Chapter 3

Stoichiometry of Formulas and Equations

3-1

Mole - Mass Relationships in Chemical Systems

3-2

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 6.022×10^{23} entities (to four significant figures).

This number is called Avogadro's number and is abbreviated as ***N.***

3-3

Figure 3.1 One mole (6.022×10^{23} entities) of some familiar substances.

3-4

Determining Molar Mass

The molar mass ($\mathcal{M}$) 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.

3-5

For **molecular elements** and for **compounds**, the formula is needed to determine the molar mass.

The molar mass of O_2 = 2 x $\mathcal{M}$ of O
= 2 x 16.00
= 32.00 g/mol

The molar mass of SO_2 = 1 x $\mathcal{M}$ of S + 2 x $\mathcal{M}$ of O
= 32.00 + 2(16.00)
= 64.00 g/mol

3-6

Table 3.1 Information Contained in the Chemical Formula of Glucose, $C_6H_{12}O_6$ ($\mathscr{M}$ = 180.16 g/mol)

	Carbon (C)	Hydrogen (H)	Oxygen (O)
Atoms/molecule of compound	6 atoms	12 atoms	6 atoms
Moles of atoms/mole of compound	6 mol of atoms	12 mol of atoms	6 mol of atoms
Atoms/mole of compound	$6(6.022\text{x}10^{23})$ atoms	$12(6.022\text{x}10^{23})$ atoms	$6(6.022\text{x}10^{23})$ atoms
Mass/molecule of compound	6(12.01 amu) = 72.06 amu	12(1.008 amu) = 12.10 amu	6(16.00 amu) = 96.00 amu
Mass/mole of compound	72.06 g	12.10 g	96.00 g

3-7

Converting Between Amount, Mass, and Number of Chemical Entities

$$\text{Mass (g)} = \text{no. of moles x } \frac{\text{no. of grams}}{\text{1 mol}} \leftarrow \boxed{g}$$

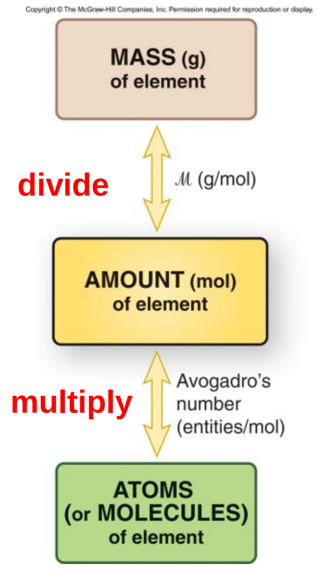
$$\text{No. of moles} = \text{mass (g) x } \frac{\text{1 mol}}{\text{no. of grams}} \leftarrow \boxed{\mathcal{M}}$$

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

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

3-8

Figure 3.2 **Mass-mole-number relationships for elements.**

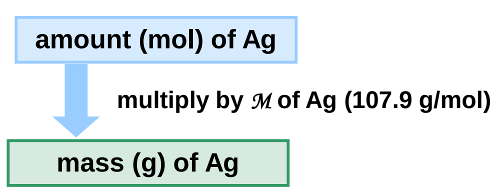


3-9

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:

$$0.0342\ \cancel{\text{mol Ag}} \times \frac{107.9\text{ g Ag}}{1\ \cancel{\text{mol Ag}}} \quad \boxed{= 3.69\text{ g Ag}}$$

3-10

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 2.85×10^{-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

↓ **multiply by 6.022×10^{23} atoms/mol**

atoms of Ga

3-11

Sample Problem 3.2

SOLUTION:

$$2.85 \times 10^{-3}\ \cancel{\text{mol Ga atoms}} \times \frac{6.022 \times 10^{23}\ \text{Ga atoms}}{1\ \cancel{\text{mol Ga atoms}}}$$

$$\boxed{= 1.72 \times 10^{21}\ \text{Ga atoms}}$$

3-12

Sample Problem 3.3 — **Calculating the Number of Entities in a Given Mass of an Element**

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 $\mathcal{M}$ of Fe (55.85 g/mol)

amount (mol) of Fe

↓ multiply by 6.022×10^{23} atoms/mol

atoms of Fe

Sample Problem 3.3

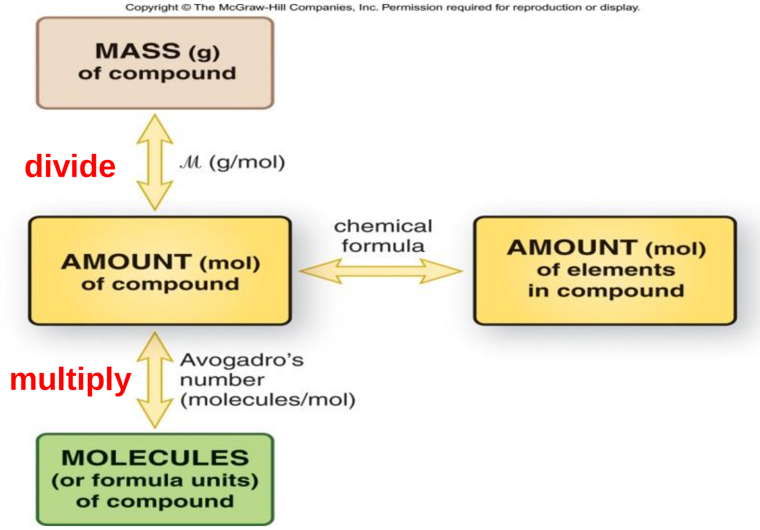
SOLUTION:

$$95.8\ \cancel{\text{g Fe}} \times \frac{1\ \text{mol Fe}}{55.85\ \cancel{\text{g Fe}}} = 1.72\ \text{mol Fe}$$

$$1.72\ \cancel{\text{mol Fe}} \times \frac{6.022\times10^{23}\ \text{atoms Fe}}{1\ \cancel{\text{mol Fe}}}$$

$$\boxed{= 1.04 \times 10^{24}\ \text{atoms Fe}}$$

Figure 3.3 Amount-mass-number relationships for compounds.



