_{12}O_6 = 6(12.0 \text{ amu}) + 12(1.0 \text{ amu}) + 6(16.0 \text{ amu}) = 180.0 \text{ amu}$$

If the chemical formula is that of a molecule, then the formula weight is also called the molecular weight..

*\*The abbreviation AW is used for atomic weight, FW for formula weight, and MW for molecular weight.*



### Sample Exercise 3.5 Calculating formula Weights

Calculate the formula weight of **(a)** sucrose,  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$  (table sugar), and **(b)** calcium nitrate,  $\text{Ca}(\text{NO}_3)_2$ .

#### Solution

**(a)** By adding the atomic weights of the atoms in sucrose, we find the formula weight to be 342.0 amu:

$$\begin{array}{rcl} 12 \text{ C atoms} & = & 12(12.0 \text{ amu}) = 144.0 \text{ amu} \\ 22 \text{ H atoms} & = & 22(1.0 \text{ amu}) = 22.0 \text{ amu} \\ 11 \text{ O atoms} & = & 11(16.0 \text{ amu}) = \underline{176.0 \text{ amu}} \\ & & 342.0 \text{ amu} \end{array}$$

**(b)** If a chemical formula has parentheses, the subscript outside the parentheses is a multiplier for all atoms inside. Thus, for  $\text{Ca}(\text{NO}_3)_2$ , we have

$$\begin{array}{rcl} 1 \text{ Ca atom} & = & 1(40.1 \text{ amu}) = 40.1 \text{ amu} \\ 2 \text{ N atoms} & = & 2(14.0 \text{ amu}) = 28.0 \text{ amu} \\ 6 \text{ O atoms} & = & 6(16.0 \text{ amu}) = \underline{96.0 \text{ amu}} \\ & & 164.1 \text{ amu} \end{array}$$

### Practice Exercise

Calculate the formula weight of **(a)**  $\text{Al}(\text{OH})_3$  and **(b)**  $\text{CH}_3\text{OH}$ .

**Answer:** **(a)** 78.0 amu, **(b)** 32.0 amu



# Percent Composition

One can find the percentage composition of the mass of a compound that comes from each of the elements in the compound by using this equation:

$$\% \text{ element} = \frac{\left( \begin{array}{c} \text{number of atoms} \\ \text{of that element} \end{array} \right) \left( \begin{array}{c} \text{atomic weight} \\ \text{of element} \end{array} \right)}{\text{formula weight of compound}} \times 100\%$$

So the percentage of carbon (C) in ethane  $\text{C}_2\text{H}_6$  is...

$$\begin{aligned} \% \text{C} &= \frac{(2)(12.0 \text{ amu})}{(30.0 \text{ amu})} \\ &= \frac{24.0 \text{ amu}}{30.0 \text{ amu}} \times 100 \\ &= 80.0\% \end{aligned}$$

## Sample Exercise 3.6 Calculating Percentage Composition

Calculate the percentage of carbon, hydrogen, and oxygen (by mass) in  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ .

### Solution

We can use Equation 3.10, relying on a periodic table to obtain the atomic weight of each component element. The atomic weights are first used to determine the formula weight of the compound. (The formula weight of  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ , 342.0 amu, was calculated in Sample Exercise 3.5.) We must then do three calculations, one for each element.

Using Equation 3.10, we have

$$\% \text{C} = \frac{(12)(12.0 \text{ amu})}{342.0 \text{ amu}} \times 100\% = 42.1\%$$

$$\% \text{H} = \frac{(22)(1.0 \text{ amu})}{342.0 \text{ amu}} \times 100\% = 6.4\%$$

$$\% \text{O} = \frac{(11)(16.0 \text{ amu})}{342.0 \text{ amu}} \times 100\% = 51.5\%$$

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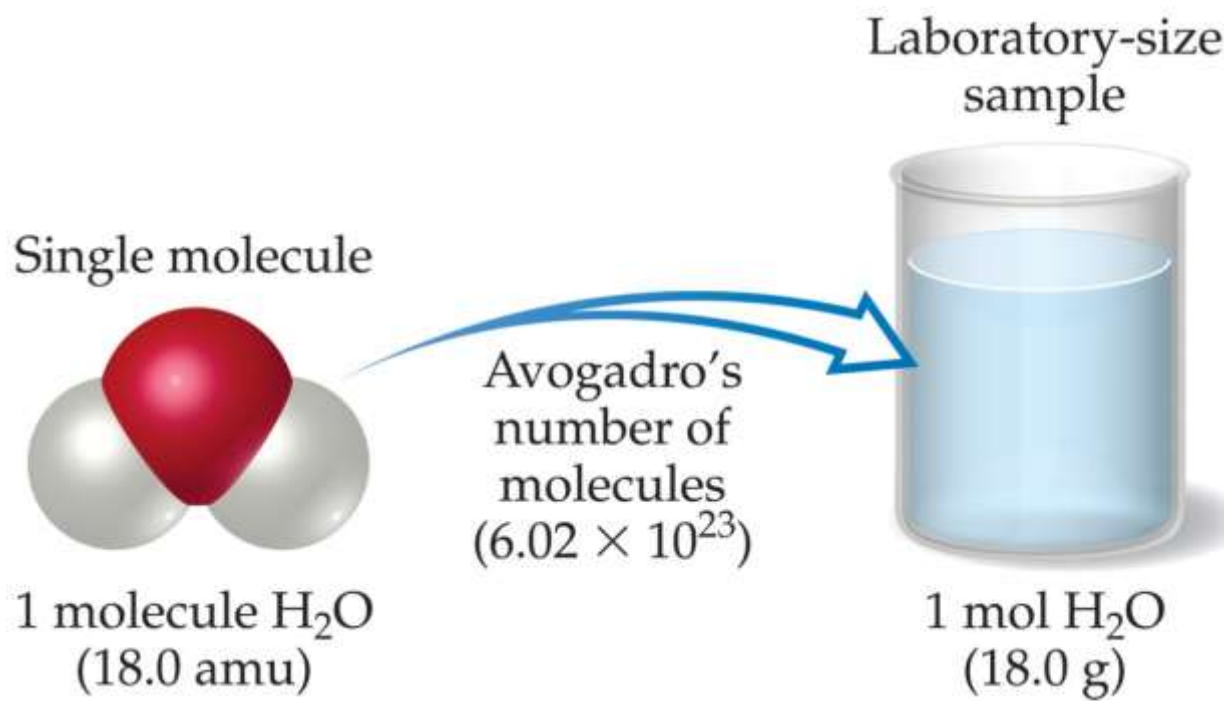
## **3.4**

# **Avogadro's Number and the Mole**

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# Avogadro's Number and the Mole

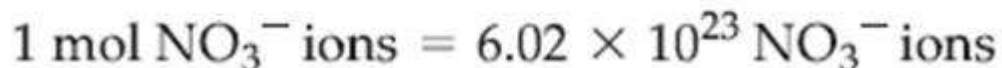
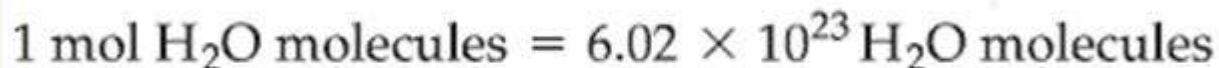
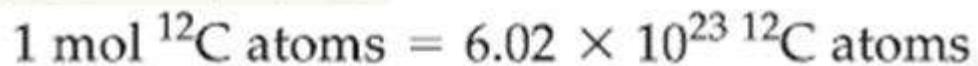
Even the smallest samples that we deal with in the laboratory contain enormous numbers of atoms, ions, or molecules. For example, a teaspoon of water (about 5 mL) contains  $2 \times 10^{23}$  water molecules.



A mole is the amount of matter that contains as many objects (atoms, molecules, or whatever objects we are considering) as the number of atoms in exactly 12 g of isotopically pure  $^{12}\text{C}$ .

From experiments, scientists have determined this number to be  $6.0221421 \times 10^{23}$ . Scientists call this number Avogadro's number, and has the symbol  $N_A$ , and round to  $6.02 \times 10^{23} \text{ mol}^{-1}$ .

A mole of atoms, a mole of molecules, or a mole of anything else all contain Avogadro's number of these objects:



## Sample Exercise 3.7 Estimating Numbers in Atoms

Without using a calculator, arrange the following samples in order of increasing numbers of carbon atoms: 12 g  $^{12}\text{C}$ , 1 mol  $\text{C}_2\text{H}_2$ ,  $9 \times 10^{23}$  molecules of  $\text{CO}_2$ .

### Solution

To determine the number of C atoms in each sample, we must convert g  $^{12}\text{C}$ , 1 mol  $\text{C}_2\text{H}_2$ , and  $9 \times 10^{23}$  molecules  $\text{CO}_2$  all to numbers of C atoms. To make these conversions, we use the definition of mole and Avogadro's number.

