Chapter 3

Stoichiometry of Formulas and Equations
Mole - Mass Relationships in Chemical Systems

3.1 The Mole

3.2 Determining the Formula of an Unknown Compound

3.3 Writing and Balancing Chemical Equations

3.4 Calculating Quantities of Reactant and Product

3.5 Fundamentals of Solution Stoichiometry
The Mole

The mole (mol) is the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12.

The term “entities” refers to atoms, ions, molecules, formula units, or electrons – in fact, any type of particle.

One mole (1 mol) contains $6.022 \times 10^{23}$ entities (to four significant figures).

This number is called Avogadro’s number and is abbreviated as $N$. 
Figure 3.1  One mole (6.022x10^{23} entities) of some familiar substances.
The molar mass ($M$) of a substance is the mass per mole of its entities (atoms, molecules or formula units).

For monatomic elements, the molar mass is the same as the atomic mass in grams per mole. The atomic mass is simply read from the Periodic Table.

The molar mass of Ne = 20.18 g/mol.
For **molecular elements** and for **compounds**, the formula is needed to determine the molar mass.

The molar mass of \( \text{O}_2 \) = \( 2 \times M \text{ of O} \)
= \( 2 \times 16.00 \)
= \( 32.00 \text{ g/mol} \)

The molar mass of \( \text{SO}_2 \) = \( 1 \times M \text{ of S} + 2 \times M \text{ of O} \)
= \( 32.00 + 2(16.00) \)
= \( 64.00 \text{ g/mol} \)
## Table 3.1 Information Contained in the Chemical Formula of Glucose $\text{C}_6\text{H}_{12}\text{O}_6$ (M = 180.16 g/mol)

<table>
<thead>
<tr>
<th></th>
<th>Carbon (C)</th>
<th>Hydrogen (H)</th>
<th>Oxygen (O)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Atoms/molecule of</td>
<td>6 atoms</td>
<td>12 atoms</td>
<td>6 atoms</td>
</tr>
<tr>
<td>compound</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Moles of atoms/mole</td>
<td>6 mol of atoms</td>
<td>12 mol of atoms</td>
<td>6 mol of atoms</td>
</tr>
<tr>
<td>of compound</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Atoms/mole of</td>
<td>6(6.022x10^{23}) atoms</td>
<td>12(6.022x10^{23}) atoms</td>
<td>6(6.022x10^{23}) atoms</td>
</tr>
<tr>
<td>compound</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass/molecule of</td>
<td>6(12.01 amu) = 72.06 amu</td>
<td>12(1.008 amu) = 12.10 amu</td>
<td>6(16.00 amu) = 96.00 amu</td>
</tr>
<tr>
<td>compound</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass/mole of</td>
<td>72.06 g</td>
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<td>96.00 g</td>
</tr>
<tr>
<td>compound</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Interconverting Moles, Mass, and Number of Chemical Entities

Mass (g) = no. of moles \times \frac{\text{no. of grams}}{1 \text{ mol}}

No. of moles = \frac{\text{mass (g)}}{\text{no. of grams}}

No. of entities = \frac{\text{no. of moles}}{6.022 \times 10^{23} \text{ entities}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ entities}}

No. of moles = \frac{\text{no. of entities}}{6.022 \times 10^{23} \text{ entities}} \times \frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ entities}}
Figure 3.2  Mass-mole-number relationships for elements.
Sample Problem 3.1  Calculating the Mass of a Given Amount of an Element

PROBLEM: Silver (Ag) is used in jewelry and tableware but no longer in U.S. coins. How many grams of Ag are in 0.0342 mol of Ag?

PLAN: To convert mol of Ag to mass of Ag in g we need the molar mass of Ag.

SOLUTION: 

\[
0.0342 \text{ mol Ag} \times \frac{107.9 \text{ g Ag}}{1 \text{ mol Ag}} = 3.69 \text{ g Ag}
\]
Sample Problem 3.2 Calculating the Number of Entities in a Given Amount of an Element

**PROBLEM:** Gallium (Ga) is a key element in solar panels, calculators and other light-sensitive electronic devices. How many Ga atoms are in 2.85 x 10^{-3} mol of gallium?

**PLAN:** To convert mol of Ga to number of Ga atoms we need to use Avogadro’s number.

mol of Ga

\[ \text{multiply by } 6.022 \times 10^{23} \text{ atoms/mol} \]

atoms of Ga
Sample Problem 3.2

SOLUTION:

\[
2.85 \times 10^{-3} \text{ mol Ga atoms} \times \frac{6.022 \times 10^{23} \text{ Ga atoms}}{1 \text{ mol Ga atoms}} = 1.72 \times 10^{21} \text{ Ga atoms}
\]
PROBLEM: Iron (Fe) is the main component of steel and is therefore the most important metal in society; it is also essential in the body. How many Fe atoms are in 95.8 g of Fe?

PLAN: The number of atoms cannot be calculated directly from the mass. We must first determine the number of moles of Fe atoms in the sample and then use Avogadro’s number.

mass (g) of Fe → divide by $M$ of Fe (55.85 g/mol) → amount (mol) of Fe → multiply by $6.022 \times 10^{23}$ atoms/mol → atoms of Fe
Sample Problem 3.3

SOLUTION:

\[
\frac{95.8 \text{ g Fe}}{55.85 \text{ g Fe}} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 1.72 \text{ mol Fe}
\]

\[
1.72 \text{ mol Fe} \times \frac{6.022 \times 10^{23} \text{ atoms Fe}}{1 \text{ mol Fe}} = 1.04 \times 10^{24} \text{ atoms Fe}
\]
Figure 3.3  Amount-mass-number relationships for compounds.

MASS (g) of compound

\[ M \text{ (g/mol)} \]

AMOUNT (mol) of compound

chemical formula

AMOUNT (mol) of elements in compound

Avogadro’s number (molecules/mol)

MOLECULES (or formula units) of compound
Sample Problem 3.4 Calculating the Number of Chemical Entities in a Given Mass of a Compound I

PROBLEM: Nitrogen dioxide is a component of urban smog that forms from the gases in car exhausts. How many molecules are in 8.92 g of nitrogen dioxide?

