Chapter 3

Stoichiometry: Calculations with Chemical Formulas and Equations

OVERVIEW OF THE CHAPTER

3.1, 3.2 CHEMICAL EQUATIONS: BALANCING AND PREDICTING PRODUCTS OF REACTIONS

Review: Elements and compounds (1.1,1.2); formulas (2.6,2.7); nomenclature (2.8).

Learning Goals: You should be able to:

1. Balance chemical equations.
2. Predict the products of a chemical reaction, having seen a suitable analogy.
3. Predict the products of the combustion reactions of hydrocarbons and simple compounds containing C, H, and O atoms.

3.3, 3.4 FORMULA WEIGHT, MOLECULAR WEIGHT, AND THE MOLE

Learning Goals: You should be able to:

1. Calculate the formula weight of a substance given its chemical formula.
2. Calculate the molecular weight of a molecular substance given its chemical formula.
3. Recognize when to use formula weights and molecular weights in calculations.
4. Calculate the molar mass of a substance from its chemical formula.
5. Interconvert the number of moles of a substance and its mass.
6. Use Avogadro's number and molar mass to calculate the number of particles making up a substance and *vice versa*.

3.5 DETERMINING EMPIRICAL AND MOLECULAR FORMULAS

Review: Empirical and molecular formulas (2.6).

Learning Goals: You should be able to:

1. Calculate the empirical formula of a compound, having been given appropriate analytical data such as elemental percentages or the quantity of CO_2 and H_2O produced by combustion.
2. Calculate the molecular formula, having been given the empirical formula and molecular weight.

3.6, 3.7 CHEMICAL EQUATIONS: MASS AND MOLE RELATIONSHIPS

Review: Dimensional analysis (1.6); rounding numbers (1.5).

Learning Goals: You should be able to:

1. Calculate the mass of a particular substance produced or used in a chemical reaction (mass–mass problem).
2. Determine the limiting reagent in a reaction.
3. Calculate the theoretical and actual yields of chemical reactions given the appropriate data.

TOPIC SUMMARIES AND EXERCISSE

CHEMICAL EQUATIONS: BALANCING AND PREDICTING PRODUCTS OF REACTIONS

Chemical equation such as $2\ C + O_2 \rightarrow 2\ CO$:

- Describe chemical processes involving **reactants** (left side of arrow) to form **products** (right side of arrow).
- Should be balanced.
- Provide a means for calculating mass relationships among products and reactants.

When you balance chemical equations, you must keep the following requirements in mind:

- Formulas of substances must be correctly written.
- The number of atoms of each type of element must be the same on both sides of the arrow.
- Only coefficients in front of substances may be adjusted to change the number of atoms on the reactant or product side. Subscripts in chemical formulas must not be changed.
- The sum of charges of ions on the left side of the arrow must be the same on the right side.

See Exercise 1 for an approach to balancing chemical reactions.

An important skill to develop is the ability to predict the products of simple chemical reactions having been given only the reactants. In this chapter we look at several classes of chemical reactions to help us develop this skill: combustion, combination, and decomposition.

Combustion reactions produce a flame and usually involve oxygen as a reactant. For example:

$$C_2H_4(g) + 3\ O_2(g) \longrightarrow 2\ CO_2(g) + 2\ H_2O(l)$$

Means a substance is a gas

Means a substance is a liquid

- The text highlights the combustion of **hydrocarbons**, carbon-, and hydrogen-containing compounds. When hydrocarbons are combusted in the presence of oxygen, carbon is converted to CO_2 and hydrogen is converted to H_2O. If oxygen is also present in a compound containing C and H, for example CH_3OH, the oxygen is used along with O_2 in balancing the equation.

Combination reactions involve forming one product from two or more reactants. For example:

$$CaO(s) + CO_2(g) \longrightarrow CaCO_3(s)$$

Decomposition reactions are the reverse of combination reactions: One reactant breaks down (decomposes) into two or more substances. For example:

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

EXERCISE 1 Balancing chemical equations

Balance the following reactions:

(a) $P_4(s) + O_2(g) \longrightarrow P_4O_{10}(s)$
(b) $SF_4(g) + H_2O(l) \longrightarrow SO_2(g) + HF(g)$

SOLUTION: *Analyze:* We are given chemical reactions which are not balanced and we are asked to determine the balancing coefficients for each substance.

Plan: A chemical reaction is balanced when the number of atoms of each type are the same on both sides of the arrow. If charged species are present, then the sum of charges on the left must also equal the sum of charges on the right. The first step is to count the number of atoms to determine if the given reaction is already balanced. If it is not balanced then we have to place coefficients in *front* of substances so that the number of atoms of each type are the same for both reactants and products.

Solve: (a) By inspection we see that the number of atoms are not balanced in the chemical equation.

$$P_4(s) + O_2(g) \longrightarrow P_4O_{10}(s)$$

An inventory of atoms shows:

	No. of P Atoms	No. of O Atoms
Reactants	4	2
Products	4	10

Although the number of phosphorus atoms is balanced, the number of oxygen atoms is not. 10 oxygen atoms, that is, 5 O_2, are needed on the reactant side to equal the 10 oxygen atoms on the product side. The balanced chemical equation is

$$P_4(s) + 5\,O_2(g) \longrightarrow P_4O_{10}(s)$$

(b) An inventory of atoms in the chemical equation

$$SF_4(g) + H_2O(l) \longrightarrow SO_2(g) + HF(g)$$

shows:

	No. of S Atoms	No. of F Atoms	No. of H Atoms	No. of O Atoms
Reactants:	1	4	2	1
Products:	1	1	1	2

Starting with the atom that is present in the greatest number on the reactant side, fluorine, we can balance the fluorine atoms on the product side with 4 HF. The inventory is now

	No. of S Atoms	No. of F Atoms	No. of H Atoms	No. of O Atoms
Reactants	1	4	2	1
Products:	1	4	4	2

The hydrogen atoms can be balanced with 2 H_2O on the reactant side. The balancing of the hydrogen atoms also causes the oxygen atoms to be balanced. The balanced equation is

$$SF_4(g) + 2\,H_2O(l) \longrightarrow SO_2(g) + 4\,HF(g)$$

Check: If we have done our tables correctly the number of atoms of each element should be the same on both sides of the arrow. However, it is advisable to look at each balanced chemical equation and recheck that this is true. For example, in (a) we can count that there are 4 P atoms on each side of the arrow and 10 oxygen atoms on each side of the arrow.

Exercise 2 Writing and balancing a combustion reaction

Write the balanced chemical equation for the combustion of butane, C_4H_{10} in air.

SOLUTION: *Analyze:* We are given butane, C_4H_{10}, and are asked to write the balanced chemical equation for its combustion reaction.

Plan: The first step is to write the skeletal equation that describes the combustion reaction for butane, a hydrocarbon. We can do this by using oxygen gas, a component of air, as a reactant. The typical products formed when a hydrocarbon is combusted are carbon dioxide and water in the gaseous state. After writing the skeletal equation we can then balance it.

Solve: The skeletal chemical equation is

$$C_4H_{10}(g) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$$

When balancing combustion reactions of simple hydrocarbons, first balance the carbon and hydrogen atoms without considering oxygen. Then balance the oxygen atoms using $O_2(g)$. There are 4 carbon atoms, therefore, place a 4 in front of CO_2; there are 10 hydrogen atoms in butane, therefore, place a 5 in front of H_2O:

$$C_4H_{10}(g) + O_2(g) \longrightarrow 4\,CO_2(g) + 5\,H_2O(l)$$

The products contain a total of 13 oxygen atoms, 8 from 4 carbon dioxide molecules and 5 from 5 water molecules. Place a 13/2 in front of $O_2(g)$ to balance the 13 oxygen atoms:

$$C_4H_{10}(g) + \frac{13}{2}\,O_2(g) \longrightarrow 4\,CO_2(g) + 5\,H_2O(l)$$

It is more convenient to deal with whole numbers; it is usually customary (but not always required) to clear the fractions by multiplying the entire equation by a number that cancels the denominators. In this case, multiply the entire equation by two to eliminate the denominator in 13/2:

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

Check: A check of the number of atoms shows that there are 8 carbon atoms, 20 hydrogen atoms, and 36 oxygen atoms on both sides of the arrow.

Exercise 3 Completing and balancing chemical reactions

Complete and balance the following reactions, and indicate the phases of each substance: (a) $SbBr_3 + H_2S \rightarrow$ (b) $LiH + H_2O \rightarrow$ (c) $CO_2 + Na \rightarrow$ (d) $FeS + O_2 \rightarrow$ Use the following reactions as examples:

(1) $3\,CO_2(g) + 4\,K(s) \rightarrow 2\,K_2CO_3(s) + C(s)$
(2) $CaH_2(s) + 2\,H_2O(l) \rightarrow Ca(OH)_2(aq) + 2\,H_2(g)$

(3) $2\,SbCl_3(s) + 3\,H_2S(g) \rightarrow Sb_2S_3(s) + 6\,HCl(g)$
(4) $2\,ZnS(s) + 3\,O_2(g) \rightarrow 2\,ZnO(s) + 2\,SO_2(g)$

SOLUTION: *Analyze:* We are given four incomplete chemical reactions and are asked to complete and balance their chemical equations.

Plan: The reactions can be completed by finding an analogous reaction among those given. Analogous compounds or elements that are from the same family or that are otherwise similar in nature often exist in the same phase.

