AP CHEMISTRY

Packet “B”

Chapter 3

“STOICHIOMETRY”

Packet B ch.3

- Read through this ditto and highlight as you Read
- Take the practice test as the end: Then check your answers.

TOPIC 3

Stoichiometry: CALCULATIONS WITH CHEMICAL FORMULAS AND EQUATIONS

This chapter introduces the concept of the mole, a quantitative model for chemical structure. It also considers how chemical structure, encoded in chemical formulas, changes through the process of chemical reactions, which are encoded in chemical equations. The chemical equation is the qualitative model for chemical change and stoichiometry is the quantitative model.

For success on the Advanced Placement Chemistry exam, students need to be able to think holistically about problem solving using the fundamental ideas of the mole and stoichiometry. Chemical problem solving is usually about the quantitative manipulation of moles and how they relate to chemical formulas and equations. Much of the content of this chapter may have been learned in the first year course, but mastery of complex problems involving limiting reactants, empirical and molecular formulas, and combustion and hydrate analyses is essential. Pay particular attention to these sections:

3.5 **Empirical Formulas and Analyses**

3.6 **Quantitative Information from Balanced Equations**

3.7 **Limiting Reactants**

Chemical Equations — Section 3.1

Stoichiometry is the area of study that examines the quantities of substances involved in chemical reactions.

A **chemical reaction** is a process by which one or more substances are converted to other substances.

Chemical equations use chemical formulas to symbolically represent chemical reactions. For example, the following chemical equation describes how butane, C_4H_{10}, burns in air:

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

reactants $\rightarrow$ products

The equation is a "chemical sentence" written in the symbolic "words" of chemical formulas and symbols. The formulas on the left are reactants, and the formulas on the right are products. The equation reads:

"Two molecules of gaseous butane react with thirteen molecules of oxygen gas to produce eight molecules of carbon dioxide gas and ten molecules of liquid water."

A **balanced chemical equation** has an equal number of atoms of each element on each side of the arrow. The coefficients preceding each formula "balance" the equation. Notice that, because of the coefficients, on each side of the arrow there are eight carbon atoms, twenty hydrogen atoms, and twenty-six oxygen atoms.

The symbols (g), (l), (s), and (aq) are used to designate the physical state of each reactant and product: gas, liquid, solid, and aqueous (dissolved in water).

Your Turn 3.1

Write a chemical sentence to illustrate how to "read" the following chemical equation:

$$CO_2(g) + 2NH_3(g) + H_2O(l) \rightarrow 2NH_4^+(aq) + CO_3^{2-}(aq)$$

Write your answer in the space provided.

To balance simple equations, start by balancing atoms other than hydrogen or oxygen. Balance hydrogen atoms next to last and balance oxygen atoms last.

Example:

Balance the following equation:

$H_2SO_4(aq) + NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

Solution:

1. *S and C are already balanced so start with Na.*

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

2. *Now balance carbon.*

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2CO_2(g) + H_2O(l) + Na_2SO_4(aq)$

3. *Next balance H.*

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2CO_2(g) + 2H_2O(l) + Na_2SO_4(aq)$

4. *Check to see that oxygen is balanced and double check all other atoms. (As in algebraic equations, the numeral 1 is usually not written.)*

Simple Patterns of Reactivity

Section 3.2

Predicting the products of chemical reactions is an essential skill to acquire in the study of chemistry. Sometimes reactions fall into simple patterns and recognizing these patterns can be helpful in predicting which products will be produced from given reactants. Topic 4 addresses predicting products of chemical reactions in more depth. For now, here are a few patterns to learn to recognize.

1. A **combination reaction** is where two elements combine to form one compound. (Other combinations are possible and Chapter 4 describes better ways to predict what will happen.)

Example:

A metal reacts with a nonmetal to produce an ionic compound. Write an equation to describe what happens when:
Solid sodium is exposed to chlorine gas.

Solution:

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

Your Turn 3.2

Write an equation to describe what happens when solid magnesium metal reacts at high temperatures with nitrogen gas. Write your answer in the space provided.

2. A **decomposition reaction** is where one reactant changes to two or more products.

Example:

Upon heating, metal carbonates decompose to yield metal oxides and carbon dioxide. Write an equation to describe what happens when:

Solid magnesium carbonate is heated.

Solution:

$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

Your Turn 3.3

Write an equation to describe what happens when liquid water is decomposed to its elements by an electrical current. Write your answer in the space provided.

3. A **combustion reaction** usually involves oxygen, often from air, reacting with hydrocarbons or other organic molecules containing carbon, hydrogen, and oxygen to produce carbon dioxide and water.

Example:

Write an equation to describe what happens when liquid ethanol burns in air.

Solution:

$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

(Recall that encoded in the name "ethanol" is a two-carbon alcohol having the −OH functional group.)

Your Turn 3.4

Write an equation to describe what happens when liquid hexane is burned in air. Write your answer in the space provided.

Avogadro's Number, the Mole, and Molar Masses

Section 3.4

Avogadro's number is 6.02×10^{23}. It represents the number of atoms in exactly 12 grams of isotopically pure ^{12}C. Because all atomic masses are based on ^{12}C, the atomic mass of any element expressed in grams represents Avogadro's number (6.02×10^{23}) of atoms of that element.

A **mole** is the amount of matter that contains 6.02×10^{23} atoms, ions, molecules, or formula units.

The **molar mass** of a substance is the mass in grams of one mole of that substance. To calculate a molar mass of any substance, add the atomic masses of all the atoms in its formula. (For convenience atomic masses are often rounded to three significant figures.) Table 3.1 shows the molar masses of various substances.