3-15

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.

mass (g) of NO_2

divide by $\mathcal{M}$

amount (mol) of NO_2

multiply by 6.022×10^{23} formula units/mol

number of NO_2 molecules

3-16

Sample Problem 3.4

SOLUTION: NO_2 is the formula for nitrogen dioxide.

$\mathcal{M}$ = (1 x $\mathcal{M}$ of N) + (2 x $\mathcal{M}$ of O)
= 14.01 g/mol + 2(16.00 g/mol)
= 46.01 g/mol

$$8.92\ \cancel{g\ NO_2} \times \frac{1\ mol\ NO_2}{46.01\ \cancel{g\ NO_2}} = 0.194\ mol\ NO_2$$

$$0.194\ \cancel{mol\ NO_2} \times \frac{6.022 \times 10^{23}\ molecules\ NO_2}{1\ \cancel{mol\ NO_2}}$$

$$\boxed{= 1.17 \times 10^{23}\ molecules\ NO_2}$$

3-17

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.

3-18

Sample Problem 3.5

mass (g) of $(NH_4)_2CO_3$

divide by $\mathcal{M}$

amount (mol) of $(NH_4)_2CO_3$

multiply by 6.022×10^{23} formula units/mol

number of $(NH_4)_2CO_3$ formula units

3 O atoms per formula unit of $(NH_4)_2CO_3$

number of O atoms

SOLUTION: $(NH_4)_2CO_3$ is the formula for ammonium carbonate.

$$\mathcal{M} = (2 \times M \text{ of N}) + (8 \times \mathcal{M} \text{ of H}) + (1 \times \mathcal{M} \text{ of C}) + (3 \times \mathcal{M} \text{ of O})$$
$$= (2 \times 14.01 \text{ g/mol}) + (8 \times 1.008 \text{ g/mol}) + (12.01 \text{ g/mol}) + (3 \times 16.00 \text{ g/mol})$$
$$\boxed{= 96.09 \text{ g/mol}}$$

3-19

Sample Problem 3.5

$$41.6\ \cancel{\text{g } (NH_4)_2CO_3} \times \frac{1 \text{ mol } (NH_4)_2CO_3}{96.09\ \cancel{\text{g } (NH_4)_2CO_3}} = 0.433 \text{ mol } (NH_4)_2CO_3$$

$$0.433\ \cancel{\text{mol } (NH_4)_2CO_3} \times \frac{6.022\text{x}10^{23} \text{ formula units } (NH_4)_2CO_3}{1\ \cancel{\text{mol } (NH_4)_2CO_3}}$$

$$\boxed{= 2.61\text{x}10^{23} \text{ formula units } (NH_4)_2CO_3}$$

$$2.61\text{x}10^{23}\ \cancel{\text{formula units } (NH_4)_2CO_3} \times \frac{3 \text{ O atoms}}{1\ \cancel{\text{formula unit of } (NH_4)_2CO_3}}$$

$$\boxed{= 7.83 \times 10^{23} \text{ O atoms}}$$

3-20

The Importance of Mass Percent

Mass % of element X =

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

Mass % of element X =

$$\frac{\text{moles of X in formula x molar mass of X (g/mol)}}{\text{mass (g) of 1 mol of compound}} \times 100$$

The individual mass percents added up to 100% (within rounding)

Sample Problem 3.6 **Calculating the Mass Percent of Each Element in a Compound from the Formula**

PROBLEM: Glucose ($C_6H_{12}O_6$) 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:

amount (mol) of element X in 1 mol compound

↓ **multiply by $\mathcal{M}$ (g/mol) of X**

mass (g) of X in 1 mol of compound

↓ **divide by mass (g) of 1 mol of compound**

mass fraction of X

↓ **multiply by 100**

mass % X in compound

3-23

Sample Problem 3.6

SOLUTION:

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

$$6 \text{ mol C} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 72.06 \text{ g C} \qquad 12 \text{ mol H} \times \frac{1.008 \text{ g H}}{1 \text{ mol H}} = 12.096 \text{ g H}$$

$$6 \text{ mol O} \times \frac{16.00 \text{ g O}}{1 \text{ mol O}} = 96.00 \text{ g O} \qquad \mathcal{M} = 180.16 \text{ g/mol}$$

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

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

$$\text{mass percent of O} = \frac{96.00 \text{ g O}}{180.16 \text{ g glucose}} = 0.5329 \times 100 = 53.29 \text{ mass \%O}$$

3-24

Determining the Mass of an Element from Its Percent

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

Mass of element X present in sample =

$$\text{mass of compound} \times \frac{\text{mass of element in 1 mol of compound}}{\text{mass of 1 mol of compound}}$$

3-25

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.

mass of glucose sample

↓ **multiply by mass percent of C in glucose**

mass of C in sample

3-26

Sample Problem 3.7

SOLUTION:

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

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

$$= 16.55 \cancel{\text{ g glucose}} \times \frac{72.06 \text{ g C}}{180.16 \cancel{\text{ g glucose}}} \quad \boxed{= 6.620 \text{ g C}}$$

3-27

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 H_2O_2.

3-28

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 $Zn_{0.21}P_{0.14}O_{0.56}$

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

$$\frac{0.21}{0.14} = 1.5 \qquad \frac{0.14}{0.14} = 1.0 \qquad \frac{0.56}{0.14} = 4.0$$

This gives $Zn_{1.5}P_{1.0}O_{4.0}$

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

1.5 x 2 = **3**   1.0 x 2 = **2**   4.0 x 2 = **8**

This gives us the empirical formula $Zn_3P_2O_8$

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 $\mathcal{M}$ (g/mol)**

amount (mol) of each element

↓ **use # of moles as subscripts**

preliminary formula

↓ **change to integer subscripts**

empirical formula

3-31

Sample Problem 3.9

SOLUTION:

$$2.82\ \cancel{\text{g Na}} \times \frac{1\ \text{mol Na}}{22.99\ \cancel{\text{g Na}}} = 0.123\ \text{mol Na}$$

$$4.35\ \cancel{\text{g Cl}} \times \frac{1\ \text{mol Cl}}{35.45\ \cancel{\text{g Cl}}} = 0.123\ \text{mol Cl}$$

$$7.83\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.489\ \text{mol O}$$

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

The empirical formula is $Na_1Cl_1O_{3.98}$ or $NaClO_4$; this compound is named sodium perchlorate.