Solve: We can complete the chemical equation given in (a) by observing that H_2S is a reactant and $H_2S(g)$ is also a reactant in the sample chemical equation (3). Furthermore the other reactant in the sample chemical equation (3) is an antimony halide as in (a). Repeating this process for each incomplete chemical equation gives the following results:

Balanced equation	**Analogous reaction**
(a) $2\,SbBr_3(s) + 3\,H_2S(g) \rightarrow Sb_2S_3(s) + 6\,HBr(g)$	(3)
(b) $LiH(s) + H_2O(l) \rightarrow LiOH(aq) + H_2(g)$	(2)
(c) $3\,CO_2(g) + 4\,Na(s) \rightarrow 2\,Na_2CO_3(s) + C(s)$	(1)
(d) $2\,FeS(s) + 3\,O_2(g) \rightarrow 2\,FeO(s) + 2\,SO_2(g)$	(4)

Comment: It takes reading and practice to develop our knowledge of chemical reactions. As this knowledge base grows we will be able to predict the products of many chemical reactions.

FORMULA WEIGHT, MOLECULAR WEIGHT, AND THE MOLE

Substances come in a variety of forms. One form is molecular: A molecule is the smallest particle of a pure substance that has the composition and properties of the pure substance and also has an independent existence. Another is ionic: An ionic substance consists of particles with charges and there is no discrete smaller unit with an independent existence. In this chapter we learn how to calculate the masses of molecular and ionic substances.

- The term **formula weight** refers to the sum of atomic weights of the atoms in a substance. It can be used with both molecular and ionic substances.
- The term **molecular weight** refers to the sum of atomic weights of the atoms in a molecular substance. For example, the molecular weight of a molecule of water, H_2O, is:

$$\begin{aligned} 2(\text{AW of H}) + 1(\text{AW of O}) &= (2\text{ atoms H})\left(\frac{1.01\text{ amu}}{1\text{ atom H}}\right) \\ &\quad + (1\text{ atom O})\left(\frac{16.00\text{ amu}}{1\text{ atom O}}\right) \\ &= 18.02\text{ amu} \end{aligned}$$

The percent composition of a substance refers to the percent by mass contributed by each element in the substance.

- The sum of percent by mass of all elements in a substance equals 100%.
- $\%\text{ by mass of an element} = \dfrac{\text{mass of an element in substance}}{\text{formula weight of substance}} \times 100$

Chemists do not ordinarily work with single molecules or atoms, but rather with trillions upon trillions of them. To facilitate the counting and weighing of such samples, a quantity called the mole has been defined.

- A **mole** of any type of particle equals the number of ^{12}C atoms in exactly 12 g of ^{12}C. Thus a mole represents a certain number of objects, just like a dozen represents 12.
- In 12 g of ^{12}C, there are 6.022×10^{23} atoms; this number is given the name **Avogadro's number** (symbol is N, and unit is g/mol).
- Thus a mole of water contains 6.022×10^{23} molecules of water, and a mole of NaCl contains 6.022×10^{23} sodium ions and 6.022×10^{23} chloride ions. Note that the total number of ions in one mole of NaCl is $2(6.022 \times 10^{23})$ ions or 1.204×10^{24} ions.
- The term **molar mass** is used to describe the mass in grams of one mole of a substance.

We can use mass–quantity relationships as conversion factors. Examples of equivalences that can be used are

1 ^{12}C atom = 12 amu	1 mol ^{12}C = 12 g
1 Cl_2 molecule = 70.90 amu	1 mol Cl_2 = 70.90 g
1 $BaCl_2$ formula = 208 amu	1 mol $BaCl_2$ = 208 g

Exercise 4 Calculating molecular and formula weights

Calculate the molecular or formula weights for: (**a**) NO_3^-; (**b**) $C_{21}H_{30}O_2$.

SOLUTION: *Analyze:* We are given the chemical formulas for an ion, (**a**), and a molecular substance, (**b**), and are asked to calculate their formula weights.

Plan: First look up the atomic weights of all elements in the substances. Then calculate the formula or molecular weight of each substance by multiplying each atomic weight by the number of atoms in the chemical formula and summing these numbers. This procedure gives the formula weight for (**a**) since it is an ion and the molecular weight for (**b**) since it is a molecular substance.

Solve:

(**a**)

$$\text{N: (1 atom N)}\left(\frac{14.01\text{ amu}}{1\text{ atom N}}\right) = 14.01\text{ amu}$$

$$\text{O: (3 atoms O)}\left(\frac{16.00\text{ amu}}{1\text{ atom O}}\right) = \underline{48.00\text{ amu}}$$

$$\text{Formula weight of } NO_3^- = 62.01\text{ amu}$$

(**b**)

$$\text{C: (21 atoms C)}\left(\frac{12.01\text{ amu}}{1\text{ atom C}}\right) = 252.21\text{ amu}$$

$$\text{H: (30 atoms H)}\left(\frac{1.01\text{ amu}}{1\text{ atom H}}\right) = 30.30\text{ amu}$$

$$\text{O: (2 atoms O)}\left(\frac{16.00\text{ amu}}{1\text{ atom O}}\right) = \underline{32.00\text{ amu}}$$

$$\text{Molecular weight of } C_{21}H_{30}O_2 = 314.51\text{ amu}$$

Check: We can estimate the formula weight for (**a**) by adding 14 amu and 3×16 amu and finding it is 62 amu, which is close to the calculated value in the solution. Similarly we can estimate the molecular formula for (**b**) by adding 20×12 amu and 30×1 amu and 2×20 amu and finding it is 310 amu, which is close to the calculated molecular weight.

EXERCISE 5 Calculating molecular weight, number of molecules and moles, and percentage of an element in a molecule

Answer the following with respect to ethanol, C_2H_6O: **(a)** What is its molecular weight? **(b)** What is the mass of 1 mol of ethanol molecules? **(c)** Calculate the number of moles of ethanol in 1.00 g. **(d)** Calculate the number of molecules in 1.00 g of ethanol. **(e)** Calculate the percentage of carbon in one molecule of ethanol.

SOLUTION: (a) *Analyze:* We are asked to calculate for ethanol, a molecular substance, its molecular weight.

Plan: We can use the same approach as in Exercise 4.

Solve: The molecular weight of ethanol is calculated as follows:

$$\text{C: (2 atoms C)}\left(\frac{12.01\text{ amu}}{1\text{ atom C}}\right) = 24.02\text{ amu}$$

$$\text{H: (6 atoms H)}\left(\frac{1.01\text{ amu}}{1\text{ atom H}}\right) = 6.06\text{ amu}$$

$$\text{O: (1 atom O)}\left(\frac{16.00\text{ amu}}{1\text{ atom O}}\right) = 16.00\text{ amu}$$

$$\text{Molecular weight of } C_2H_6O = 46.08\text{ amu}$$

Check: We can estimate the molecular weight by adding 2×12 amu, 6×1 amu, and 1×16 amu and finding it is 46 amu, which is close to the calculated molecular weight.

(b) *Analyze:* We are asked to calculate the mass of one mole of ethanol.

Plan: Use the definition of molar mass: The mass of one mole of a substance is its molecular weight expressed in grams.

Solve: The mass of 1 mol of ethanol is the weight of one molecule expressed in grams, 46.08 g.

(c) *Analyze*: You are given a mass of ethanol; you are asked to solve for the number of moles.

Plan: You need a conversion factor that will change 1.00 g of ethanol to moles. The molar mass relationship yields the necessary conversion factor:

$$46.08\text{ g ethanol} = 1\text{ mol ethanol}$$

Since you want to carry out the conversion *grams* → *moles* (the symbol → indicates a unit conversion), the conversion factor you need to use is 1 mol ethanol/46.08 g ethanol.

Solve: Moles ethanol $= (1.00\text{ g ethanol})\left(\dfrac{1\text{ mol ethanol}}{46.08\text{ g ethanol}}\right) = 0.0217\text{ mol ethanol}$

Check: The answer is significantly less than one mole, which is rational because the given mass of ethanol is less than the mass of one mole.

(d) *Analyze*: You are asked for the number of molecules in 1.00 g of ethanol. This will require the number of moles of ethanol which is calculated in **(b)**.

Plan: Most problems that ask for the number of particles, such as molecules or atoms, will require you at some point in the calculation to use Avogadro's number to convert the number of moles of the substance to number of particles. The two relationships needed for the conversion of *grams* → *moles* → *molecules* are:

46.08 g ethanol = 1 mol ethanol [*Molar Mass*]

1 mol ethanol = 6.022×10^{23} molecules ethanol [Avogadro's number]

Solve: $$\text{Molecules ethanol} = (1.00\text{ g ethanol})\left(\frac{1\text{ mol ethanol}}{46.08\text{ g ethanol}}\right) \times \left(\frac{6.022 \times 10^{23}\text{ molecules ethanol}}{1\text{ mol ethanol}}\right) = 1.31 \times 10^{22}\text{ molecules ethanol}$$

Check: The answer is less than 6.02×10^{23}, which is rational given that the number of moles is less than one.

(e) *Analyze*: You are asked for the percentage of carbon in one molecule of ethanol. This will require the use of the molecular weight of ethanol and the number of atoms of carbon in a molecule of ethanol.

Plan: The percentage of any element in a molecular compound is the mass of that element in one molecule divided by the molecular weight of the molecule and multiplied by 100:

$$\%\text{ C} = \frac{\left(\begin{matrix}\text{number C atoms}\\ \text{per molecule}\end{matrix}\right)(\text{AW of C})}{(\text{mass of one } C_2H_6O \text{ molecule})} \times 100$$

Solve: $$\%\text{ C} = \frac{\left(\dfrac{2\text{ atoms C}}{1\text{ molecule } C_2H_6O}\right)\left(\dfrac{12.01\text{ amu}}{1\text{ atom C}}\right)}{\dfrac{46.08\text{ amu}}{1\text{ molecule } C_2H_6O}} \times 100 = 52.13\%$$

Alternatively, the percentage of carbon in C_2H_6O can be solved for as follows:

$$\left(\frac{2\text{ mol C}}{1\text{ mol } C_2H_6O}\right)\left(\frac{12.01\text{ g C}}{1\text{ mol C}}\right)\left(\frac{1\text{ mol } C_2H_6O}{46.08\text{ g } C_2H_6O}\right) = \frac{0.5213\text{ g C}}{1\text{ g } C_2H_6O} = \frac{52.13\text{ g C}}{100\text{ g } C_2H_6O}$$

The last ratio is equivalent to saying that the percentage of carbon in C_2H_6O is 52.13%.