Common misconception: "Molar mass" is a universal term that is often used to replace the terms atomic mass, molecular mass, and formula mass. Molar mass is used to express the mass of one mole of any substance, whether it is an atom (atomic mass), a molecule (molecular mass), an ion (formula mass), or an ionic compound (formula mass).

Table 3.1. The molar mass of any substance is the mass in grams of 6.02×10^{23} particles or one mole of that substance.

Formula	Number of particles	Representative particles	Molar mass	Alternate term
Ar	6.02×10^{23}	atoms	39.9 g/mol	atomic mass
CO_2	6.02×10^{23}	molecules	44.0 g/mol	molecular mass
NaBr	6.02×10^{23}	formula units	103 g/mol	formula mass
CO_3^{2-}	6.02×10^{23}	ions	60.0 g/mol	formula mass

Grams, moles and representative particles (atoms, molecules, ions or formula units) are converted one to another using the "Mole Road" described in Figure 3.1.

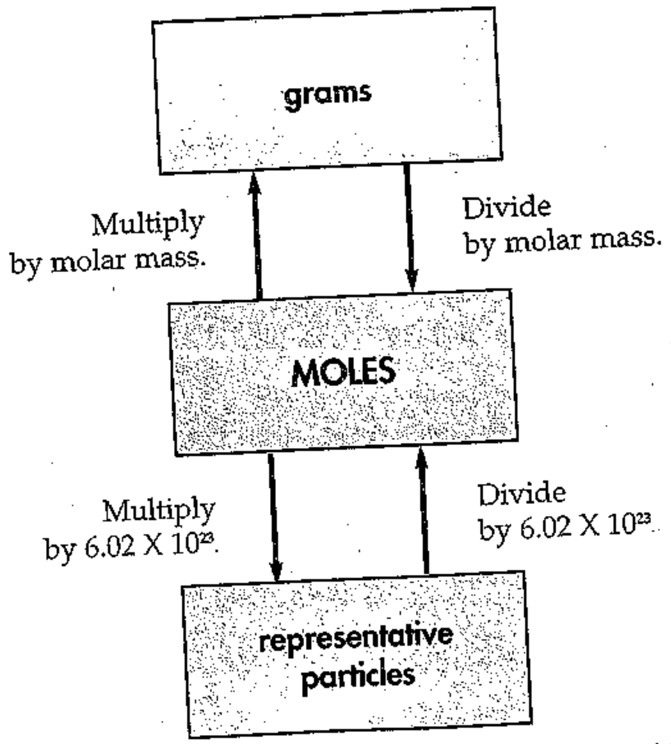


Figure 3.1. The "Mole Road". Divide to convert to moles. Multiply to convert from moles.

Calculating Percentage Composition of a Compound

The **percentage composition** of a compound is the percentage by mass contributed by each element in the compound. To calculate the percentage composition of an element in any formula, divide the molar mass of the element multiplied by the number of times it appears in the formula, by the molar mass of the formula, and multiply by 100.

% composition = 100 × (molar mass of element × subscript for element)/(molar mass of substance)

Example:

What is the % composition of Na_2CO_3?

Solution:

$\% \, Na = 100 \times 2(23.0) \, g \, Na/[2(23.0) + 12.0 + 3(16.0) \, g]$
$= 43.4\% \, Na$

$\% \, C = 100 \times 12.0 \, g \, C/[2(23.0) + 12.0 + 3(16.0) \, g] = 11.3\% \, C$

$\% \, O = 100 \times 3(16.0) \, g/[2(23.0) + 12.0 + 3(16.0) \, g] = 45.3\% \, O$

Calculating An Empirical Formula from Percentage Composition

Section 3.5

The **empirical formula** for a compound expresses the simplest ratio of atoms in the formula. The percentage composition of a compound can be determined experimentally by chemical analysis and the empirical formula can be calculated from the percentage composition.

Example:

What is the empirical formula of a compound containing 68.4% chromium and 31.6% oxygen?

Solution:

In chemistry, percentage always means mass percentage, unless specified otherwise. The data means that for every 100 grams of compound, there are 68.4 g of Cr and 31.6 g of O.

1. *Write the formula in terms of grams.*

Cr *68.4* **O** *31.6*

2. Convert grams to moles by dividing by the molar mass of each element.

Cr 68.4/52.0 **O** 31.6/16.0 = **Cr** 1.315 **O** 1.975

3. Convert to small numbers by dividing each mole quantity by the smaller mole quantity.

Cr 1.315/1.315 **O** 1.975/1.315 = **Cr** 1 **O** 1.5

4. If necessary, multiply each mole quantity by a small whole number that converts all quantities to whole numbers.

Cr 1 *x* 2 **O** 1.5 *x* 2 = Cr_2O_3

Molecular Formulas From Empirical Formulas

A **molecular formula** tells exactly how many atoms are in one molecule of the compound. The subscripts in a molecular formula are always whole number multiples of the subscripts in the empirical formula. Molecular formulas can be determined from empirical formulas if the molar mass of the compound is known.

Example:

A compound containing only carbon, hydrogen, and oxygen is 63.16% C and 8.77% H. It has a molar mass of 114 g/mol. What is its empirical formula and molecular formula?

Solution:

1. Write the formula in terms of grams.

C *63.16* **H** *8.77* **O** *28.07*

(Calculate the grams of oxygen by subtracting the grams of carbon and grams of hydrogen from 100g.
g O = 100 *g* − 63.16 *g* C − 8.77 *g H* = 28.07 *g* O.)