3-32

Molecular Formulas

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.

$$\frac{\text{molar mass (g/mol)}}{\text{empirical formula mass (g/mol)}} = \text{whole-number multiple}$$

3-33

Sample Problem 3.10 **Determining a Molecular Formula from Elemental Analysis and Molar Mass**

PROBLEM: Elemental analysis of lactic acid ($\mathcal{M}$ = 90.08 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:

assume 100 g lactic acid; then mass % = mass in grams

↓ **divide each mass by $\mathcal{M}$**

amount (mol) of each element

↓ **use # mols as subscripts; convert to integers**

empirical formula

↓ **divide $\mathcal{M}$ by the molar mass for the empirical formula; multiply empirical formula by this number**

molecular formula

3-34

Sample Problem 3.10

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

$$40.0\ \cancel{g\ C} \times \frac{1\ \text{mol C}}{12.01\ \cancel{g\ C}} = 3.33\ \text{mol C} \qquad 6.71\ \cancel{g\ H} \times \frac{1\ \text{mol H}}{1.008\ \cancel{g\ H}} = 6.66\ \text{mol H} \qquad 53.3\ \cancel{g\ O} \times \frac{1\ \text{mol O}}{16.00\ \cancel{g\ O}} = 3.33\ \text{mol O}$$

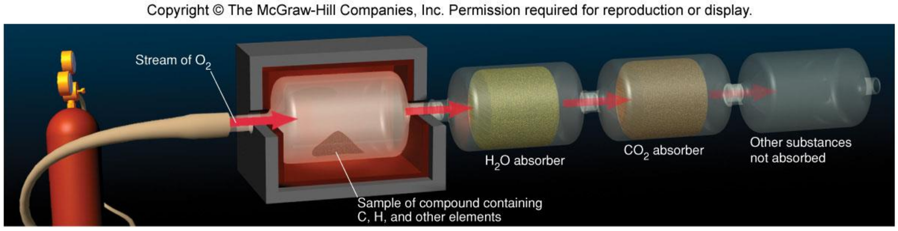
$$C_{\frac{3.33}{3.33}}\ H_{\frac{6.66}{3.33}}\ O_{\frac{3.33}{3.33}} \longrightarrow CH_2O \quad \text{empirical formula}$$

$$\frac{\text{molar mass of lactate}}{\text{mass of } CH_2O} \longrightarrow \frac{90.08\ \text{g/mol}}{30.03\ \text{g/mol}} = \mathbf{3}$$

$\longrightarrow$ $\boxed{\textbf{C}_3\textbf{H}_6\textbf{O}_3 \textbf{ is the molecular formula}}$

3-35

Figure 3.4 **Combustion apparatus for determining formulas of organic compounds.**



$$C_nH_m + (n+\frac{m}{4})\ O_2 = n\ CO2(g) + \frac{m}{2}\ H_2O(g)$$

3-36

Sample Problem 3.11 **Determining a Molecular Formula from Combustion Analysis**

PROBLEM:

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

mass of CO_2 absorber after combustion = 85.35 g
mass of CO_2 absorber before combustion = 83.85 g
mass of H_2O absorber after combustion = 37.96 g
mass of H_2O absorber before combustion = 37.55 g

What is the molecular formula of vitamin C?

PLAN: The masses of CO_2 and H_2O produced will give us the masses of C and H present in the original sample. From this we can determine the mass of O.

3-37

Sample Problem 3.11

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

↓ **mass of each compound x mass % of oxidized element**

mass of each oxidized element

↓ **mass of vitamin C – (mass of C + H)**

mass of O

↓ **divide each mass by $\mathcal{M}$**

mol of C, H, and O

↓ **use # mols as subscripts; convert to integers**

empirical formula → **molecular formula**

3-38

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_2O: 37.96 g - 37.55 g = 0.41 g

$$0.41 \text{ g } H_2O \times \frac{2.016 \text{ g H}}{18.02 \text{ g } H_2O} = 0.046 \text{ g H}$$

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

= 1.000 g - (0.409 + 0.046) g = 0.545 g O

3-39

Sample Problem 3.11

Convert mass to moles:

$$\frac{0.409 \text{ g C}}{12.01 \text{ g/mol C}} = 0.0341 \text{ mol C} \qquad \frac{0.046 \text{ g H}}{1.008 \text{ g/mol H}} = 0.0456 \text{ mol H}$$

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

Divide by smallest to get the preliminary formula:

$$\text{C } \frac{0.0341}{0.0341} = 1 \qquad \text{H } \frac{0.0456}{0.0341} = 1.34 \qquad \text{O } \frac{0.0341}{0.0341} = 1$$

$$C_1H_{1.34}O_1 = C_3H_{4.01}O_3 \longrightarrow C_3H_4O_3$$

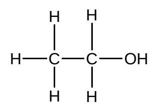
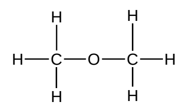
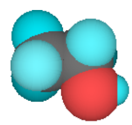
Divide molar mass by mass of empirical formula:

$$\frac{176.12 \text{ g/mol}}{88.06 \text{ g}} = 2.000 \text{ mol} \longrightarrow \boxed{\mathbf{C_6H_8O_6}}$$