Check: Carbon is present in the greatest total mass in the molecule and thus it should represent a significant percentage. Also the percentage is less than 100. Except for an element by itself, the percentage of an element in a compound must be less than 100%.

EXERCISE 6 Calculating the mass of a substance given the mass of an element in its chemical formula

A sample of $Na_2B_4O_7$ contains 0.3478 g of sodium. What is the mass of this sample?

SOLUTION: *Analyze*: We are asked for the mass of a salt, $Na_2B_4O_7$, and we are given the mass of sodium in the sample.

Plan: We can calculate the mass of $Na_2B_4O_7$ by recognizing that the mass of the sample is related to the mass of the sodium in the following manner:

$$\text{Mass of sample} = (\text{mass of Na}) \times \left(\begin{matrix}\text{conversion factors that}\\ \text{change g Na to g } Na_2B_4O_7\end{matrix}\right)$$

To convert from grams Na to grams $Na_2B_4O_7$, we will need to go through the following series of conversions:

$$\textit{Grams}\text{ Na} \longrightarrow \textit{moles}\text{ Na} \longrightarrow \textit{moles}\ Na_2B_4O_7 \longrightarrow \textit{grams}\ Na_2B_4O_7$$

The required equivalences are

$1 \text{ mol } Na_2B_4O_7 = 2 \text{ mol Na}$ [from formula of $Na_2B_4O_7$]*

$1 \text{ mol } Na_2B_4O_7 = 201.24 \text{ g } Na_2B_4O_7$ [from molar mass of $Na_2B_4O_7$]

$1 \text{ mol Na} = 23.00 \text{ g Na}$ [from molar mass of Na]

Solve: $$\text{Mass of sample} = (0.3478 \text{ g Na})\left(\frac{1 \text{ mol Na}}{23.00 \text{ g Na}}\right)$$

$$\times \left(\frac{1 \text{ mol } Na_2B_4O_7}{2 \text{ mol Na}}\right)\left(\frac{201.24 \text{ g } Na_2B_4O_7}{1 \text{ mol } Na_2B_4O_7}\right)$$

$$= 1.522 \text{ g } Na_2B_4O_7$$

Check: The answer is larger than the mass of sodium, and this is rational given that the mass of any element in a compound is less than the mass of the compound.

DETERMINING EMPIRICAL AND MOLECULAR FORMULAS

In Chapter 2 we learned that an empirical formula shows the simplest whole-number ratio of atoms. A molecular formula shows the actual number of atoms. For example, CH_2 is the empirical formula for C_2H_4. An empirical formula is typically determined from percent composition data. *The subscripts in an empirical formula are calculated as follows:*

- Convert the mass percent of each element to grams using an arbitrarily chosen sample size, such as 100 g.
- Determine the number of moles of each element.
- Determine the simplest whole-number ratio of atoms in the compound by dividing the number of moles of each element by the number of moles of the element having the smallest number of moles.
- If the ratios are not whole numbers, multiply the ratios by an integer that clears the denominators of the fractions. For example, the numbers 0.50 and 1.75 are expressed as fractions: $\frac{1}{2}$ and $\frac{7}{4}\left(1\frac{3}{4}\right)$. If they are multiplied by four, the ratios are converted to whole numbers: $4 \times \frac{1}{2} = 2$ and $4 \times \frac{7}{4} = 7$. If a ratio is very near a whole number or fraction, such as 1.05 or 1.55, assume that they can be expressed as 1.00 and 1.50 because of experimental error.

To determine the molecular formula of a compound, we need its molecular weight:

- Calculate the number of empirical formula units making up the molecular formula by dividing the mass of one mole of the substance by the mass of one empirical formula.
- To determine the molecular formula, multiply the subscripts of the empirical formula by the number of empirical formula units making up the molecular formula.

EXERCISE 7 Determining the empirical and molecular formulas of a compound

A compound contains only the elements Al and O. Its elemental composition is determined to be 53.0% aluminum and 47.0% oxygen. The mass of one mole of the compound is 102 g. What is the empirical formula of the compound? What is the molecular formula?

*The conversion factor $1 \text{ mol } Na_2B_4O_7/2 \text{ mol Na}$ is included in the problem to reflect the fact that there are two sodium atoms per $Na_2B_4O_7$ formula unit.

SOLUTION: *Analyze*: We are asked to calculate the empirical and molecular formula of a compound containing Al and O. The percent composition data provides information for determining the empirical formula, and molar mass permits calculation of the molar formula.

Plan: Follow the outline for determining empirical formulas from percent composition data. First convert the mass percent of each element to grams in an arbitrarily chosen sample size, for example, 100 g. Then determine the number of moles of each element in the 100 gram sample and the ratio of moles. This gives the empirical formula. To determine the molecular formula, divide the molar mass by the empirical mass and multiply the subscripts of the empirical formula by this number.

Solve:

$$\text{Grams Al} = (\text{100-g sample})\left(\frac{53.0 \text{ g Al}}{100 \text{ g}}\right) = 53.0 \text{ g Al}$$

$$\text{Grams O} = (\text{100-g sample})\left(\frac{47.0 \text{ g O}}{100 \text{ g}}\right) = 47.0 \text{ g O}$$

Next determine the number of moles of each element in 100 g of the sample:

$$\text{Moles Al} = (53.0 \text{ g Al})\left(\frac{1 \text{ mol Al}}{27.0 \text{ g Al}}\right) = 1.96 \text{ mol Al}$$

$$\text{Moles O} = (47.0 \text{ g O})\left(\frac{1 \text{ mol O}}{16.0 \text{ g O}}\right) = 2.94 \text{ mol O}$$

To determine the empirical formula of the compound, calculate the simplest whole-number ratio of atoms in the compound. This is done by dividing the number of moles of each element by the number of moles of the element having the *smallest* number of moles. In this case Al has the fewer number of moles; thus you divide by 1.96:

$$Al_{\frac{1.96}{1.96}}O_{\frac{2.94}{1.96}} = Al_{1.00}O_{1.50}$$

The subscripts are not all integers. You must multiply them by an integer that will convert 1.50 into an integer. Inspection should convince you that if you multiply by 2 you will convert 1.50 to 3.00 and 1.00 to 2.00. The empirical formula is therefore Al_2O_3.

To determine the molecular formula, divide the mass of one mole by the mass of one empirical formula unit. The mass of one empirical formula unit for Al_2O_3 is 102 g, the same as the mass of one mole. Thus, the empirical and molecular formulas are identical, Al_2O_3.

Check: The formula Al_2O_3 is reasonable because aluminum has a charge of 3+ as an ion and oxygen has a charge of 2− as an ion; the sum of charges adds to zero for the formula and is consistent with observation about charges.

CHEMICAL EQUATIONS: MASS AND MOLE RELATIONSHIPS

A balanced chemical equation gives us information about the relative number of moles of reactants and products. The masses of substances in the reaction are determined from the mode concept. The branch of chemistry that deals with these quantitative relationships is termed **stoichiometry**.

To understand the above statements, examine the following balanced chemical reaction and see what mass–mole relationships can be derived from it:

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(l)$$

- On the atomic-molecular level, the equation states:

$$1 \text{ molecule } CH_4 + 2 \text{ molecules } O_2 \longrightarrow 1 \text{ molecule } CO_2 + 2 \text{ molecules } H_2O$$

- Or, since an Avogadro's number of molecules is equivalent to a mole of molecules:

$$1 \text{ mol } CH_4 + 2 \text{ mol } O_2 \longrightarrow 1 \text{ mol } CO_2 + 2 \text{ mol } H_2O$$

- The numerical coefficients in front of the reactants and products show the ratio of moles in which the chemical substances react. For example, because 2 mol of O_2 is required to react with 1 mol of CH_4, then we know that 4 mol O_2, are required to react with 2 mol of CH_4—that is, O_2 always reacts with CH_4 in a 2:1 mole ratio. You can represent this stoichiometry by the statement

$$1 \text{ mol } CH_4 \simeq 2 \text{ mol O}$$

where the symbol $\simeq$ means a stoichiometrically equivalent quantity *in terms of the given reaction*. Similarly, you can represent the stoichiometric ratio for the formation or products as

$$1 \text{ mol } CO_2 \simeq 2 \text{ mol } H_2O$$

That is, 1 mole of CO_2 forms for every 2 moles of water that forms.

- Other stoichiometric ratios can be derived from the balanced chemical reaction. For example:

$$\begin{aligned} 1 \text{ mol } CH_4 &\simeq 1 \text{ mol } CO_2 \\ 2 \text{ mol } O_2 &\simeq 1 \text{ mol } CO_2 \\ 1 \text{ mol } CH_4 &\simeq 2 \text{ mol } H_2O \\ 2 \text{ mol } O_2 &\simeq 2 \text{ mol } H_2O \end{aligned}$$

- All of the above stoichiometrically equivalent statements can be converted to mass equivalences by converting a mole of a substance to its molar mass.

Various kinds of chemical problems involving stoichiometry are encountered in chemical practice.

- One important type involves a **limiting reactant**. In many reactions, one or more substances are in excess and therefore some will be left over when the reaction is completed. The substance that is completely consumed determines the amount of product formed and is called the limiting reactant. Exercise 11 explores how to solve limiting reactant problems.
- Most chemical reactions are not 100% efficient; they do not produce as much product as expected from the stoichiometry. The extent of a reaction is given by **percent yield**:

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

See Exercise 11.