2. Convert grams to moles by dividing by the molar mass of each element.

C 63.16/12.0 **H** 8.77/1.00 **O** 28.07/16.0 = **C** 5.263 **H** 8.77 **O** 1.754

3. Convert to small numbers by dividing each mole quantity by the smallest mole quantity.

C 5.263/1.754 **H** 8.77/1.754 **O** 1.754/1.754 = $C_3H_5O_1$

empirical formula

4. All quantities are whole numbers.

5. *Divide the known molar mass by the mass of one mole of the empirical formula. The result produces the integer by which you multiply the empirical formula to obtain the molecular formula.*

$(114\ g/mol)/(57.0\ g/mol) = 2$

$\mathbf{C}3 \times 2\ \mathbf{H}5 \times 2\ \mathbf{O}1 \times 2 = \mathbf{C_6H_{10}O_2}$

molecular formula

Empirical Formulas From Combustion Analysis

When a compound containing carbon, hydrogen, and oxygen is completely combusted, all the carbon is converted to carbon dioxide, and all the hydrogen becomes water. The empirical formula of the compound can be calculated from the measured masses of the products.

Example:

A 3.489 g sample of a compound containing C, H, and O yields 7.832 g of carbon dioxide, and 1.922 grams of water upon combustion. What is the simplest formula of the compound?
The unbalanced chemical equation is:

$CxHyOz + O_2 \rightarrow CO_2 + H_2O$

X *is the number of moles of carbon in the compound because all the carbon in the compound is converted to* CO_2. *X = the number of moles of* CO_2 *because there is one mole of carbon in one mole of* CO_2.

$\mathbf{X} = mol\ C = 7.832\ g\ CO_2/44.0\ g/mol = \mathbf{0.178\ mol\ C}$

Y *is the number of moles of hydrogen in the compound because all the hydrogen becomes water. Y = twice the number of moles of water because there are two moles of hydrogen in one mole of water.*

$\mathbf{Y} = mol\ H = (1.922\ g\ H_2O/18.0\ g/mol) \times 2 = \mathbf{0.2136\ mol\ H}$

Z *is the number of mol of O. To obtain the number of grams of O in the compound, subtract the number of grams of C (X mol × 12.0 g/mol) and the number of grams of H (Y mol × 1.00 g/mol) from the total grams of the compound. Convert the result to moles of O by dividing by 16.0 g/mol.*

$\mathbf{Z} = mol\ O = (3.489 - 12.0\ X - 1.00\ Y)/16 = \mathbf{0.0712\ mol\ O}$

Convert to small numbers by dividing each mole quantity by the smallest mole quantity.

$\mathbf{C_{0.178}H_{0.2136}O_{0.0712}} =$

$\mathbf{C_{0.178/0.0712}H_{0.2136/0.0712}O_{0.0712/0.0712} = C_{2.5}H_3O_1}$

Finally, multiply each mole quantity by a small whole number that converts all quantities to whole numbers.

$\mathbf{C_{2.5\times 2}H_{3\times 2}O_{1\times 2} = C_5H_6O_2}$

Formulas of Hydrates

Ionic compounds often form crystal structures called **hydrates** by acquiring one or more water molecules per formula unit. For example, solid sodium thiosulfate decahydrate, $Na_2S_2O_3 \cdot 10H_2O$, has ten water molecules per formula unit of sodium thiosulfate. Heating a sample of hydrate causes it to lose water. The number of water molecules per formula unit can be calculated from the mass difference before and after heating.

Example:

When 21.91 g of a hydrate of copper(II) sulfate is heated to drive off the water, 14.00 g of anhydrous copper(II) sulfate remain. What is the formula of the hydrate?

Solution:

This is a variation of an empirical formula problem. The solution lies in calculating the ratio of H_2O moles to the moles of $CuSO_4$. The chemical equation is:

$$\underset{21.91\ g}{CuSO_4 \cdot X\ H_2O(s)} \rightarrow \underset{14.00\ g}{CuSO_4(s)} + X\ H_2O(g)$$

Calculate the grams of water by subtracting the grams of copper(II) sulfate from the grams of the hydrate:

$g\ H_2O = 21.91g - 14.00g = 7.91\ g\ H_2O$

$mol\ H_2O = 7.91\ g\ H_2O/18.0\ g/mol = 0.439\ mol\ H_2O$

$mol\ CuSO_4 = 14.00\ g/159.5\ g/mol = 0.08777\ mol\ CuSO_4.$

$mol\ H_2O/mol\ CuSO_4 = 0.439\ mol/0.08777\ mol = 5.00$

The formula has five moles of water per mole of copper(II) sulfate:

$CuSO_4 \cdot 5H_2O$

Section 3.6 Stoichiometry

Stoichiometry is the area of study that examines the quantities of substances involved in chemical reactions. The coefficients in a balanced chemical equation indicate both the relative number of molecules (or formula units) involved in the reaction, and the relative number of moles. For example, the equation for the combustion of octane, C_8H_{18}, a component of gasoline is:

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

The four coefficients that balance the equation are proportional to one another and can be used to relate mole quantities of reactants and/or products.

Example:

How many moles of octane will burn in the presence of 37.0 moles of oxygen gas?

Solution:

The balanced equation tells us that two moles of octane will burn in 25 moles of oxygen gas, so the answer to the question involves the ratio 2 mol C_8H_{18} per 25 mol O_2.

$\times$ *mol* $C_8H_{18} = 37.0\ mol\ O_2\ (2\ mol\ C_8H_{18}/25\ mol\ O_2) = 37.0 \times 2/25$
$= 2.96\ mol\ C_8H_{18}$.