3-40

Isomers

Table 3.2 Two Constitutional Isomers of C_2H_6O

Property	C_2H_6O Ethanol	C_2H_6O Dimethyl Ether
$\mathcal{M}$ (g/mol)	46.07	46.07
Boiling Point	78.5˚C	-25ºC
Density at 20ºC	0.789 g/mL (liquid)	0.00195 g/mL (gas)
Structural formula		
Space-filling model		

3-41

Writing and Balancing Chemical Equations

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

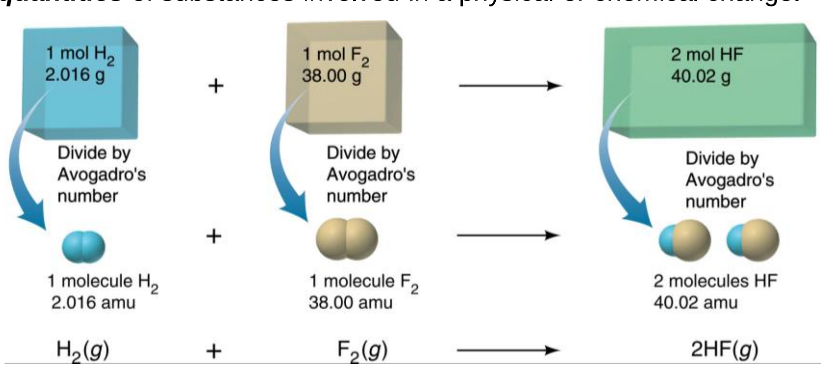
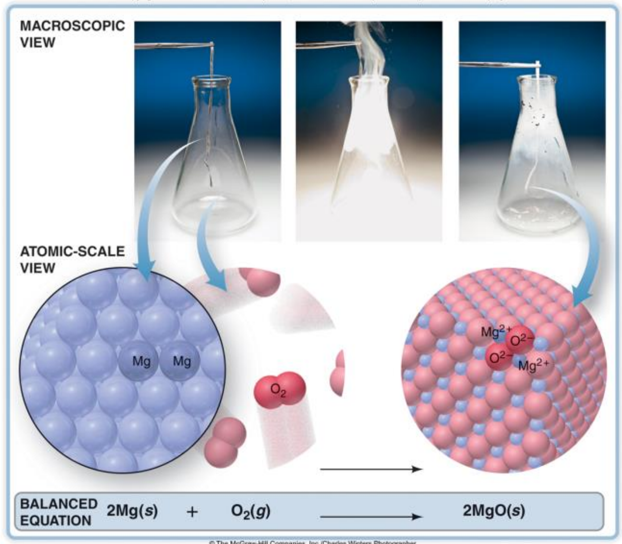


Figure 3.5

The formation of HF gas on the macroscopic and molecular levels.

3-42

Figure 3.6 A three-level view of the reaction between magnesium and oxygen.



3-43

Features of Chemical Equations

A **yield arrow** points from reactants to products.

$$\underbrace{Mg \quad + \quad O_2}_{} \longrightarrow MgO$$

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.

3-44

Balancing a Chemical Equation

translate the statement

magnesium and oxygen gas react to give magnesium oxide:
Mg + O_2 → MgO

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

2Mg + O_2 → **2**MgO

adjust coefficients if necessary

check that all atoms balance → **specify states of matter**

2Mg (*s*) + O_2 (*g*) → **2**MgO (*s*)

3-45

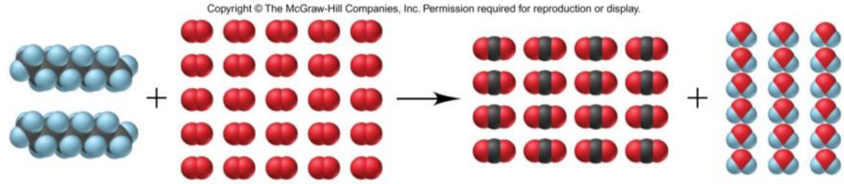
Sample Problem 3.12 **Balancing Chemical Equations**

PROBLEM: Within the cylinders of a car's engine, the hydrocarbon octane (C_8H_{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:	SOLUTION:
translate the statement	$C_8H_{18} + O_2 \longrightarrow CO_2 + H_2O$
balance the atoms	$C_8H_{18} + \frac{25}{2}O_2 \longrightarrow 8\,CO_2 + 9\,H_2O$
adjust the coefficients	$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$
check the atoms balance	$2C_8H_{18} + 25O_2 \longrightarrow 16CO_2 + 18H_2O$
specify states of matter	$2C_8H_{18}(l) + 25O_2\,(g) \longrightarrow 16CO_2\,(g) + 18H_2O\,(g)$

3-46

Molecular Scene
Combustion of Octane

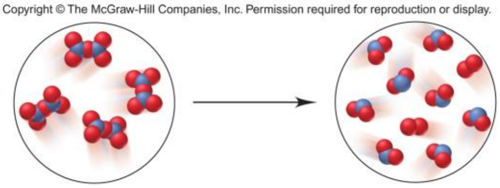


$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

3-47

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.

3-48

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 N_2O_5. There are 4 N_2O_5 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 NO_2 and O_2. There are 8 NO_2 molecules and 2 O_2 molecules shown.

The reaction depicted is $\mathbf{4\ N_2O_5 \rightarrow 8\ NO_2 + 2\ O_2}$.