A mole is defined as the amount of matter that contains as many units of the matter as there are C atoms in exactly 12 g of  $^{12}\text{C}$ . Thus, 12 g of  $^{12}\text{C}$  contains 1 mol of C atoms (that is,  $6.02 \times 10^{23}$  C atoms). One mol of  $\text{C}_2\text{H}_2$  contains  $6 \times 10^{23}$   $\text{C}_2\text{H}_2$  molecules. Because there are two C atoms in each  $\text{C}_2\text{H}_2$  molecule, this sample contains  $12 \times 10^{23}$  C atoms. Because each  $\text{CO}_2$  molecule contains one C atom, the sample of  $\text{CO}_2$  contains  $9 \times 10^{23}$  C atoms. Hence, the order is 12 g  $^{12}\text{C}$  ( $6 \times 10^{23}$  C atoms)  $<$   $9 \times 10^{23}$   $\text{CO}_2$  molecules ( $9 \times 10^{23}$  C atoms)  $<$  1 mol  $\text{C}_2\text{H}_2$  ( $12 \times 10^{23}$  C atoms).

### Practice Exercise

Without using a calculator, arrange the following samples in order of increasing number of O atoms: 1 mol  $\text{H}_2\text{O}$ , 1 mol  $\text{CO}_2$ ,  $3 \times 10^{23}$  molecules  $\text{O}_3$ .

**Answer:** 1 mol  $\text{H}_2\text{O}$  ( $6 \times 10^{23}$  O atoms)  $3 \times 10^{23}$  molecules  $\text{O}_3$  ( $9 \times 10^{23}$  O atoms) 1 mol  $\text{CO}_2$  ( $12 \times 10^{23}$  O atoms)

### Sample Exercise 3.8 Converting Moles to Atoms

Calculate the number of H atoms in 0.350 mol of  $\text{C}_6\text{H}_{12}\text{O}_6$ .

#### Solution

Avogadro's number provides the conversion factor between the number of moles of  $\text{C}_6\text{H}_{12}\text{O}_6$  and the number of molecules of  $\text{C}_6\text{H}_{12}\text{O}_6$ . Once we know the number of molecules of  $\text{C}_6\text{H}_{12}\text{O}_6$ , we can use the chemical formula, which tells us that each molecule of  $\text{C}_6\text{H}_{12}\text{O}_6$  contains 12 H atoms. Thus, we convert moles of  $\text{C}_6\text{H}_{12}\text{O}_6$  to molecules of  $\text{C}_6\text{H}_{12}\text{O}_6$  and then determine the number of atoms of H from the number of molecules of  $\text{C}_6\text{H}_{12}\text{O}_6$ :

Moles  $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow$  molecules  $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow$  atoms H

#### Solve

$$\begin{aligned}\text{H atoms} &= (0.350 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6) \left( \frac{6.02 \times 10^{23} \text{ molecules } \text{C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6} \right) \left( \frac{12 \text{ H atoms}}{1 \text{ molecule } \text{C}_6\text{H}_{12}\text{O}_6} \right) \\ &= 2.53 \times 10^{24} \text{ H atoms}\end{aligned}$$

### Practice Exercise

How many oxygen atoms are in (a) 0.25 mol  $\text{Ca}(\text{NO}_3)_2$  and (b) 1.50 mol of sodium carbonate?

**Answer:** (a)  $9.0 \times 10^{23}$ , (b)  $2.71 \times 10^{24}$

# Molar Mass

By definition, a molar mass is the mass of 1 mol of a substance (i.e., g/mol).

- The molar mass of an element is the mass number for the element that we find on the periodic table.
- The formula weight (in amu's) will be the same number as the molar mass (in g/mol).

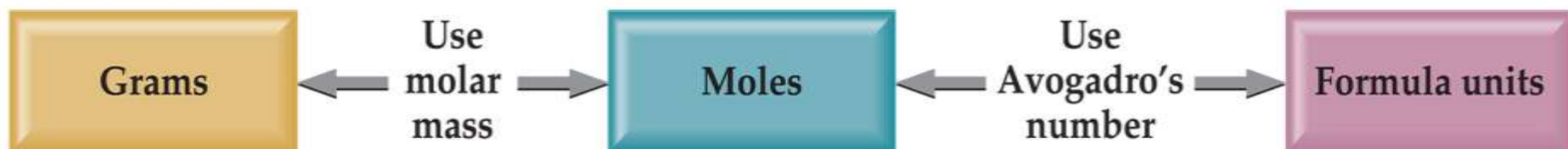
$$\text{Avogadro's No (N}_A\text{)} = \frac{\text{No. of atoms or molecules or ions or particles}}{\text{No. of mole}}$$

$(6.022 \times 10^{23})$

$$\text{No. of moles} = \frac{\text{Mass (g)}}{\text{MW (g/mol)}}$$



# Using Moles



Moles provide a bridge between mass and the number of particles (from the molecular scale to the real-world scale).

# Mole Relationships

Name of Substance	Formula	Formula Weight (amu)	Molar Mass (g/mol)	Number and Kind of Particles in One Mole
Atomic nitrogen	N	14.0	14.0	$6.02 \times 10^{23}$ N atoms
Molecular nitrogen	N <sub>2</sub>	28.0	28.0	$\left\{ \begin{array}{l} 6.02 \times 10^{23} \text{ N}_2 \text{ molecules} \\ 2(6.02 \times 10^{23}) \text{ N atoms} \end{array} \right.$
Silver	Ag	107.9	107.9	$6.02 \times 10^{23}$ Ag atoms
Silver ions	Ag <sup>+</sup>	107.9 <sup>a</sup>	107.9	$6.02 \times 10^{23}$ Ag <sup>+</sup> ions
Barium chloride	BaCl <sub>2</sub>	208.2	208.2	$\left\{ \begin{array}{l} 6.02 \times 10^{23} \text{ BaCl}_2 \text{ units} \\ 6.02 \times 10^{23} \text{ Ba}^{2+} \text{ ions} \\ 2(6.02 \times 10^{23}) \text{ Cl}^- \text{ ions} \end{array} \right.$

<sup>a</sup>Recall that the electron has negligible mass; thus, ions and atoms have essentially the same mass.

- One mole of atoms, ions, or molecules contains Avogadro's number of those particles.
- One mole of molecules or formula units contains Avogadro's number times the number of atoms or ions of each element in the compound.

calculate the number of copper atoms in an old copper penny. Such a penny weighs about 3 g, and we will assume that it is 100% copper:

$$\begin{aligned}\text{Cu atoms} &= (3 \text{ g Cu}) \left( \frac{1 \text{ mol Cu}}{63.5 \text{ g Cu}} \right) \left( \frac{6.02 \times 10^{23} \text{ Cu atoms}}{1 \text{ mol Cu}} \right) \\ &= 3 \times 10^{22} \text{ Cu atoms}\end{aligned}$$

### Sample Exercise 3.9 Calculating Molar Mass

What is the mass in grams of 1.000 mol of glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$ ?

#### Solution

The molar mass of a substance is found by adding the atomic weights of its component atoms.

6 C atoms	=	6(12.0 amu)	=	72.0 amu
12 H atoms	=	12(1.0 amu)	=	12.0 amu
6 O atoms	=	6(16.0 amu)	=	96.0 amu
				<hr/> 180.0 amu

Because glucose has a formula weight of 180.0 amu, one mole of this substance has a mass of 180.0 g. In other words,  $\text{C}_6\text{H}_{12}\text{O}_6$  has a molar mass of 180.0 g/mol.

### Practice Exercise

Calculate the molar mass of  $\text{Ca}(\text{NO}_3)_2$ .

**Answer:** 164.1 g/mol

### Sample Exercise 3.10 Converting Grams to Moles

Calculate the number of moles of glucose ( $\text{C}_6\text{H}_{12}\text{O}_6$ ) in 5.380 g of  $\text{C}_6\text{H}_{12}\text{O}_6$ .

#### Solution

The molar mass of a substance provides the factor for converting grams to moles. The molar mass of  $\text{C}_6\text{H}_{12}\text{O}_6$  is 180.0 g/mol.

Using 1 mol  $\text{C}_6\text{H}_{12}\text{O}_6 = 180.0$  g  $\text{C}_6\text{H}_{12}\text{O}_6$  to write the appropriate conversion factor, we have

$$\text{Moles } \text{C}_6\text{H}_{12}\text{O}_6 = (5.380 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6}) \left( \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.0 \text{ g } \cancel{\text{C}_6\text{H}_{12}\text{O}_6}} \right) = 0.02989 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6$$

#### Practice Exercise

How many moles of sodium bicarbonate ( $\text{NaHCO}_3$ ) are in 508 g of  $\text{NaHCO}_3$ ?

**Answer:** 6.05 mol  $\text{NaHCO}_3$

### Sample Exercise 3.11 Converting Moles to Grams

Calculate the mass, in grams, of 0.433 mol of calcium nitrate.

#### Solution

To convert moles to grams, we need the molar mass, which we can calculate using the chemical formula and atomic weights.