Alternatively, we can use the mole road:
To convert mol O_2 to mol C_8H_{18}, multiply by 2/25

$mol\ C_8H_{18} = 37.0\ mol\ (2/25) = 2.96\ mol\ C_8H_{18}$

(Always multiply by the coefficient at the head of the arrow and divide by the coefficient at the tail of the arrow.)

$$mol\ C_8H_{18} \underset{x\,25/2}{\overset{x\,2/25}{\leftrightarrows}} mol\ O_2$$

To convert mol C_8H_{18} to mol O_2, multiply by 25/2

Figure 3.2 shows an expanded mole road to include the relationship between the moles of any reactant or product in a balanced chemical equation.

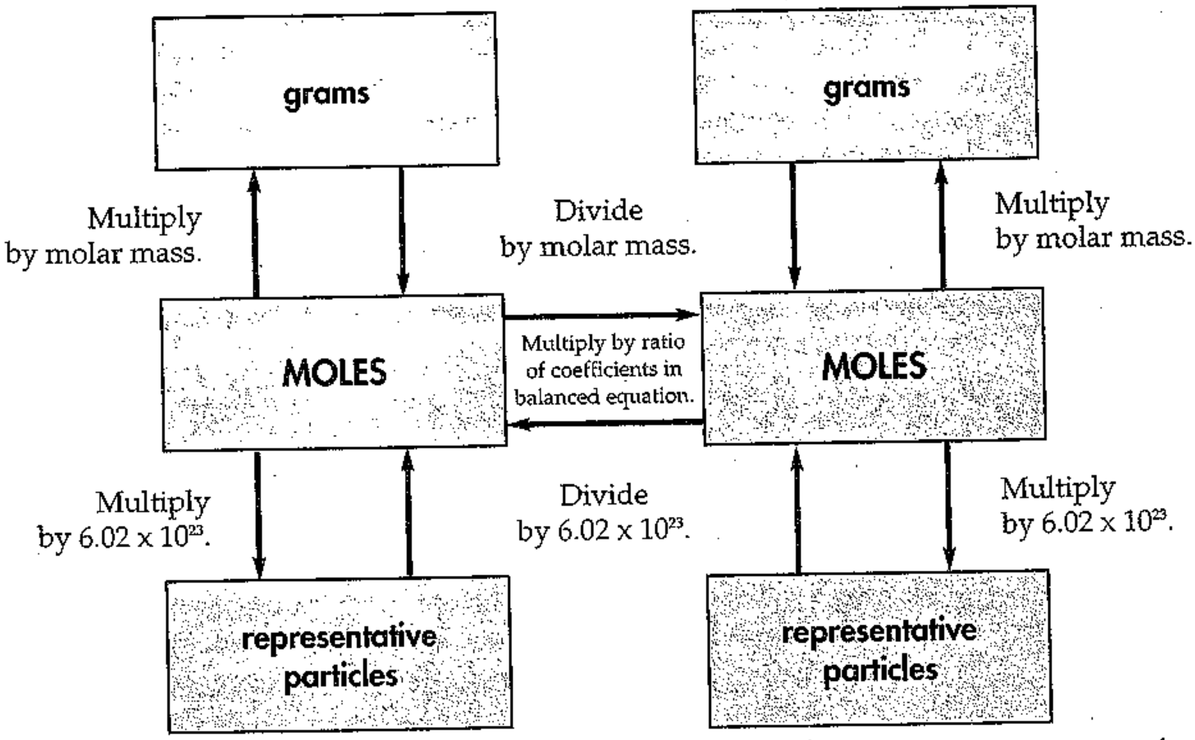


Figure 3.2. The stoichiometry "Mole Road". Divide to convert to moles. Multiply to convert from moles. Multiply by the ratio of coefficients in a balanced equation to convert moles of one substance to moles of another substance.

This stoichiometry mole road allows us to relate any quantity in a balanced equation to any other quantity in the same equation.

Example:

How many grams of carbon dioxide are obtained when 695 g (about one gallon) of octane are burned in oxygen?

Solution:

Follow the road map:

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

$$\begin{array}{ccc} \mathbf{695g\ C_8H_{18}} & & \mathbf{?g\ CO_2} \\ \downarrow & & \uparrow \\ mol\ C_8H_{18} & \xrightarrow[\mathbf{2\ mol}]{\mathbf{16\ mol}} & mol\ CO_2 \end{array}$$

Divide by molar mass of octane:
÷114g/mol

Multiply by ratio of coefficients:

Multiply by molar mass of CO_2:
x 44.0g/mol

$$x\ g\ CO_2 = (695/114)(16/2)(44.0) = 2150\ g\ CO_2$$

Section 3.7 Limiting Reactants

The **limiting reactant** is the reactant that is completely consumed in a chemical reaction. The limiting reactant limits the amount of products formed.

The **excess reactant** is usually the other reactant. Some of the excess reactant is left un-reacted when the limiting reactant is completely consumed.

A **stoichiometric mixture** means that both reactants are limiting and both are completely consumed by the reaction. There is an excess of neither. Because of their subtle nature, quantitative limiting reactant problems are among the most difficult to master in general chemistry. The mole road is a useful tool in solving limiting reactant problems.

Common misconception: Stoichiometry calculations can calculate only how much reactants react or how much products are formed. Stoichiometry cannot calculate how much excess reactant is left unreacted. To calculate how much excess reactant remains after the reaction is complete, first calculate how much is consumed and then subtract that amount from how much total reactant was initially present.

Example:

255 grams of octane and 1510 grams of oxygen gas are present at the beginning of a reaction that goes to completion and forms carbon dioxide and water according to the following equation.