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

$$\mathbf{2\ N_2O_5\ (\mathit{g}) \rightarrow 4\ NO_2\ (\mathit{g}) + O_2\ (\mathit{g})}$$

3-49

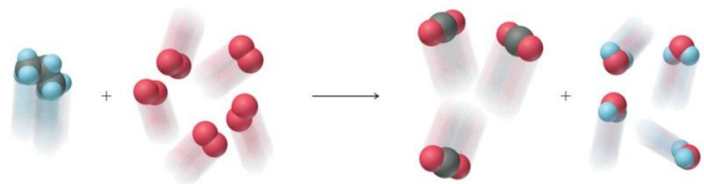
Stoichiometric Calculations

- 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.

$$\mathbf{2\ N_2O_5\ (\mathit{g}) \rightarrow 4\ NO_2\ (\mathit{g}) + O_2\ (\mathit{g})}$$

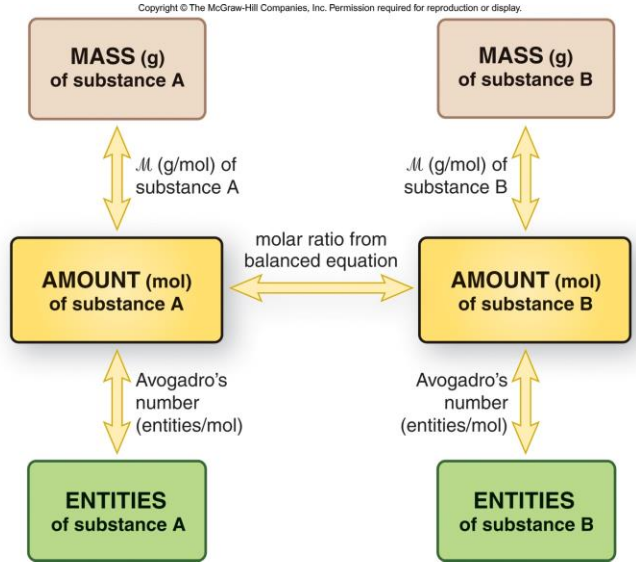
3-50

Table 3.3 Information Contained in a Balanced Equation

Viewed in Terms of	Reactants $C_3H_8(g)$ + 5 $O_2(g)$	→	Products 3 $CO_2(g)$ + 4 $H_2O(g)$
Molecules	1 molecule C_3H_8 + 5 molecules O_2	→	3 molecules CO_2 + 4 molecules H_2O
			
Amount (mol)	1 mol C_3H_8 + 5 mol O_2	→	3 mol CO_2 + 4 mol H_2O
Mass (amu)	44.09 amu C_3H_8 + 160.00 amu O_2	→	132.03 amu CO_2 + 72.06 amu H_2O
Mass (g)	44.09 g C_3H_8 + 160.00 g O_2	→	132.03 g CO_2 + 72.06 g H_2O
Total Mass (g)	204.09 g	→	204.09 g

3-51

Figure 3.7 **Summary of amount-mass-number relationships in a chemical equation.**



3-52

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**

moles of oxygen

SOLUTION: $2\ Cu_2S\ (s) + 3\ O_2\ (g) \rightarrow 2\ Cu_2O\ (s) + 2\ SO_2\ (g)$

$$10.0\ \cancel{\text{mol } Cu_2S} \times \frac{3\ \text{mol } O_2}{2\ \cancel{\text{mol } Cu_2S}}\ \boxed{= 15.0\ \text{mol } O_2}$$

3-53

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.

mol of copper(I) sulfide

↓ **use the mole ratio as a conversion factor**

mol of sulfur dioxide

↓ **multiply by $\mathcal{M}$ of sulfur dioxide**

mass of sulfur dioxide

3-54

Sample Problem 3.15

SOLUTION: $2\ Cu_2S\ (s) + 3\ O_2\ (g) \rightarrow 2\ Cu_2O\ (s) + 2\ SO_2\ (g)$

$$10.0\ \cancel{\text{mol } Cu_2S} \times \frac{2\ \cancel{\text{mol } SO_2}}{2\ \cancel{\text{mol } Cu_2S}} \times \frac{64.07\ \text{g } SO_2}{1\ \cancel{\text{mol } SO_2}} \quad \boxed{= 641\ \text{g } SO_2}$$

3-55

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:

mass of copper(I) oxide (g)

↓ **divide by $\mathcal{M}$ of copper(I) oxide**

mol of copper(I) oxide

↓ **use mole ratio as conversion factor**

mol of oxygen

↓ **multiply by $\mathcal{M}$ of copper(I) oxide**

mass of oxygen (g)

3-56

Sample Problem 3.16

SOLUTION: $2\ Cu_2S\ (s) + 3\ O_2\ (g) \rightarrow 2\ Cu_2O\ (s) + 2\ SO_2\ (g)$

$$2.86\ \text{kg}\ \cancel{Cu_2O} \times \frac{10^3\ \text{g}}{1\ \cancel{\text{kg}}} \times \frac{1\ \text{mol}\ Cu_2O}{143.10\ \cancel{\text{g}\ Cu_2O}} = 20.0\ \text{mol}\ Cu_2O$$

$$20.0\ \cancel{\text{mol}\ Cu_2O} \times \frac{3\ \cancel{\text{mol}\ O_2}}{2\ \cancel{\text{mol}\ Cu_2O}} \times \frac{32.00\ \cancel{\text{g}}\ O_2}{1\ \cancel{\text{mol}\ O_2}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}} \quad \boxed{= 0.959\ \text{kg}\ O_2}$$

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.

$$\begin{array}{ccc} A & \rightarrow & C \\ \underline{C} & \underline{\rightarrow} & \underline{B} \\ A & \rightarrow & B \end{array}$$

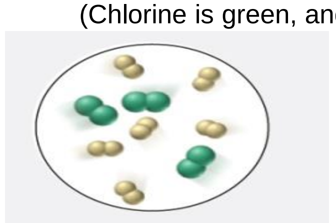
Limiting Reactants

- 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.

3-59

Sample Problem 3.17 **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.)

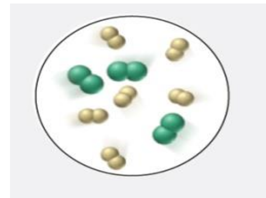


(a) Find the limiting reactant.
(b) Write a reaction table for the process.
(c) Draw a representative portion of the mixture after the reaction is complete. (Hint: The ClF_3 molecule has 1 Cl atom bonded to 3 individual F atoms).

