Chemistry, The Central Science, 11th edition
Theodore L. Brown; H. Eugene LeMay, Jr.;
Bruce E. Bursten; Catherine J. Murphy

Chapter 3

Stoichiometry

Calculations with Chemical Formulas and Equations

Ahmad Aqel Ifseisi
Assistant Professor of Analytical Chemistry
College of Science, Department of Chemistry
King Saud University
P.O. Box 2455 Riyadh 11541 Saudi Arabia
Office: AA53
Tel. 014674198, Fax: 014675992
Web site: http://fac.ksu.edu.sa/aifseisi
E-mail: ahmad3qel@yahoo.com
aifseisi@ksu.edu.sa

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.

3.1
Chemical Equations

Chemical Equations

Chemical equations are concise representations of chemical reactions.

When hydrogen gas H_2 burns, it reacts with oxygen O_2 in the air to form water H_2O.

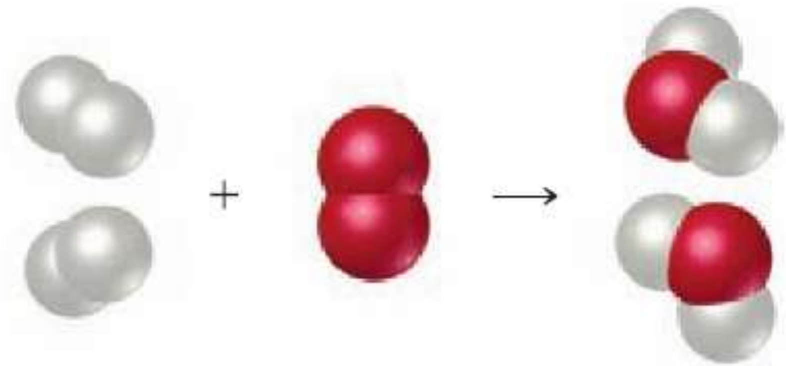
$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

(Reactants) (Products)

+ reacts with

$\longrightarrow$ produces

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
H_2O	One molecule of water:		Two H atoms and one O atom
$2\,H_2O$	Two molecules of water:		Four H atoms and two O atoms
H_2O_2	One molecule of hydrogen peroxide:		Two H atoms and two O atoms

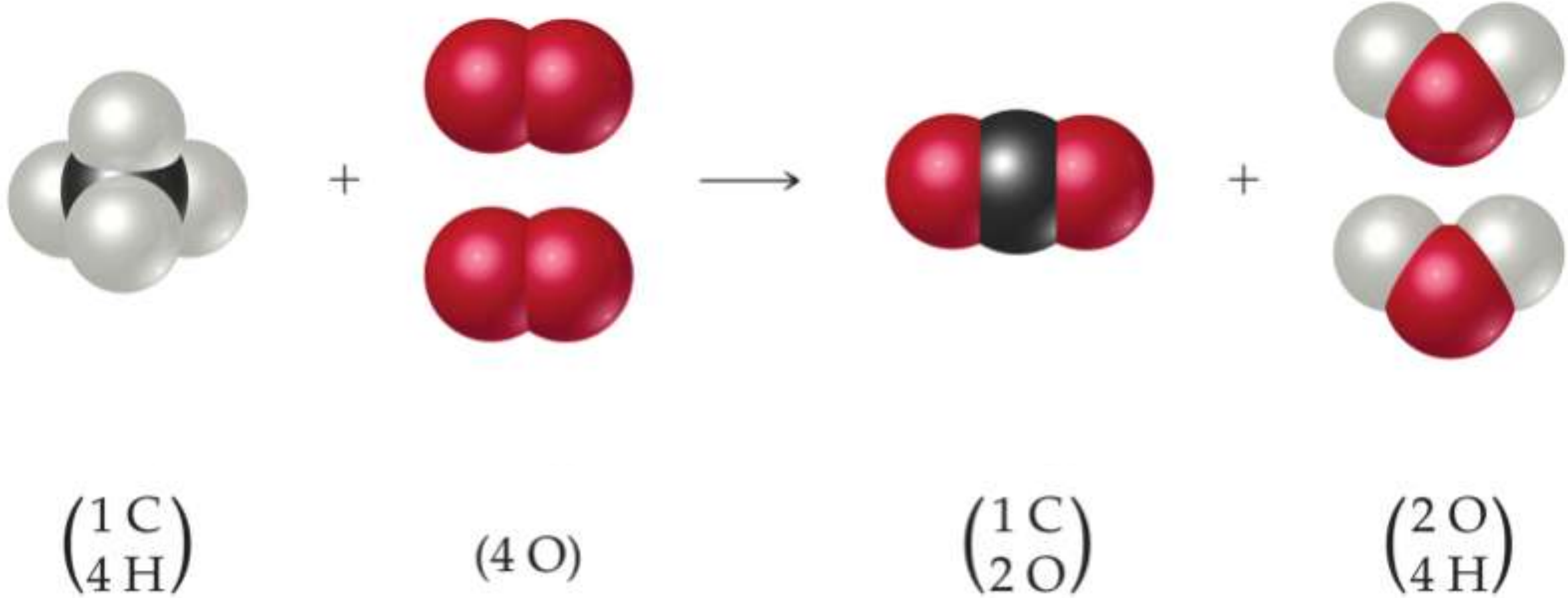
Subscripts tell the number of atoms of each element in a molecule.
Coefficients tell the number of molecules.

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

3 atoms Mg, 6 atoms O, 6 atoms H

Anatomy of a Chemical Equation

$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$



$$CH_{4\,(g)} + 2\,O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$

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.

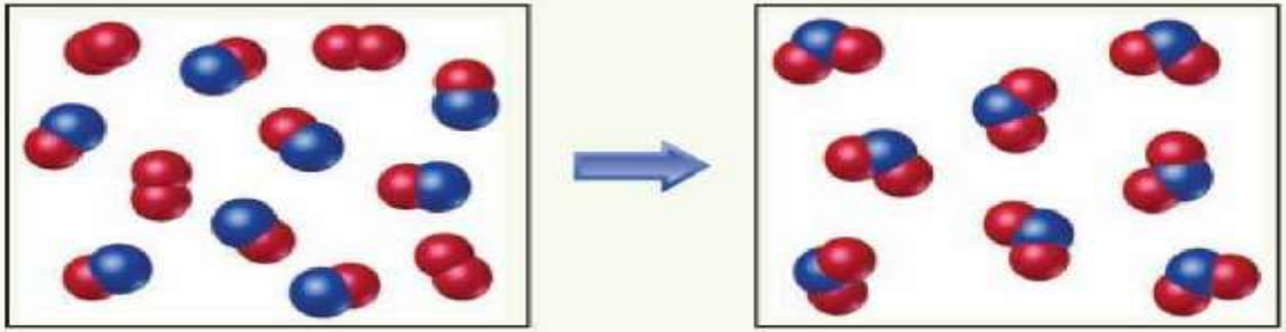
Subscripts present within a formula and tell the number of atoms of each element in a molecule.

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. Δ 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

$$O_2 + NO \longrightarrow NO_2 \quad \text{(unbalanced)}$$

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:

$$O_2 + NO \longrightarrow 2\,NO_2 \quad \text{(unbalanced)}$$

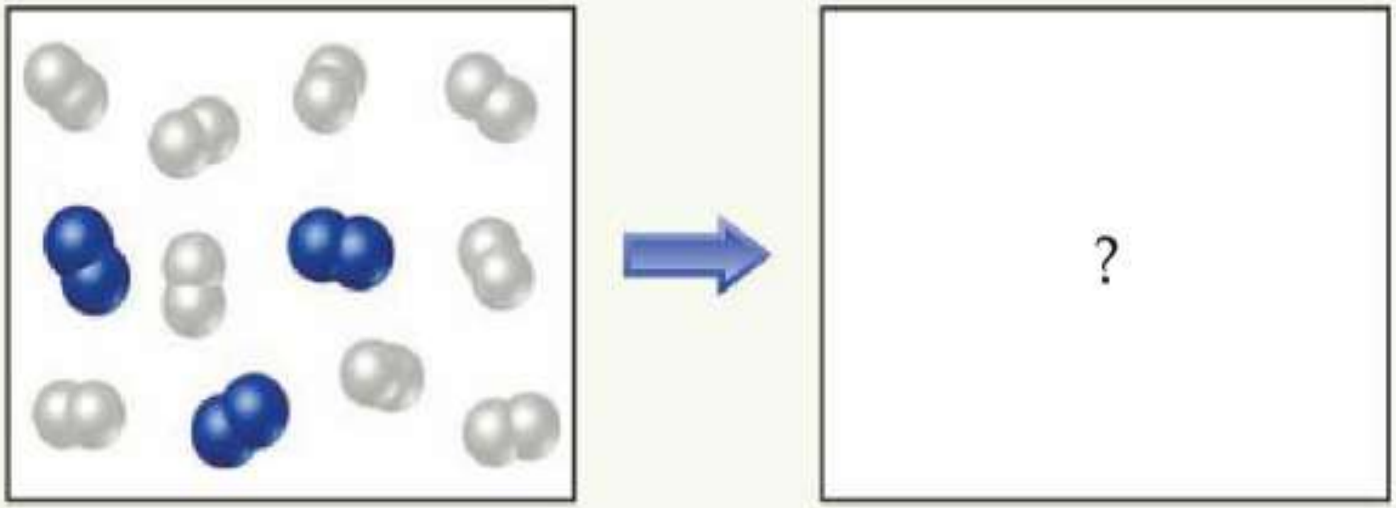
Now there are two N atoms and four O atoms on the right. Placing the coefficient 2 in front of NO balances both the number of N atoms and O atoms:

$$O_2 + 2\,NO \longrightarrow 2\,NO_2 \quad \text{(balanced)}$$

(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 NH_3 molecules should be shown in the right box?