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

a. *What is the limiting reactant?*

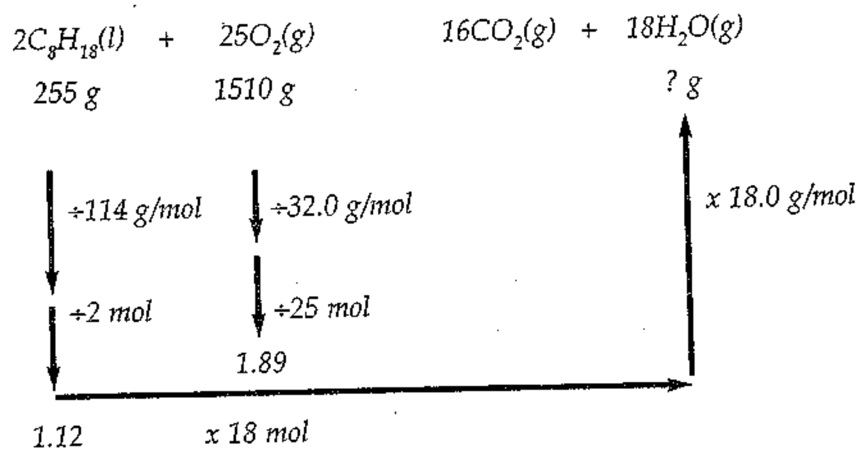
b. *How many grams of water are formed when the limiting reactant is completely consumed?*

c. *How many grams of excess reactant is consumed?*

d. *How many grams of excess reactant is left un-reacted?*

Solution:

a. To find the limiting reactant, first compare the number of moles of reactants relative to the ratio in which they react. To do this, divide the number of moles of each reactant by its corresponding coefficient that balances the equation. The resulting lower number always identifies the limiting reactant. Then use the mole road to solve the problem based on the identified limiting reactant.



a. C_8H_{18} *is the limiting reactant because* 1.12 < 1.89.

b. $\times$ *mol* H_2O = (255 *g*/114 *g*/*mol*)(18 *mol*/2 *mol*)(18.0 *g*/*mol*)
= 362 *g* H_2O

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

255 *g*	1510 *g*		? *g*

$$255\ g \xrightarrow{\div 114\ g/mol} \xrightarrow{\times\ 25\ mol/2\ mol} \xrightarrow{\times\ 32.0\ g/mol} ?\ g\ O_2\ react$$

x mol O_2 = (255 *g*)/114 *g*/*mol*)(25 *mol*/2 *mol*)(32.0 *g*/*mol*)
= 895 *g* O_2 *react*

d. *g* O_2 *left un-reacted* = 1510 *g* − 895 *g* = 615 *g* O_2 *un-reacted*

Theoretical, Actual, and Percent Yields

The **theoretical yield** of a reaction is the quantity of product that is calculated to form.

The **actual yield** is the amount of product actually obtained, and is usually less than the theoretical yield.

The **percent yield** relates the actual yield to the theoretical yield:

Percent yield = (actual yield)/(theoretical yield) × 100

Example:

In the previous example, the theoretical yield of water is calculated to be 362 g. What is the percent yield if the actual yield of water is only 312 g?

Solution:

Percent yield = 312*g*/362*g* × 100 = 86.2%

Multiple Choice Questions

1. *Ammonia forms when hydrogen gas reacts with nitrogen gas. If equal number of moles of nitrogen and hydrogen are combined, the maximum number of moles of ammonia that could be formed will be equal to:*

   A) *the number of moles of hydrogen.*

   B) *the number of moles of nitrogen.*

   C) *twice the number of moles of hydrogen.*

   D) *twice the number of moles of nitrogen.*

   E) *two thirds the number of moles of hydrogen.*

2. *If $C_4H_{10}O$ undergoes complete combustion, what is the sum of the co-efficients when the equation is completed and balanced using smallest whole numbers?*

   A) 8

   B) 16

   C) 22

   D) 25

   E) 32

3. *What are the products when lithium carbonate is heated?*

   A) $LiOH + CO_2$

   B) $Li_2O + CO_2$

   C) $LiO + CO_2$

   D) $LiC + O_2$

   E) $LiO + CO$

4. *Beginning with 48 moles of H_2, how many moles of $Cu(NH_3)_4Cl_2(aq)$ can be obtained if the synthesis of $Cu(NH_3)_4Cl_2(aq)$ is carried out through the following sequential reactions? Assume that a 50% yield of product(s) is(are) obtained in each reaction.*

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

   II. $4NH_3(g) + CuSO_4(aq) \longrightarrow Cu(NH_3)_4SO_4(aq)$

   III. $Cu(NH_3)_4SO_4 + 2NaCl \longrightarrow Cu(NH_3)_4Cl_2(aq) + Na_2SO_4(aq)$

A) 1

B) 2

C) 4

D) 8

E) 12

5. *What mass of water can be obtained from 4.0 g of H_2 and 16 g of O_2?*

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

A) 9 g

B) 18 g

C) 36g

D) 54 g

E) 72 g

6. *The empirical formula of pyrogallol is C_2H_2O and its molar mass is 126. Its molecular formula is:*

A) C_2H_2O

B) $C_4H_4O_2$

C) $C_2H_6O_3$

D) $C_6H_6O_3$

E) $C_2H_6O_6$

7. *What is the maximum amount of water that can be prepared from the reaction of 20.0 g of HBr with 20.0 g of $Ca(OH)_2$?*

$2HBr + Ca(OH)_2 \dashrightarrow CaBr_2 + 2H_2O$

A) (20/81)(2/2)(18) g

B) (20/74)(2/2)(18) g

C) (20/81)(2/1)(18) g

D) (20/74)(2/1)(18) g

E) (20/74)(1/2)(18) g

8. *How many moles of ozone, O_3, could be formed from 96.0 g of oxygen gas, O_2?*