EXERCISE 8 Using stoichiometry to determine the amount of a product formed in a chemical reaction I

How many moles of Al_2O_3 are produced when 0.50 mol Al reacts with an excess of PbO_2? The balanced chemical equation is

$$4\ Al(s) + 3\ PbO_2(s) \rightarrow 2\ Al_2O_3(s) + 3\ Pb(s)$$

SOLUTION: *Analyze*: We are given the number of moles of Al and asked how many moles of Al_2O_3 are produced. This will require a stoichiometric conversion using the chemical reaction.

Plan: Determine the stoichiometric equivalences that allow us to make the transformation:

$$\textit{Moles}\ Al \longrightarrow \textit{moles}\ Al_2O_3$$

The stoichiometric equivalence derived from the balanced chemical equation that relates moles of Al to moles of Al_2O_3 is:

$$4\ \text{mol Al} \simeq 2\ \text{mol}\ Al_2O_3$$

The problem states that there is sufficient PbO_2 to react with 0.50 mol Al. Thus the amount of PbO_2 present does not have to be considered.

Solve:
$$\text{Moles}\ Al_2O_3 = (0.50\ \text{mol Al})\left(\frac{2\ \text{mol}\ Al_2O_3}{4\ \text{mol Al}}\right) = 0.25\ \text{mol}\ Al_2O_3$$

Thus when 0.50 mol Al reacts with an excess of PbO_2, 0.25 mol Al_2O_3 is produced.

Check: Note that the calculation uses units with the name of the substance. This ensures that the calculated value is associated with the correct substance. If you, for example, inadvertently invert the stoichiometric ratio, you will find that the units with substances do not properly cancel. This is one of the best ways of checking when doing stoichiometric problems.

EXERCISE 9 Using stoichiometry to determine the amount of a reactant needed in a chemical reaction

Determine how many grams of HI are required to form 1.20 moles of H_2 when HI reacts according to the balanced chemical equation

$$2\ HI(g) \rightarrow H_2(g) + I_2(g)$$

SOLUTION: *Analyze*: We are given the number of moles of a product, H_2, that forms and asked to determine how many grams of HI are required. Again, as in the previous problem, we will need stoichiometric ratios.

Plan: This problem requires you to work from products to reactants as follows:

$$\textit{Moles}\ H_2 \longrightarrow \textit{moles}\ HI \longrightarrow \textit{grams}\ HI$$

From this sequence of proposed conversions, we see that you need both mole and mass stoichiometric equivalences. These are

$$2\ \text{mol HI} \simeq 1\ \text{mol}\ H_2\ \text{(from balanced equation)}$$

$$127.91\ \text{g HI} = 1\ \text{mol HI (from molar mass of HI)}$$

Solve: The problem can be solved in two steps:

$$\textit{Moles}\ H_2 \longrightarrow \textit{moles}\ HI:$$

$$\text{moles HI} = (1.20\ \text{mol}\ H_2)\left(\frac{2\ \text{mol HI}}{1\ \text{mol}\ H_2}\right) = 2.40\ \text{mol HI}$$

and

$$\textit{Moles}\ \text{HI} \longrightarrow \textit{grams}\ \text{HI}:$$

$$\text{grams HI} = (2.40\ \text{mol HI})\left(\frac{127.91\ \text{g HI}}{1\ \text{mol HI}}\right) = 307\ \text{g HI}$$

Alternatively, the two steps can be combined into one as follows:

$$\text{Grams HI} = (1.20\ \text{mol}\ H_2)\left(\frac{2\ \text{mol HI}}{1\ \text{mol}\ H_2}\right)\left(\frac{127.91\ \text{g HI}}{1\ \text{mol HI}}\right) = 307\ \text{g HI}$$

In the *Student's Guide* the latter approach will be the one usually followed when solving stoichiometry problems.

Check: The units associated with the final answer have the correct substance indicated. An estimate of 2.5×128 in the last step gives about 300, which is the magnitude of the final answer.

EXERCISE 10 Using stoichiometry to determine the amount of a product formed in a chemical reaction II

Determine how many grams of PCl_3 are produced when 2.80 g of Cl_2 reacts with a sufficient quantity of P_4 according to the chemical reaction

$$P_4(s) + 6\ Cl_2(g) \longrightarrow 4\ PCl_3(l)$$

SOLUTION: The *Analysis* and *Plan* for this problem follow the same format as in the previous problems. We are given the grams of chlorine gas and asked to calculate the grams of product formed. We will need stoichiometric ratios (equivalences) as follows:

$$\textit{Grams}\ Cl_2 \longrightarrow \textit{moles}\ Cl_2 \longrightarrow \textit{moles}\ PCl_3 \longrightarrow \textit{grams}\ PCl_3$$

The necessary stoichiometric equivalence is

$$6\ \text{mol}\ Cl_2 \simeq 4\ \text{mol}\ PCl_3$$

Also, we will need the following molar mass equivalences:

$$70.90\ \text{g}\ Cl_2 = 1\ \text{mol}\ Cl_2$$

$$137.32\ \text{g}\ PCl_3 = 1\ \text{mol}\ PCl_3$$

Solve: The grams of PCl_3 produced are calculated using the sequence of conversions previously given:

$$\text{Grams}\ PCl_3 = (2.80\ \text{g}\ Cl_2)\left(\frac{1\ \text{mol}\ Cl_2}{70.90\ \text{g}\ Cl_2}\right)$$

$$\times\left(\frac{4\ \text{mol}\ PCl_3}{6\ \text{mol}\ Cl_2}\right)\left(\frac{137.32\ \text{g}\ PCl_3}{1\ \text{mol}\ PCl_3}\right) = 3.62\ \text{g}\ PCl_3$$

Check: The final unit has the correct substance, PCl_3, identified. This is the unknown in the question.

EXERCISE 11 Determining the limiting reactant in a chemical reaction and the percent yield

(a) What is the limiting reactant when 10.0 g of C_2H_6 reacts with 50.0 g of O_2 according to the chemical equation

$$2\ C_2H_6(g) + 7\ O_2(g) \longrightarrow 4\ CO_2(g) + 6\ H_2O(l)$$

(b) Calculate the percent yield of the reaction

$$2\,Al(OH)_3(s) + 3\,H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(s) + 6\,H_2O(l)$$

given that 205 g of $Al(OH)_3$ reacts with 751 g of H_2SO_4 to yield 252 g of $Al_2(SO_4)_3$.

SOLUTION: (a) *Analyze*: We are given masses of the two reactants and asked to determine the limiting reactant. This means we have to identify the limiting reactant and provide an explanation.

Plan: How do we know when a problem involves a limiting reactant? If the quantities of two or more reactants are given, then we should assume it is a limiting reactant problem until we show otherwise. If the quantity of only one reactant is given, then we can assume all other reactants are in excess. There are two general methods for doing limiting reactant problems:

1. Choose one of the reactants and calculate the stoichiometric quantities required for the other reactants to react with it. We then compare the calculated quantities to the given quantities to determine the limiting reactant. If a given quantity is smaller than the calculated quantity, then that reactant is the limiting reactant. If all calculated quantities are equal to or smaller than the given quantities, then the reference reactant used to calculate all other quantities is the limiting reactant.
2. Choose a product in the reaction. If there is a question about theoretical yield, choose the product needed to calculate the theoretical yield. Do not change the chosen product in the following calculations—it must be fixed for proper interpretation of the results. Use each reactant and separately calculate the theoretical amount of the chosen product. *The reactant yielding the smallest quantity of the chosen product is the limiting reactant.*

Procedure (1) above will be used to solve this problem. Calculate the exact mass of O_2 required to react with 10.0 g of C_2H_6. If the available mass of oxygen is greater than the calculated mass of oxygen, then C_2H_6 is the limiting reactant. Conversely, oxygen is the limiting reactant if its available mass is less than the calculated mass. The sequence of conversions required to make this determination is

$$\textit{Grams}\ C_2H_6 \rightarrow \textit{moles}\ C_2H_6 \rightarrow \textit{moles}\ O_2 \rightarrow \textit{grams}\ O_2$$

The required stoichiometric and molar mass equivalences needed for these conversions are

$$2\text{ mol }C_2H_6 \simeq 7\text{ mol }O_2$$

$$30.08\text{ g }C_2H_6 = 1\text{ mol }C_2H_6$$

$$32.00\text{ g }O_2 = 1\text{ mol }O_2$$

Solve: Following the sequence of steps shown above,

$$\text{Grams }O_2 = (10.0\text{ g }C_2H_6)\left(\frac{1\text{ mol }C_2H_6}{30.08\text{ g }C_2H_6}\right)$$

$$\times\left(\frac{7\text{ mol }O_2}{2\text{ mol }C_2H_6}\right)\left(\frac{32.00\text{ g }O_2}{1\text{ mol }O_2}\right) = 37.2\text{ g }O_2$$

The available mass of O_2, 50.0 g, is greater than the calculated mass (37.2 g) of O_2 required to react completely with 10.0 g of C_2H_6. Therefore, O_2 is in excess, and C_2H_6 is the limiting reagent.

Check: The magnitude of the answer is in agreement with a rough estimation for the calculation: $10\left(\frac{1}{30}\right)\left(\frac{7}{2}\right)\left(\frac{30}{1}\right) = 35$. The units are also correct.

(b) *Analyze* and *Plan*: This problem asks for the percent yield of one of the products; thus, we will need to calculate the theoretical amount of aluminum sulfate. It also may be a limiting reactant problem because we are given the masses of the reactants. To calculate the theoretical amount of aluminum sulfate and also to determine the limiting reactant, we will use procedure (2) because it requires us to calculate the amount of a product. Each reactant is assumed to be the limiting reactant and the theoretical amount of aluminum sulfate is calculated. The limiting reactant will produce the *fewer* number of moles of aluminum sulfate.