Because the calcium ion is  $\text{Ca}^{2+}$  and the nitrate ion is  $\text{NO}_3^-$ , calcium nitrate is  $\text{Ca}(\text{NO}_3)_2$ . Adding the atomic weights of the elements in the compound gives a formula weight of 164.1 amu. Using 1 mol  $\text{Ca}(\text{NO}_3)_2 = 164.1 \text{ g Ca}(\text{NO}_3)_2$  to write the appropriate conversion factor, we have

$$\text{Grams Ca}(\text{NO}_3)_2 = (0.433 \text{ mol Ca}(\text{NO}_3)_2) \left( \frac{164.1 \text{ g Ca}(\text{NO}_3)_2}{1 \text{ mol Ca}(\text{NO}_3)_2} \right) = 71.1 \text{ g Ca}(\text{NO}_3)_2$$

#### Practice Exercise

What is the mass, in grams, of **(a)** 6.33 mol of  $\text{NaHCO}_3$  and **(b)**  $3.0 \times 10^{-5}$  mol of sulfuric acid?

**Answer:** **(a)** 532 g, **(b)**  $2.9 \times 10^{-3}$  g

### Sample Exercise 3.12 Calculating the Number of Molecules and Number of Atoms from Mass

**(a)** How many glucose molecules are in 5.23 g of  $\text{C}_6\text{H}_{12}\text{O}_6$ ? **(b)** How many oxygen atoms are in this sample?

#### Solution

**(a)** The strategy for determining the number of molecules in a given quantity of a substance is summarized in Figure 3.10. We must convert 5.23 g  $\text{C}_6\text{H}_{12}\text{O}_6$  to moles  $\text{C}_6\text{H}_{12}\text{O}_6$ , which can then be converted to molecules  $\text{C}_6\text{H}_{12}\text{O}_6$ . The first conversion uses the molar mass of  $\text{C}_6\text{H}_{12}\text{O}_6$ :

1 mol  $\text{C}_6\text{H}_{12}\text{O}_6$  = 180.0 g  $\text{C}_6\text{H}_{12}\text{O}_6$ . The second conversion uses Avogadro's number.

Molecules  $\text{C}_6\text{H}_{12}\text{O}_6$

$$\begin{aligned} &= (5.23 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.0 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6} \right) \left( \frac{6.02 \times 10^{23} \text{ molecules } \text{C}_6\text{H}_{12}\text{O}_6}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6} \right) \\ &= 1.75 \times 10^{22} \text{ molecules } \text{C}_6\text{H}_{12}\text{O}_6 \end{aligned}$$

**(b)** To determine the number of O atoms, we use the fact that there are six O atoms in each molecule of  $\text{C}_6\text{H}_{12}\text{O}_6$ . Thus, multiplying the number of molecules  $\text{C}_6\text{H}_{12}\text{O}_6$  by the factor (6 atoms O/1 molecule  $\text{C}_6\text{H}_{12}\text{O}_6$ ) gives the number of O atoms.

$$\begin{aligned} \text{Atoms O} &= (1.75 \times 10^{22} \text{ molecules } \text{C}_6\text{H}_{12}\text{O}_6) \left( \frac{6 \text{ atoms O}}{1 \text{ molecule } \text{C}_6\text{H}_{12}\text{O}_6} \right) \\ &= 1.05 \times 10^{23} \text{ atoms O} \end{aligned}$$

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# **3.5**

## **Empirical Formulas from Analyses**

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# Empirical Formulas from Analyses

From Section 2.6;

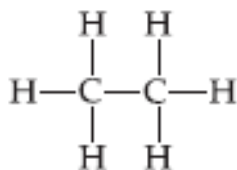
**Molecular formula:** chemical formula that indicate the actual numbers and types of atoms in a molecule.

**Empirical formula:** chemical formula that give only the relative number of atoms of each type in a molecule.

The subscripts in an empirical formula are always the smallest possible whole-number ratios.

For example; the molecular formula of hydrogen peroxide is  $\text{H}_2\text{O}_2$ , where as its empirical formula is  $\text{HO}$ . The molecular formula of ethylene is  $\text{C}_2\text{H}_4$ , where as its empirical formula is  $\text{CH}_2$ . For water,  $\text{H}_2\text{O}$  the molecular and the empirical formulas are identical.

The structural formula for the substance ethane is shown here:



(a)  $\text{C}_2\text{H}_6$ , (b)  $\text{CH}_3$

(a) What is the molecular formula for ethane? (b) What is its empirical formula?

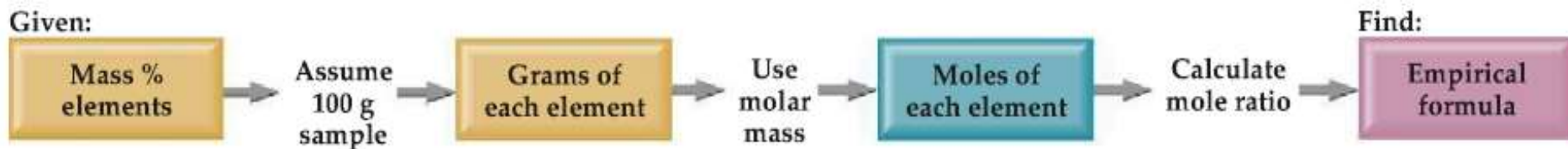


The empirical formula for a substance tell us the relative number of atoms of each element it contains.

The empirical formula of  $\text{H}_2\text{O}$  shows that water contains two H atoms for each O atom. This ratio also applies on the molar level: 1 mole of  $\text{H}_2\text{O}$  contains 2 mole of H atoms and 1 mole of O atoms.

Conversely, the ratio of the number of moles of each element in a compound gives the subscripts in a compounds empirical formula. Therefore, the mole concept provides a way of calculating the empirical formulas of chemical substances.

**One can calculate the empirical formula from the percent composition.**



### **SAMPLE EXERCISE 3.13** | Calculating an Empirical Formula

Ascorbic acid (vitamin C) contains 40.92% C, 4.58% H, and 54.50% O by mass.

What is the empirical formula of ascorbic acid?

**Solve** We *first* assume, for simplicity, that we have exactly 100 g of material (although any mass can be used). In 100 g of ascorbic acid, therefore, we have

*Second*, we calculate the number of moles of each element:

*Third*, we determine the simplest whole-number ratio of moles by dividing each number of moles by the smallest number of moles, 3.406:

The ratio for H is too far from 1 to attribute the difference to experimental error; in fact, it is quite close to  $1\frac{1}{3}$ . This suggests that if we multiply the ratio by 3, we will obtain whole numbers:

The whole-number mole ratio gives us the subscripts for the empirical formula:

40.92 g C, 4.58 g H, and 54.50 g O.

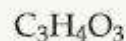
$$\text{Moles C} = (40.92 \text{ g } \cancel{\text{C}}) \left( \frac{1 \text{ mol C}}{12.01 \text{ g } \cancel{\text{C}}} \right) = 3.407 \text{ mol C}$$

$$\text{Moles H} = (4.58 \text{ g } \cancel{\text{H}}) \left( \frac{1 \text{ mol H}}{1.008 \text{ g } \cancel{\text{H}}} \right) = 4.54 \text{ mol H}$$

$$\text{Moles O} = (54.50 \text{ g } \cancel{\text{O}}) \left( \frac{1 \text{ mol O}}{16.00 \text{ g } \cancel{\text{O}}} \right) = 3.406 \text{ mol O}$$

$$\text{C: } \frac{3.407}{3.406} = 1.000 \quad \text{H: } \frac{4.54}{3.406} = 1.33 \quad \text{O: } \frac{3.406}{3.406} = 1.000$$

$$\text{C:H:O} = 3(1:1.33:1) = 3:4:3$$



## Practice Exercise

A 5.325-g sample of methyl benzoate, a compound used in the manufacture of perfumes, contains 3.758 g of carbon, 0.316 g of hydrogen, and 1.251 g of oxygen. What is the empirical formula of this substance?

*Answer:*  $\text{C}_4\text{H}_4\text{O}$

## Practice Exercise

The compound of *para*-aminobenzoic acid (you may have seen it listed as PABA on your bottle of sunscreen as a UV filter) is composed of carbon (61.31%), hydrogen (5.14%), nitrogen (10.21%), and oxygen (23.33%). Find the empirical formula of PABA.

*Answer:*  $\text{C}_7\text{H}_7\text{NO}_2$

## Molecular formula from empirical formula

From percentage compositions we can obtain the empirical formula.

We can obtain the molecular formula from the empirical formula if we are given the molecular weight.

This whole number multiple is the ratio between the molecular and empirical formulas weight.

$$\text{Whole-number multiple} = \frac{\text{molecular weight}}{\text{empirical formula weight}}$$

In the vitamin-C example, the empirical formula is  $\text{C}_3\text{H}_4\text{O}_3$

So the empirical formula weight is

$$3(12) + 4(1) + 3(16) = 88 \text{ amu.}$$

The experimentally determined molecular formula weight is 176 amu.

The molecular weight is 2 times empirical weight ( $176/88 = 2$ ).

Then 2 ( $\text{C}_3\text{H}_4\text{O}_3$ )

The molecular formula is  $\text{C}_6\text{H}_8\text{O}_6$

### **SAMPLE EXERCISE 3.14** | Determining a Molecular Formula

Mesitylene, a hydrocarbon that occurs in small amounts in crude oil, has an empirical formula of  $\text{C}_3\text{H}_4$ . The experimentally determined molecular weight of this substance is 121 amu. What is the molecular formula of mesitylene?