Answer: Six NH_3 molecules

SAMPLE EXERCISE 3.2 | Balancing Chemical Equations

Balance this equation:

$$Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$$

SOLUTION

$$2\,Na(s) + 2\,H_2O(l) \longrightarrow 2\,NaOH(aq) + H_2(g)$$

PRACTICE EXERCISE

Balance the following equations by providing the missing coefficients:

(a) $\_\_Fe(s) + \_\_O_2(g) \longrightarrow \_\_Fe_2O_3(s)$

(b) $\_\_C_2H_4(g) + \_\_O_2(g) \longrightarrow \_\_CO_2(g) + \_\_H_2O(g)$

(c) $\_\_Al(s) + \_\_HCl(aq) \longrightarrow \_\_AlCl_3(aq) + \_\_H_2(g)$

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

3.2
Some Simple Patterns of Chemical Reactivity

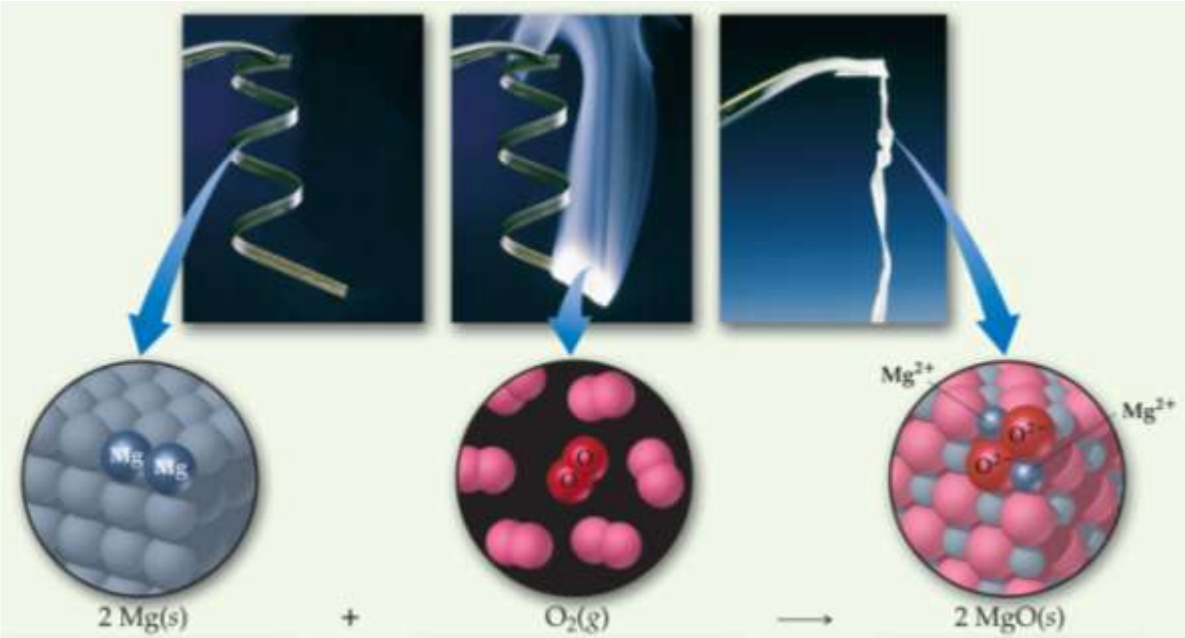
Reaction Types

(some simple patterns of chemical reactivity)

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

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

Combination Reactions

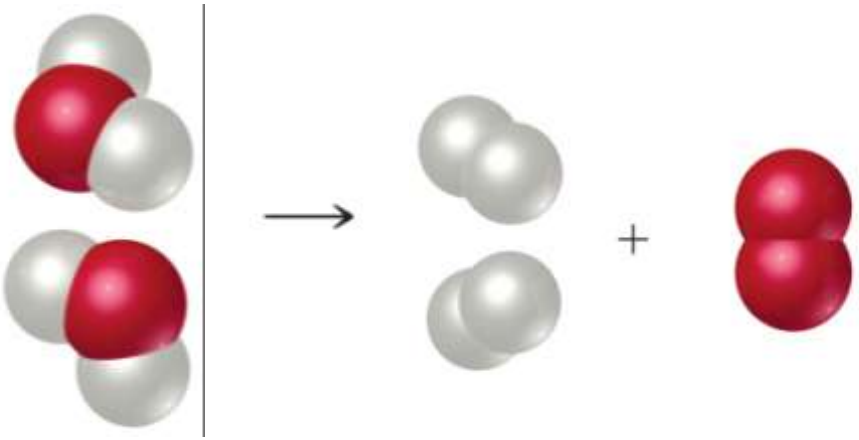


- In this type of reaction two or more substances react to form one product.
- A combination reaction between a metal and a nonmetal produce ionic solid.

A + B ⟶ C

- **Examples:**
  - $2\ Mg_{(s)} + O_{2\,(g)} \longrightarrow 2\ MgO_{(s)}$
  - $C_{(s)} + O_{2\,(g)} \longrightarrow CO_{2\,(g)}$
  - $N_{2\,(g)} + 3\ H_{2\,(g)} \longrightarrow 2\ NH_{3\,(g)}$
  - $CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2\,(s)}$
  - $C_3H_{6\,(g)} + Br_{2\,(l)} \longrightarrow C_3H_6Br_{2\,(l)}$

Decomposition Reactions



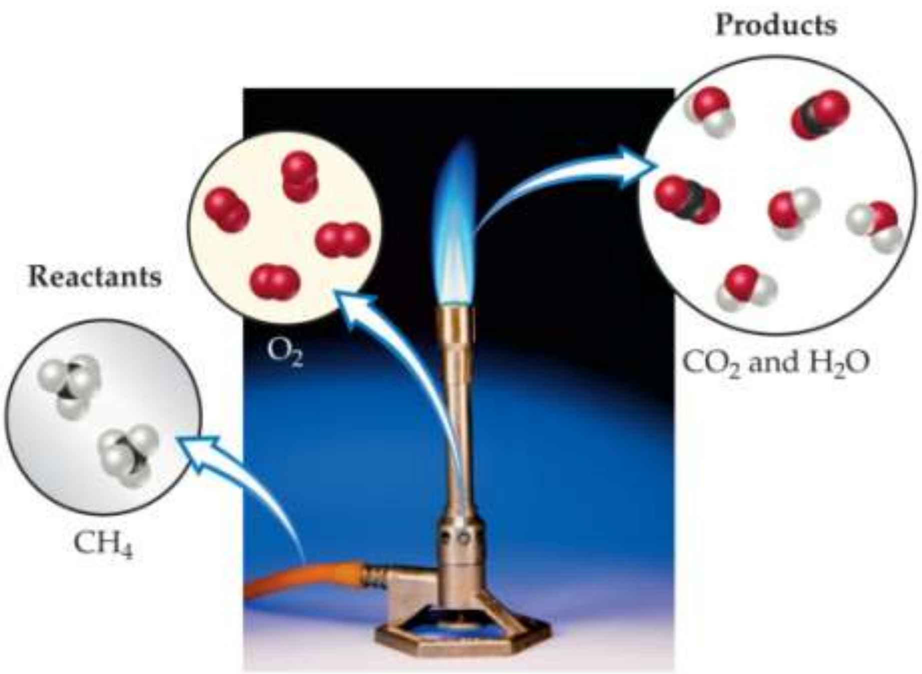
- In this type one substance breaks down into two or more substances.

- Many compounds undergo decomposition reactions when heated.

A $\longrightarrow$ B + C

- **Examples:**
  - $CaCO_{3\,(s)} \longrightarrow CaO_{\,(s)} + CO_{2\,(g)}$
  - $2\ KClO_{3\,(s)} \longrightarrow 2\ KCl_{\,(s)} + 3\ O_{2\,(g)}$
  - $Cu(OH)_{2\,(s)} \longrightarrow CuO_{\,(s)} + H_2O_{\,(l)}$
  - $PbCO_{3\,(s)} \longrightarrow PbO_{\,(s)} + CO_{2\,(g)}$
  - $2\ NaN_{3\,(s)} \longrightarrow 2\ Na_{\,(s)} + 3\ N_{2\,(g)}$