PLAN: Write the formula for the compound and calculate its molar mass. Use the given mass to calculate first the number of moles and then the number of molecules.

\[
\text{mass (g) of NO}_2 \quad \text{divide by } M \quad \text{amount (mol) of NO}_2 \quad \text{multiply by } 6.022 \times 10^{23} \text{ formula units/mol} \quad \text{number of NO}_2 \text{ molecules}
\]
Sample Problem 3.4

**SOLUTION:** NO$_2$ is the formula for nitrogen dioxide.

$$ \mathcal{M} = (1 \times M \text{ of N}) + (2 \times M \text{ of O}) $$

$$ = 14.01 \text{ g/mol} + 2(16.00 \text{ g/mol}) $$

$$ = 46.01 \text{ g/mol} $$

$$ 8.92 \text{ g NO}_2 \times \frac{1 \text{ mol NO}_2}{46.01 \text{ g NO}_2} = 0.194 \text{ mol NO}_2 $$

$$ 0.194 \text{ mol NO}_2 \times \frac{6.022 \times 10^{23} \text{ molecules NO}_2}{1 \text{ mol NO}_2} = 1.17 \times 10^{23} \text{ molecules NO}_2 $$
Sample Problem 3.5  Calculating the Number of Chemical Entities in a Given Mass of a Compound II

PROBLEM: Ammonium carbonate, a white solid that decomposes on warming, is a component of baking powder.

a) How many formula units are in 41.6 g of ammonium carbonate?

b) How many O atoms are in this sample?

PLAN:
Write the formula for the compound and calculate its molar mass. Use the given mass to calculate first the number of moles and then the number of formula units.

The number of O atoms can be determined using the formula and the number of formula units.
Sample Problem 3.5

\[
\text{mass (g) of (NH}_4\text{)}_2\text{CO}_3 \quad \text{divide by } M \\
\text{amount (mol) of (NH}_4\text{)}_2\text{CO}_3 \\
\text{multiply by } 6.022 \times 10^{23} \text{ formula units/mol} \\
\text{number of (NH}_4\text{)}_2\text{CO}_3 \text{ formula units} \\
3 \text{ O atoms per formula unit of (NH}_4\text{)}_2\text{CO}_3 \\
\text{number of O atoms}
\]

**SOLUTION:** (NH\textsubscript{4})\textsubscript{2}CO\textsubscript{3} is the formula for ammonium carbonate.

\[
M = (2 \times M \text{ of N}) + (8 \times M \text{ of H}) + (1 \times M \text{ of C}) + (3 \times M \text{ of O}) \\
= (2 \times 14.01 \text{ g/mol}) + (8 \times 1.008 \text{ g/mol}) \\
+ (12.01 \text{ g/mol}) + (3 \times 16.00 \text{ g/mol}) \\
= 96.09 \text{ g/mol}
\]
Sample Problem 3.5

41.6 g \((\text{NH}_4)_2\text{CO}_3\) x \(\frac{1 \text{ mol} (\text{NH}_4)_2\text{CO}_3}{96.09 \text{ g} (\text{NH}_4)_2\text{CO}_3}\) = 0.433 mol \((\text{NH}_4)_2\text{CO}_3\)

0.433 mol \((\text{NH}_4)_2\text{CO}_3\) x \(\frac{6.022 \times 10^{23} \text{ formula units} (\text{NH}_4)_2\text{CO}_3}{1 \text{ mol} (\text{NH}_4)_2\text{CO}_3}\) = 2.61\times 10^{23} \text{ formula units (NH}_4)_2\text{CO}_3\)

2.61\times 10^{23} \text{ formula units (NH}_4)_2\text{CO}_3\) x \(\frac{3 \text{ O atoms}}{1 \text{ formula unit of (NH}_4)_2\text{CO}_3}\) = 7.83 x 10^{23} \text{ O atoms}
Mass Percent from the Chemical Formula

Mass % of element X =

\[
\frac{\text{atoms of } X \text{ in formula} \times \text{atomic mass of } X \text{ (amu)}}{\text{molecular (or formula) mass of compound (amu)}} \times 100
\]

Mass % of element X =

\[
\frac{\text{moles of } X \text{ in formula} \times \text{molar mass of } X \text{ (g/mol)}}{\text{mass (g) of 1 mol of compound}} \times 100
\]
Sample Problem 3.6  Calculating the Mass Percent of Each Element in a Compound from the Formula

PROBLEM: Glucose (C$_6$H$_{12}$O$_6$) is a key nutrient for generating chemical potential energy in biological systems. What is the mass percent of each element in glucose?

PLAN: Find the molar mass of glucose, which is the mass of 1 mole of glucose. Find the mass of each element in 1 mole of glucose, using the molecular formula.

The mass % for each element is calculated by dividing the mass of that element in 1 mole of glucose by the total mass of 1 mole of glucose, multiplied by 100.
Sample Problem 3.6

PLAN:

1. amount (mol) of element X in 1 mol compound
2. multiply by $M$ (g/mol) of X
3. mass (g) of X in 1 mol of compound
4. divide by mass (g) of 1 mol of compound
5. mass fraction of X
6. multiply by 100
7. mass % X in compound
Sample Problem 3.6

SOLUTION:

In 1 mole of glucose there are 6 moles of C, 12 moles H, and 6 moles O.

\[
\begin{align*}
6 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} &= 72.06 \text{ g C} \\
12 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} &= 12.096 \text{ g H} \\
6 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} &= 96.00 \text{ g O}
\end{align*}
\]

\( M = 180.16 \text{ g/mol} \)

mass percent of C = \( \frac{72.06 \text{ g C}}{180.16 \text{ g glucose}} \times 100 = 0.3999 \times 100 = 39.99 \text{ mass } \% \text{C} \)

mass percent of H = \( \frac{12.096 \text{ g H}}{180.16 \text{ g glucose}} \times 100 = 0.06714 \times 100 = 6.714 \text{ mass } \% \text{H} \)

mass percent of O = \( \frac{96.00 \text{ g O}}{180.16 \text{ g glucose}} \times 100 = 0.5329 \times 100 = 53.29 \text{ mass } \% \text{O} \)
Mass Percent and the Mass of an Element

Mass percent can also be used to calculate the mass of a particular element in any mass of a compound.

\[
\text{Mass of element X present in sample} = \frac{\text{mass of element in 1 mol of compound}}{\text{mass of compound}} \times \frac{\text{mass of compound}}{\text{mass of 1 mol of compound}}
\]
Sample Problem 3.7 Calculating the Mass of an Element in a Compound

PROBLEM: Use the information from Sample Problem 3.6 to determine the mass (g) of carbon in 16.55 g of glucose.

PLAN: The mass percent of carbon in glucose gives us the relative mass of carbon in 1 mole of glucose. We can use this information to find the mass of carbon in any sample of glucose.

\[
\text{mass of glucose sample} \times \text{mass percent of C in glucose} = \text{mass of C in sample}
\]
Sample Problem 3.7

SOLUTION:

Each mol of glucose contains 6 mol of C, or 72.06 g of C.

\[
\text{Mass (g) of C} = \text{mass (g) of glucose} \times \frac{6 \text{ mol} \times \text{M of C (g/mol)}}{\text{mass (g) of 1 mol of glucose}}
\]

\[
= \frac{16.55 \text{ g glucose}}{180.16 \text{ g glucose}} \times 72.06 \text{ g C} = 6.620 \text{ g C}
\]
Empirical and Molecular Formulas

The empirical formula is the simplest formula for a compound that agrees with the elemental analysis. It shows the lowest whole number of moles and gives the relative number of atoms of each element present.