Solve: First, assume aluminum hydroxide is the limiting reactant and convert the grams of aluminum hydroxide to grams of aluminum sulfate:

$$\text{g } Al_2(SO_4)_3 = (205 \text{ g } Al(OH)_3)\left(\frac{1 \text{ mol } Al(OH)_3}{78.0 \text{ g}}\right)$$

$$\times\left(\frac{1 \text{ mol } Al_2(SO_4)_3}{2 \text{ mol } Al(OH)_3}\right)\left(\frac{342.1 \text{ g } Al_2(SO_4)_3}{1 \text{ mol } Al_2(SO_4)_3}\right) = 450 \text{ g } Al_2(SO_4)_3$$

Then assume sulfuric acid is the limiting reactant and determine the grams of aluminum sulfate produced:

$$\text{g } Al_2(SO_4)_3 = (751 \text{ g } H_2SO_4)\left(\frac{1 \text{ mol } H_2SO_4}{98.1 \text{ g}}\right)$$

$$\times\left(\frac{1 \text{ mol } Al_2(SO_4)_3}{3 \text{ mol } H_2SO_4}\right)\left(\frac{342.1 \text{ g } Al_2(SO_4)_3}{1 \text{ mol } Al_2(SO_4)_3}\right) = 873 \text{ g } Al_2(SO_4)_3$$

$Al(OH)_3$ is the limiting reagent because it yields the smaller amount of $Al_2(SO_4)_3$ and thus the theoretical amount of $Al_2(SO_4)_3$ is 450 g. The percent yield is

$$\text{Percent yield} = \frac{252 \text{ g}}{450 \text{ g}} \times 100 = 56.0\%$$

Check: The magnitudes of the two answers agree with a rough calculation for each step, $200\left(\frac{1}{80}\right)\left(\frac{1}{2}\right)\left(\frac{350}{1}\right) \simeq 400$ and $750\left(\frac{1}{100}\right)\left(\frac{1}{3}\right)\left(\frac{350}{1}\right) \simeq 800$. The units are also correct.

Note: In Exercises 8 through 11 every sequence of conversions involved moles. The unit mole is the central focus point for most stoichiometric conversions. The reason is that if you know the number of moles of a substance, you can transform it to either grams or number of particles.

$$\boxed{\text{Number of particles}} \underset{\text{Avogadro's number}}{\rightleftarrows} \boxed{\text{Moles}} \underset{\text{Molar mass}}{\rightleftarrows} \boxed{\text{Grams}}$$

Avogadro's number is used to convert between number of particles and number of moles. The mass of one mole of a substance (molar mass) is used to convert between number of moles and grams. Keep the above picture in mind when you do stoichiometric problems.

SELF-TEST QUESTIONS

Key Terms

Having reviewed key terms in Chapter 3, match key terms with phrases and identify statements as true or false. If a statement is false, indicate why it is incorrect.

Match each phrase with the best term:

3.1 Equal numbers of atoms of each element occur on both sides of a reaction arrow.

3.2 Oxygen is typically a reactant in this type of reaction.

3.3 6.022×10^{23} is given this historical name.

3.4 The name given to the mass in amu of a compound containing all nonmetals.

3.5 The name given to the mass in amu of any compound.

3.6 In the following reaction, when 6 mol of Cu and 18 mol of HNO_3 are initially present, we give this term to Cu.

$$3\,Cu + 8\,HNO_3 \longrightarrow 3\,Cu(NO_3)_2 + 2\,NO + 4\,H_2O$$

3.7 The total mass of materials in a chemical reaction does not change.

3.8 The name given to the maximum amount of a product obtainable in a chemical reaction.

Terms:

(a) Avogadro's number
(b) balanced chemical reaction
(c) combustion reaction
(d) formula weight
(e) law of conservation of mass
(f) limiting reactant
(g) molecular weight
(h) theoretical yield

True-False Statements:

3.9 The area of study of *stoichiometry* involves measuring concentrations and percent yield, for example.

3.10 *Reactants* can be gases and liquids.

3.11 *Products* can be liquids and solids.

3.12 To determine if a reaction is a *combination reaction*, you would look to see if two or more products are formed.

3.13 To determine if a reaction is a *decomposition reaction*, you would look to see if two or more reactants exist.

3.14 A *mole* is a direct measure of mass.

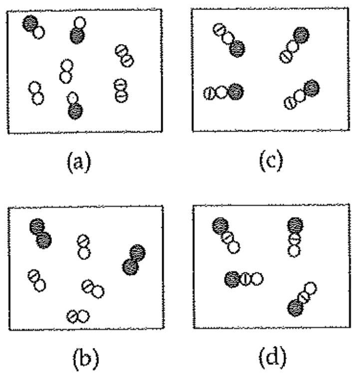
3.15 The mass of one mole of a substance is known as *molar mass*.

3.16 Percent yield is defined as

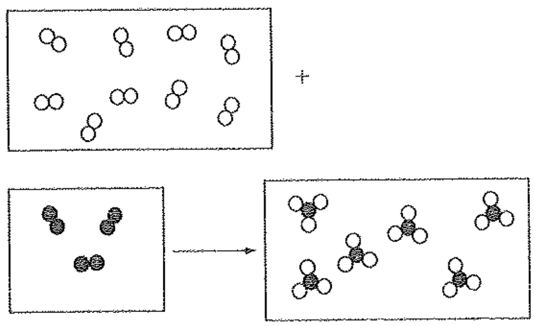
$$\frac{\text{amount of product formed}}{\text{initial amount of reactant}} \times 100$$

Problems and Short-Answer Questions

3.17 In the following boxes unshaded spheres represent oxygen atoms, spheres with a line represent nitrogen atoms and fully shaded spheres represent bromine. Which box contains spheres that represent reactants and which box represents the product for the following reaction: $2\,NO(g) + Br_2(g) \longrightarrow 2\,NOBr(g)$?



3.18 Unshaded spheres represent atom X and shaded spheres represent atom Y in the following diagram. The diagram shows reactants converted to product in the correct stoichiometric ratio. How many moles of product form if 2.0 mol X_2 react with 1.0 mol Y_2? If there is a limiting reactant, identify it.



3.19 Balance the following equations:

(a) $AgNO_3(aq) + CaCl_2(aq) \longrightarrow AgCl(s) + Ca(NO_3)_2(aq)$
(b) $VO(s) + Fe_2O_3(s) \longrightarrow FeO(s) + V_2O_5(s)$
(c) $Na(s) + H_2O(l) \longrightarrow NaOH(aq) + H_2(g)$
(d) $NH_4NO_3(s) \longrightarrow N_2O(g) + H_2O(l)$
(e) $MnO_2(s) + HCl(ag) \longrightarrow Cl_2(g) + MnCl_2(aq) + H_2O(l)$

3.20 Write balanced equations for the following reactions:

(a) aqueous silver nitrate reacts with aqueous copper(II) chloride to form insoluble silver chloride and aqueous copper(II) nitrate;
(b) metallic aluminum reacts with oxygen gas to form solid aluminum oxide;
(c) aqueous barium chloride reacts with aqueous potassium sulfate to form solid barium sulfate and aqueous potassium chloride;
(d) solid magnesium chloride reacts with aqueous sodium hydroxide to yield insoluble magnesium hydroxide and aqueous sodium chloride;

(e) solid potassium chlorate decomposes to solid potassium chloride and oxygen gas.

3.21 A sample of a B,H-containing compound is analyzed and it contains 78.2% boron. Is this sufficient information to determine the empirical and molecular formulas of the compound? Explain without doing calculations. In your explanation identify any information that is needed and why.

3.22 Freon-12 has the formula CCl_2F_2. If you are told that a sample has 3.00×10^{20} molecules of freon-12, is this sufficient information to determine the mass of the sample? Explain without doing calculations. In your explanation identify any information that is needed and why.

3.23 You are told that the percent yield for the following reaction is 79.1%: $C_6H_{12}O_6(s) \longrightarrow 2\,C_6H_6O(l) + CO_2(g)$. Is this sufficient information to determine the amount of $C_6H_{12}O_6(s)$ that reacted? Explain without doing calculations. In your explanation identify any information that is needed and why.

3.24 Ethylene, C_2H_4, is used to make the plastic polyethylene.

(a) What are its molecular weight and formula weight?

(b) How many moles of C_2H_4 are there in 3.20×10^{-2} g?

(c) How many molecules of C_2H_4 are there in 3.20×10^{-2} g?

3.25 If a certain quantity of NO(*g*) has a mass of 5.00 g, what is the quantity of H_2O containing the same number of molecules?

3.26 What is the density of $F_2(g)$ in grams per cubic centimeter if at 0 °C one mole of it occupies 22.4 L?

3.27 Answer the following questions using the reaction

$$2\,KClO_3(s) \longrightarrow 2\,KCl(s) + 3\,O_2(g)$$

(a) How many moles of $KClO_3$ are required to produce 9 moles of O_2?

(b) How many moles of KCl are produced if 4.0 moles of O_2 form?

(c) How many moles of KCl form from the decomposition of 3.0 moles of $KClO_3$?

3.28 The main constituent of gallstones is cholesterol. Cholesterol may have a role in heart attacks and blood clot formation. Its elemental percentage composition is 83.87% C, 11.99% H, and 4.14% O. It has a molecular weight of 386.64 amu. Calculate its empirical and molecular formulas.

3.29 What is the empirical formula of compound such that 200.0 g of it contains 87.2 g of P and 112.8 g of O?

3.30 A stimulant of the nervous system found in coffee, tea, and cola is caffeine. Caffeine contains the following weight percentage of elements: 49.5% C; 28.9% N; 16.5% O; 5.2% H. What is the empirical formula of caffeine?

3.31 $H_2C_2O_4$ and $KMnO_4$ react according to the equation

$$5\,H_2C_2O_4 + 2\,KMnO_4 \longrightarrow 4\,H_2O + 10\,CO_2 + 2\,MnO_2 + 2\,KOH$$

How many grams of CO_2 are formed when 10.05 g of $H_2C_2O_4$ and 26.72 g of $KMnO_4$ are mixed together?