Combustion Reactions

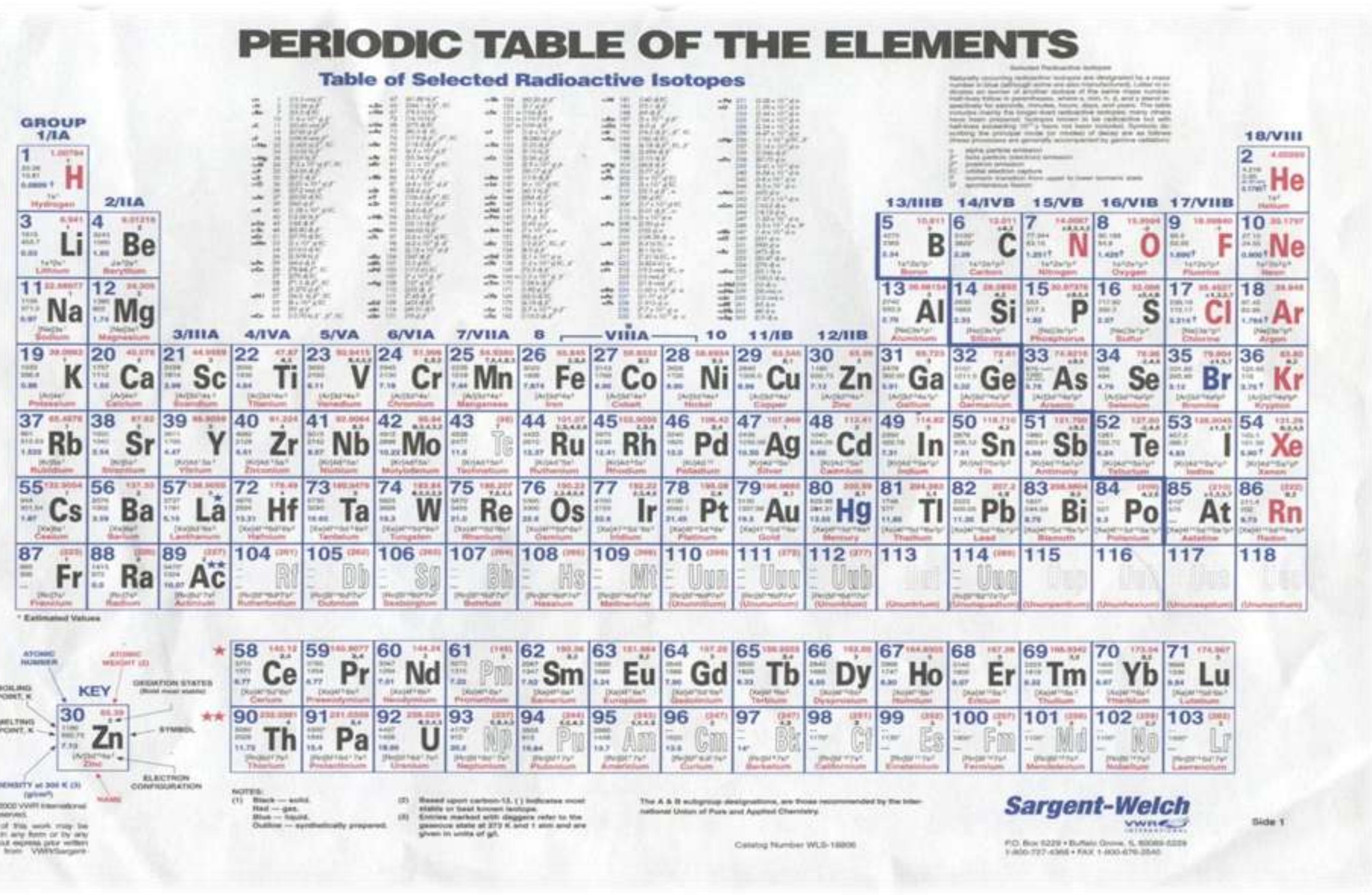


- These are generally rapid reactions that produce a flame.
- Most often involve hydrocarbons reacting with oxygen in the air.
- Hydrocarbon compounds contain only Carbons (**C**) and Hydrogen (**H**)).

- Examples:
  - $CH_{4\,(g)} + 2\ O_{2\,(g)} \longrightarrow CO_{2\,(g)} + 2\ H_2O_{\,(g)}$
  - $C_3H_{8\,(g)} + 5\ O_{2\,(g)} \longrightarrow 3\ CO_{2\,(g)} + 4\ H_2O_{\,(g)}$

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

Na_2S



SELECTED POLYATOMIC IONS

Hg_2^{2+}	dimercury (I)	CrO_4^{2-}	chromate
NH_4^+	ammonium	$Cr_2O_7^{2-}$	dichromate
$C_2H_3O_2^-$ / CH_3COO^-	acetate	MnO_4^-	permanganate
		MnO_4^{2-}	manganate
CN^-	cyanide	NO_2^-	nitrite
CO_3^{2-}	carbonate	NO_3^-	nitrate
HCO_3^-	hydrogen carbonate	OH^-	hydroxide
		PO_4^{3-}	phosphate
$C_2O_4^{2-}$	oxalate	SCN^-	thiocyanate
ClO^-	hypochlorite	SO_3^{2-}	sulfite
ClO_2^-	chlorite	SO_4^{2-}	sulfate
ClO_3^-	chlorate	HSO_4^-	hydrogen sulfate
ClO_4^-	perchlorate	$S_2O_3^{2-}$	thiosulfate

Examples:

Na_2SO_4	$CaSO_4$
Na_3PO_4	$Ca_3(PO_4)_2$
$NaNO_3$	$Ca(NO_3)_2$
$NaOH$	$Ca(OH)_2$

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.)

$$2\,Li(s) + F_2(g) \longrightarrow 2\,LiF(s)$$

$$BaCO_3(s) \longrightarrow BaO(s) + CO_2(g)$$

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)** $HgS(s) \longrightarrow Hg(l) + S(s)$; **(b)** $4\,Al(s) + 3\,O_2(g) \longrightarrow 2\,Al_2O_3(s)$

SAMPLE EXERCISE 3.4 | Writing Balanced Equations for Combustion Reactions

Write the balanced equation for the reaction that occurs when methanol, $CH_3OH(l)$, is burned in air.

$$CH_3OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(g)$$

$$2\,CH_3OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 4\,H_2O(g)$$

PRACTICE EXERCISE

Write the balanced equation for the reaction that occurs when ethanol, $C_2H_5OH(l)$, is burned in air.

Answer: $C_2H_5OH(l) + 3\,O_2(g) \longrightarrow 2\,CO_2(g) + 3\,H_2O(g)$

3.3
Formula Weights

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, $CaCl_2$, would be

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

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

$$\begin{aligned} \text{FW of } H_2SO_4 &= 2(\text{AW of H}) + (\text{AW of S}) + 4(\text{AW of O}) \\ &= 2(1.0\text{ amu}) + 32.1\text{ amu} + 4(16.0\text{ amu}) \\ &= 98.1\text{ amu} \end{aligned}$$

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, the would be

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

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

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, $C_{12}H_{22}O_{11}$ (table sugar), and **(b)** calcium nitrate, $Ca(NO_3)_2$.

SOLUTION

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

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

(b) If a chemical formula has parentheses, the subscript outside the parentheses is a multiplier for all atoms inside. Thus, for $Ca(NO_3)_2$, we have

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

PRACTICE EXERCISE

Calculate the formula weight of **(a)** $Al(OH)_3$ and **(b)** CH_3OH.
Answers: **(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 C_2H_6 is…

$$\%C = \frac{(2)(12.0\ \text{amu})}{(30.0\ \text{amu})}$$

$$= \frac{24.0\ \text{amu}}{30.0\ \text{amu}} \times 100$$

$$= 80.0\%$$

SAMPLE EXERCISE 3.6 | Calculating Percentage Composition

Calculate the percentage of carbon, hydrogen, and oxygen (by mass) in $C_{12}H_{22}O_{11}$.

SOLUTION

Let's examine this question using the problem-solving steps in the "Strategies in Chemistry: Problem Solving" essay that appears on the next page.

Analyze We are given a chemical formula, $C_{12}H_{22}O_{11}$, and asked to calculate the percentage by mass of its component elements (C, H, and O).

Plan 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 $C_{12}H_{22}O_{11}$, 342.0 amu, was calculated in Sample Exercise 3.5.) We must then do three calculations, one for each element.

Solve Using Equation 3.10, we have

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

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

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

Check The percentages of the individual elements must add up to 100%, which they do in this case. We could have used more significant figures for our atomic weights, giving more significant figures for our percentage composition, but we have adhered to our suggested guideline of rounding atomic weights to one digit beyond the decimal point.

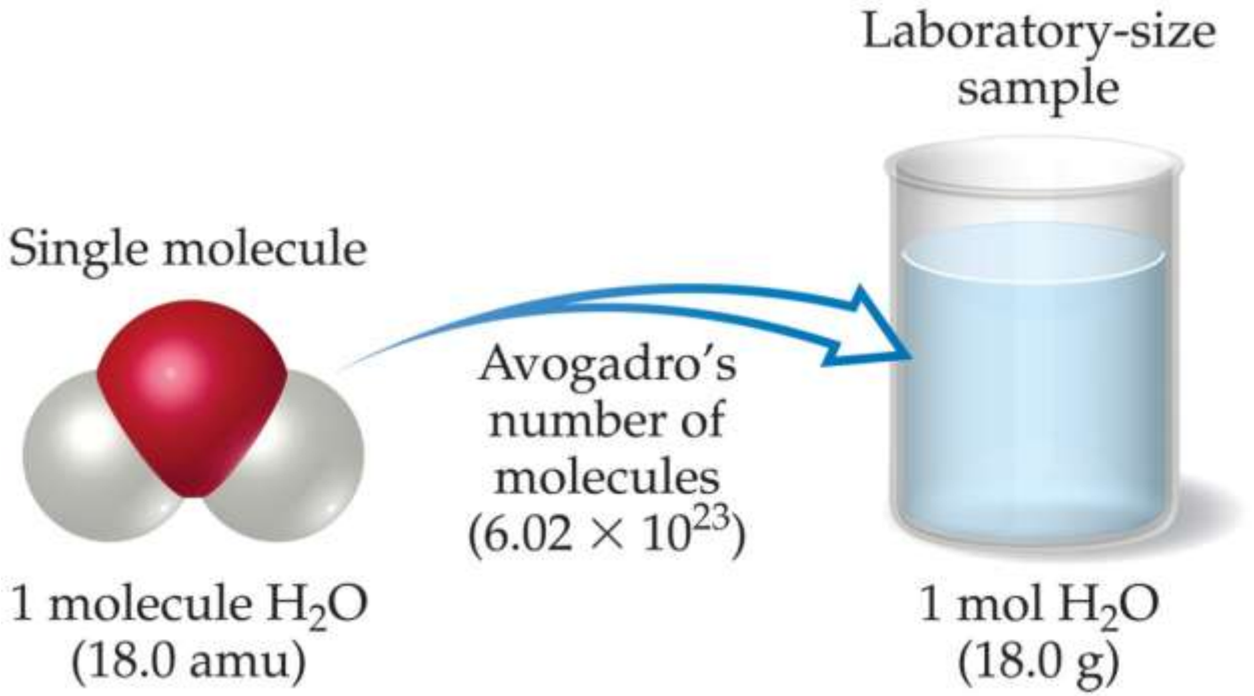
PRACTICE EXERCISE

Calculate the percentage of nitrogen, by mass, in $Ca(NO_3)_2$.
Answer: 17.1%

3.4
Avogadro's Number and the Mole

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 2×10^{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}C.