The empirical formula for hydrogen peroxide is HO.

The molecular formula shows the actual number of atoms of each element in a molecule of the compound.

The molecular formula for hydrogen peroxide is H₂O₂.
Sample Problem 3.8  Determining an Empirical Formula from Amounts of Elements

PROBLEM:  A sample of an unknown compound contains 0.21 mol of zinc, 0.14 mol of phosphorus, and 0.56 mol of oxygen. What is its empirical formula?

PLAN:  Find the relative number of moles of each element. Divide by the lowest mol amount to find the relative mol ratios (empirical formula).

- amount (mol) of each element
- use # of moles as subscripts
- preliminary formula
- change to integer subscripts
- empirical formula
Sample Problem 3.8

**SOLUTION:** Using the numbers of moles of each element given, we write the preliminary formula $\text{Zn}_{0.21}\text{P}_{0.14}\text{O}_{0.56}$

Next we divide each fraction by the smallest one; in this case 0.14:

\[
\frac{0.21}{0.14} = 1.5 \quad \frac{0.14}{0.14} = 1.0 \quad \frac{0.56}{0.14} = 4.0
\]

This gives $\text{Zn}_{1.5}\text{P}_{1.0}\text{O}_{4.0}$

We convert to whole numbers by multiplying by the *smallest integer* that gives whole numbers; in this case 2:

\[
1.5 \times 2 = 3 \quad 1.0 \times 2 = 2 \quad 4.0 \times 2 = 8
\]

This gives us the empirical formula $\text{Zn}_3\text{P}_2\text{O}_8$
Sample Problem 3.9  Determining an Empirical Formula from Masses of Elements

PROBLEM:  Analysis of a sample of an ionic compound yields 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What is the empirical formula and the name of the compound?

PLAN:  Find the relative number of moles of each element. Divide by the lowest mol amount to find the relative mol ratios (empirical formula).

- mass (g) of each element
- divide by \( M \) (g/mol)
- amount (mol) of each element
- use # of moles as subscripts
- preliminary formula
- change to integer subscripts
- empirical formula
SOLUTION:  
\[
2.82 \text{ g Na} \times \frac{1 \text{ mol Na}}{22.99 \text{ g Na}} = 0.123 \text{ mol Na}
\]
\[
4.35 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.123 \text{ mol Cl}
\]
\[
7.83 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.489 \text{ mol O}
\]

\[
\text{Na and Cl} = \frac{0.123}{0.123} = 1 \quad \text{and} \quad \text{O} = \frac{0.489}{0.123} = 3.98
\]

The empirical formula is \( \text{Na}_1\text{Cl}_1\text{O}_{3.98} \) or \( \text{NaClO}_4 \); this compound is named sodium perchlorate.
Determining the Molecular Formula

The molecular formula gives the *actual* numbers of moles of each element present in 1 mol of compound.

The molecular formula is a *whole-number multiple* of the empirical formula.

\[
\frac{\text{molar mass (g/mol)}}{\text{empirical formula mass (g/mol)}} = \text{whole-number multiple}
\]
Sample Problem 3.10 Determining a Molecular Formula from Elemental Analysis and Molar Mass

**PROBLEM:** Elemental analysis of lactic acid \((M = 90.08 \text{ g/mol})\) shows it contains 40.0 mass % C, 6.71 mass % H, and 53.3 mass % O. Determine the empirical formula and the molecular formula for lactic acid.

**PLAN:**

1. Assume 100 g lactic acid; then mass % = mass in grams
2. Divide each mass by \(M\)
3. Amount (mol) of each element
4. Use # mols as subscripts; convert to integers
5. Empirical formula
6. Divide \(M\) by the molar mass for the empirical formula; multiply empirical formula by this number
7. Molecular formula
Sample Problem 3.10

SOLUTION: Assuming there are 100. g of lactic acid;

\[
\begin{align*}
40.0 \text{ g} \text{C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} &= 3.33 \text{ mol C} \\
6.71 \text{ g} \text{H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} &= 6.66 \text{ mol H} \\
53.3 \text{ g} \text{O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} &= 3.33 \text{ mol O}
\end{align*}
\]

\[
\begin{align*}
\text{C}_{3.33} &\quad \text{H}_{6.66} &\quad \text{O}_{3.33} \\
\text{CH}_2\text{O} &\quad \text{empirical formula}
\end{align*}
\]

\[
\frac{90.08 \text{ g/mol}}{30.03 \text{ g/mol}} = 3
\]

\[
\begin{align*}
\text{molar mass of lactate} &\quad \text{mass of CH}_2\text{O}
\end{align*}
\]

\[
\begin{align*}
\text{C}_3\text{H}_6\text{O}_3 \text{ is the molecular formula}
\end{align*}
\]
Figure 3.4  Combustion apparatus for determining formulas of organic compounds.

\[ C_nH_m + (n+ \frac{m}{2}) O_2 = n \text{CO}(g) + \frac{m}{2} \text{H}_2\text{O}(g) \]
Sample Problem 3.11 Determining a Molecular Formula from Combustion Analysis

PROBLEM:
When a 1.000 g sample of vitamin C (M = 176.12 g/mol) is placed in a combustion chamber and burned, the following data are obtained:

- mass of CO₂ absorber after combustion = 85.35 g
- mass of CO₂ absorber before combustion = 83.85 g
- mass of H₂O absorber after combustion = 37.96 g
- mass of H₂O absorber before combustion = 37.55 g

What is the molecular formula of vitamin C?