3.32 Silicon carbide, SiC, is an important industrial abrasive. It is formed by the reaction of SiO_2 and carbon at high temperatures:

$$SiO_2 + 3\,C \longrightarrow SiC + 2\,CO$$

(a) Calculate the number of moles of silicon carbide formed when 5.00 g of carbon reacts with an excess of SiO_2.

(b) What is the minimum amount of carbon required to react with 25.0 g of SiO_2?

3.33 The reaction for the production of iron from the reduction of the ore hematite, Fe_2O_3, is as follows:

$$Fe_2O_3 + 3\,CO \longrightarrow 2\,Fe + 3\,CO_2$$

(a) If the reaction yields 4.52 g of CO_2, how many grams of Fe are also formed?

(b) How many grams of Fe are formed from 7.25 g of Fe_2O_3 and 6.00 g of CO?

3.34 Answer the following questions using the reaction

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

(a) If 1.0 mol Na and 1.0 mol H_2O react and the yield is 85.0%, is there a limiting reactant? Does the reaction go to completion?

(b) If 4.0 mol Na reacts with 6 mol H_2O, what is the limiting reactant and how many moles of the other reactant remain?

(c) If in problem (b) you had been given 5.0 g Na and 10.0 g H_2O and you were asked to solve for the number of grams of excess reactant, how would the problem have been solved? Explain in words without doing a numerical calculation.

Integrative Exercises

3.35 A 3.15 g sample of KCl contains chlorine-35 and chlorine-37. The fractional abundance of chlorine-35 is 0.75771. This sample is reacted with excess $AgNO_3$ in water to form AgCl in a yield of 85.5%. (a) What is the mass of AgCl? (b) What is the mass of the chlorine-35 atoms in the AgCl?

3.36 A student prepared a sample of an unknown substance containing only nitrogen and oxygen. After studying the composition of the substance the student named it nitrogen (IV) oxide. Other students also studied the substance and found that it contained 30.45% N and 69.55% O and had a molar mass of 91.98 g. Did the original student properly name the compound? Explain.

3.37 Urea is an important compound because it is used as a fertilizer. It is commercially produced from two gases, each containing two different elements; water is the other product formed. The carbon-containing gas is colorless, a major component of exhaled air, and extinguishes a flame. The percentage composition of urea is 20.0% C, 46.6% N, 26.6% O, and 6.7% H, and it has a molar mass of 60.1 g. Given this information and your knowledge of gases obtained from the text, determine the two gases used in the production of urea and write the balanced chemical equation.

3.38 How do the formula or molecular weights of the following differ: $C^{16}O_2$, $C^{18}O_2$, and $K_2^{18}O$? Explain.

Multiple-Choice Questions

3.39 Which of the following equations does not obey the law of conservation of mass?

1. $C_6H_{12}O_6 + 6\,O_2 \rightarrow 6\,CO_2 + 6\,H_2O$
2. $C_2H_6 + 7\,O_2 \rightarrow 4\,CO_2 + 6\,H_2O$
3. $BCl_3 + H_2O \rightarrow H_3BO_3 + 3\,HCl$

(a) (1) (d) (1) and (2)
(b) (2) (e) (2) and (3)
(c) (3)

3.40 0.400 moles of a substance weighs 17.6 g. What is its molar mass?

(a) 53.2 g/mol (d) 7.04 g/mol
(b) 44.0 g/mol (e) 3.52 g/mol
(c) 17.6 g/mol

3.41 What is the weight percentage of Al in Al_2O_3?

(a) 26.46% (d) 47.08%
(b) 20.93% (e) 53.21%
(c) 52.92%

3.42 In photosynthesis, $CO_2(g)$ and $H_2O(l)$ are converted into glucose, $C_6H_{12}O_6$ (a sugar), and O_2. If 0.256 mol of $C_6H_{12}O_6$ is formed by the reaction of CO_2 with water, how many grams of CO_2 would be needed?

(a) 67.6 g (d) 76.6 g
(b) 11.3 g (e) 0.256 g
(c) 0.0349 g

3.43 Which of the following pairs of substances, with their indicated quantities, contain(s) the same number of atoms: (1) 0.50 mol HCl, 0.50 mol He; (2) 0.20 mol H_3PO_4, 0.80 mol N_2; (3) 0.45 mol HNO_3, 0.45 mol HNO_2?

(a) (1) (d) all of them
(b) (2) (e) none of them
(c) (3)

3.44 What is the mass in grams of 4.25×10^{20} molecules of H_2O?

(a) 0.127 g (d) 142 g
(b) 0.0127 g (e) 4.25×10^{20} g
(c) 1420 g

3.45 Which of the following is the limiting reagent if three moles of H_2O and eight moles of NO_2 are available in the reaction:

$$3\,NO_2(g) + H_2O(l) \longrightarrow 2\,HNO_3(l) + NO$$

(a) NO_2 (c) HNO_3
(b) H_2O (d) NO

3.46 What is the percentage yield of C_2H_2 when 50.0 g of $CaC_2(s)$(molar mass = 64.01 g) reacts with an excess of water to yield 13.5 g of C_2H_2O(molar mass = 26.04 g) according to the following reaction:

$$CaC_2(s) + 2\,H_2(l) \longrightarrow Ca(OH)_2(s) + C_2H_2(g)$$

(a) 27.0% (c) 66.5%
(b) 51.8% (d) 82.5%

3.47 What is the empirical formula of benzoic acid if it has a mass percentage composition of 69% carbon, 5% hydrogen, and 26% oxygen?

(a) $C_7H_6O_2$ (c) C_5H_6O
(b) $C_6H_7O_2$ (d) C_3H_4O

3.48 What is the coefficient in front of CO_2 when the following chemical equation is balanced?

$$C_3H_7OH(l) + O_2(g) \longrightarrow CO_2(g) + H_2O(l)$$

(a) 2 (c) 6 (e) 9
(b) 4 (d) 7

3.49 Which characteristic of a chemical formula typically informs you that it is *not* a molecular formula?

(a) All elements have no subscripts (that is, a subscript of one).
(b) All elements have subscripts greater than one.
(c) All elements are the same type.
(d) One of the elements is a metal.
(e) One of the elements is a nonmetal.

SELF-TEST SOLUTIONS

3.1 (b) **3.2** (c) **3.3** (a) **3.4** (g) **3.5** (d) **3.6** (f) **3.7** (e) **3.8** (h) **3.9** True. **3.10** True. Also, solids. **3.11** True. Also gases. **3.12** False. A single product is formed from two reacting substances. **3.13** False. A single reactant decomposes. **3.14** False. It is a counting number—like a dozen. It is not a measure of mass, although if you know the number of moles of a chemical substance and its formula, you can calculate its mass. **3.15** True. **3.16** False. It is

$$\frac{\text{amount of product formed}}{\text{theoretical amount of product}} \times 100$$

3.17 Box **b** represents the reactants and box **c** represents the product. Note in box **d** that molecules are of the type BrNO, which is not the correct arrangement of atoms shown in the chemical formula of the product, NOBr.

3.18 The first box has nine molecules of X_2 and the second box has three molecules of Y_2. These react to form six molecules of YX_3 in the third box. This can be written as $9X_2 + 3Y_2 \rightarrow 6YX_3$. Dividing by three gives the balanced chemical equation: $3X_2 + Y_2 \rightarrow 2YX_3$. The limiting reactant is X_2 because only 2.0 moles of X_2 are available to react with 1.0 mole of Y_2. The stoichiometry of the chemical reaction requires 3 mol of X_2 to react with 1 mol of Y_2. The number of moles of YX_3 formed is $(2.0\text{ mol } X_2)(2\text{ mol } YX_3/3\text{ mol } X_2) = 1.3\text{ mol } YX_3$.

3.19 (a) $2\,AgNO_3(aq) + CaCl_2(aq) \rightarrow 2\,AgCl(s) + Ca(NO_3)_2(aq)$

(b) $2\,VO(s) + 3\,Fe_2O_3(s) \rightarrow 6\,FeO(s) + V_2O_5(s)$
(c) $2\,Na(s) + 2\,H_2O(l) \rightarrow 2\,NaOH(aq) + H_2(g)$
(d) $NH_4NO_3(s) \rightarrow N_2O(g) + 2\,H_2O(l)$

(e) $MnO_2(s) + 4\,HCl(aq) \rightarrow Cl_2(g) + MnCl_2(aq) + 2\,H_2O(l)$

3.20 (a) $2\,AgNO_3(aq) + CuCl_2(aq) \rightarrow 2\,AgCl(s) + Cu(NO_3)_2(aq)$

(b) $4\,Al(s) + 3\,O_2(g) \rightarrow 2\,Al_2O_3(s)$

(c) $BaCl_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,KCl(aq)$

(d) $MgCl_2(s) + 2\,NaOH(aq) \rightarrow Mg(OH)_2(s) + 2\,NaCl\,(aq)$

(e) $2\,KClO_3(s) \rightarrow 2\,KCl(s) + 3\,O_2(g)$

3.21 The empirical formula can be determined by (1) calculating the percent oxygen in the compound by subtracting the 78.2% of boron from 100%, (2) determining the mass of each element in an arbitrary sample size from the percent composition data, (3) calculating the relative number of moles of each element in the sample and (4) finally the empirical formula. However, the molecular formula cannot be determined without knowing the molar mass of the compound.

3.22 There is sufficient information. Avogadro's number is used to calculate the number of moles of freon-12 $[(\text{no. molecules})\left(\frac{1\text{ mol}}{6.02 \times 10^{23}\text{ molecules}}\right)]$. The molar mass is calculated from the chemical formula and the mass of the sample is therefore (no. moles)(molar mass).