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

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

$$1 \text{ mol } ^{12}\text{C atoms} = 6.02 \times 10^{23}\ ^{12}\text{C atoms}$$

$$1 \text{ mol } H_2O \text{ molecules} = 6.02 \times 10^{23}\ H_2O \text{ molecules}$$

$$1 \text{ mol } NO_3^- \text{ ions} = 6.02 \times 10^{23}\ NO_3^- \text{ ions}$$

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}C, 1 mol C_2H_2, 9×10^{23} molecules of CO_2.

Solution

Analyze We are given amounts of different substances expressed in grams, moles, and number of molecules and asked to arrange the samples in order of increasing numbers of C atoms.

Plan To determine the number of C atoms in each sample, we must convert g ^{12}C, 1 mol C_2H_2, and 9×10^{23} molecules CO_2 all to numbers of C atoms. To make these conversions, we use the definition of mole and Avogadro's number.

Solve 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}C. Thus, 12 g of ^{12}C contains 1 mol of C atoms (that is, 6.02×10^{23} C atoms). One mol of C_2H_2 contains 6×10^{23} C_2H_2 molecules. Because there are two C atoms in each C_2H_2 molecule, this sample contains 12×10^{23} C atoms. Because each CO_2 molecule contains one C atom, the sample of CO_2 contains 9×10^{23} C atoms. Hence, the order is 12 g ^{12}C (6×10^{23} C atoms) < 9×10^{23} CO_2 molecules (9×10^{23} C atoms) < 1 mol C_2H_2 (12×10^{23} C atoms).

Check We can check our results by comparing the number of moles of C atoms in each sample because the number of moles is proportional to the number of atoms. Thus, 12 g of ^{12}C is 1 mol C; 1 mol of C_2H_2 contains 2 mol C, and 9×10^{23} molecules of CO_2 contain 1.5 mol C, giving the same order as above: 12 g ^{12}C (1 mol C) < 9×10^{23} CO_2 molecules (1.5 mol C) < 1 mol C_2H_2 (2 mol C).

PRACTICE EXERCISE

Without using a calculator, arrange the following samples in order of increasing number of O atoms: 1 mol H_2O, 1 mol CO_2, 3×10^{23} molecules O_3.

Answer: 1 mol H_2O (6×10^{23} O atoms) < 3×10^{23} molecules O_3 (9×10^{23} O atoms) < 1 mol CO_2 (12×10^{23} O atoms)

Sample Exercise 3.8 Converting Moles to Atoms

Calculate the number of H atoms in 0.350 mol of $C_6H_{12}O_6$.

Solution

Analyze We are given both the amount of a substance (0.350 mol) and its chemical formula ($C_6H_{12}O_6$). The unknown is the number of H atoms in the sample.

Plan Avogadro's number provides the conversion factor between the number of moles of $C_6H_{12}O_6$ and the number of molecules of $C_6H_{12}O_6$. Once we know the number of molecules of $C_6H_{12}O_6$, we can use the chemical formula, which tells us that each molecule of $C_6H_{12}O_6$ contains 12 H atoms. Thus, we convert moles of $C_6H_{12}O_6$ to molecules of $C_6H_{12}O_6$ and then determine the number of atoms of H from the number of molecules of $C_6H_{12}O_6$:

$$\text{Moles } C_6H_{12}O_6 \rightarrow \text{molecules } C_6H_{12}O_6 \rightarrow \text{atoms H}$$

Solve

$$\begin{aligned}\text{H atoms} &= (0.350\ \cancel{\text{mol } C_6H_{12}O_6})\left(\frac{6.02 \times 10^{23}\ \cancel{\text{molecules } C_6H_{12}O_6}}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)\left(\frac{12\ \text{H atoms}}{1\ \cancel{\text{molecule } C_6H_{12}O_6}}\right)\\ &= 2.53 \times 10^{24}\ \text{H atoms}\end{aligned}$$

PRACTICE EXERCISE

How many oxygen atoms are in **(a)** 0.25 mol $Ca(NO_3)_2$ and **(b)** 1.50 mol of sodium carbonate?

Answers: **(a)** 9.0×10^{23}, **(b)** 2.71×10^{24}

Molar Mass

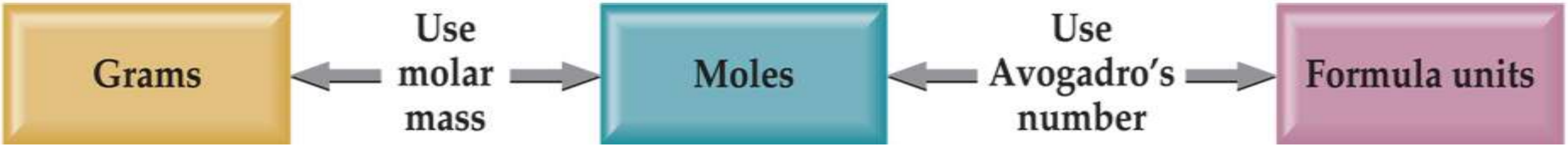
By definition, a molar mass is the mass of 1 mol of a substance (i.e., g/mol).

- 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).

$$\textbf{Avogadro's No } (N_A) \; (6.022 \times 10^{23}) = \frac{\text{No. of atoms or molecules or ions or particles}}{\text{No. of mole}}$$

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

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	6.02×10^{23} N atoms
Molecular nitrogen	N_2	28.0	28.0	6.02×10^{23} N_2 molecules $2(6.02 \times 10^{23})$ N atoms
Silver	Ag	107.9	107.9	6.02×10^{23} Ag atoms
Silver ions	Ag^+	107.9[a]	107.9	6.02×10^{23} Ag^+ ions
Barium chloride	$BaCl_2$	208.2	208.2	6.02×10^{23} $BaCl_2$ units 6.02×10^{23} Ba^{2+} ions $2(6.02 \times 10^{23})$ Cl^- ions

[a] 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:

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

Sample Exercise 3.9 Calculating Molar Mass

What is the mass in grams of 1.000 mol of glucose, $C_6H_{12}O_6$?

Solution

Analyze We are given a chemical formula and asked to determine its molar mass.
Plan The molar mass of a substance is found by adding the atomic weights of its component atoms.
Solve

$$\begin{aligned} 6\ \text{C atoms} &= 6(12.0\ \text{amu}) = 72.0\ \text{amu} \\ 12\ \text{H atoms} &= 12(1.0\ \text{amu}) = 12.0\ \text{amu} \\ 6\ \text{O atoms} &= 6(16.0\ \text{amu}) = \underline{96.0\ \text{amu}} \\ & \qquad\qquad\qquad\qquad 180.0\ \text{amu} \end{aligned}$$

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, $C_6H_{12}O_6$ has a molar mass of 180.0 g/mol.

PRACTICE EXERCISE

Calculate the molar mass of $Ca(NO_3)_2$.
Answer: 164.1 g/mol

SAMPLE EXERCISE 3.10 | Converting Grams to Moles

Calculate the number of moles of glucose ($C_6H_{12}O_6$) in 5.380 g of $C_6H_{12}O_6$.

SOLUTION

Analyze We are given the number of grams of a substance and its chemical formula and asked to calculate the number of moles.

Plan The molar mass of a substance provides the factor for converting grams to moles. The molar mass of $C_6H_{12}O_6$ is 180.0 g/mol (Sample Exercise 3.9).

Solve Using 1 mol $C_6H_{12}O_6$ = 180.0 g $C_6H_{12}O_6$ to write the appropriate conversion factor, we have

$$\text{Moles } C_6H_{12}O_6 = (5.380 \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1 \text{ mol } C_6H_{12}O_6}{180.0 \cancel{\text{g } C_6H_{12}O_6}}\right) = 0.02989 \text{ mol } C_6H_{12}O_6$$

Check Because 5.380 g is less than the molar mass, a reasonable answer is less than one mole. The units of our answer (mol) are appropriate. The original data had four significant figures, so our answer has four significant figures.

PRACTICE EXERCISE

How many moles of sodium bicarbonate ($NaHCO_3$) are in 508 g of $NaHCO_3$?
Answer: 6.05 mol $NaHCO_3$

SAMPLE EXERCISE 3.11 | Converting Moles to Grams

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

SOLUTION

Analyze We are given the number of moles and the name of a substance and asked to calculate the number of grams in the sample.