PLAN: The masses of CO₂ and H₂O produced will give us the masses of C and H present in the original sample. From this we can determine the mass of O.
Sample Problem 3.11

(mass after combustion – mass before) for each absorber
= mass of compound in each absorber

mass of each compound × mass % of oxidized element

mass of each oxidized element

mass of vitamin C – (mass of C + H)

mass of O

divide each mass by \( M \)

mol of C, H, and O

use # mols as subscripts; convert to integers

empirical formula

molecular formula
Sample Problem 3.11

SOLUTION: For CO$_2$: 85.35 g - 83.85 g = 1.50 g

\[
1.50 \text{ g CO}_2 \times \frac{12.01 \text{ g C}}{44.01 \text{ g CO}_2} = 0.409 \text{ g C}
\]

For H$_2$O: 37.96 g - 37.55 g = 0.41 g

\[
0.41 \text{ g H}_2\text{O} \times \frac{2.016 \text{ g H}}{18.02 \text{ g H}_2\text{O}} = 0.046 \text{ g H}
\]

Mass of O = mass of vitamin C – (mass of C + mass of H)

\[
= 1.000 \text{ g} - (0.409 + 0.046) \text{ g} = 0.545 \text{ g O}
\]
Sample Problem 3.11

Convert mass to moles:

\[
\frac{0.409 \text{ g C}}{12.01 \text{ g/mol C}} = 0.0341 \text{ mol C} \quad \frac{0.046 \text{ g H}}{1.008 \text{ g/mol H}} = 0.0456 \text{ mol H}
\]

\[
\frac{0.545 \text{ g O}}{16.00 \text{ g/mol O}} = 0.0341 \text{ mol O}
\]

Divide by smallest to get the preliminary formula:

\[
\frac{0.0341}{0.0341} = 1 \quad \frac{0.0456}{0.0341} = 1.34 \quad \frac{0.0341}{0.0341} = 1
\]

\[
C_{1.34}H_{1.34}O_1 = C_3H_{4.01}O_3 \quad \rightarrow \quad C_3H_4O_3
\]

Divide molar mass by mass of empirical formula:

\[
\frac{176.12 \text{ g/mol}}{88.06 \text{ g}} = 2.000 \text{ mol} \quad \rightarrow \quad \boxed{C_6H_8O_6}
\]
Table 3.2  Some Compounds with Empirical Formula CH$_2$O
(Composition by Mass: 40.0% C, 6.71% H, 53.3% O)

<table>
<thead>
<tr>
<th>Name</th>
<th>Molecular Formula</th>
<th>Whole-Number Multiple</th>
<th>$M$  (g/mol)</th>
<th>Use or Function</th>
</tr>
</thead>
<tbody>
<tr>
<td>formaldehyde</td>
<td>CH$_2$O</td>
<td>1</td>
<td>30.03</td>
<td>disinfectant; biological preservative</td>
</tr>
<tr>
<td>acetic acid</td>
<td>C$_2$H$_4$O$_2$</td>
<td>2</td>
<td>60.05</td>
<td>acetate polymers; vinegar (5% soln)</td>
</tr>
<tr>
<td>lactic acid</td>
<td>C$_3$H$_6$O$_3$</td>
<td>3</td>
<td>90.09</td>
<td>sour milk; forms in exercising muscle</td>
</tr>
<tr>
<td>erythrose</td>
<td>C$_4$H$_8$O$_4$</td>
<td>4</td>
<td>120.10</td>
<td>part of sugar metabolism</td>
</tr>
<tr>
<td>ribose</td>
<td>C$<em>5$H$</em>{10}$O$_5$</td>
<td>5</td>
<td>150.13</td>
<td>component of nucleic acids and B$_2$</td>
</tr>
<tr>
<td>glucose</td>
<td>C$<em>6$H$</em>{12}$O$_6$</td>
<td>6</td>
<td>180.16</td>
<td>major energy source of the cell</td>
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</tbody>
</table>
## Table 3.3 Two Pairs of Constitutional Isomers

<table>
<thead>
<tr>
<th>Property</th>
<th>C\textsubscript{4}H\textsubscript{10}</th>
<th>C\textsubscript{2}H\textsubscript{6}O</th>
</tr>
</thead>
<tbody>
<tr>
<td>M (g/mol)</td>
<td>Butane 58.12</td>
<td>Ethanol 46.07</td>
</tr>
<tr>
<td>Boiling Point</td>
<td>2-Methylpropane 58.12</td>
<td>Dimethyl Ether 46.07</td>
</tr>
<tr>
<td></td>
<td>-0.5\textdegree C</td>
<td>78.5\textdegree C</td>
</tr>
<tr>
<td></td>
<td>-11.06\textdegree C</td>
<td>-25\textdegree C</td>
</tr>
<tr>
<td>Density at 20\textdegree C</td>
<td>0.579 g/mL (gas)</td>
<td>0.789 g/mL (liquid)</td>
</tr>
<tr>
<td></td>
<td>0.549 g/mL (gas)</td>
<td>0.00195 g/mL (gas)</td>
</tr>
<tr>
<td>Structural formula</td>
<td></td>
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</tr>
<tr>
<td></td>
<td>Butane</td>
<td>Ethanol</td>
</tr>
<tr>
<td></td>
<td>2-Methylpropane</td>
<td>Dimethyl Ether</td>
</tr>
<tr>
<td>Space-filling model</td>
<td></td>
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</tr>
</tbody>
</table>
Chemical Equations

A chemical equation uses formulas to express the identities and quantities of substances involved in a physical or chemical change.

Figure 3.6
The formation of HF gas on the macroscopic and molecular levels.
Figure 3.7 A three-level view of the reaction between magnesium and oxygen.
Features of Chemical Equations

Reactants are written on the left.

Products are written on the right.

The equation must be balanced; the same number and type of each atom must appear on both sides.

A yield arrow points from reactants to products.

\[
\text{Mg} + \text{O}_2 \rightarrow \text{MgO}
\]
Balancing a Chemical Equation

- translate the statement: magnesium and oxygen gas react to give magnesium oxide: \( \text{Mg} + \text{O}_2 \rightarrow \text{MgO} \)

- balance the atoms using *coefficients*; formulas cannot be changed

  \[ 2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO} \]

- adjust coefficients if necessary

- check that all atoms balance

- specify states of matter

  \[ 2\text{Mg} \ (s) + \text{O}_2 \ (g) \rightarrow 2\text{MgO} \ (s) \]
Sample Problem 3.12

Balancing Chemical Equations

**PROBLEM:** Within the cylinders of a car’s engine, the hydrocarbon octane (C\textsubscript{8}H\textsubscript{18}), one of many components of gasoline, mixes with oxygen from the air and burns to form carbon dioxide and water vapor. Write a balanced equation for this reaction.

**PLAN:**
- Translate the statement
- Balance the atoms
- Adjust the coefficients
- Check the atoms balance
- Specify states of matter

**SOLUTION:**

\[ \text{C}_8\text{H}_{18} + \frac{25}{2} \text{O}_2 \rightarrow 8 \text{CO}_2 + 9 \text{H}_2\text{O} \]

\[ 2\text{C}_8\text{H}_{18} + 25\text{O}_2 \rightarrow 16\text{CO}_2 + 18\text{H}_2\text{O} \]

\[ 2\text{C}_8\text{H}_{18}(l) + 25\text{O}_2(g) \rightarrow 16\text{CO}_2(g) + 18\text{H}_2\text{O}(g) \]
Molecular Scene
Combustion of Octane

\[ 2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g) \]
Sample Problem 3.13  Balancing an Equation from a Molecular Scene

PROBLEM: The following molecular scenes depict an important reaction in nitrogen chemistry. The blue spheres represent nitrogen while the red spheres represent oxygen. Write a balanced equation for this reaction.