A) 0.500 mol

B) 1.00 mol

C) 2.00 mol

D) 3.00 mol

E) 1/16 mol

9. *The percentage of oxygen in $C_8H_{12}O_2$ is:*

A) (16/140)(100)

B) (32/140)(100)

C) (16/124)(100)

D) (140/32)(100)

E) (32/124)(100)

10. *A compound contains 48% O, 40.0% Ca, and the remainder is C. What is its empirical formula?*

A) $O_3C_2Ca_2$

B) O_3CCa_2

C) O_3CCa

D) O_3CCa_2

E) O_2CCa

Free Response Question

1. *Combustion of 8.652 grams of a compound containing C, H, O, and N yields 11.088 g of CO_2, 3.780 grams of H_2O and 3.864 grams of NO_2.*

a. *How many moles of C, H, and N are contained in the sample?*

b. *How many grams of oxygen are contained in the sample?*

c. *What is the simplest formula of the compound?*

d. *If the molar mass of the compound lies between 200 and 300, what is its molecular formula?*

e. *Write and balance a chemical equation for the combustion of the compound.*

Additional Practice in *Chemistry the Central Science*

For more practice answering questions in preparation for the Advanced Placement examination try these problems in Chapter 3 of Chemistry the Central Science:

Additional Exercises: 3.81, 3.84, 3.86, 3.88, 3.89, 3.90, 3.91, 3.96, 3.98, 3.99, 3.100.

Integrative Exercises: 3.103, 3.104, 3.105, 3.107.

Multiple Choice Answers and Explanations

1. *E. The equation is* $3H_2(g) + N_2(g) \dashrightarrow 2NH_3(g)$. *If equal number of moles of reactants are mixed, the limiting reactant is* H_2 *because the reaction requires three times as much* H_2 *as it does* N_2. *The number of moles of ammonia depends on the limiting reactant,* H_2, *and is 2/3 the amount of* H_2 *present.*

2. *B. The products of a complete combustion of* $C_4H_{10}O$ *are carbon dioxide and water. The correctly completed and balanced equation is:*

   $C_4H_{10}O + 6O_2 \dashrightarrow 4CO_2 + 5H_2O$

   The sum of the coefficients is $1 + 6 + 4 + 5 = 16$.

3. *B. Upon heating, a metal carbonate decomposes to a metal oxide and carbon dioxide gas.* $Li_2CO_3(s) \dashrightarrow Li_2O(s) + CO_2(g)$

4. *A. Reaction 1 will yield* $48(2/3) \times 0.50 = 16$ *mol of* NH_3. *Reaction 2 will yield* $16(1/4) \times 0.50 = 2$ *moles of* $Cu(NH_3)_4SO_4(aq)$. *Reaction 3 will yield* $2(1/1) \times 0.50 = 1$ *mol of* $Cu(NH_3)_4Cl_2(aq)$.

5. *B. 4.0 g* H_2 *is 2 mol. 16 g* O_2 *is 0.5 mol.* O_2 *is the limiting reactant. 0.5 mol of* O_2 *will produce 1.0 mol* H_2O, *which is 18 g.*

6. *D. The ratio of the molar mass to the mass of the empirical formula is* $126/42 = 3$. *This means that the molecular formula has three times as many atoms as the empirical formula.*

7. *A. The limiting reactant is HBr.*

   $(20.0\,g)/(81.0\,g/mol)/2 < (20.0\,g)/(74.0\,g/mol)/1.$

   The 20.0 grams of HBr will limit how much water will be formed.

   $$\times\ g\ H_2O = 20.0\ g\ HBr(1\ mol/81.0\ g)(2\ mol\ H_2O/2\ mol\ HBr)(18.0\ g/mol)$$

8. *C. The balanced equation is:* $3O_2(g) \dashrightarrow 2O_3(g)$

   $$\times\ mol\ O_3 = 96.0\ g\ O_2\ (1\ mol/32.0\ g)(2\ mol\ O_3/3\ mol\ O_2) = 2.00\ mol$$

9. B. $\%\ O = [(2 \times \text{atomic mass of } O)/(\text{molar mass of } C_8H_{12}O_2)] \times 100 = (2 \times 16)/(140) \times 100 = (32/140)(100)$

10. C. *The percentage composition is the number of grams of each element contained in 100 grams of compound. The % of C is 100% − 48.0% − 40.0% = 12.0% C. Convert grams of each atom to moles by dividing by the respective atomic mass.*

$$O_{48.0/16.0}C_{12.0/12.0}Ca_{40.0/40.0} = O_3C_1Ca_1$$

This is calcium carbonate, $CaCO_3$.