3.23 There is insufficient information. Percent yield $= 79.1\% = \frac{\text{amount of product formed}}{\text{theoretical amount of product that should form}} \times 100.$

There are two unknowns in the relationship and neither is provided. If you are given the actual amount of product that forms, you can then calculate the theoretical amount of product that should form. You can then do a stoichiometric conversion to calculate the original quantity of the reactant using the theoretical amount of product:

$$\textit{g product} \longrightarrow \textit{mol product} \longrightarrow \textit{mol reactant} \longrightarrow \textit{g reactant.}$$

3.24 (a) Molecular and formula weights are the same; 28.04 amu.

(b) $$\text{Moles } C_2H_4 = (3.20 \times 10^{-2}\text{ g } C_2H_4) \times \left(\frac{1\text{ mol } C_2H_4}{28.04\text{ g } C_2H_4}\right) = 1.14 \times 10^{-3}\text{ mol } C_2H_4$$

(c) $$\text{Molecules } C_2H_4 = (1.14 \times 10^{-3}\text{ mol } C_2H_4) \times \left(\frac{6.022 \times 10^{23}\text{ molecules } C_2H_4}{1\text{ mol } C_2H_4}\right) = 6.87 \times 10^{20}\text{ molecules } C_2H_4$$

3.25 Since equal numbers of moles contain equal numbers of molecules, we want a quantity of H_2O such that it is equivalent in number of moles to the number of moles of NO in the given sample weight. The conversions required are *grams NO → moles NO → moles H_2O → grams H_2O*. The required equivalences are 1 mol H_2O = 1 mole NO; 30.01 g NO = 1 mol NO; 18.02 g H_2O = 1 mol H_2O. Hence,

$$\text{Grams } H_2O = (5.00\text{ g NO})\left(\frac{1\text{ mol NO}}{30.01\text{ g NO}}\right)\left(\frac{1\text{ mol } H_2O}{1\text{ mol NO}}\right) \times \left(\frac{18.02\text{ g } H_2O}{1\text{ mol } H_2O}\right) = 3.00\text{ g } H_2O$$

3.26 The definition of density is: density = mass/volume. The mass of 1 mol of F_2 occupying 22.4 L is the molecular weight of F_2 in grams, which equals 38.0 g. The density of $F_2(g)$ is

$$\text{Density} = \frac{38.0\text{ g}}{(22.4\text{ L})\left(\frac{1000\text{ mL}}{1\text{ L}}\right)\left(\frac{1\text{ cm}^3}{1\text{ mL}}\right)} = 1.70 \times 10^{-3}\,\frac{\text{g}}{\text{cm}^3}$$

3.27 (a) $$\text{mol } KClO_3 = (9.0\text{ mol } O_2)\left(\frac{2\text{ mol } KClO_3}{3\text{ mol } O_2}\right) = 6.0\text{ mol } KClO_3$$

(b) $$\text{mol KCl} = (4.0\text{ mol } O_2)\left(\frac{2\text{ mol } KCl_3}{3\text{ mol } O_2}\right) = 2.7\text{ mol KCl}$$

(c) $$\text{mol KCl} = (3.0\text{ mol } KClO_3)\left(\frac{2\text{ mol KCl}}{2\text{ mol } KClO_3}\right) = 3.0\text{ mol KCl}$$

3.28 The number of moles of each element in an arbitrary sample size of 100 g is as follows:

$$\text{Moles C} = (83.37\text{ g C})\left(\frac{1\text{ mol C}}{12.01\text{ g C}}\right) = 6.983\text{ mol C}$$

$$\text{Moles H} = (11.99\text{ g H})\left(\frac{1\text{ mol H}}{1.008\text{ g H}}\right) = 11.89\text{ mol H}$$

$$\text{Moles O} = (4.14\text{ g O})\left(\frac{1\text{ mol O}}{16.00\text{ g O}}\right) = 0.260\text{ mol O}$$

The mole ratio of O:H:C is 0.260/0.260:11:89/0.260:6.983/0.260, or 1:45.7:26.9. The empirical formula is $C_{27}H_{46}O$. The empirical formula weight is 386.64 amu. This is the same as the actual molecular weight: thus the molecular formula is $C_{27}H_{46}O$.

3.29 From the masses of P and O in the sample, we can calculate the number of moles of each in the sample:

$$\text{Moles P} = (87.2\text{ g P})\left(\frac{1\text{ mol P}}{30.97\text{ g P}}\right) = 2.82\text{ mol P}$$

$$\text{Moles O} = (112.8\text{ g O})\left(\frac{1\text{ mol O}}{16.00\text{ g O}}\right) = 7.05\text{ mol O}$$

The mole ratio of P to O is 2.82/2.82:7.05/2.82, or 1:2.5, or 1: $2\frac{1}{2}$ or 2:5. The empirical formula is P_2O_5.

3.30 The moles of each element in an arbitrary sample size of 100 g are as follows:

$$\text{Moles C} = (49.5\text{ g C})\left(\frac{1\text{ mol C}}{12.01\text{ g C}}\right) = 4.12\text{ mol C}$$

$$\text{Moles N} = (28.9\text{ g N})\left(\frac{1\text{ mol N}}{14.01\text{ g N}}\right) = 2.06\text{ mol N}$$

$$\text{Moles O} = (16.5\text{ g O})\left(\frac{1\text{ mol O}}{16.0\text{ g O}}\right) = 1.03\text{ mol O}$$

$$\text{Moles H} = (5.2\text{ g H})\left(\frac{1\text{ mol H}}{1.01\text{ g H}}\right) = 5.15\text{ mol H}$$

The ratio of moles of O:N:C:H is 1.03/1.03:2.03/1.03:4.12/1.03: 5.15/1.03, or 1:2:4:5. The empirical formula is $C_4H_5N_2O$.

3.31 Reaction equation: $5\,H_2C_2O_4 + 2\,KMnO_4 \rightarrow 4\,H_2O + 10\,CO_2 + 2\,MnO_2 + 2\,KOH$. Let us determine if $KMnO_4$ is in excess. Conversions required: *Grams* $H_2C_2O_4 \rightarrow$ *moles* $H_2C_2O_4 \rightarrow$ *moles* $KMnO_4 \rightarrow$ *grams* $KMnO_4$. Equivalences needed for this calculation: $5\text{ mol }H_2C_2O_4 \hat{=} 2\text{ mol }KMnO_4$; $90.4\text{ g }H_2C_2O_4 = 1\text{ mol }H_2C_2O_4$; $158.0\text{ g }KMnO_4 = 1\text{ mol }KMnO_4$. Calculating the amount of $KMnO_4$ required to react with $10.05\text{ g }H_2C_2O_4$:

$$\text{Grams }KMnO_4 = (10.05\text{ g }H_2C_2O_4)\left(\frac{1\text{ mol }H_2C_2O_4}{90.04\text{ g }H_2C_2O_4}\right) \times \left(\frac{2\text{ mol }KMnO_4}{5\text{ mol }H_2C_2O_4}\right)\left(\frac{158.0\text{ g }KMnO_4}{1\text{ mol }KMnO_4}\right) = 7.06\text{ g }KMnO_4$$

$KMnO_4$ is in excess because 26.72 g is available but only 7.06 g is required. To calculate the mass of CO_2 formed, we must make the following conversions: *grams* $H_2C_2O_4 \rightarrow$ *moles* $H_2C_2O_4 \rightarrow$ *moles* $CO_2 \rightarrow$ *grams* CO_2. Additional equivalences needed are $5\text{ mol }H_2C_2O_4 \hat{=} 10\text{ mol }CO_2$; $44.01\text{ g }CO_2 = 1\text{ mol }CO_2$. Hence,

$$\text{Grams }CO_2 = (10.05\text{ g }C_2H_2O_4)\left(\frac{1\text{ mol }H_2C_2O_4}{90.05\text{ g }H_2C_2O_4}\right) \times \left(\frac{10\text{ mol }CO_2}{5\text{ mol }H_2C_2O_4}\right)\left(\frac{44.1\text{ g }CO_2}{1\text{ mol }CO_2}\right) = 9.825\text{ g }CO_2$$

3.32 Reaction equation: $SiO_2 + 3\,C \rightarrow SiC + 2\,CO$.

(a) Conversions: *grams C* $\rightarrow$ *moles C* $\rightarrow$ *moles SiC*. Equivalences needed: $12.01\text{ g C} = 1\text{ mol C}$; $3\text{ mol C} \hat{=} 1\text{ mol SiC}$.

$$\text{Moles SiC} = (5.00\text{ g C})\left(\frac{1\text{ mol C}}{12.01\text{ g C}}\right)\left(\frac{1\text{ mol SiC}}{3\text{ mol C}}\right) = 0.139\text{ mol SiC}$$

(b) Conversions: *Grams* $SiO_2 \rightarrow$ *moles* $SiO_2 \rightarrow$ *moles C* $\rightarrow$ *grams C*. Additional equivalences needed: $1\text{ mol }SiO_2 \hat{=} 3\text{ mol C}$; $60.09\text{ g }SiO_2 = 1\text{ mol }SiO_2$.

$$\text{Grams C} = (25.00\text{ g }SiO_2)\left(\frac{1\text{ mol }SiO_2}{60.09\text{ g }SiO_2}\right) \times \left(\frac{3\text{ mol C}}{1\text{ mol }SiO_2}\right)\left(\frac{12.01\text{ g C}}{1\text{ mol C}}\right) = 14.99\text{ g C}$$

3.33 Reaction equation: $Fe_2O_3 + 3\,CO \rightarrow 2\,Fe + 3\,CO_2$.

(a) Conversions: *Grams* $CO_2 \rightarrow$ *moles* $CO_2 \rightarrow$ *moles Fe* $\rightarrow$ *grams Fe*. Equivalences needed: $44.01\text{ g }CO_2 = 1\text{ mol }CO_2$; $2\text{ mol Fe} \hat{=} 3\text{ mol }CO_2$; $55.85\text{ g Fe} = 1\text{ mol Fe}$.