**Solve** First, we calculate the formula weight of the empirical formula,  $\text{C}_3\text{H}_4$ :

$$3(12.0 \text{ amu}) + 4(1.0 \text{ amu}) = 40.0 \text{ amu}$$

Next, we divide the molecular weight by the empirical formula weight to obtain the multiple used to multiply the subscripts in  $\text{C}_3\text{H}_4$ :

$$\frac{\text{Molecular weight}}{\text{Empirical formula weight}} = \frac{121}{40.0} = 3.02$$

molecular formula:  $\text{C}_9\text{H}_{12}$ .

### Practice Exercise

Ethylene glycol, the substance used in automobile antifreeze, is composed of 38.7% C, 9.7% H, and 51.6% O by mass. Its molar mass is 62.1 g/mol. **(a)** What is the empirical formula of ethylene glycol? **(b)** What is its molecular formula?

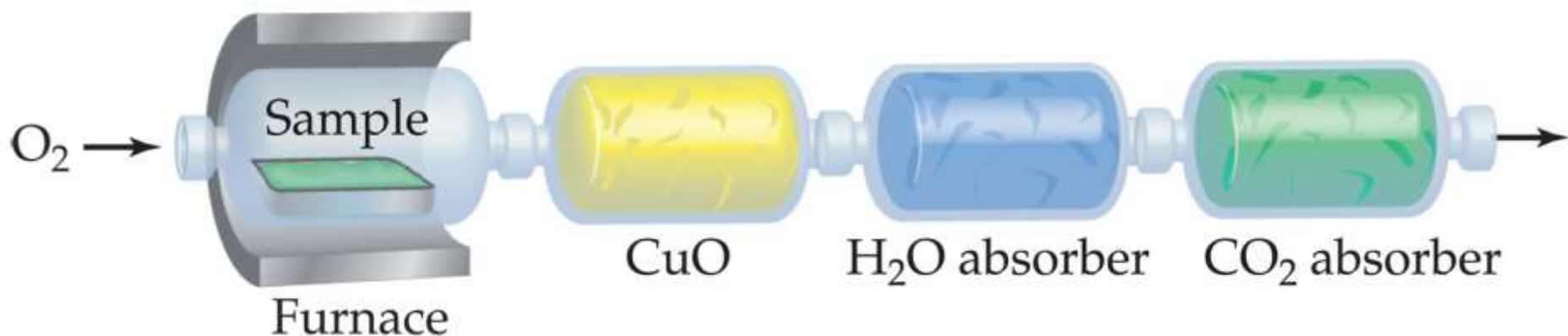
**Answers:** **(a)**  $\text{CH}_3\text{O}$ , **(b)**  $\text{C}_2\text{H}_6\text{O}_2$

The word “**Empirical Formula**” means “based on observation and experiment”.

Chemists have devised a number of experimental techniques to determine empirical formulas.

- **Combustion Analysis**
- **Elemental Analysis**

# Combustion Analysis



Compounds containing C, H and O are routinely analyzed through combustion in a chamber.

- **C** is determined from the mass of  $CO_2$  produced.
- **H** is determined from the mass of  $H_2O$  produced.
- **O** is determined by difference after the C and H have been determined.

### Sample Exercise 3.15 Determining Empirical Formula by Combustion Analysis

Isopropyl alcohol, a substance sold as rubbing alcohol, is composed of C, H, and O. Combustion of 0.255 g of isopropyl alcohol produces 0.561 g of CO<sub>2</sub> and 0.306 g of H<sub>2</sub>O. Determine the empirical formula of isopropyl alcohol.

#### Solution

To calculate the number of grams of C, we first use the molar mass of CO<sub>2</sub>, 1 mol CO<sub>2</sub> = 44.0 g CO<sub>2</sub>, to convert grams of CO<sub>2</sub> to moles of CO<sub>2</sub>. Because each CO<sub>2</sub> molecule has only 1 C atom, there is 1 mol of C atoms per mole of CO<sub>2</sub> molecules. This fact allows us to convert the moles of CO<sub>2</sub> to moles of C. Finally, we use the molar mass of C, 1 mol C = 12.0 g C, to convert moles of C to grams of C. Combining the three conversion factors, we have

$$\text{Grams C} = (0.561 \text{ g CO}_2) \left( \frac{1 \text{ mol CO}_2}{44.0 \text{ g CO}_2} \right) \left( \frac{1 \text{ mol C}}{1 \text{ mol CO}_2} \right) \left( \frac{12.0 \text{ g C}}{1 \text{ mol C}} \right) = 0.153 \text{ g C}$$



## Solution (continued)

The calculation of the number of grams of H from the grams of H<sub>2</sub>O is similar, although we must remember that there are 2 mol of H atoms per 1 mol of H<sub>2</sub>O molecules:

$$\text{Grams H} = (0.306 \text{ g H}_2\text{O}) \left( \frac{1 \text{ mol H}_2\text{O}}{18.0 \text{ g H}_2\text{O}} \right) \left( \frac{2 \text{ mol H}}{1 \text{ mol H}_2\text{O}} \right) \left( \frac{1.01 \text{ g H}}{1 \text{ mol H}} \right) = 0.0343 \text{ g H}$$

The total mass of the sample, 0.255 g, is the sum of the masses of the C, H, and O. Thus, we can calculate the mass of O as follows:

$$\begin{aligned} \text{Mass of O} &= \text{mass of sample} - (\text{mass of C} + \text{mass of H}) \\ &= 0.255 \text{ g} - (0.153 \text{ g} + 0.0343 \text{ g}) = 0.068 \text{ g O} \end{aligned}$$

We then calculate the number of moles of C, H, and O in the sample:

$$\begin{aligned} \text{Moles C} &= (0.153 \text{ g C}) \left( \frac{1 \text{ mol C}}{12.0 \text{ g C}} \right) = 0.0128 \text{ mol C} \\ \text{Moles H} &= (0.0343 \text{ g H}) \left( \frac{1 \text{ mol H}}{1.01 \text{ g H}} \right) = 0.0340 \text{ mol H} \\ \text{Moles O} &= (0.068 \text{ g O}) \left( \frac{1 \text{ mol O}}{16.0 \text{ g O}} \right) = 0.0043 \text{ mol O} \end{aligned}$$

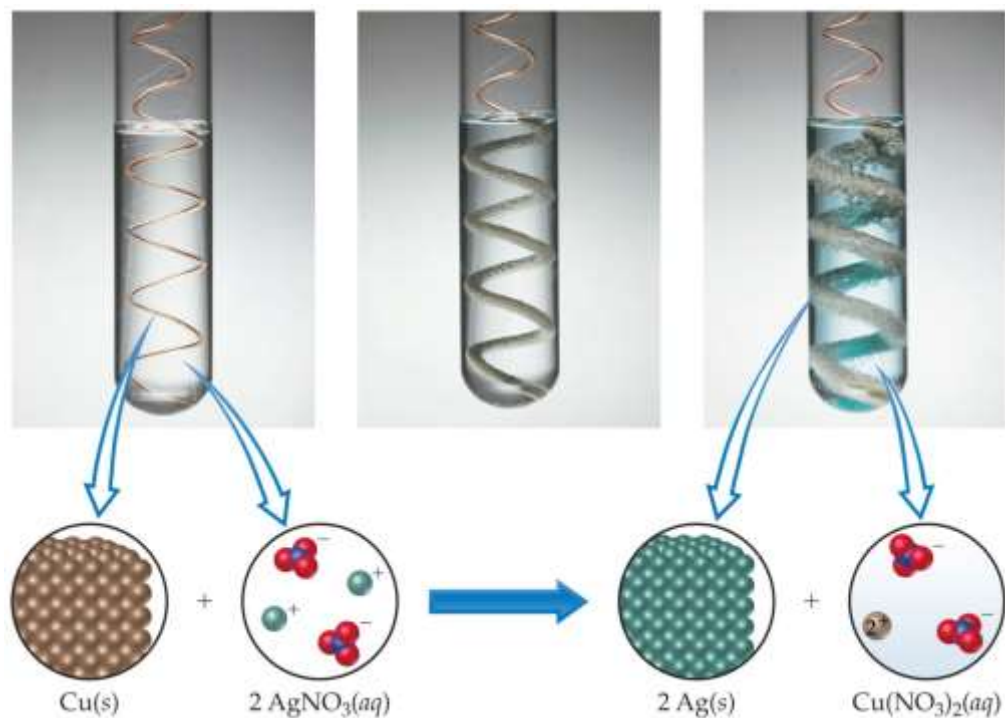
To find the empirical formula, we must compare the relative number of moles of each element in the sample. The relative number of moles of each element is found by dividing each number by the smallest number, 0.0043. The mole ratio of C:H:O so obtained is 2.98:7.91:1.00. The first two numbers are very close to the whole numbers 3 and 8, giving the empirical formula C<sub>3</sub>H<sub>8</sub>O.