3-60

Sample Problem 3.17

SOLUTION: The balanced equation is Cl_2 (*g*) + $3F_2$ (*g*) → $2ClF_3$ (*g*)



There are 3 molecules of Cl_2 and 6 molecules of F_2 depicted:

$$3\ \cancel{\text{molecules}\ Cl_2} \times \frac{2\ \text{molecules}\ ClF_3}{1\ \cancel{\text{molecule}\ Cl_2}} = 6\ \text{molecules}\ ClF_3$$

$$6\ \cancel{\text{molecules}\ F_2} \times \frac{2\ \text{molecules}\ ClF_3}{3\ \cancel{\text{molecule}\ Cl_2}} = 4\ \text{molecules}\ ClF_3$$

Since the given amount of F_2 can form less product, it is the limiting reactant.

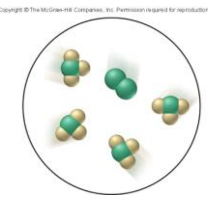
3-61

Sample Problem 3.17

We use the amount of F_2 to determine the "change" in the reaction table, since F_2 is the limiting reactant:

Molecules	Cl_2 (*g*) +	$3F_2$ (*g*) →	$2ClF_3$ (*g*)
Initial	3	6	0
Change	-2	**-6**	+4
Final	1	0	4

The final reaction scene shows that all the F_2 has reacted and that there is Cl_2 left over. 4 molecules of ClF_2 have formed:



3-62

Sample Problem 3.18 **Calculating Quantities in a Limiting-Reactant Problem: Amount to Amount**

PROBLEM: In another preparation of ClF_3, 0.750 mol of Cl_2 reacts with 3.00 mol of F_2.

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

PLAN: Find the limiting reactant by calculating the amount (mol) of ClF_3 that can be formed from each given amount of reactant. Use this information to construct a reaction table.

SOLUTION: The balanced equation is $Cl_2\ (g) + 3F_2\ (g) \rightarrow 2ClF_3\ (g)$

$$0.750\ \cancel{\text{mol Cl}_2} \times \frac{2\ \text{mol ClF}_3}{1\ \cancel{\text{mol Cl}_2}} = 1.50\ \text{mol ClF}_3$$

$$3.00\ \cancel{\text{mol F}_2} \times \frac{2\ \text{mol ClF}_3}{3\ \cancel{\text{mol F}_2}} = 2.00\ \text{mol ClF}_3$$

Cl_2 is limiting, because it yields less ClF_3.

Sample Problem 3.18

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.

Moles	$Cl_2\ (g)$ +	$3F_2\ (g)$ →	$2ClF_3\ (g)$
Initial	0.750	3.00	0
Change	**-0.750**	- 2.25	+1.50
Final	0	0.75	1.50

Sample Problem 3.19 — **Calculating Quantities in a Limiting-Reactant Problem: Mass to Mass**

PROBLEM: A fuel mixture used in the early days of rocketry consisted of two liquids, hydrazine (N_2H_4) and dinitrogen tetraoxide (N_2O_4), which ignite on contact to form nitrogen gas and water vapor.

(a) How many grams of nitrogen gas form when 1.00×10^2 g of N_2H_4 and 2.00×10^2 g of N_2O_4 are mixed?

(b) Write a reaction table for this process.

PLAN: Find the limiting reactant by calculating the amount (mol) of ClF_3 that can be formed from each given mass of reactant. Use this information to construct a reaction table.

Sample Problem 3.19

mass (g) of N_2H_4 $\xrightarrow{\text{divide by } \mathcal{M}}$ **mol of N_2H_4** $\xrightarrow{\text{mole ratio}}$ **mol of N_2**

mass (g) of N_2O_4 $\xrightarrow{\text{divide by } \mathcal{M}}$ **mol of N_2O_4** $\xrightarrow{\text{mole ratio}}$ **mol of N_2**

select lower number of moles of N_2
multiply by $\mathcal{M}$

mass of N_2

Sample Problem 3.19

SOLUTION: $2N_2H_4\ (l) + N_2O_4\ (l) \rightarrow 3N_2\ (g) + 4H_2O\ (g)$

For N_2H_4: $1.00\text{x } 10^2\ \cancel{g\ N_2H_4} \times \dfrac{1\ mol\ N_2H_4}{32.05\ \cancel{g\ N_2H_4}} = 3.12\ mol\ N_2H_4$

$3.12\ \cancel{mol\ N_2H_4} \times \dfrac{3\ mol\ N_2}{2\ \cancel{mol\ N_2H_4}} = 4.68\ mol\ N_2$

For N_2O_4: $2.00\text{x } 10^2\ \cancel{g\ N_2O_4} \times \dfrac{1\ mol\ N_2O_4}{92.02\ \cancel{g\ N_2O_4}} = 2.17\ mol\ N_2$

$2.17\ \cancel{mol\ N_2O_4} \times \dfrac{3\ mol\ N_2}{1\ \cancel{mol\ N_2O_4}} = 6.51\ mol\ N_2$

N_2H_4 is limiting and only 4.68 mol of N_2 can be produced:

$4.68\ \cancel{mol\ N_2} \times \dfrac{28.02\ g\ N_2}{1\ \cancel{mol\ N_2}}\ \boxed{= 131\ g\ N_2}$

3-67

Sample Problem 3.19

All the N_2H_4 reacts since it is the limiting reactant. For every 2 moles of N_2H_4 that react 1 mol of N_2O_4 reacts and 3 mol of N_2 form:

$3.12\ \cancel{mol\ N_2H_4} \times \dfrac{1\ mol\ N_2O_4}{2\ \cancel{mol\ N_2H_4}} = 1.56\ mol\ N_2O_4$ reacts

Moles	**$2N_2H_4$ (l) +**	**N_2O_4 (*l*) →**	**$3N_2$ (*g*) +**	**$4H_2O$ (*g*)**
Initial	3.12	2.17	0	0
Change	**-3.12**	**- 1.56**	**+4.68**	+6.24
Final	0	0.61	4.68	6.24

3-68

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.