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

Solve Because the calcium ion is Ca^{2+} and the nitrate ion is NO_3^-, calcium nitrate is $Ca(NO_3)_2$. Adding the atomic weights of the elements in the compound gives a formula weight of 164.1 amu. Using 1 mol $Ca(NO_3)_2$ = 164.1 g $Ca(NO_3)_2$ to write the appropriate conversion factor, we have

$$\text{Grams } Ca(NO_3)_2 = (0.433 \cancel{\text{ mol } Ca(NO_3)_2})\left(\frac{164.1 \text{ g } Ca(NO_3)_2}{1 \cancel{\text{ mol } Ca(NO_3)_2}}\right) = 71.1 \text{ g } Ca(NO_3)_2$$

Check The number of moles is less than 1, so the number of grams must be less than the molar mass, 164.1 g. Using rounded numbers to estimate, we have $0.5 \times 150 = 75$ g. The magnitude of our answer is reasonable. Both the units (g) and the number of significant figures (3) are correct.

PRACTICE EXERCISE

What is the mass, in grams, of **(a)** 6.33 mol of $NaHCO_3$ and **(b)** 3.0×10^{-5} mol of sulfuric acid?

Answers: **(a)** 532 g, **(b)** 2.9×10^{-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 $C_6H_{12}O_6$? **(b)** How many oxygen atoms are in this sample?

(a) Plan 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 $C_6H_{12}O_6$ to moles $C_6H_{12}O_6$, which can then be converted to molecules $C_6H_{12}O_6$. The first conversion uses the molar mass of $C_6H_{12}O_6$: 1 mol $C_6H_{12}O_6$ = 180.0 g $C_6H_{12}O_6$. The second conversion uses Avogadro's number.

Solve

Molecules $C_6H_{12}O_6$

$$= (5.23\ \cancel{\text{g } C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{180.0\ \cancel{\text{g } C_6H_{12}O_6}}\right)\left(\frac{6.02 \times 10^{23}\ \text{molecules } C_6H_{12}O_6}{1\ \cancel{\text{mol } C_6H_{12}O_6}}\right)$$

$$= 1.75 \times 10^{22}\ \text{molecules } C_6H_{12}O_6$$

(b) Plan To determine the number of O atoms, we use the fact that there are six O atoms in each molecule of $C_6H_{12}O_6$. Thus, multiplying the number of molecules $C_6H_{12}O_6$ by the factor (6 atoms O/1 molecule $C_6H_{12}O_6$) gives the number of O atoms.

Solve

$$\text{Atoms O} = (1.75 \times 10^{22}\ \cancel{\text{molecules } C_6H_{12}O_6})\left(\frac{6\ \text{atoms O}}{1\ \cancel{\text{molecule } C_6H_{12}O_6}}\right)$$

$$= 1.05 \times 10^{23}\ \text{atoms O}$$

PRACTICE EXERCISE

(a) How many nitric acid molecules are in 4.20 g of HNO_3? **(b)** How many O atoms are in this sample?

Answers: **(a)** 4.01×10^{22} molecules HNO_3, **(b)** 1.20×10^{23} atoms O

3.5
Empirical Formulas from Analyses

Empirical Formulas from Analyses

From Section 2.6;

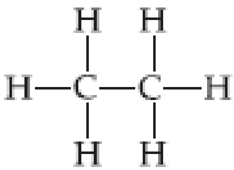
Molecular formula: chemical formula that indicate the actual numbers and types of atoms in a molecule.

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 H_2O_2, where as its empirical formula is HO. The molecular formula of ethylene is C_2H_4, where as its empirical formula is CH_2. For water, H_2O the molecular and the empirical formulas are identical.

The structural formula for the substance ethane is shown here:



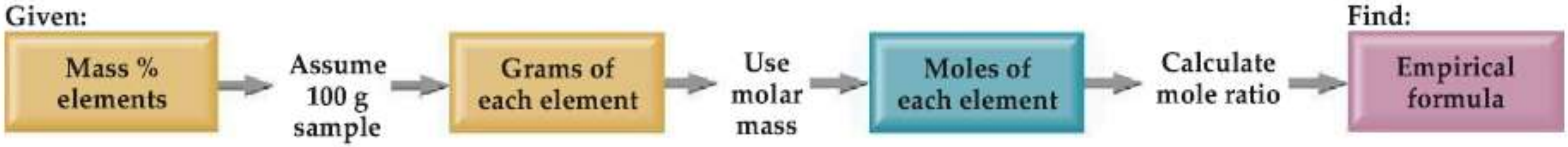
(a) C_2H_6, (b) CH_3

(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 H_2O shows that water contains two H atoms for each O atom. This ratio also applies on the molar level: 1 mole of H_2O 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

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

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

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

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

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

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

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

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:

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

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

$C_3H_4O_3$

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: C_4H_4O

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: $\mathbf{C_7H_7NO_2}$

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 $C_3H_4O_3$
So the empirical formula weight is
3(12) + 4(1) + 3(16) = 88 amu.
The experimentally determined molecular formula weight is 176 amu.
The molecular weight is 2 times empirical weight (176/88 = 2).
Then 2 ($C_3H_4O_3$)
The molecular formula is $C_6H_8O_6$

SAMPLE EXERCISE 3.14 | Determining a Molecular Formula

Mesitylene, a hydrocarbon that occurs in small amounts in crude oil, has an empirical formula of C_3H_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, C_3H_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 C_3H_4:

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

molecular formula: C_9H_{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)** CH_3O, **(b)** $C_2H_6O_2$

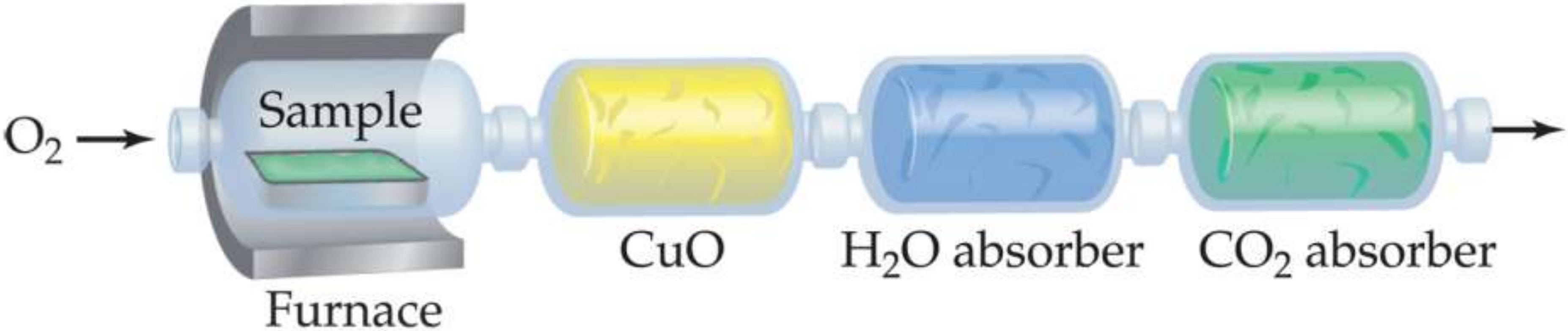
The word “**Empirical Formula**” means “based on observation and experiment”.

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_2 and 0.306 g of H_2O. Determine the empirical formula of isopropyl alcohol.

Solution

To calculate the number of grams of C, we first use the molar mass of CO_2, 1 mol CO_2 = 44.0 g CO_2, to convert grams of CO_2 to moles of CO_2. Because each CO_2 molecule has only 1 C atom, there is 1 mol of C atoms per mole of CO_2 molecules. This fact allows us to convert the moles of CO_2 to moles of C. Finally, we use the molar mass of C, 1 mol C =1 2.0 g C, to convert moles of C to grams of C. Combining the three conversion factors, we have

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

Solution (continued)

The calculation of the number of grams of H from the grams of H_2O is similar, although we must remember that there are 2 mol of H atoms per 1 mol of H_2O molecules:

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

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:

$$\text{Moles C} = (0.153\ \cancel{\text{g C}})\left(\frac{1\ \text{mol C}}{12.0\ \cancel{\text{g C}}}\right) = 0.0128\ \text{mol C}$$

$$\text{Moles H} = (0.0343\ \cancel{\text{g H}})\left(\frac{1\ \text{mol H}}{1.01\ \cancel{\text{g H}}}\right) = 0.0340\ \text{mol H}$$

$$\text{Moles O} = (0.068\ \cancel{\text{g O}})\left(\frac{1\ \text{mol O}}{16.0\ \cancel{\text{g O}}}\right) = 0.0043\ \text{mol O}$$