## Practice Exercise

(a) Caproic acid, which is responsible for the foul odor of dirty socks, is composed of C, H, and O atoms. Combustion of a 0.225-g sample of this compound produces 0.512 g CO<sub>2</sub> and 0.209 g H<sub>2</sub>O. What is the empirical formula of caproic acid? (b) Caproic acid has a molar mass of 116 g/mol. What is its molecular formula?

**Answers:** (a) C<sub>3</sub>H<sub>6</sub>O, (b) C<sub>6</sub>H<sub>12</sub>O<sub>2</sub>

# Elemental Analyses






Compounds containing other elements are analyzed using methods analogous to those used for C, H and O.

## **3.6**

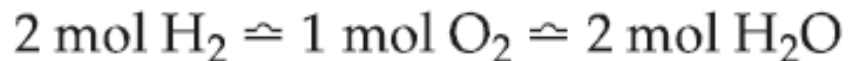
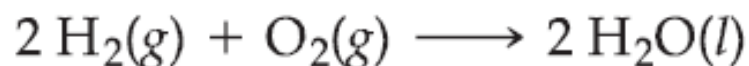
# **Quantitative Information from Balanced Equations**

# Quantitative Information from Balanced Equations

The coefficients in the balanced equation give the ratio of *moles* of reactants and products. Therefore, the mole concept allows us to convert this information to the masses of the substances.

Equation:	$2 \text{H}_2(\text{g})$	+	$\text{O}_2(\text{g})$	$\longrightarrow$	$2 \text{H}_2\text{O}(\text{l})$
Molecules:	2 molecules $\text{H}_2$	+	1 molecule $\text{O}_2$	$\longrightarrow$	2 molecules $\text{H}_2\text{O}$
					
Mass (amu):	4.0 amu $\text{H}_2$	+	32.0 amu $\text{O}_2$	$\longrightarrow$	36.0 amu $\text{H}_2\text{O}$
Amount (mol):	2 mol $\text{H}_2$	+	1 mol $\text{O}_2$	$\longrightarrow$	2 mol $\text{H}_2\text{O}$
Mass (g):	4.0 g $\text{H}_2$	+	32.0 g $\text{O}_2$	$\longrightarrow$	36.0 g $\text{H}_2\text{O}$

The coefficients in the balanced chemical equation indicate both the relative numbers of molecules and the relative numbers of moles in the reaction.



The quantities 2 mole  $\text{H}_2$ , 1 mole  $\text{O}_2$  and 2 mole  $\text{H}_2\text{O}$  are called stoichiometrically equivalent quantities.

These stoichiometric relations can be used to convert between quantities of reactants and products in a chemical equation.

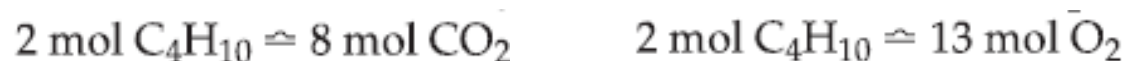
For example, the number of moles of  $\text{H}_2\text{O}$  produced from 1.57 mole of  $\text{O}_2$  can be calculated as follows:

$$\text{Moles H}_2\text{O} = (1.57 \text{ mol O}_2) \left( \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \right) = 3.14 \text{ mol H}_2\text{O}$$

When 1.57 mol  $\text{O}_2$  reacts with  $\text{H}_2$  to form  $\text{H}_2\text{O}$ , how many moles of  $\text{H}_2$  are consumed in the process?



Calculate the mass of  $\text{CO}_2$  and  $\text{O}_2$  produced when 1.00 g of  $\text{C}_4\text{H}_{10}$  is burned?

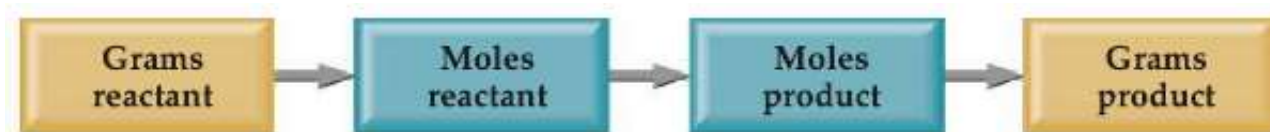


$$\begin{aligned} \text{Moles C}_4\text{H}_{10} &= (1.00 \text{ g C}_4\text{H}_{10}) \left( \frac{1 \text{ mol C}_4\text{H}_{10}}{58.0 \text{ g C}_4\text{H}_{10}} \right) \\ &= 1.72 \times 10^{-2} \text{ mol C}_4\text{H}_{10} \end{aligned}$$

$$\begin{aligned} \text{Moles CO}_2 &= (1.72 \times 10^{-2} \text{ mol C}_4\text{H}_{10}) \left( \frac{8 \text{ mol CO}_2}{2 \text{ mol C}_4\text{H}_{10}} \right) \\ &= 6.88 \times 10^{-2} \text{ mol CO}_2 \end{aligned}$$

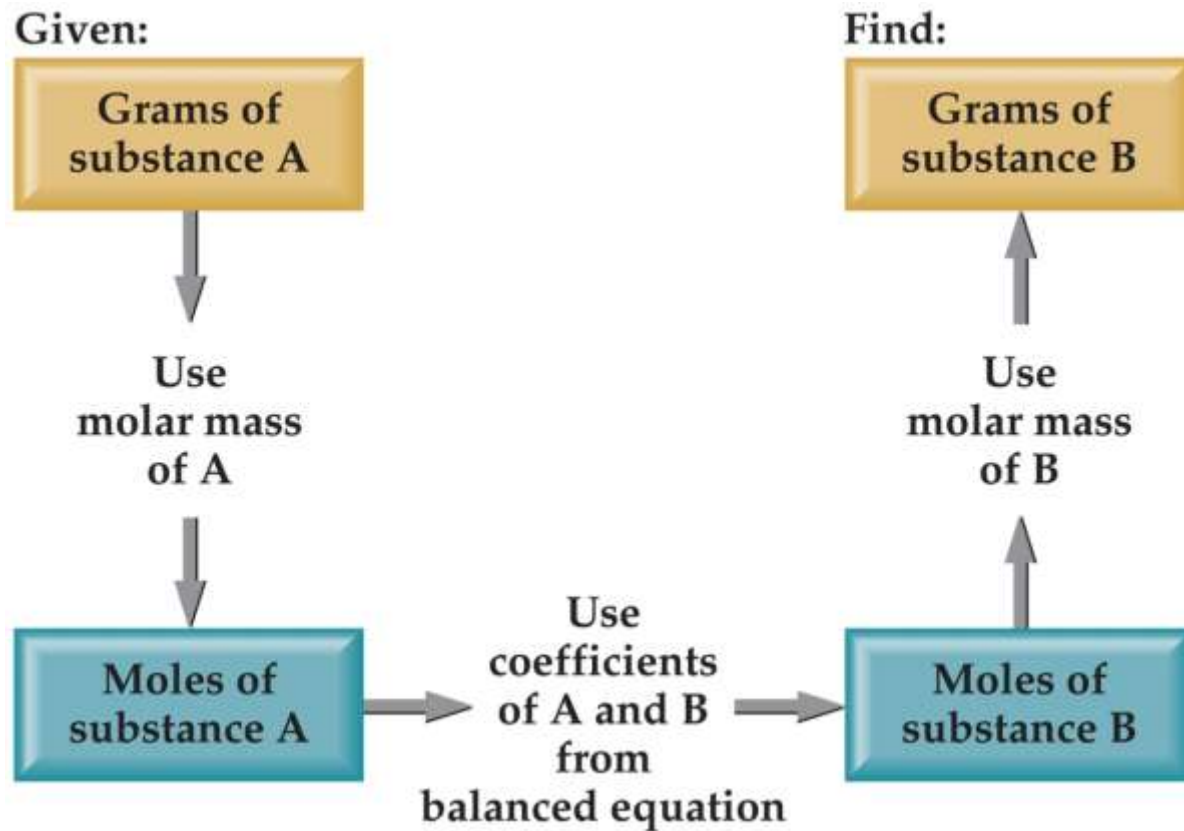
$$\text{Grams CO}_2 = (6.88 \times 10^{-2} \text{ mol CO}_2) \left( \frac{44.0 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = \mathbf{3.03 \text{ g CO}_2}$$

$$\begin{aligned} \text{Grams CO}_2 &= (1.00 \text{ g C}_4\text{H}_{10}) \left( \frac{1 \text{ mol C}_4\text{H}_{10}}{58.0 \text{ g C}_4\text{H}_{10}} \right) \left( \frac{8 \text{ mol CO}_2}{2 \text{ mol C}_4\text{H}_{10}} \right) \left( \frac{44.0 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) \\ &= 3.03 \text{ g CO}_2 \end{aligned}$$



$$\begin{aligned} \text{Grams O}_2 &= (1.00 \text{ g C}_4\text{H}_{10}) \left( \frac{1 \text{ mol C}_4\text{H}_{10}}{58.0 \text{ g C}_4\text{H}_{10}} \right) \left( \frac{13 \text{ mol O}_2}{2 \text{ mol C}_4\text{H}_{10}} \right) \left( \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} \right) \\ &= 3.59 \text{ g O}_2 \end{aligned}$$

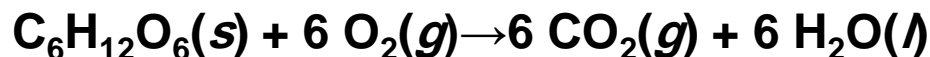
# Stoichiometric Calculations



Starting with the mass of Substance **A** you can use the ratio of the coefficients of **A** and **B** to calculate the mass of Substance **B** formed (if it's a product) or used (if it's a reactant).