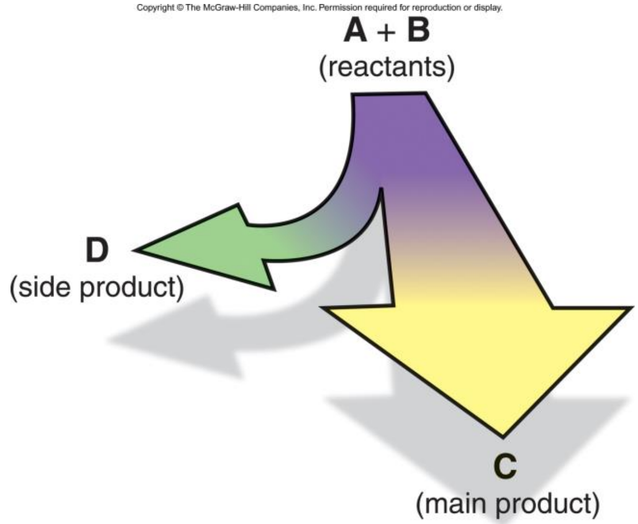
The actual yield is usually less than the theoretical yield.

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

Can expressed in moles or grams

3-69

Figure 3.8 **The effect of side reactions on the yield of the main product.**



3-70

Sample Problem 3.20 **Calculating Percent Yield**

PROBLEM: Silicon carbide (SiC) is made by reacting sand(silicon dioxide, SiO_2) 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**

3-71

Sample Problem 3.20

SOLUTION: $SiO_2(s) + 3C(s) \rightarrow SiC(s) + 2CO(g)$

What is the percent yield if 51.4 kg of SiC is recovered from processing 100.0 kg of sand?

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

mol SiO_2 = mol SiC = 1664 mol SiC

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

$$\frac{51.4\ \text{kg}}{66.73\ \text{kg}} \times 100 \quad \boxed{= 77.0\%}$$

3-72

Solution Stoichiometry

- 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.

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters of solution}} \qquad M = \frac{\text{mol solute}}{\text{L soln}}$$

3-73

Sample Problem 3.21 **Calculating the Molarity of a Solution**

PROBLEM: What is the molarity of an aqueous solution that contains 0.715 mol of glycine (H_2NCH_2COOH) in 495 mL?

PLAN:

Molarity is the number of moles of solute per liter of solution.

mol of glycine

↓ **divide by volume**

concentration in mol/mL

↓ **10^3 mL = 1 L**

molarity of glycine

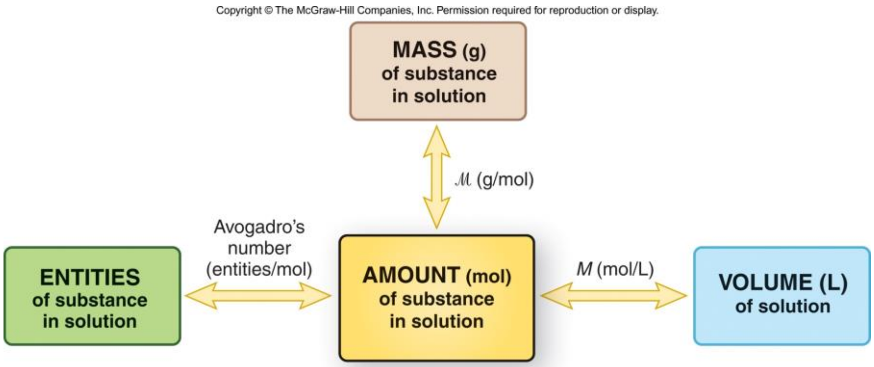
SOLUTION:

$$\frac{0.715 \text{ mol glycine}}{495\ \cancel{\text{mL}}\ \text{soln}} \times \frac{1000\ \cancel{\text{mL}}}{1 \text{ L}}$$

$$\boxed{= 1.44\ M \text{ glycine}}$$

3-74

Figure 3.9 **Summary of mass-mole-number-volume relationships in solution.**



M X V = # of moles

Sample Problem 3.22 **Calculating Mass of Solute in a Given Volume of 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.

volume of solution

multiply by *M*

moles of solute

multiply by $\mathcal{M}$

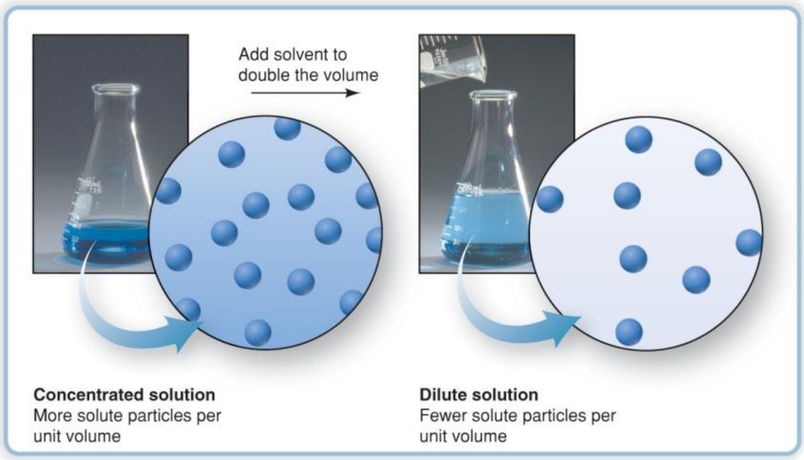
grams of solute

SOLUTION:

$$1.75\ \cancel{L} \times \frac{0.460\ \text{moles}}{1\ \cancel{L}} = 0.805\ \text{mol}\ Na_2HPO_4$$

$$0.805\ \cancel{\text{mol}\ Na_2HPO_4} \times \frac{141.96\ \text{g}\ Na_2HPO_4}{1\ \cancel{\text{mol}\ Na_2HPO_4}} \quad \boxed{= 114\ \text{g}\ Na_2HPO_4}$$

Figure 3.10 Converting a concentrated solution to a dilute solution.