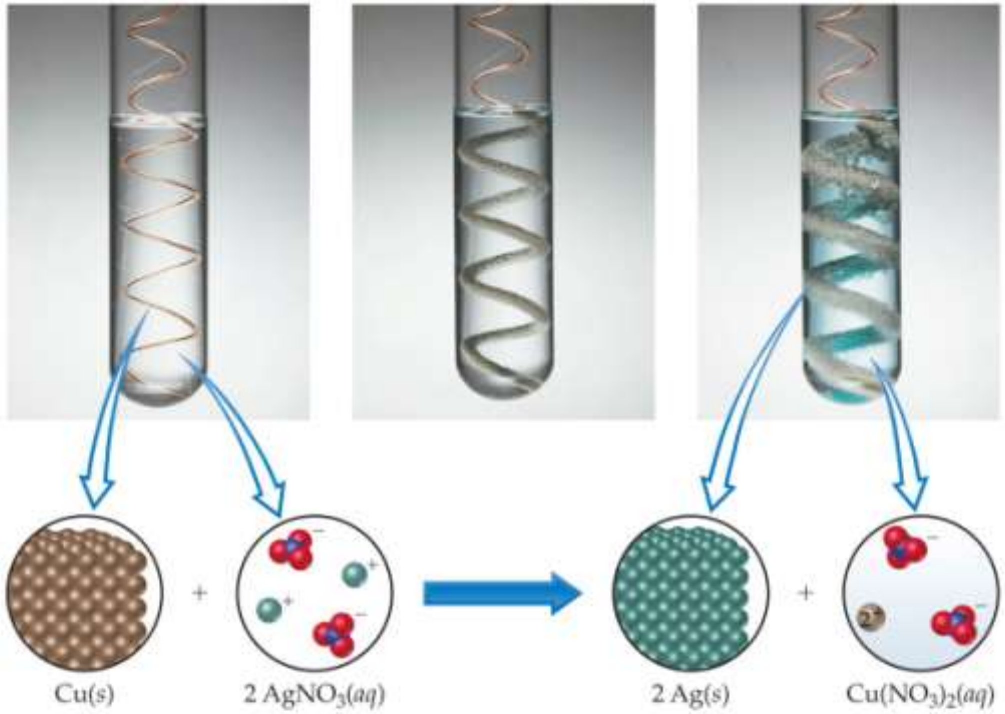
To find the empirical formula, we must compare the relative number of moles of each element in the sample. The relative number of moles of each element is found by dividing each number by the smallest number, 0.0043. The mole ratio of C:H:O so obtained is 2.98:7.91:1.00. The first two numbers are very close to the whole numbers 3 and 8, giving the empirical formula C_3H_8O.

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_2 and 0.209 g H_2O. 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_3H_6O, **(b)** $C_6H_{12}O_2$

Elemental Analyses

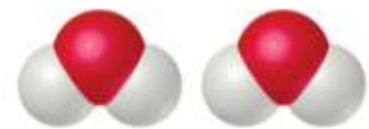


Compounds containing other elements are analyzed using methods analogous to those used for C, H and O.

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:	$2\,H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2\,H_2O(l)$
Molecules:	2 molecules H_2	+	1 molecule O_2	$\longrightarrow$	2 molecules H_2O
					
Mass (amu):	4.0 amu H_2	+	32.0 amu O_2	$\longrightarrow$	36.0 amu H_2O
Amount (mol):	2 mol H_2	+	1 mol O_2	$\longrightarrow$	2 mol H_2O
Mass (g):	4.0 g H_2	+	32.0 g O_2	$\longrightarrow$	36.0 g H_2O

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

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(l)$$

$$2\ \text{mol}\ H_2 \simeq 1\ \text{mol}\ O_2 \simeq 2\ \text{mol}\ H_2O$$

The quantities 2 mole H_2, 1 mole O_2 and 2 mole H_2O 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 H_2O produced from 1.57 mole of O_2 can be calculated as follows:

$$\text{Moles } H_2O = (1.57\ \cancel{\text{mol } O_2})\left(\frac{2\ \text{mol } H_2O}{1\ \cancel{\text{mol } O_2}}\right) = 3.14\ \text{mol } H_2O$$

When 1.57 mol O_2 reacts with H_2 to form H_2O, how many moles of H_2 are consumed in the process?

$$2\,C_4H_{10}(l) + 13\,O_2(g) \longrightarrow 8\,CO_2(g) + 10\,H_2O(g)$$

Calculate the mass of CO_2 and O_2 produced when 1.00 g of C_4H_{10} is burned?

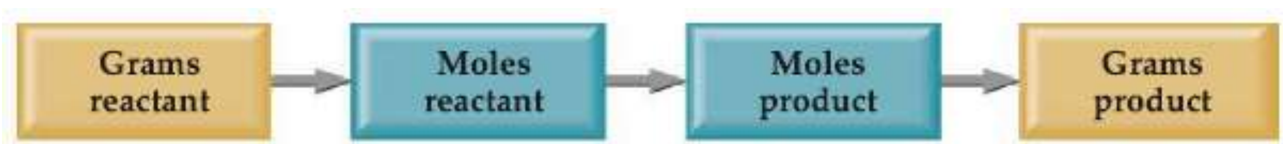
$2\text{ mol } C_4H_{10} \mathrel{\hat{=}} 8\text{ mol } CO_2$ $\qquad$ $2\text{ mol } C_4H_{10} \mathrel{\hat{=}} 13\text{ mol } O_2$

$$\text{Moles } C_4H_{10} = (1.00\ \cancel{\text{g } C_4H_{10}})\left(\frac{1\text{ mol } C_4H_{10}}{58.0\ \cancel{\text{g } C_4H_{10}}}\right) = 1.72 \times 10^{-2}\text{ mol } C_4H_{10}$$

$$\text{Moles } CO_2 = (1.72 \times 10^{-2}\ \cancel{\text{mol } C_4H_{10}})\left(\frac{8\text{ mol } CO_2}{2\ \cancel{\text{mol } C_4H_{10}}}\right) = 6.88 \times 10^{-2}\text{ mol } CO_2$$

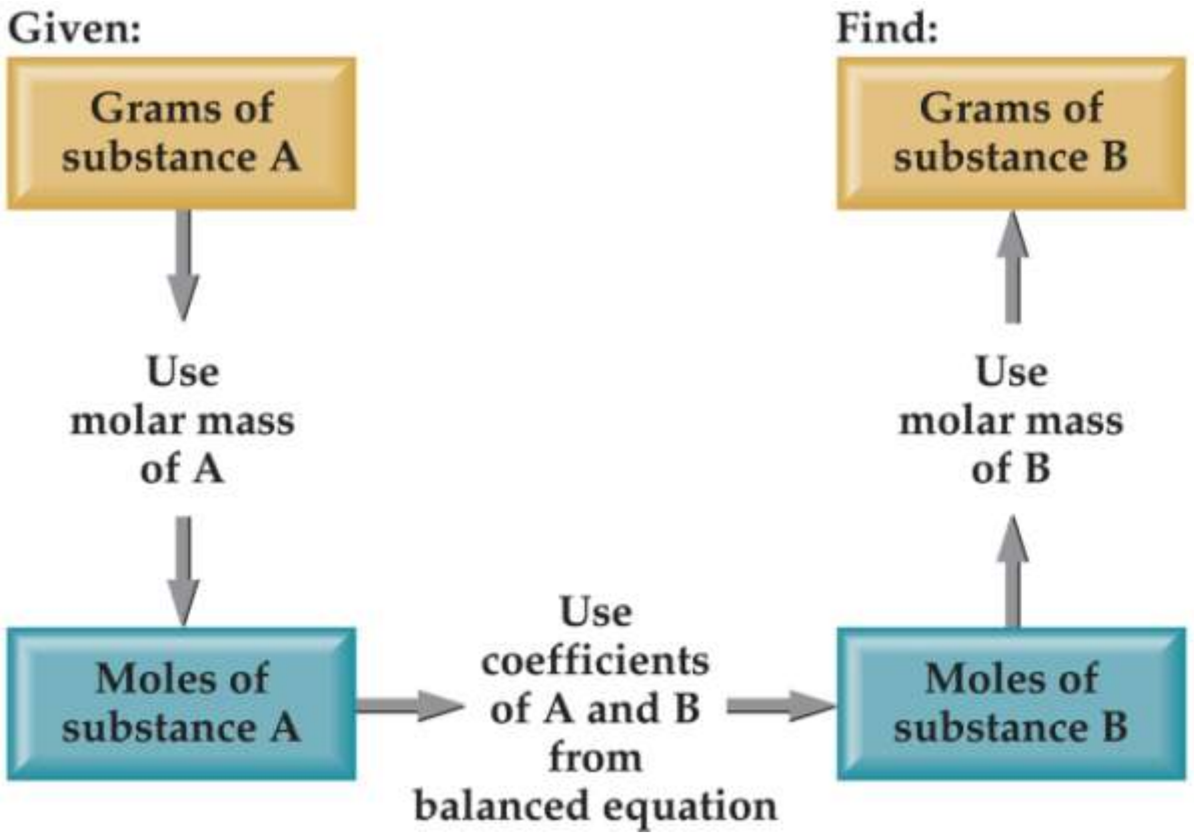
$$\text{Grams } CO_2 = (6.88 \times 10^{-2}\ \cancel{\text{mol } CO_2})\left(\frac{44.0\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right)$$ **= 3.03 g CO_2**

$$\text{Grams } CO_2 = (1.00\ \cancel{\text{g } C_4H_{10}})\left(\frac{1\ \cancel{\text{mol } C_4H_{10}}}{58.0\ \cancel{\text{g } C_4H_{10}}}\right)\left(\frac{8\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_4H_{10}}}\right)\left(\frac{44.0\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right) = 3.03\text{ g } CO_2$$



$$\text{Grams } O_2 = (1.00\ \cancel{\text{g } C_4H_{10}})\left(\frac{1\ \cancel{\text{mol } C_4H_{10}}}{58.0\ \cancel{\text{g } C_4H_{10}}}\right)\left(\frac{13\ \cancel{\text{mol } O_2}}{2\ \cancel{\text{mol } C_4H_{10}}}\right)\left(\frac{32.0\text{ g } O_2}{1\ \cancel{\text{mol } O_2}}\right) = 3.59\text{ g } O_2$$

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, $C_6H_{12}O_6$?