### Sample Exercise 3.16 Calculating Amounts of Reactants and Products

How many grams of water are produced in the oxidation of 1.00 g of glucose,  $\text{C}_6\text{H}_{12}\text{O}_6$ ?



#### Solution

First, use the molar mass of  $\text{C}_6\text{H}_{12}\text{O}_6$  to convert from grams  $\text{C}_6\text{H}_{12}\text{O}_6$  to moles  $\text{C}_6\text{H}_{12}\text{O}_6$ :

Second, use the balanced equation to convert moles of  $\text{C}_6\text{H}_{12}\text{O}_6$  to moles of  $\text{H}_2\text{O}$ :

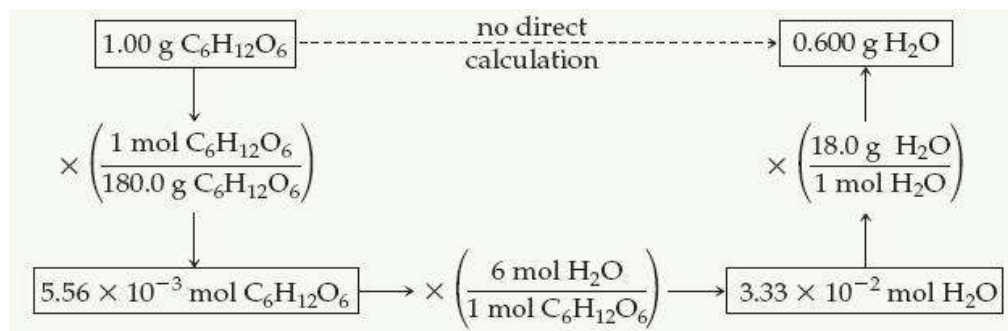
Third, use the molar mass of  $\text{H}_2\text{O}$  to convert from moles of  $\text{H}_2\text{O}$  to grams of  $\text{H}_2\text{O}$ :

The steps can be summarized in a diagram like that in Figure 3.13:

$$\text{Moles } \text{C}_6\text{H}_{12}\text{O}_6 = (1.00 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.0 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6} \right)$$

$$\text{Moles } \text{H}_2\text{O} = (1.00 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.0 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6} \right) \left( \frac{6 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6} \right)$$

$$\begin{aligned} \text{Grams } \text{H}_2\text{O} &= (1.00 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6) \left( \frac{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6}{180.0 \text{ g } \text{C}_6\text{H}_{12}\text{O}_6} \right) \left( \frac{6 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{C}_6\text{H}_{12}\text{O}_6} \right) \left( \frac{18.0 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} \right) \\ &= 0.600 \text{ g } \text{H}_2\text{O} \end{aligned}$$



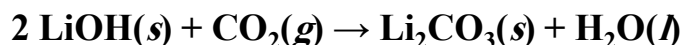


### Sample Exercise 3.17 Calculating Amounts of Reactants and Products

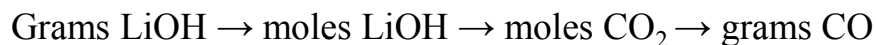
Solid lithium hydroxide is used in space vehicles to remove the carbon dioxide exhaled by astronauts. The lithium hydroxide reacts with gaseous carbon dioxide to form solid lithium carbonate and liquid water. How many grams of carbon dioxide can be absorbed by 1.00 g of lithium hydroxide?

#### Solution

**Plan** The verbal description of the reaction can be used to write a balanced equation:



We are given the grams of LiOH and asked to calculate grams of CO<sub>2</sub>. We can accomplish this task by using the following sequence of conversions:



The conversion from grams of LiOH to moles of LiOH requires the molar mass of LiOH (6.94 + 16.00 + 1.01 = 23.95 g/mol). The conversion of moles of LiOH to moles of CO<sub>2</sub> is based on the balanced chemical equation: 2 mol LiOH → 1 mol CO<sub>2</sub>. To convert the number of moles of CO<sub>2</sub> to grams, we must use the molar mass of CO<sub>2</sub>: 12.01 + 2(16.00) = 44.01 g/mol.

#### Solve

$$(1.00 \text{ g LiOH}) \left( \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} \right) \left( \frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}} \right) \left( \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 0.919 \text{ g CO}_2$$

## Practice Exercise

The decomposition of  $\text{KClO}_3$  is commonly used to prepare small amounts of  $\text{O}_2$  in the laboratory:



How many grams of  $\text{O}_2$  can be prepared from 4.50 g of  $\text{KClO}_3$ ?

**Answer:** 1.77 g

## Practice Exercise

Propane,  $\text{C}_3\text{H}_8$ , is a common fuel used for cooking and home heating. What mass of  $\text{O}_2$  is consumed in the combustion of 1.00 g of propane?

**Answer:** 3.64 g

---

# **3.7**

## **Limiting Reactants**

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# Limiting Reactants

How Many Cheese Sandwiches Can I Make?



+



The amount of available bread limits the number of sandwiches.

# How Many Cookies Can I Make?



- You can make cookies until you run out of one of the ingredients.
- Once this family runs out of sugar, they will stop making cookies.

In this example the sugar would be the **limiting reactant**, because it will limit the amount of cookies you can make.

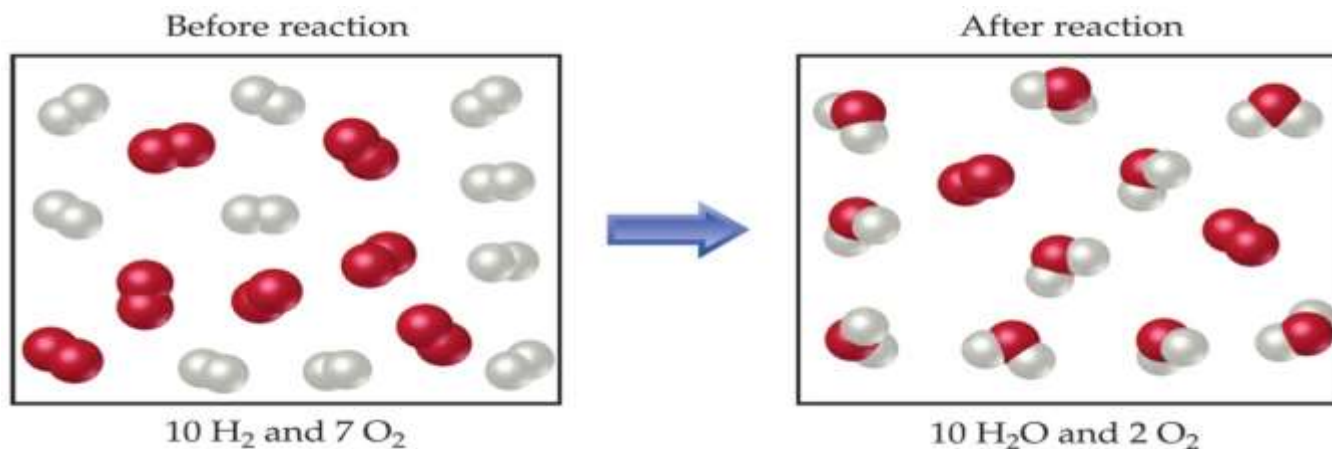
An analogous situation occurs in chemical reactions when one of the reactants is used up before the others. The reaction stops as soon as any one of the reactants is totally consumed, leaving the excess reactants as leftovers.



Suppose, for example, mixture of 10 mole  $\text{H}_2$  and 7 mole  $\text{O}_2$  react to form water.

The number of  $\text{O}_2$  needed to react with all the  $\text{H}_2$  is:

$$\text{Moles O}_2 = (10 \text{ mol H}_2) \left( \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \right) = 5 \text{ mol O}_2$$



In this example,  $\text{H}_2$  would be the **limiting reactant**, which means that once all the  $\text{H}_2$  has been consumed the reaction stops. And  $\text{O}_2$  would be the **excess reactant**, and some is left over when the reaction stops.

Summary of the example data.

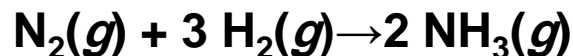
	$2 \text{H}_2(\text{g})$	+	$\text{O}_2(\text{g})$	$\longrightarrow$	$2 \text{H}_2\text{O}(\text{g})$
Initial quantities:	10 mol		7 mol		0 mol
Change (reaction):	-10 mol		-5 mol		+10 mol
Final quantities:	0 mol		2 mol		10 mol

The **limiting reactant** (or limiting reagent) is the reactant that is completely consumed in a reaction (present in the smallest stoichiometric amount).