Free Response Answer

a. *The moles of carbon are equal to the moles of carbon dioxide because there is one mole of C in every mole of CO_2. Divide the grams of CO_2 produced by the molar mass of CO_2 to obtain the moles of carbon. Divide the mass of water by the molar mass of water and multiply by two to obtain the moles of hydrogen. Repeat for the moles of N.*

$$\begin{aligned} mol\ C &= 11.088\ g\ CO_2/44.0\ g/mol = 0.252\ mol\ CO_2 \\ &= 0.252\ mol\ C \end{aligned}$$

$$\begin{aligned} mol\ H &= 3.780\ g\ H_2O/18.0\ g/mol = 0.210\ mol\ H_2O \times 2 \\ &= 0.420\ mol\ H \end{aligned}$$

$$\begin{aligned} mol\ N &= 3.864\ g\ NO_2/46.0\ g/mol = 0.0840\ mol\ NO_2 \\ &= 0.0840\ mol\ N \end{aligned}$$

b. *To obtain the number of grams of oxygen in the formula, subtract from the given sample mass, the number of grams of each of the other elements, C, H, and N. To obtain their masses, multiply the number of moles calculated in Part a by the atomic masses of the respective elements.*

$$g\ O = 8.652\ g\ compound - g\ C - g\ H - g\ N$$

$$\begin{aligned} g\ O &= 8.652\ g - (0.252\ mol\ C \times 12.0\ g/mol) - (0.420\ mol\ H \times \\ &\quad 1.00\ g/mol) - (0.0840\ mol\ N \times 14.0\ g/mol) = 4.032\ g\ O \end{aligned}$$

c. *To obtain the empirical formula convert the mass of O calculated in Part b to moles by dividing by the atomic mass of oxygen. Then convert each mole quantity to small whole numbers by dividing each by the lowest of the mole quantities.*

$$Mol\ O = 4.032\ g/16.0\ g/mol = 0.252\ mol\ O$$

$$C_{0.252}H_{0.420}N_{0.0840}O_{0.252}$$

$$C_{0.252/0.0840}H_{0.420/0.0840}N_{0.0840/0.0840}O_{0.252/0.0840}$$

$$C_3H_5N_1O_3$$

d. *To obtain the molecular formula, calculate the molar mass of the empirical formula determined in Part c. Multiply the result by small integers until an integer multiple of the molar mass is a number between 200 and 300. Multiply each subscript in the empirical formula by this integer. Check your work by calculating the molar mass of the resulting molecular formula to see that it lies between 200 and 300 g/mol.*

Molar mass of $C_3H_5N_1O_3$ = 3(12.0) + 5(1.00) + 1(14.0) + 3(16.0) = 103 *g/mol*

2 × 103 *g/mol* = 206 *g/mol.*

$C_{3\times2}H_{5\times2}N_{1\times2}O_{3\times2} = C_6H_{10}N_2O_6$

e. *Balance in order, C, N, H, and O:*

$2C_6H_{10}N_2O_6 + 15O_2(g) \rightarrow 12CO_2(g) + 10H_2O(g) + 4NO_2(g)$

Answers to Your Turn

3.1. *One molecule of gaseous carbon dioxide reacts with two molecules of ammonia gas and one molecule of liquid water to produce two aqueous ammonium ions and one aqueous carbonate ion.*

3.2. $3Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

3.3. $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$

3.4. $2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(l)$

Alternatively: $C_6H_{14}(l) + 19/2\ O_2(g) \rightarrow 6CO_2(g) + 7H_2O(l)$

AP CHEMISTRY

Packet "B"

Chapter 3 TEST

"STOICHIOMETRY"

1. When the reaction below is balanced, the coefficients are __________.

$$NH_3 + O_2 \rightarrow NO_2 + H_2O$$

a. 1,1,1,1
b. 4,7,4,6
c. 2,3,2,3
d. 1,3,1,2
e. none of these

2. What is the coefficient of Fe_3O_4 when the following equation is balanced?

$$Al + Fe_3O_4 \rightarrow Al_2O_3 + Fe$$

a. 2
b. 3
c. 4
d. 5
e. 1

3. What is the coefficient of $FeCl_3$ when the following equation is balanced?

$$FeCl_3 + Na_2CO_3 \rightarrow Fe_2(CO_3)_3 + NaCl$$

a. 1
b. 2
c. 3
d. 5
e. 4

4. Consider the following reaction:

$$C_3H_8O_3 + O_2 \rightarrow CO_2 + H_2O$$

The coefficient of $C_3H_8O_3$ when the equation is balanced is __________.
a. 1
b. 2
c. 3
d. 7
e. 5

5. What is the coefficient of H_2SO_4 when the following equation is balanced?

$$H_2SO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$$

a. 1
b. 2
c. 3
d. 4
e. 0.5

6. When the equation $K(s) + H_2O(l) \rightarrow KOH(aq) + H_2(g)$ is balanced, the coefficient of water is ___________.
a. 1
b. 2
c. 3
d. 4
e. 5

7. When the equation $N_2O_5(g) + H_2O(l) \rightarrow HNO_3(aq)$ is balanced, the sum of the coefficients is ___________.
a. 5
b. 3
c. 4
d. 6
e. 8

8. When the equation $N_2O_5(g) + H_2O(l) \rightarrow HNO_3(aq)$ is balanced, the coefficient for water is ___________.
a. 5
b. 2
c. 3
d. 4
e. 1

9. Of the following compounds, which one will form something other than carbon dioxide and water when combusted?
a. C_2H_6
b. CH_4
c. $C_{32}H_{44}O_9$
d. CH_3OCH_3
e. C_5H_5N

10. When a hydrocarbon is combusted in air, what component of air reacts?
a. oxygen
b. nitrogen
c. carbon dioxide
d. water
e. argon

11. Of the reactions below, which one is a decomposition reaction?
a. $NH_4Cl \rightarrow NH_3 + HCl$
b. $2Mg + O_2 \rightarrow 2MgO$
c. $2N_2 + 3H_2 \rightarrow 2NH_3$
d. $Zn + Cu(NO_3)_2 \rightarrow Cu + Zn(NO_3)_2$
e. $Cd(NO_3)_2 + Na_2S \rightarrow CdS + 2NaNO_3$