PLAN: Determine the formulas of the reactants and products from their composition. Arrange this information in the correct equation format and balance correctly, including the states of matter.
Sample Problem 3.13

SOLUTION:

The reactant circle shows only one type of molecule, composed of 2 N and 5 O atoms. The formula is thus N₂O₅. There are 4 N₂O₅ molecules depicted.

The product circle shows two types of molecule; one has 1 N and 2 O atoms while the other has 2 O atoms. The products are NO₂ and O₂. There are 8 NO₂ molecules and 2 O₂ molecules shown.

The reaction depicted is \( 4\, \text{N}_2\text{O}_5 \rightarrow 8\, \text{NO}_2 + 2\, \text{O}_2 \).

Writing the equation with the smallest whole-number coefficients and states of matter included:

\( 2\, \text{N}_2\text{O}_5 (g) \rightarrow 4\, \text{NO}_2 (g) + \text{O}_2 (g) \)
Stoichiometric Calculations

• The coefficients in a balanced chemical equation
  – represent the relative number of reactant and product particles
  – and the relative number of moles of each.

• Since moles are related to mass
  – the equation can be used to calculate masses of reactants
    and/or products for a given reaction.

• The mole ratios from the balanced equation are used as
  conversion factors.
### Table 3.4 Information Contained in a Balanced Equation

<table>
<thead>
<tr>
<th>Viewed in Terms of</th>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{C}_3\text{H}_8(g) + 5 \text{O}_2(g)$</td>
<td>$3 \text{CO}_2(g) + 4 \text{H}_2\text{O}(g)$</td>
<td></td>
</tr>
</tbody>
</table>

- **Molecules**: $1 \text{ molecule C}_3\text{H}_8 + 5 \text{ molecules O}_2 \rightarrow 3 \text{ molecules CO}_2 + 4 \text{ molecules H}_2\text{O}$

- **Amount (mol)**: $1 \text{ mol C}_3\text{H}_8 + 5 \text{ mol O}_2 \rightarrow 3 \text{ mol CO}_2 + 4 \text{ mol H}_2\text{O}$

- **Mass (amu)**: $44.09 \text{ amu C}_3\text{H}_8 + 160.00 \text{ amu O}_2 \rightarrow 132.03 \text{ amu CO}_2 + 72.06 \text{ amu H}_2\text{O}$

- **Mass (g)**: $44.09 \text{ g C}_3\text{H}_8 + 160.00 \text{ g O}_2 \rightarrow 132.03 \text{ g CO}_2 + 72.06 \text{ g H}_2\text{O}$

- **Total Mass (g)**: $204.09 \text{ g}$
Summary of amount-mass-number relationships in a chemical equation.

Figure 3.8
Sample Problem 3.14 Calculating Quantities of Reactants and Products: Amount (mol) to Amount (mol)

PROBLEM: Copper is obtained from copper(I) sulfide by roasting it in the presence of oxygen gas) to form powdered copper(I) oxide and gaseous sulfur dioxide.

How many moles of oxygen are required to roast 10.0 mol of copper(I) sulfide?

PLAN: write and balance the equation
use the mole ratio as a conversion factor

SOLUTION: 2 \( \text{Cu}_2\text{S} \) (s) + 3 \( \text{O}_2 \) (g) → 2 \( \text{Cu}_2\text{O} \) (s) + 2 \( \text{SO}_2 \) (g)

\[
\frac{10.0 \text{ mol Cu}_2\text{S}}{2 \text{ mol Cu}_2\text{S}} \times \frac{3 \text{ mol O}_2}{2 \text{ mol Cu}_2\text{S}} = 15.0 \text{ mol O}_2
\]
Sample Problem 3.15  Calculating Quantities of Reactants and Products: Amount (mol) to Mass (g)

PROBLEM: During the process of roasting copper(I) sulfide, how many grams of sulfur dioxide form when 10.0 mol of copper(I) sulfide reacts?

PLAN: Using the balanced equation from the previous problem, we again use the mole ratio as a conversion factor.

\[
\text{mol of copper(I) sulfide} \quad \rightarrow \quad \text{mol of sulfur dioxide} \quad \rightarrow \quad \text{mass of sulfur dioxide}
\]

\begin{align*}
\text{use the mole ratio as a conversion factor} \\
\text{multiply by } \mathcal{M} \text{ of sulfur dioxide}
\end{align*}
Sample Problem 3.15

**SOLUTION:** 

\[ 2 \text{Cu}_2\text{S} \ (s) + 3 \text{O}_2 \ (g) \rightarrow 2 \text{Cu}_2\text{O} \ (s) + 2 \text{SO}_2 \ (g) \]

\[
rac{10.0 \text{ mol Cu}_2\text{S}}{2 \text{ mol Cu}_2\text{S}} \times \frac{2 \text{ mol SO}_2}{1 \text{ mol SO}_2} \times 64.07 \text{ g SO}_2 = 641 \text{ g SO}_2
\]
Sample Problem 3.16 Calculating Quantities of Reactants and Products: Mass to Mass

PROBLEM: During the roasting of copper(I) sulfide, how many kilograms of oxygen are required to form 2.86 kg of copper(I) oxide?

PLAN:

1. **mass of oxygen**
   - divide by $M$ of oxygen
2. **mol of oxygen**
3. **mol of copper(I) oxide**
   - use mole ratio as conversion factor
4. **multiply by $M$ of copper(I) oxide**
5. **mass of copper(I) oxide**
Sample Problem 3.16

**SOLUTION:**

\[2 \text{ Cu}_2\text{S} (s) + 3 \text{ O}_2 (g) \rightarrow 2 \text{ Cu}_2\text{O} (s) + 2 \text{ SO}_2 (g)\]

\[
\frac{2.86 \text{ kg Cu}_2\text{O} \times 10^3 \text{ g}}{1 \text{ kg}} \times \frac{1 \text{ mol Cu}_2\text{O}}{143.10 \text{ g Cu}_2\text{O}} = 20.0 \text{ mol Cu}_2\text{O}
\]

\[
\frac{20.0 \text{ mol Cu}_2\text{O} \times 3 \text{ mol O}_2}{2 \text{ mol Cu}_2\text{O}} \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \times \frac{1 \text{ kg}}{10^3 \text{ g}} = 0.959 \text{ kg O}_2
\]
Reactions in Sequence

- Reactions often occur in sequence.
- The product of one reaction becomes a reactant in the next.
- An overall reaction is written by combining the reactions;
  - any substance that forms in one reaction and reacts in the next can be eliminated.
PROBLEM: Roasting is the first step in extracting copper from chalcocite, the ore used in the previous problem. In the next step, copper(I) oxide reacts with powdered carbon to yield copper metal and carbon monoxide gas. Write a balanced overall equation for the two-step process.