3-77

Sample Problem 3.23 **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.

volume of dilute soln

↓ **multiply by *M* of dilute soln**

moles of NaCl in dilute soln = mol NaCl in concentrated soln

↓ **divide by *M* of concentrated soln**

L of concentrated soln

3-78

Sample Problem 3.23

$$\boxed{M_{dil} \times V_{dil} = \#\ \text{mol solute} = M_{conc} \times V_{conc}}$$

SOLUTION:

Using the volume and molarity for the dilute solution:

$$0.80\ \cancel{\text{L soln}} \times \frac{0.15\ \text{mol NaCl}}{1\ \cancel{\text{L soln}}} = 0.12\ \text{mol NaCl}$$

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

$$0.12\ \cancel{\text{mol NaCl}} \times \frac{1\ \text{L soln}}{6.0\ \cancel{\text{mol NaCl}}} \ \boxed{= 0.020\ \text{L soln}}$$

A 0.020 L portion of the concentrated solution must be diluted to a final volume of 0.80 L.

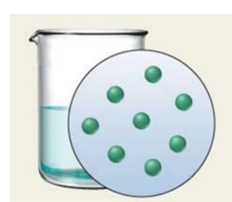
3-79

Sample Problem 3.24 **Visualizing Changes in Concentration**

PROBLEM: The beaker and circle represent 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.

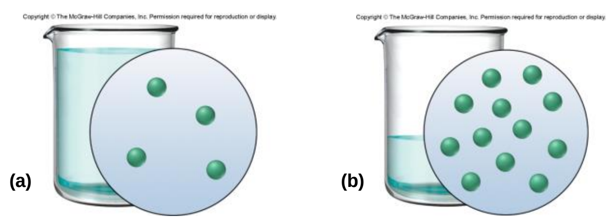
3-80

Sample Problem 3.24

SOLUTION: $N_{dil} \times V_{dil} = N_{conc} \times V_{conc}$
where N is the number of particles.

(a) $N_{dil} = N_{conc} \times \dfrac{V_{conc}}{V_{dil}} = 8 \text{ particles} \times \dfrac{1 \text{ mL}}{2 \text{ mL}} = 4 \text{ particles}$

(b) $N_{conc} = N_{dil} \times \dfrac{V_{dil}}{V_{conc}} = 8 \text{ particles} \times \dfrac{1 \text{ mL}}{\frac{2}{3} \text{ mL}} = 12 \text{ particles}$

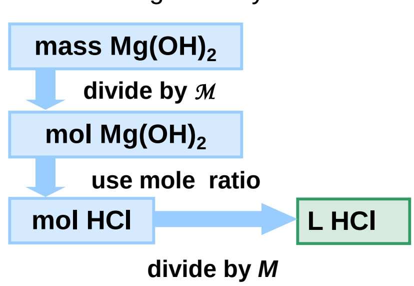


3-81

Sample Problem 3.25 **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.



3-82

Sample Problem 3.25

SOLUTION:

$$Mg(OH)_2\ (s) + 2HCl\ (aq) \rightarrow MgCl_2\ (aq) + 2H_2O\ (l)$$

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

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

$$3.4\text{x}10^{-3}\ \cancel{mol\ HCl} \times \frac{1L\ HCl\ soln}{0.10\ \cancel{mol\ HCl}} \quad \boxed{= 3.4\text{x}10^{-2}\ L\ HCl}$$

3-83

Sample Problem 3.26 **Solving Limiting-Reactant Problems for Reactions in Solution**

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.

volume of $Hg(NO_3)_2$ soln
↓ **multiply by *M***
mol of $Hg(NO_3)_2$
↓ **mole ratio**
mol of HgS

volume of Na_2S soln
↓ **multiply by *M***
mol of Na_2S
↓ **mole ratio**
mol of HgS

↓ **select lower number of moles of HgS multiply by $\mathcal{M}$**

mass of HgS

3-84

Sample Problem 3.26

SOLUTION: $Hg(NO_3)_2\ (aq) + Na_2S\ (aq) \rightarrow HgS\ (s) + 2NaNO_3\ (aq)$

$$0.050\ \cancel{L\ Hg(NO_3)_2} \times \frac{0.010\ \cancel{mol\ Hg(NO_3)_2}}{1\ \cancel{L\ Hg(NO_3)_2}} \times \frac{1\ mol\ HgS}{1\ \cancel{mol\ Hg(NO_3)_2}}$$

$= 5.0\text{x}10^{-4}$ mol HgS

$$0.020\ \cancel{L\ Na_2S} \times \frac{0.10\ \cancel{mol\ Na_2S}}{1\ \cancel{L\ Na_2S}} \times \frac{1\ mol\ HgS}{1\ \cancel{mol\ Na_2S}} = 2.0\text{x}10^{-3}\ mol\ HgS$$

$Hg(NO_3)_2$ is the limiting reactant because it yields less HgS.

$$5.0 \times 10^{-4}\ \cancel{mol\ HgS} \times \frac{232.7\ g\ HgS}{1\ \cancel{mol\ HgS}}\ \boxed{= 0.12\ g\ HgS}$$

3-85

Sample Problem 3.26

The reaction table is constructed using the amount of $Hg(NO_3)_2$ to determine the changes, since it is the limiting reactant.

Amount	**$Hg(NO_3)_2$ (*aq*) +**	**Na_2S (*aq*) →**	**HgS (*s*) +**	**$2NaNO_3$ (*aq*)**
Initial	5.0×10^{-4}	2.0×10^{-3}	0	0
Change	$\mathbf{-5.0 \times 10^{-4}}$	$\mathbf{-5.0 \times 10^{-4}}$	$\mathbf{+5.0 \times 10^{-4}}$	$\mathbf{+1.0 \times 10^{-3}}$
Final	0	1.5×10^{-3}	5.0×10^{-4}	$+1.0 \times 10^{-3}$

3-86