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

Solution

First, use the molar mass of $C_6H_{12}O_6$ to convert from grams $C_6H_{12}O_6$ to moles $C_6H_{12}O_6$:
Second, use the balanced equation to convert moles of $C_6H_{12}O_6$ to moles of H_2O:
Third, use the molar mass of H_2O to convert from moles of H_2O to grams of H_2O:
The steps can be summarized in a diagram like that in Figure 3.13:

$$\text{Moles } C_6H_{12}O_6 = (1.00\ \text{g}\ \cancel{C_6H_{12}O_6})\left(\frac{1\ \text{mol}\ C_6H_{12}O_6}{180.0\ \text{g}\ \cancel{C_6H_{12}O_6}}\right)$$

$$\text{Moles } H_2O = (1.00\ \text{g}\ \cancel{C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}{180.0\ \text{g}\ \cancel{C_6H_{12}O_6}}\right)\left(\frac{6\ \text{mol}\ H_2O}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}\right)$$

$$\text{Grams } H_2O = (1.00\ \text{g}\ \cancel{C_6H_{12}O_6})\left(\frac{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}{180.0\ \text{g}\ \cancel{C_6H_{12}O_6}}\right)\left(\frac{6\ \cancel{\text{mol}\ H_2O}}{1\ \cancel{\text{mol}\ C_6H_{12}O_6}}\right)\left(\frac{18.0\ \text{g}\ H_2O}{1\ \cancel{\text{mol}\ H_2O}}\right)$$
$$= 0.600\ \text{g}\ H_2O$$

$$\begin{array}{ccccc}
\boxed{1.00\ \text{g}\ C_6H_{12}O_6} & \xrightarrow[\text{calculation}]{\text{no direct}} & & & \boxed{0.600\ \text{g}\ H_2O} \\
\downarrow & & & & \uparrow \\
\times\left(\dfrac{1\ \text{mol}\ C_6H_{12}O_6}{180.0\ \text{g}\ C_6H_{12}O_6}\right) & & & & \times\left(\dfrac{18.0\ \text{g}\ H_2O}{1\ \text{mol}\ H_2O}\right) \\
\downarrow & & & & \uparrow \\
\boxed{5.56\times10^{-3}\ \text{mol}\ C_6H_{12}O_6} & \longrightarrow & \times\left(\dfrac{6\ \text{mol}\ H_2O}{1\ \text{mol}\ C_6H_{12}O_6}\right) & \longrightarrow & \boxed{3.33\times10^{-2}\ \text{mol}\ H_2O}
\end{array}$$

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:

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

We are given the grams of LiOH and asked to calculate grams of CO_2. We can accomplish this task by using the following sequence of conversions:

$$\text{Grams LiOH} \rightarrow \text{moles LiOH} \rightarrow \text{moles } CO_2 \rightarrow \text{grams CO}$$

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 LiOH (6.94 + 16.00 + 1.01 = 23.95 g/mol) moles of CO_2 is based on the balanced chemical equation: 2 mol LiOH   1 mol CO_2. To convert the number of moles of CO_2 to grams, we must use the molar mass of CO_2: 12.01 + 2(16.00) = 44.01 g/mol.

Solve

$$(1.00\ \cancel{\text{g LiOH}})\left(\frac{1\ \cancel{\text{mol LiOH}}}{23.95\ \cancel{\text{g LiOH}}}\right)\left(\frac{1\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol LiOH}}}\right)\left(\frac{44.01\ \text{g } CO_2}{1\ \cancel{\text{mol } CO_2}}\right) = 0.919\ \text{g } CO_2$$

Practice Exercise

The decomposition of $KClO_3$ is commonly used to prepare small amounts of O_2 in the laboratory:

$$\mathbf{2\ KClO_3(\mathit{s}) \rightarrow 2\ KCl(\mathit{s}) + 3\ O_2(\mathit{g})}$$

How many grams of O_2 can be prepared from 4.50 g of $KClO_3$?
Answer: 1.77 g

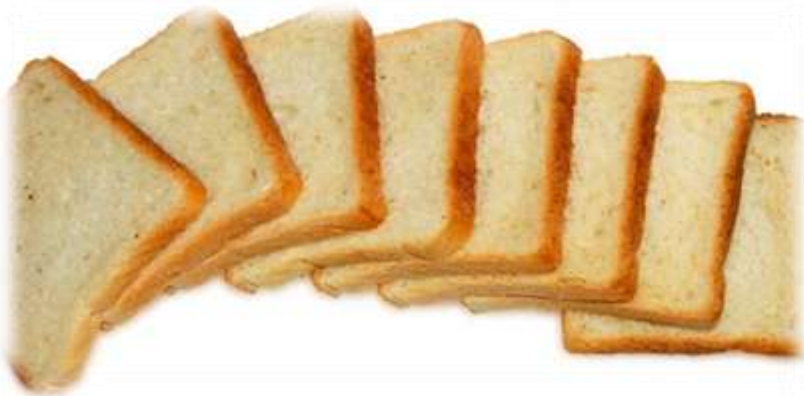
Practice Exercise

Propane, C_3H_8, is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?
Answer: 3.64 g

3.7
Limiting Reactants

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.

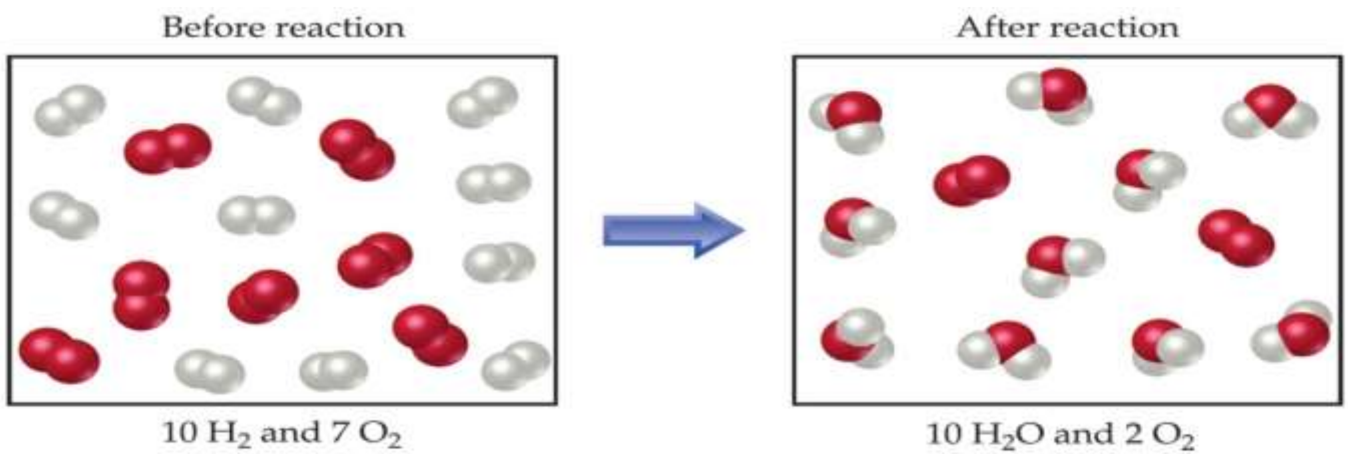
An analogous situation occurs in chemical reactions when one of the reactants is used up before the others. The reaction stops as soon as any one of the reactants is totally consumed, leaving the excess reactants as leftovers.

$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g) \qquad\qquad 2\text{ mol } H_2 \simeq 1\text{ mol } O_2$$

Suppose, for example, mixture of 10 mole H_2 and 7 mole O_2 react to form water.

The number of O_2 needed to react with all the H_2 is:

$$\text{Moles } O_2 = (10\ \cancel{\text{mol } H_2})\left(\frac{1\text{ mol } O_2}{2\ \cancel{\text{mol } H_2}}\right) = 5\text{ mol } O_2$$



In this example, H_2 would be the **limiting reactant**, which means that once all the H_2 has been consumed the reaction stops. And O_2 would be the **excess reactant**, and some is left over when the reaction stops.

Summary of the example data.

	$2\,H_2(g)$	+	$O_2(g)$	$\longrightarrow$	$2\,H_2O(g)$
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 regent 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 N_2 from the air into nitrogen-containing compounds is based on the reaction of N_2 and H_2 to form ammonia (NH_3):

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

How many moles of NH_3 can be formed from 3.0 mol of N_2 and 6.0 mol of H_2?

Solution

Solve The number of moles of H_2 needed for complete consumption of 3.0 mol of N_2 is:

$$\text{Moles } H_2 = (3.0\ \cancel{\text{mol } N_2})\left(\frac{3\ \text{mol } H_2}{1\ \cancel{\text{mol } N_2}}\right) = 9.0\ \text{mol } H_2$$

Because only 6.0 mol H_2 is available, we will run out of H_2 before the N_2 is gone, and H_2 will be the limiting reactant. We use the quantity of the limiting reactant, H_2, to calculate the quantity of NH_3 produced:

$$\text{Moles } NH_3 = (6.0\ \cancel{\text{mol } H_2})\left(\frac{2\ \text{mol } NH_3}{3\ \cancel{\text{mol } H_2}}\right) = 4.0\ \text{mol } NH_3$$

	$N_2(g)$	+	$3\ H_2(g)$	$\longrightarrow$	$2\ NH_3(g)$
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:

$$2\ H_2(g) + O_2\ (g) \rightarrow 2\ H_2O\ (g)$$

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

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

$$2\ \text{mol}\ H_2 \simeq 1\ \text{mol}\ O_2 \simeq 2\ \text{mol}\ H_2O$$

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

$$\text{Moles}\ H_2 = (150\ \cancel{\text{g}\ H_2})\left(\frac{1\ \text{mol}\ H_2}{2.00\ \cancel{\text{g}\ H_2}}\right) = 75\ \text{mol}\ H_2$$

$$\text{Moles}\ O_2 = (1500\ \cancel{\text{g}\ O_2})\left(\frac{1\ \text{mol}\ O_2}{32.0\ \cancel{\text{g}\ O_2}}\right) = 47\ \text{mol}\ O_2$$