PLAN: Write individual balanced equations for each step.
Adjust the coefficients so that any common substances can be canceled.
Add the adjusted equations together to obtain the overall equation.
Sample Problem 3.17

**SOLUTION:** Write individual balanced equations for each step:

\[ 2\text{Cu}_2\text{S} (s) + 3\text{O}_2 (g) \rightarrow 2\text{Cu}_2\text{O} (s) + 2\text{SO}_2 (g) \]

\[ \text{Cu}_2\text{O} (s) + \text{C} (s) \rightarrow 2\text{Cu} (s) + \text{CO} (g) \]

Adjust the coefficients so that the 2 moles of Cu₂O formed in reaction 1 are used up in reaction 2:

\[ 2\text{Cu}_2\text{S} (s) + 3\text{O}_2 (g) \rightarrow 2\text{Cu}_2\text{O} (s) + 2\text{SO}_2 (g) \]

\[ 2\text{Cu}_2\text{O} (s) + 2\text{C} (s) \rightarrow 4\text{Cu} (s) + 2\text{CO} (g) \]

Add the equations together:

\[ 2\text{Cu}_2\text{S} (s) + 3\text{O}_2 (g) + 2\text{C} (s) \rightarrow 2\text{SO}_2 (g) + 4\text{Cu} (s) + 2\text{CO} (g) \]
Limiting Reactants

- So far we have assumed that reactants are present in the correct amounts to react completely.
- In reality, one reactant may limit the amount of product that can form.
- The limiting reactant will be completely used up in the reaction.
- The reactant that is not limiting is in excess – some of this reactant will be left over.
Figure 3.10  An ice cream sundae analogy for limiting reactions.
Sample Problem 3.18  Using Molecular Depictions in a Limiting-Reactant Problem

PROBLEM: Chlorine trifluoride, an extremely reactive substance, is formed as a gas by the reaction of elemental chlorine and fluorine. The molecular scene shows a representative portion of the reaction mixture before the reaction starts. (Chlorine is green, and fluorine is yellow.)

(a) Find the limiting reactant.
(b) Write a reaction table for the process.
(c) Draw a representative portion of the mixture after the reaction is complete. (Hint: The ClF₃ molecule has 1 Cl atom bonded to 3 individual F atoms).
Sample Problem 3.18

**PLAN:** Write a balanced chemical equation. To determine the limiting reactant, find the number of molecules of product that would form from the given numbers of molecules of each reactant. Use these numbers to write a reaction table and use the reaction table to draw the final reaction scene.

**SOLUTION:** The balanced equation is \( \text{Cl}_2 (g) + 3\text{F}_2 (g) \rightarrow 2\text{ClF}_3 (g) \)

There are 3 molecules of \( \text{Cl}_2 \) and 6 molecules of \( \text{F}_2 \) depicted:

\[
\begin{align*}
3 \text{ molecules } \text{Cl}_2 \times \frac{2 \text{ molecules } \text{ClF}_3}{1 \text{ molecule } \text{Cl}_2} &= 6 \text{ molecules } \text{ClF}_3 \\
6 \text{ molecules } \text{F}_2 \times \frac{2 \text{ molecules } \text{ClF}_3}{3 \text{ molecules } \text{Cl}_2} &= 4 \text{ molecules } \text{ClF}_3
\end{align*}
\]

Since the given amount of \( \text{F}_2 \) can form less product, it is the limiting reactant.
Sample Problem 3.18

We use the amount of \( F_2 \) to determine the “change” in the reaction table, since \( F_2 \) is the limiting reactant:

<table>
<thead>
<tr>
<th>Molecules</th>
<th>( \text{Cl}_2 \ (g) )</th>
<th>+</th>
<th>( 3\text{F}_2 \ (g) )</th>
<th>( \rightarrow )</th>
<th>( 2\text{ClF}_3 \ (g) )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>3</td>
<td></td>
<td>6</td>
<td>0</td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>-2</td>
<td></td>
<td>-6</td>
<td>+4</td>
<td></td>
</tr>
<tr>
<td>Final</td>
<td>1</td>
<td></td>
<td>0</td>
<td>4</td>
<td></td>
</tr>
</tbody>
</table>

The final reaction scene shows that all the \( F_2 \) has reacted and that there is \( \text{Cl}_2 \) left over. 4 molecules of \( \text{ClF}_2 \) have formed:
Sample Problem 3.19 Calculating Quantities in a Limiting-Reactant Problem: Amount to Amount

PROBLEM: In another preparation of ClF₃, 0.750 mol of Cl₂ reacts with 3.00 mol of F₂.

(a) Find the limiting reactant.
(b) Write a reaction table.

PLAN: Find the limiting reactant by calculating the amount (mol) of ClF₃ that can be formed from each given amount of reactant. Use this information to construct a reaction table.

SOLUTION: The balanced equation is Cl₂ (g) + 3F₂ (g) → 2ClF₃ (g)

\[
\begin{align*}
0.750 \text{ mol Cl}_2 & \times \frac{2 \text{ mol ClF}_3}{1 \text{ mol Cl}_2} = 1.50 \text{ mol ClF}_3 \\
3.00 \text{ mol F}_2 & \times \frac{2 \text{ mol ClF}_3}{3 \text{ mol F}_2} = 2.00 \text{ mol ClF}_3
\end{align*}
\]

Cl₂ is limiting, because it yields less ClF₃.
All the Cl$_2$ reacts since this is the limiting reactant. For every 1 Cl$_2$ that reacts, 3 F$_2$ will react, so $3(0.750)$ or 2.25 moles of F$_2$ reacts.

<table>
<thead>
<tr>
<th>Moles</th>
<th>Cl$_2$ (g)</th>
<th>+</th>
<th>3F$_2$ (g)</th>
<th>→</th>
<th>2ClF$_3$ (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>0.750</td>
<td>3.00</td>
<td>0</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change</td>
<td>-0.750</td>
<td>-2.25</td>
<td>+1.50</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final</td>
<td>0</td>
<td>0.75</td>
<td>1.50</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
PROBLEM: A fuel mixture used in the early days of rocketry consisted of two liquids, hydrazine \( \text{N}_2\text{H}_4 \) and dinitrogen tetraoxide \( \text{N}_2\text{O}_4 \), which ignite on contact to form nitrogen gas and water vapor.

(a) How many grams of nitrogen gas form when \( 1.00 \times 10^2 \) g of \( \text{N}_2\text{H}_4 \) and \( 2.00 \times 10^2 \) g of \( \text{N}_2\text{O}_4 \) are mixed?