$$\text{Grams Fe} = (4.52\text{ g }CO_2)\left(\frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2}\right) \times \left(\frac{2\text{ mol Fe}}{3\text{ mol }CO_2}\right)\left(\frac{55.85\text{ g Fe}}{1\text{ mol Fe}}\right) = 3.82\text{ g Fe}$$

(b) This is another limiting reactant problem. Let us calculate the amount of CO that reacts with 7.25 g of Fe_2O_3. The required conversions are: *Grams* $Fe_2O_3 \rightarrow$ *moles* $Fe_2O_3 \rightarrow$ *moles CO* $\rightarrow$ *grams CO*. Equivalences needed: $159.7\text{ g }Fe_2O_3 = 1\text{ mol }Fe_2O_3$; $1\text{ mol }Fe_2O_3 \hat{=} 3\text{ mol CO}$; $28.01\text{ g CO} = 1\text{ mol CO}$.

$$\text{Grams CO} = (7.25\text{ g }Fe_2O_3)\left(\frac{1\text{ mol }Fe_2O_3}{159.7\text{ g }Fe_2O_3}\right) \times \left(\frac{3\text{ mol CO}}{1\text{ mol }Fe_2O_3}\right)\left(\frac{28.01\text{ g CO}}{1\text{ mol CO}}\right) = 3.81\text{ g CO}$$

Since 6.00 g of CO is available and only 3.82 g is required, CO is in excess. To calculate the amount of Fe formed, we need to make the following conversions. *Grams* $Fe_2O_3 \rightarrow$ *moles* $Fe_2O_3 \rightarrow$ *moles Fe* $\rightarrow$ *grams Fe*. Additional equivalences needed: $1\text{ mol }Fe_2O_3 \hat{=} 2\text{ mol Fe}$; $55.85\text{ g Fe} = 1\text{ mol Fe}$.

$$\text{Grams Fe} = (7.25\text{ g }Fe_2O_3)\left(\frac{1\text{ mol }Fe_2O_3}{159.7\text{ g }Fe_2O_3}\right) \times \left(\frac{2\text{ mol Fe}}{1\text{ mol }Fe_2O_3}\right)\left(\frac{55.85\text{ g Fe}}{1\text{ mol Fe}}\right) = 5.07\text{ g Fe}$$

3.34 (a) The reactants combine in a one-to-one mole ratio, and the given ratio of moles of the reactants is 1.0, thus there is no limiting reactant. The percent yield tells you that only 85% of the theoretical maximum amount of the products formed; the reaction did not go to completion.

(b) The ratio of the given number of moles is $\frac{6.0\text{ mol }H_2O}{4.0\text{ mol Na}}$ or 1.5. This is larger than the stoichiometric ratio of 1.0. The larger ratio occurs because the numerator is relatively larger than the denominator term. Thus, water must be present in a larger quantity than needed. Na is the limiting reactant. The number of moles of water that reacts is also 4.0 mol because the stoichiometric equivalence is $\frac{2.0\text{ mol }H_2O}{2.0\text{ mol Na}}$. Therefore, the number of moles of water remaining unreacted is $6.0\text{ mol }H_2O - 4.0\text{ mol }H_2O = 2.0\text{ mol }H_2O$.

(c) The only difference would be an extra step in converting the number of grams of each reactant to number of moles. You would do the same analysis using the number of moles of reactants to determine the limiting reactant and the number of moles of excess reactant. Finally, you would convert the number of moles of excess reactant to grams of excess reactant.

3.35 (a) The number of moles of KCl is (3.15 g) (1 mol/74.55 *g*) = 0.0423 mol KCl. The number of moles of Cl in the sample is the same because there is a 1:1 mole ratio of K:Cl. The number of grams of Cl in the sample is (0.0423 mol)(35.45 g/1 mol) = 1.50 g Cl. The theoretical quantity of AgCl is the amount produced assuming the reaction yields 100%. Therefore, the number of moles of Cl in the theoretical sample is 0.0423 mol Cl assuming all of the chlorine atoms in KCl are combined with silver atoms. However, since the yield is less than 100%, not all chlorine atoms are combined with silver atoms. The number of moles of Cl combined with AgCl is equal to: (percent yield in fractional form)(theoretical number of moles of Cl in AgCl) = (0.855)(0.0423 mol Cl) = 0.0362 mol Cl. This number also represents the number of moles of AgCl formed since the Ag:Cl mole ratio is 1:1. The number of grams of AgCl formed is thus: (0.0362 molAgCl) (143.32 *g*/1 mol) = 5.19 g AgCl.

(b) The number of grams of chlorine-35 in the sample of AgCl is: (fractional abundance of Cl-35) × (mass Cl in *AgCl sample*) = (0.75771)(0.0362 mol Cl)(35.45 *g*/1 mol) = 0.972 *g Cl*-35.

3.36 To determine if the correct name is used, you need to determine the molecular formula of the unknown substance. First calculate the empirical formula by assuming a 100.0 g sample:

$$\text{mol N} = (100.0\text{ g})\ (0.3045)\ (1\text{ mol N}/14.0067\text{ g}) = 2.165\text{ mol N}$$

$$\text{mol O} = (100.0\text{ g})\ (0.6955)\ (1\text{ mol O}/15.994\text{ g}) = 4.349\text{ mol O}$$

Thus, the empirical formula is $N_{2.165}O_{4.349}$ or NO_2. The mass of this empirical formula is 45.99 g. The number of empirical units in one molecular formula is (91.98 g)/(45.99 g) = 2.000 or simply 2. Therefore, the molecular formula is N_2O_4. The proper name of this substance is dinitrogen tetraoxide. The name nitrogen(IV) oxide, or NO_2, only gives the name of the empirical formula.

3.37 First determine the empirical formula of urea from the composition data. Assume a 100 g sample to determine the relative number of moles of each element. C: (20.0 g)(1 mol/12.01 g) = 1.67 mol C. N: (46.6 g) (1 mol /14.01 g) = 3.33 mol N. O: (26.6 g)(1 mol/16.0 g) = 1.66 mol O. H: (6.7 g)(1 mol/1.01 g) = 6.6 mol H. Divide by the smallest number of moles to determine the ratio of moles of atoms in the empirical formula: 1.66 mol is the smallest. CN_2OH_4 is the result. The molar mass of this empirical formula is 60.1 g, which is the same as the molar mass. Therefore, this formula also represents the molecular formula. The reactants used in the chemical reaction to form urea are two gases in which each contain two different elements; thus the possible combinations of elements are CN, CH, CO, NO, HO, and HN. From your reading of the text, three carbon-containing gases that you may have encountered are CH_4 (methane), CO (carbon monoxide), and CO_2 (carbon dioxide). Methane is a combustible gas and would not extinguish a flame. Of the two carbon oxides, carbon dioxide is the logical choice since it is a gas produced in the lungs, is exhaled during breathing, and can extinguish flames. If carbon dioxide is one reactant, then the other reactant must contain the elements N and H so that urea can be produced. The most likely gas is ammonia, NH_3. Other combinations such as N_2H_4 are not commonly used gases. Thus, the balanced chemical reaction must be:

$$2\,NH_3 + CO_2 \longrightarrow CO(NH_2)_2 + H_2O$$

3.38 The formula weight (and the molecular weight) of $C^{18}O_2$ is larger than that of $C^{16}O_2$ because the oxygen-18 isotope has a greater isotopic mass than that of the oxygen-16 isotope. $K_2{}^{18}O$ has a greater formula weight compared to the other two substances because of the larger mass of the potassium atom coupled with two atoms of it present in the substance. The term molecular weight is not used with $K_2{}^{18}O$ because it is ionic and no molecular unit exists.

3.39 (e) $2\,C_2H_6 + 7\,O_2 \longrightarrow 4\,CO_2 + 6\,H_2O$

$BCl_3 + 3\,H_2O \longrightarrow H_3BO_3 + 3\,HCl$

3.40 (b) Molar mass = 17.6 g/0.400 mol = 44.0 g/mol

3.41 (c)

$$\%\,\text{Al} = \left(\frac{(2\text{ mol Al})\left(\dfrac{26.98\text{ g Al}}{1\text{ mol Al}}\right)}{101.96\text{ g }Al_2O_3}\right) \times 100 = 52.92\%$$

3.42 (a) Chemical equation:

$$6\,CO_2 + 6\,H_2O \longrightarrow C_6H_{12}O_6 + 6\,O_2.$$

$$\text{Grams }CO_2 = (0.256\text{ mol }C_6H_{12}O_2) \times \frac{(6\text{ mol }CO_2)}{(1\text{ mol }C_6H_{12}O_2)} \times \frac{(44.01\text{ g }CO_2)}{(1\text{ mol }CO_2)} = 67.6\text{ g }CO_2$$

3.43 (b)

$$\left(\frac{8\text{ mol of atoms in }H_3PO_4}{1\text{ mol }H_3PO_4}\right) \times (0.20\text{ mol }H_3PO_4) = 1.6\text{ mol of atoms}$$

$$\left(\frac{2\text{ mol of atoms in }N_2}{1\text{ mol }N_2}\right)(0.80\text{ mol }N_2) = 1.6\text{ mol of atoms}$$

3.44 (b)

$$\text{Grams }H_2O = (4.25 \times 10^{20}\text{ molecules}) \times \left(\frac{1\text{ mol}}{6.022 \times 10^{23}\text{ molecules}}\right)\left(\frac{18.02\text{ g }H_2O}{1\text{ mol }H_2O}\right) = 0.0127\text{ g }H_2O$$

3.45 (a) NO_2 and H_2O react in a 3:1 ratio of moles. The given moles are in the ratio 8:3, or $2\frac{2}{3}$:1. The last ratio shows that there is an insufficient quantity of NO_2 to react completely with water.

3.46 (c) 50.0 g × (1 mol/64.1 g) × (26.04 g C_2H_2/1 mol) = 20.3 g. (13.5 g/20.3 g) × 100 = 66.5%

3.47 (a)

3.48 (c)

3.49 (d)