12. Which one of the following substances would be the product of this combination reaction?

$$Al(s) + I_2(s) \rightarrow \_\_\_\_\_\_\_\_\_\_$$

a. AlI_2
b. AlI
c. AlI_3
d. Al_2I_3
e. Al_3I_2

13. The balanced equation for the decomposition of sodium azide is ___________.
a. $2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$
b. $2NaN_3(s) \rightarrow Na_2(s) + 3N_2(g)$
c. $NaN_3(s) \rightarrow Na(s) + N_2(g)$
d. $NaN_3(s) \rightarrow Na(s) + N_2(g) + N(g)$
e. $2NaN_3(s) \rightarrow 2Na(s) + 2N_2(g)$

3

14. Which of the following are combination reactions?
    1) $CH_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
    2) $CaO(s) + CO_2(g) \rightarrow CaCO_3(s)$
    3) $Mg(s) + O_2(g) \rightarrow MgO(s)$
    4) $PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$
    a. 1, 2, and 3
    b. 2 and 3
    c. 1, 2, 3, and 4
    d. 4 only
    e. 2, 3, and 4

15. Currently, the atomic mass unit is based on what element?
    a. hydrogen
    b. oxygen
    c. sodium
    d. carbon
    e. helium

16. The formula of nitrobenzene is $C_6H_5NO_2$. The molecular mass of this compound is _____ amu.
    a. 107.11
    b. 43.03
    c. 109.10
    d. 123.11
    e. 3.06

17. The formula weight (amu) of potassium phosphate is __________.
    a. 173.17
    b. 251.37
    c. 212.27
    d. 196.27
    e. 86.07

18. What is the mass % of carbon in dimethylsulfoxide, C_2H_6SO?
    a. 60.0
    b. 20.6
    c. 30.7
    d. 7.74
    e. 79.8

19. The formula weight (amu) of $Ca(NO_3)_2$ is ____________.
    a. 102.083
    b. 164.088
    c. 204.166
    d. 150.081
    e. 116.090

20. Which of the following, strictly speaking, does not have a molecular weight?
    a. SO_2
    b. CH_4
    c. H_2O
    d. Li_2S
    e. each of these has a molecular weight

28. How many oxygen atoms are there in 52.06 g of carbon dioxide?
   a. 1.424×10^{24}
   b. 6.022×10^{23}
   c. 1.204×10^{24}
   d. 5.088×10^{23}
   e. 1.018×10^{24}

29. What is the empirical formula for a compound that contains 29% Na, 41% S, and 30% O by mass?
   a. $Na_2S_2O_3$
   b. $NaSO_2$
   c. $NaSO$
   d. $NaSO_3$
   e. $Na_2S_2O_6$

30. The empirical formula of a compound with the molecular formula C_2H_6 and a molecular weight of 30.07 amu is __________.
   a. CH_2
   b. CH_3
   c. C_2H_6
   d. C_3H
   e. C_8H_{24}

31. A compound containing only carbon and hydrogen is 80.0% by mass carbon and 20.0% by mass hydrogen. What is the empirical formula of the compound?
   a. $C_{20}H_{60}$
   b. C_7H_{20}
   c. CH_3
   d. C_2H_6
   e. CH_4

32. Combustion analysis of 0.9835 g of a compound containing carbon, hydrogen, and oxygen yielded 1.900 g CO_2 and 1.070 g H_2O. What is the empirical formula of the compound?
   a. C_2H_5O
   b. $C_4H_{10}O_2$
   c. $C_4H_{11}O_2$
   d. $C_4H_{10}O$
   e. $C_2H_5O_2$

33. The combustion of C_3H_8 produces CO_2 and H_2O:

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$$

The reaction of 2.5 mol of O_2 will produce __________ mol of H_2O.
   a. 4.0
   b. 3.0
   c. 2.5
   d. 2.0
   e. 1.0

34. How many moles of NH_3 are produced by the complete reaction of 6.0 mol of H_2O according to the following equation?

$$Li_3N + 3H_2O \rightarrow 3LiOH + NH_3$$

a. 2.0
b. 1.0
c. 6.0
d. 3.0
e. 18

35. How many grams of N_2 are required to completely react with 9.3 g of H_2 according to the following equation?

$$N_2 + 3H_2 \rightarrow 2NH_3$$

a. 1.3×10^2
b. 2.0
c. 43
d. 3.9×10^2
e. 4.6

36. What mass (g) of carbon dioxide will be produced by the complete combustion of 0.4560 g of $C_5H_{19}O$?
a. 1.054
b. 0.02395
c. 0.004790
d. 2.108
e. 0.5270

37. Why is carbon dioxide called a greenhouse gas?
a. It is needed in large quantities by plants growing in greenhouses.
b. It is a major product of greenhouses.
c. Lead carbonate is a common pigment used in green house paint.
d. It absorbs heat energy radiated away from the surface of the earth, effectively trapping it in the atmosphere, much like the glass of a greenhouse.
e. Bacterial degradation of fertilizers in a greenhouse environment produce large quantities of carbon dioxide.

38. Silver nitrate and aluminum chloride react with each other by exchanging anions:

$$AgNO_3 + AlCl_3 \longrightarrow Al(NO_3)_3 + AgCl$$

What mass (g) of AgCl is produced when 4.22 g of $AgNo_3$ react with 7.73 g of $AlCl_3$. The reaction must be balanced.
a. 17.6
b. 4.22
c. 24.9
d. 3.56
e. 11.9

39. C_3H_8 is ____% carbon by mass.
a. 81.71
b. 83.24
c. 83.91
d. 84.28
e. 84.52

40. How many moles of H_2SO_4 are in 61 grams of H_2SO_4?
a. 1.0503
b. 1.0133×10^{24}
c. 3.6722×10^{25}
d. 0.6219
e. 5982.88