(b) Write a reaction table for this process.

PLAN: Find the limiting reactant by calculating the amount (mol) of \( \text{ClF}_3 \) that can be formed from each given mass of reactant. Use this information to construct a reaction table.
Sample Problem 3.20

- Mass (g) of N₂H₄ divide by \( M \)
  - Mol of N₂H₄
    - Mole ratio
      - Mol of N₂
        - Select lower number of moles of N₂
          - Multiply by \( M \)
            - Mass of N₂

- Mass (g) of N₂O₄ divide by \( M \)
  - Mol of N₂O₄
    - Mole ratio
      - Mol of N₂
        - Mass of N₂
Sample Problem 3.20

**SOLUTION:** \(2\text{N}_2\text{H}_4 (l) + \text{N}_2\text{O} (l) \rightarrow 3\text{N}_2 (g) + 4\text{H}_2\text{O} (g)\)

For \(\text{N}_2\text{H}_4\): \(1.00 \times 10^2 \text{ g N}_2\text{H}_4 \times \frac{1 \text{ mol N}_2\text{H}_4}{32.05 \text{ g N}_2\text{H}_4} = 3.12 \text{ mol N}_2\text{H}_4\)

\(3.12 \text{ mol N}_2\text{H}_4 \times \frac{3 \text{ mol N}_2}{2 \text{ mol N}_2\text{H}_4} = 4.68 \text{ mol N}_2\)

For \(\text{N}_2\text{O}_4\): \(2.00 \times 10^2 \text{ g N}_2\text{O}_4 \times \frac{1 \text{ mol N}_2\text{O}_4}{92.02 \text{ g N}_2\text{O}_4} = 2.17 \text{ mol N}_2\)

\(2.17 \text{ mol N}_2\text{O}_4 \times \frac{3 \text{ mol N}_2}{1 \text{ mol N}_2\text{O}_4} = 6.51 \text{ mol N}_2\)

\(\text{N}_2\text{H}_4\) is limiting and only 4.68 mol of \(\text{N}_2\) can be produced:

\(4.68 \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 131 \text{ g N}_2\)
Sample Problem 3.20

All the N₂H₄ reacts since it is the limiting reactant. For every 2 moles of N₂H₄ that react 1 mol of N₂O₄ reacts and 3 mol of N₂ form:

\[
3.12 \text{ mol N}_2\text{H}_4 \times \frac{1 \text{ mol N}_2\text{O}_4}{2 \text{ mol N}_2\text{H}_4} = 1.56 \text{ mol N}_2\text{O}_4 \text{ reacts}
\]

<table>
<thead>
<tr>
<th>Moles</th>
<th>2N₂H₄ (l) +</th>
<th>N₂O₄ (l) →</th>
<th>3N₂ (g) +</th>
<th>4H₂O (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>3.12</td>
<td>2.17</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-3.12</td>
<td>-1.56</td>
<td>+4.68</td>
<td>+6.24</td>
</tr>
<tr>
<td>Final</td>
<td>0</td>
<td>0.61</td>
<td>4.68</td>
<td>6.24</td>
</tr>
</tbody>
</table>
Reaction Yields

The **theoretical yield** is the amount of product calculated using the molar ratios from the balanced equation.

The **actual yield** is the amount of product actually obtained.

The actual yield is usually less than the theoretical yield.

\[
\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100
\]
Figure 3.11  The effect of side reactions on the yield of the main product.
PROBLEM: Silicon carbide (SiC) is made by reacting sand (silicon dioxide, SiO2) with powdered carbon at high temperature. Carbon monoxide is also formed. What is the percent yield if 51.4 kg of SiC is recovered from processing 100.0 kg of sand?

PLAN:

- write balanced equation
- find mol reactant
- find mol product
- find g product predicted
- percent yield
Sample Problem 3.21

SOLUTION: \[ \text{SiO}_2(s) + 3\text{C}(s) \rightarrow \text{SiC}(s) + 2\text{CO}(g) \]

\[
100.0 \text{ kg} \text{SiO}_2 \times \frac{10^3 \text{g}}{1 \text{ kg}} \times \frac{1 \text{ mol SiO}_2}{60.09 \text{ g SiO}_2} = 1664 \text{ mol SiO}_2
\]

\[
\text{mol SiO}_2 = \text{mol SiC} = 1664 \text{ mol SiC}
\]

\[
1664 \text{ mol SiC} \times \frac{40.10 \text{ g SiC}}{1 \text{ mol SiC}} \times \frac{1 \text{ kg}}{10^3 \text{g}} = 66.73 \text{ kg}
\]

\[
\frac{51.4 \text{ kg}}{66.73 \text{ kg}} \times 100 = 77.0\%
\]
Many reactions occur in solution.
A solution consists of one or more solutes dissolved in a solvent.
The concentration of a solution is given by the quantity of solute present in a given quantity of solution.
Molarity ($M$) is often used to express concentration.

\[
M = \frac{\text{mol solute}}{\text{L soln}}
\]

Molarity = \frac{\text{moles solute}}{\text{liters of solution}}
Sample Problem 3.22 Calculating the Molarity of a Solution

**PROBLEM:** What is the molarity of an aqueous solution that contains 0.715 mol of glycine (H₂NCH₂COOH) in 495 mL?

**PLAN:**
Molarity is the number of moles of solute per liter of solution.

**SOLUTION:**
\[
\text{molarity of glycine} = \frac{0.715 \text{ mol glycine}}{495 \text{ mL soln}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.44 \text{ M glycine}
\]
Figure 3.13  Summary of mass-mole-number-volume relationships in solution.
PROBLEM:  How many grams of solute are in 1.75 L of 0.460 M sodium monohydrogen phosphate buffer solution?

PLAN:  Calculate the moles of solute using the given molarity and volume. Convert moles to mass using the molar mass of the solute.

SOLUTION:

\[
1.75 \text{ L} \times \frac{0.460 \text{ moles}}{1 \text{ L}} = 0.805 \text{ mol } \text{Na}_2\text{HPO}_4
\]

\[
0.805 \text{ mol } \text{Na}_2\text{HPO}_4 \times \frac{141.96 \text{ g } \text{Na}_2\text{HPO}_4}{1 \text{ mol } \text{Na}_2\text{HPO}_4} = 114 \text{ g } \text{Na}_2\text{HPO}_4
\]
Laboratory preparation of molar solutions.

A • Weigh the solid needed.
• Transfer the solid to a volumetric flask that contains about half the final volume of solvent.

B Dissolve the solid thoroughly by swirling.