Limiting reagent because it determines or limits the amount of product formed.

## Sample Exercise 3.18 Calculating the Amount of Product Formed from a Limiting Reactant

The most important commercial process for converting  $\text{N}_2$  from the air into nitrogen-containing compounds is based on the reaction of  $\text{N}_2$  and  $\text{H}_2$  to form ammonia ( $\text{NH}_3$ ):



How many moles of  $\text{NH}_3$  can be formed from 3.0 mol of  $\text{N}_2$  and 6.0 mol of  $\text{H}_2$ ?

### Solution

The number of moles of  $\text{H}_2$  needed for complete consumption of 3.0 mol of  $\text{N}_2$  is:

$$\text{Moles H}_2 = (3.0 \text{ mol N}_2) \left( \frac{3 \text{ mol H}_2}{1 \text{ mol N}_2} \right) = 9.0 \text{ mol H}_2$$

Because only 6.0 mol  $\text{H}_2$  is available, we will run out of  $\text{H}_2$  before the  $\text{N}_2$  is gone, and  $\text{H}_2$  will be the limiting reactant. We use the quantity of the limiting reactant,  $\text{H}_2$ , to calculate the quantity of  $\text{NH}_3$  produced:

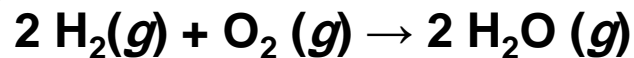
$$\text{Moles NH}_3 = (6.0 \text{ mol H}_2) \left( \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} \right) = 4.0 \text{ mol NH}_3$$

	$\text{N}_2(g)$	+	$3 \text{H}_2(g)$	$\longrightarrow$	$2 \text{NH}_3(g)$
Initial quantities:	3.0 mol		6.0 mol		0 mol
Change (reaction):	-2.0 mol		-6.0 mol		+4.0 mol
Final quantities:	1.0 mol		0 mol		4.0 mol



### Sample Exercise 3.19 Calculating the Amount of Product Formed from a Limiting Reactant

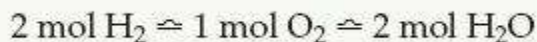
Consider the following reaction that occurs in a fuel cell:



This reaction, properly done, produces energy in the form of electricity and water. Suppose a fuel cell is set up with 150 g of hydrogen gas and 1500 grams of oxygen gas (each measurement is given with two significant figures). How many grams of water can be formed?

#### Solution

From the balanced equation, we have the following stoichiometric relations:



Using the molar mass of each substance, we can calculate the number of moles of each reactant:

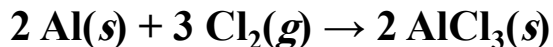
$$\begin{aligned}\text{Moles H}_2 &= (150 \text{ g H}_2) \left( \frac{1 \text{ mol H}_2}{2.00 \text{ g H}_2} \right) = 75 \text{ mol H}_2 \\ \text{Moles O}_2 &= (1500 \text{ g O}_2) \left( \frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \right) = 47 \text{ mol O}_2\end{aligned}$$

Thus, there are more moles of  $\text{H}_2$  than  $\text{O}_2$ . The coefficients in the balanced equation indicate, however, that the reaction requires 2 moles of  $\text{H}_2$  for every 1 mole of  $\text{O}_2$ . Therefore, to completely react all the  $\text{O}_2$ , we would need  $2 \times 47 = 94$  moles of  $\text{H}_2$ . Since there are only 75 moles of  $\text{H}_2$ ,  $\text{H}_2$  is the limiting reagent. We therefore use the quantity of  $\text{H}_2$  to calculate the quantity of product formed. We can begin this calculation with the grams of  $\text{H}_2$ , but we can save a step by starting with the moles of  $\text{H}_2$  that were calculated previously in the exercise:

$$\begin{aligned}\text{Grams H}_2\text{O} &= (75 \text{ moles H}_2) \left( \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \right) \left( \frac{18.0 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} \right) \\ &= 1400 \text{ g H}_2\text{O} \text{ (to two significant figures)}\end{aligned}$$

## Practice Exercise

Consider the reaction:

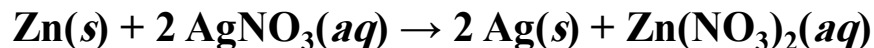


A mixture of 1.50 mol of Al and 3.00 mol of  $\text{Cl}_2$  is allowed to react. **(a)** Which is the limiting reactant? **(b)** How many moles of  $\text{AlCl}_3$  are formed? **(c)** How many moles of the excess reactant remain at the end of the reaction?

**Answers:** **(a)** Al, **(b)** 1.50 mol, **(c)** 0.75 mol  $\text{Cl}_2$

## Practice Exercise

A strip of zinc metal with a mass of 2.00 g is placed in an aqueous solution containing 2.50 g of silver nitrate, causing the following reaction to occur:



**(a)** Which reactant is limiting? **(b)** How many grams of Ag will form? **(c)** How many grams of  $\text{Zn}(\text{NO}_3)_2$  will form? **(d)** How many grams of the excess reactant will be left at the end of the reaction?

**Answers:** **(a)**  $\text{AgNO}_3$ , **(b)** 1.59 g, **(c)** 1.39 g, **(d)** 1.52 g Zn

# Theoretical Yields

- The **theoretical yield** is the maximum amount of product that can be made (when all of the limiting reactant reacts).
  - In other words it's the amount of product possible as calculated through the stoichiometry problem.
- This is different from the **actual yield**, which is the amount one actually produces in a reaction and measures.
  - The actual yield is almost always less than (and can never be greater than) the theoretical yield.
  - Reasons because part of the reactants may not react, or may react in a way different from the desired (side reactions), or its not always possible to recover all of the product from the reaction mixture.

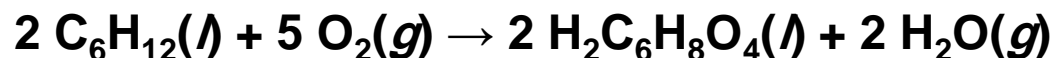
# Percent Yield

One finds the percent yield by comparing the amount actually obtained (actual yield) to the amount it was possible to make (theoretical yield).

$$\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

### Sample Exercise 3.20 Calculating the Theoretical Yield and the Percent Yield for a Reaction

Adipic acid,  $\text{H}_2\text{C}_6\text{H}_8\text{O}_4$ , is used to produce nylon. The acid is made commercially by a controlled reaction between cyclohexane ( $\text{C}_6\text{H}_{12}$ ) and  $\text{O}_2$ :



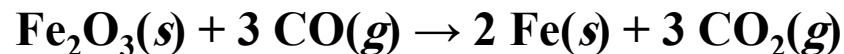
- (a) Assume that you carry out this reaction starting with 25.0 g of cyclohexane and that cyclohexane is the limiting reactant. What is the theoretical yield of adipic acid?
- (b) If you obtain 33.5 g of adipic acid from your reaction, what is the percent yield of adipic acid?

#### Solution

$$\begin{aligned} \text{(a) Grams } \text{H}_2\text{C}_6\text{H}_8\text{O}_4 &= (25.0 \text{ g } \text{C}_6\text{H}_{12}) \left( \frac{1 \text{ mol } \text{C}_6\text{H}_{12}}{84.0 \text{ g } \text{C}_6\text{H}_{12}} \right) \left( \frac{2 \text{ mol } \text{H}_2\text{C}_6\text{H}_8\text{O}_4}{2 \text{ mol } \text{C}_6\text{H}_{12}} \right) \left( \frac{146.0 \text{ g } \text{H}_2\text{C}_6\text{H}_8\text{O}_4}{1 \text{ mol } \text{H}_2\text{C}_6\text{H}_8\text{O}_4} \right) \\ &= 43.5 \text{ g } \text{H}_2\text{C}_6\text{H}_8\text{O}_4 \end{aligned}$$
$$\text{(b) Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{33.5 \text{ g}}{43.5 \text{ g}} \times 100\% = 77.0\%$$

## Practice Exercise

Imagine that you are working on ways to improve the process by which iron ore containing  $\text{Fe}_2\text{O}_3$  is converted into iron. In your tests you carry out the following reaction on a small scale:



(a) If you start with 150 g of  $\text{Fe}_2\text{O}_3$  as the limiting reagent, what is the theoretical yield of Fe? (b) If the actual yield of Fe in your test was 87.9 g, what was the percent yield?

**Answers:** (a) 105 g Fe, (b) 83.7%



Q & A



When hydrocarbons are burned in air, they form:

- a. water and carbon dioxide.**
- b. charcoal.
- c. methane.
- d. oxygen and water.