Thus, there are more moles of H_2 than O_2. The coefficients in the balanced equation indicate, however, that the reaction requires 2 moles of H_2 for every 1 mole of O_2. Therefore, to completely react all the O_2, we would need $2 \times 47 = 94$ moles of H_2. Since there are only 75 moles of H_2, H_2 is the limiting reagent. We therefore use the quantity of H_2 to calculate the quantity of product formed. We can begin this calculation with the grams of H_2, but we can save a step by starting with the moles of H_2 that were calculated previously in the exercise:

$$\text{Grams}\ H_2O = (75\ \cancel{\text{moles}\ H_2})\left(\frac{2\ \cancel{\text{mol}\ H_2O}}{2\ \cancel{\text{mol}\ H_2}}\right)\left(\frac{18.0\ \text{g}\ H_2O}{1\ \cancel{\text{mol}\ H_2O}}\right)$$

$$= 1400\ \text{g}\ H_2O\ \text{(to two significant figures)}$$

Practice Exercise

Consider the reaction:

$$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$$

A mixture of 1.50 mol of Al and 3.00 mol of Cl_2 is allowed to react. **(a)** Which is the limiting reactant? **(b)** How many moles of $AlCl_3$ 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 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:

$$Zn(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Zn(NO_3)_2(aq)$$

(a) Which reactant is limiting? **(b)** How many grams of Ag will form? **(c)** How many grams of $Zn(NO_3)_2$ will form? **(d)** How many grams of the excess reactant will be left at the end of the reaction?
Answers: **(a)** $AgNO_3$, **(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, $H_2C_6H_8O_4$, is used to produce nylon. The acid is made commercially by a controlled reaction between cyclohexane (C_6H_{12}) and O_2:

$$2\ C_6H_{12}(l) + 5\ O_2(g) \rightarrow 2\ H_2C_6H_8O_4(l) + 2\ H_2O(g)$$

(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

(a) $\text{Grams } H_2C_6H_8O_4 = (25.0\ \cancel{g\ C_6H_{12}})\left(\frac{1\ \cancel{mol\ C_6H_{12}}}{84.0\ \cancel{g\ C_6H_{12}}}\right)\left(\frac{2\ \cancel{mol\ H_2C_6H_8O_4}}{2\ \cancel{mol\ C_6H_{12}}}\right)\left(\frac{146.0\ g\ H_2C_6H_8O_4}{1\ \cancel{mol\ H_2C_6H_8O_4}}\right)$

$= 43.5\ g\ H_2C_6H_8O_4$

(b) $\text{Percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% = \frac{33.5\ g}{43.5\ g} \times 100\% = 77.0\%$

Practice Exercise

Imagine that you are working on ways to improve the process by which iron ore containing Fe_2O_3 is converted into iron. In your tests you carry out the following reaction on a small scale:

$$\mathbf{Fe_2O_3(\mathit{s}) + 3\ CO(\mathit{g}) \rightarrow 2\ Fe(\mathit{s}) + 3\ CO_2(\mathit{g})}$$

(a) If you start with 150 g of Fe_2O_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 $\mathbf{Na_3PO_4}$ is:

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

The percentage by mass of phosphorus in $\mathbf{Na_3PO_4}$ 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. $C_2H_5O_2$
b. C_2H_6O
c. $C_2H_6O_2$
d. $C_3H_4O_2$

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 CO_2 and 2.26 mg of water. What is methyl methacrylate's molecular formula?

a. C_7H_{16}
b. $C_6H_{12}O$
c. $C_5H_8O_2$
d. $C_4H_4O_3$

$$2\,Fe + 3\,Cl_2 \rightarrow 2\,FeCl_3$$

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

$$C_3H_4O_4 + 2\,C_2H_6O \rightarrow C_7H_{12}O_4 + 2\,H_2O$$

When 15.0 grams of each reactant were mixed together, the yield of $C_7H_{12}O_4$ 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 $\mathbf{MgSO_4 \cdot 7\,H_2O}$?

- 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 $\mathbf{Al_2(SO_4)_3}$?

- 1 sulfur atom
- 3 sulfur atoms
- 4 sulfur atoms
- 6.0×10^{23} sulfur atoms
- 1.8×10^{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

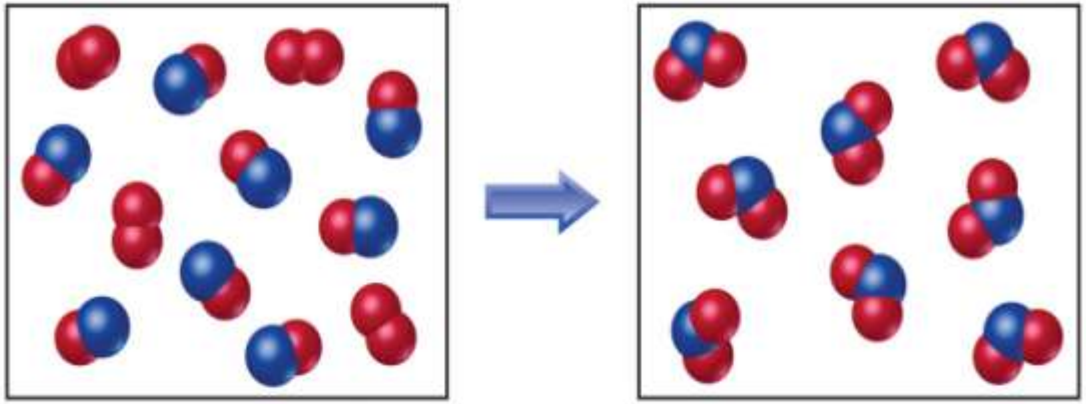

An alkali metal


Ca in H_2O

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

$$2\,NO\,(g) + O_2\,(g) \longrightarrow 2\,NO_2\,(g)$$

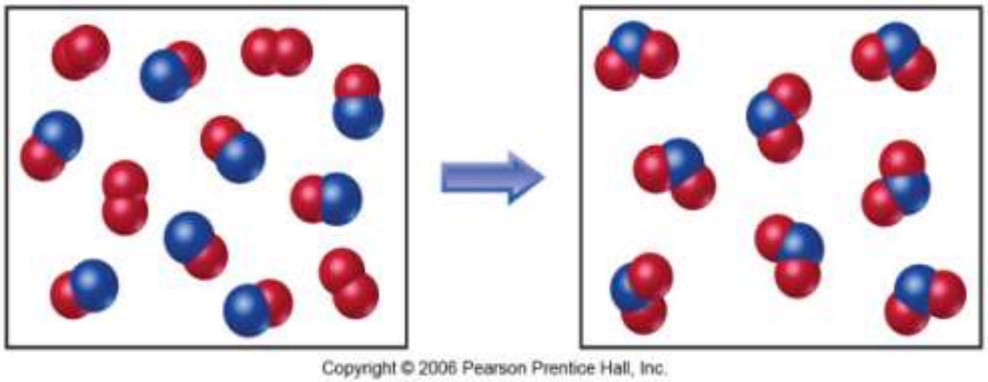
1. 0.5 mol O_2
2. 1.0 mol O_2
3. 1.5 mol O_2
4. 2.0 mol O_2
5. 2.5 mol O_2



If 10.0 moles of NO are reacted with 6.0 moles O_2, how many moles NO_2 are produced?

$$2\,NO\,(g) \quad + \quad O_2\,(g) \longrightarrow 2\,NO_2\,(g)$$

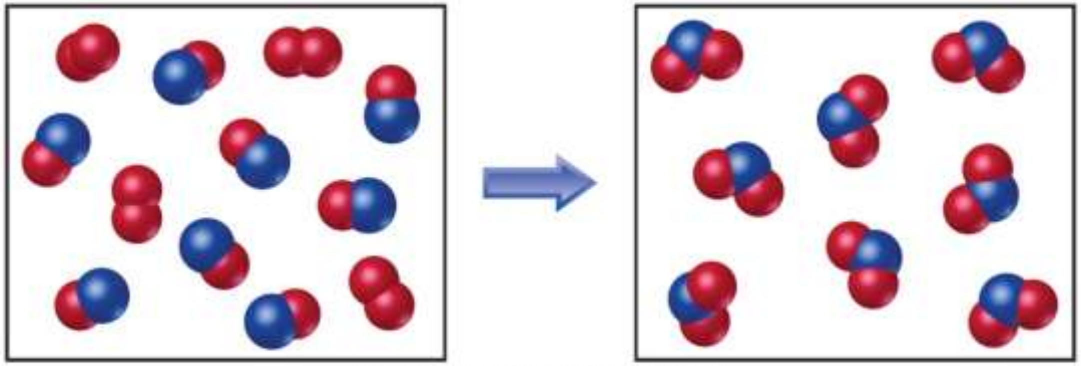
1. 2.0 mol NO_2
2. 6.0 mol NO_2
3. 10.0 mol NO_2
4. 16.0 mol NO_2
5. 32.0 mol NO_2



If 10.0 moles of CO are reacted with 6.0 moles O_2, how many moles of the excess reagent remain?

$$2\ NO\ (g) \quad + \quad O_2\ (g) \longrightarrow 2\ NO_2\ (g)$$

1. 1.0 mol O_2
2. 5.0 mol O_2
3. 4.0 mol NO
4. 8.0 mol NO
5. None of the above




Thank You!