C Add solvent until the solution reaches its final volume.
Figure 3.15  Converting a concentrated solution to a dilute solution.
Sample Problem 3.24 Preparing a Dilute Solution from a Concentrated Solution

PROBLEM: “Isotonic saline” is a 0.15 \( M \) aqueous solution of NaCl. How would you prepare 0.80 L of isotonic saline from a 6.0 \( M \) stock solution?

PLAN: To dilute a concentrated solution, we add only solvent, so the moles of solute are the same in both solutions. The volume and molarity of the dilute solution gives us the moles of solute. Then we calculate the volume of concentrated solution that contains the same number of moles.

\[
\text{volume of dilute soln} \cdot M_{\text{dilute soln}} \cdot \text{mol NaCl in dilute soln} = \text{mol NaCl in concentrated soln}
\]

\[
\text{divide by } M_{\text{concentrated soln}}
\]

L of concentrated soln
Sample Problem 3.24

\[ M_{\text{dil}} \times V_{\text{dil}} = \# \text{ mol solute} = M_{\text{conc}} \times V_{\text{conc}} \]

**SOLUTION:**

Using the volume and molarity for the dilute solution:

\[ 0.80 \text{ L soln} \times \frac{0.15 \text{ mol NaCl}}{1 \text{ L soln}} = 0.12 \text{ mol NaCl} \]

Using the moles of solute and molarity for the concentrated solution:

\[ \frac{0.12 \text{ mol NaCl}}{6.0 \text{ mol NaCl}} \times \frac{1 \text{ L soln}}{1 \text{ L soln}} = 0.020 \text{ L soln} \]

A 0.020 L portion of the concentrated solution must be diluted to a final volume of 0.80 L.
Sample Problem 3.25  Visualizing Changes in Concentration

PROBLEM: The beaker and circle represents a unit volume of solution. Draw the solution after each of these changes:

(a) For every 1 mL of solution, 1 mL of solvent is added.

(b) One third of the volume of the solution is boiled off.

PLAN: Only the volume of the solution changes; the total number of moles of solute remains the same. Find the new volume and calculate the number of moles of solute per unit volume.
Sample Problem 3.25

**SOLUTION:**  
\[ N_{\text{dil}} \times V_{\text{dil}} = N_{\text{conc}} \times V_{\text{conc}} \]
where \( N \) is the number of particles.

(a)  
\[ N_{\text{dil}} = N_{\text{conc}} \times \frac{V_{\text{conc}}}{V_{\text{dil}}} = 8 \text{ particles} \times \frac{1 \text{ mL}}{2 \text{ mL}} = 4 \text{ particles} \]

(b)  
\[ N_{\text{conc}} = N_{\text{dil}} \times \frac{V_{\text{dil}}}{V_{\text{conc}}} = 8 \text{ particles} \times \frac{1 \text{ mL}}{\frac{2}{3} \text{ mL}} = 12 \text{ particles} \]
Sample Problem 3.26

Calculating Quantities of Reactants and Products for a Reaction in Solution

PROBLEM: A 0.10 M HCl solution is used to simulate the acid concentration of the stomach. How many liters of “stomach acid” react with a tablet containing 0.10 g of magnesium hydroxide?

PLAN: Write a balanced equation and convert the mass of Mg(OH)₂ to moles. Use the mole ratio to determine the moles of HCl, then convert to volume using molarity.
Sample Problem 3.26

**SOLUTION:**

\[
\text{Mg(OH)}_2 (s) + 2\text{HCl} (aq) \rightarrow \text{MgCl}_2 (aq) + 2\text{H}_2\text{O} (l)
\]

\[
0.10 \text{ g Mg(OH)}_2 \times \frac{1 \text{ mol Mg(OH)}_2}{58.33 \text{ g Mg(OH)}_2} = 1.7 \times 10^{-3} \text{ mol Mg(OH)}_2
\]

\[
= 1.7 \times 10^{-3} \text{ mol Mg(OH)}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg(OH)}_2} = 3.4 \times 10^{-3} \text{ mol HCl}
\]

\[
3.4 \times 10^{-3} \text{ mol HCl} \times \frac{1 \text{ L HCl soln}}{0.10 \text{ mol HCl}} = 3.4 \times 10^{-2} \text{ L HCl}
\]
PROBLEM: In a simulation mercury removal from industrial wastewater, 0.050 L of 0.010 M mercury(II) nitrate reacts with 0.020 L of 0.10 M sodium sulfide. How many grams of mercury(II) sulfide form? Write a reaction table for this process.

PLAN: Write a balanced chemical reaction. Determine limiting reactant. Calculate the grams of mercury(II) sulfide product.
Sample Problem 3.27

**SOLUTION:** \( \text{Hg(NO}_3\text{)}_2 \ (aq) + \text{Na}_2\text{S} \ (aq) \rightarrow \text{HgS} \ (s) + 2\text{NaNO}_3 \ (aq) \)

\[
0.050 \text{ L Hg(NO}_3\text{)}_2 \times 0.010 \text{ mol Hg(NO}_3\text{)}_2 \times \frac{1 \text{ mol HgS}}{1 \text{ mol Hg(NO}_3\text{)}_2} = 5.0 \times 10^{-4} \text{ mol HgS}
\]

\[
0.020 \text{ L Na}_2\text{S} \times 0.10 \text{ mol Na}_2\text{S} \times \frac{1 \text{ mol HgS}}{1 \text{ mol Na}_2\text{S}} = 2.0 \times 10^{-3} \text{ mol HgS}
\]

\[\text{Hg(NO}_3\text{)}_2 \text{ is the limiting reactant because it yields less HgS.}\]

\[
5.0 \times 10^{-4} \text{ mol HgS} \times \frac{232.7 \text{ g HgS}}{1 \text{ mol HgS}} = 0.12 \text{ g HgS}
\]
Sample Problem 3.27

The reaction table is constructed using the amount of Hg(NO$_3$)$_2$ to determine the changes, since it is the limiting reactant.

<table>
<thead>
<tr>
<th>Amount</th>
<th>Hg(NO$_3$)$_2$ (aq) +</th>
<th>Na$_2$S (aq) →</th>
<th>HgS (s) +</th>
<th>2NaNO$_3$ (aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td>5.0 x 10$^{-4}$</td>
<td>2.0 x 10$^{-3}$</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change</td>
<td>-5.0 x 10$^{-4}$</td>
<td>-5.0 x 10$^{-4}$</td>
<td>+5.0 x 10$^{-4}$</td>
<td>+1.0 x 10$^{-3}$</td>
</tr>
<tr>
<td>Final</td>
<td>0</td>
<td>1.5 x 10$^{-3}$</td>
<td>5.0 x 10$^{-4}$</td>
<td>+1.0 x 10$^{-3}$</td>
</tr>
</tbody>
</table>
Figure 3.16  An overview of amount-mass-number stoichiometric relationships.