The formula weight of  $\text{Na}_3\text{PO}_4$  is:

- a. 70 grams/mole
- b. 164 grams/mole
- c. 265 grams/mole
- d. 116 grams/mole

The percentage by mass of phosphorus in  $\text{Na}_3\text{PO}_4$  is:

- a. 44.0%
- b. 11.7%
- c. 26.7%
- d. 18.9%

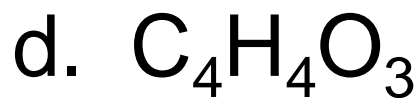
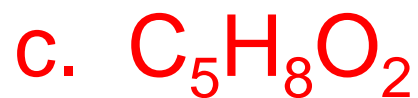
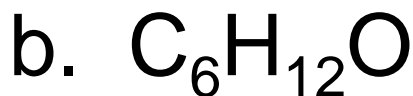
The formula weight of any substance is equal to:

- a. Avogadro's number.
- b. its atomic weight.
- c. its density.
- d. its molar mass.

Ethyl alcohol contains 52.2% C, 13.0% H, and 34.8 % O by mass. What is the empirical formula of ethyl alcohol?

- a.  $\text{C}_2\text{H}_5\text{O}_2$
- b.  $\text{C}_2\text{H}_6\text{O}$
- c.  $\text{C}_2\text{H}_6\text{O}_2$
- d.  $\text{C}_3\text{H}_4\text{O}_2$

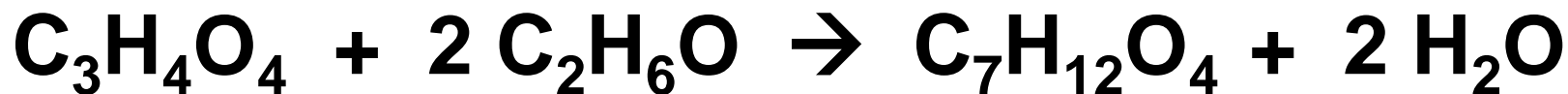
Methyl methacrylate has a molar mass of 100 g/mole. When a sample of methyl methacrylate weighing 3.14 mg was completely combusted, the only products formed were 6.91 mg of  $\text{CO}_2$  and 2.26 mg of water. What is methyl methacrylate's molecular formula?





If 10.0 grams of iron and 20.0 grams of chlorine react as shown, what is the theoretical yield of ferric chloride?

- a. 10.0 grams
- b. 20.0 grams
- c. 29.0 grams
- d. 30.0 grams



When 15.0 grams of each reactant were mixed together, the yield of  $\text{C}_7\text{H}_{12}\text{O}_4$  was 15.0 grams. What was the percentage yield?

- a. 100.0%
- b. 75.0%
- c. 65.0%
- d. 50.0%

The percentage yield of a reaction is  $(100.0\%)(X)$ . Which of the following is  $X$ ?

- a. theoretical yield / actual yield
- b. calculated yield / actual yield
- c. calculated yield / theoretical yield
- d. actual yield / theoretical yield



How many oxygen atoms are present in  **$\text{MgSO}_4 \cdot 7 \text{H}_2\text{O}$** ?

- 4 oxygen atoms
- 5 oxygen atoms
- 7 oxygen atoms
- 11 oxygen atoms
- 18 oxygen atoms

How many sulfur atoms are present in 1.0 mole of  $\text{Al}_2(\text{SO}_4)_3$ ?

- 1 sulfur atom
- 3 sulfur atoms
- 4 sulfur atoms
- $6.0 \times 10^{23}$  sulfur atoms
- $1.8 \times 10^{24}$  sulfur atoms

If you have equal masses of the following metals, which will have the most number of atoms?

1. lithium

2. sodium

3. potassium

4. rubidium

5. calcium



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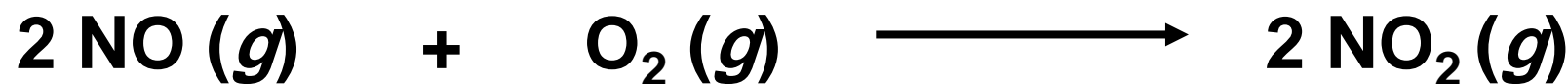
An alkali metal



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Ca in H<sub>2</sub>O

How many moles of oxygen gas are required to react completely with 1.0 mole NO?



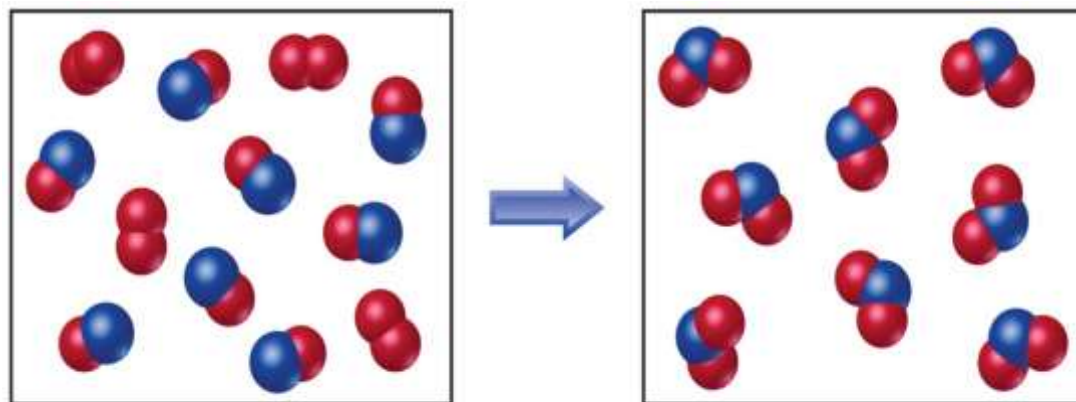
1. 0.5 mol O<sub>2</sub>

2. 1.0 mol O<sub>2</sub>

3. 1.5 mol O<sub>2</sub>

4. 2.0 mol O<sub>2</sub>

5. 2.5 mol O<sub>2</sub>

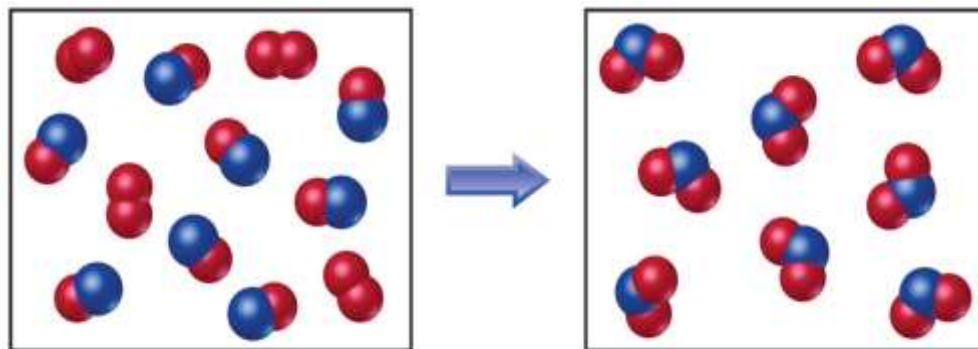


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If 10.0 moles of NO are reacted with 6.0 moles O<sub>2</sub>, how many moles NO<sub>2</sub> are produced?



1. 2.0 mol NO<sub>2</sub>
2. 6.0 mol NO<sub>2</sub>
3. 10.0 mol NO<sub>2</sub>
4. 16.0 mol NO<sub>2</sub>
5. 32.0 mol NO<sub>2</sub>



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If 10.0 moles of NO are reacted with 6.0 moles O<sub>2</sub>, how many moles of the excess reagent remain?



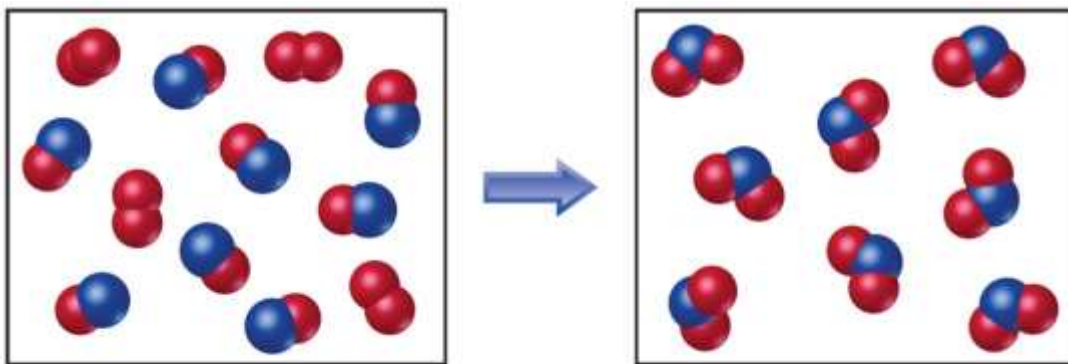
1. 1.0 mol O<sub>2</sub>

2. 5.0 mol O<sub>2</sub>

3. 4.0 mol NO

4. 8.0 mol NO

5. None of the above



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Calculate the percentage of nitrogen, by mass, in  $\text{Ca}(\text{NO}_3)_2$ .

***Answer: 17.1%***

**(a)** How many nitric acid molecules are in 4.20 g of  $\text{HNO}_3$ ? **(b)** How many O atoms are in this sample?

***Answer:*** **(a)**  $4.01 \times 10^{22}$  molecules  $\text{HNO}_3$ , **(b)**  $1.20 \times 10^{23}$  atoms O



