Welcome to AP Chemistry!

Welcome to Advanced Placement Chemistry! Some of you have taken an Advanced Placement course before, while for others, this will be your first experience with AP. This course requires hard work and good study habits. Therefore, now and in the future, it is important to keep focused on the advantages offered in this course. A few advantages are listed below:

1. One of the most important benefits of this course is that if you take and pass the AP exam given in May, you will be able to receive college credit at many universities and colleges.

2. Whether or not a student performs well on the AP exam, she may choose to take freshman chemistry in college. Students who have done this have found that they have a tremendous advantage over others who have not taken AP Chemistry. Most of the material covered in a college chemistry course will be a review for students who have taken AP Chemistry and increases the chances of getting the "A" in college chem. There are several reasons for this.

   a. High school classes are smaller than college classes. The opportunity for individual extra help is much greater. It is not unusual for a freshman college chemistry class to have 200 students.

   b. Your AP Chemistry grade will be cushioned by Laboratory Report grades and homework assignments as well as the 1.1 weighting.

   c. Freshman chemistry is often used to "weed out" students in to prevent large numbers of students moving into upper division classes. In AP Chem, your only competition is yourself, not your colleagues.

3. Having AP Chemistry on your transcript reflects very favorably on you. Many universities are looking for ways to distinguish the thousands of students that apply for admission from each other. The decision to take one of the most rigorous AP courses reveals your willingness to accept a challenge and work hard.

4. AP Chemistry will teach you to think critically and synthesize concepts at higher levels. This is excellent preparation for the higher level thinking required at the college level.

So let's get started. The Summer assignment will review many of the concepts you have learned in sophomore chemistry. It will provide the foundation for the rest of the year and reinforce skills you will rely upon throughout the course. While we will review briefly during the first week, the first exam will cover the material in these three chapters.

I look forward to meeting all of you and working with you next year!
Dr. Feinman

AP Chem Summer Assignment

The following 3 Chapter reviews are material covered during sophomore Chem. There will be a test during the first week of school on the information in this packet.

Chapter 1

Introduction: Matter and Measurement

OVERVIEW OF THE CHAPTER

1.1, 1.2, 1.3 MATTER: ELEMENTS, COMPOUNDS, AND MIXTURES

Learning Goals: You should be able to:

1. Distinguish between physical and chemical properties and also between simple physical and chemical changes.
2. Differentiate between the three states of matter.
3. Distinguish between elements, compounds, and mixtures.
4. Give the symbols for the elements discussed in this chapter.

1.4 PHYSICAL QUANTITIES AND UNITS

Review: Concept of fraction; exponential notation (see text: Appendix A).

Learning Goal: You should be able to list the basic SI and metric units and the commonly used prefixes in scientific measurements.

1.5 UNCERTAINTY IN MEASUREMENT: SIGNIFICANT FIGURES

Review: Exponential notation (see text Appendix A).

Learning Goals: You should be able to:

1. Determine the number of significant figures in a measured quantity.
2. Express the result of a calculation with the proper number of significant figures.

1.4 TEMPERATURE AND DENSITY: INTENSIVE PROPERTIES

Learning Goals: You should be able to:

1. Convert temperatures among the ~~Fahrenheit~~, Celsius, and Kelvin scales.
2. Perform calculations involving density.

1.6 DIMENSIONAL ANALYSIS

Review: Concepts of fraction and ratio.

Learning Goal: You should be able to convert between units by using dimensional analysis.

Compounds - chemically combined

mixtures - physically combined

TOPIC SUMMARIES AND EXERCISES

MATTER: ELEMENTS, COMPOUNDS, AND MIXTURES

Matter is any material that occupies space and has mass. Three phases (states) of matter exist: gas, liquid, and solid.

- A sample of matter is either a substance or a mixture.
- **Substances** are either elements or compounds. **Elements** cannot be separated into simpler new substances. **Compounds** consist of two or more elements chemically combined in a definite ratio. A compound can be chemically decomposed into its elements.
- **Mixtures** are physical combinations of two or more substances and are either homogeneous or heterogeneous. Note that a *heterogeneous* mixture exhibits more than one phase and possesses a nonuniform distribution of substances. A *homogeneous* mixture consists of one phase and a uniform distribution of substances. ie. Solutions, Air, Alloys,
- Mixtures can be separated into substances by physical means.

An aqueous solution is a homogeneous* mixture because it is (1) one phase (2) uniform in density *

Alterations in matter can involve chemical or physical changes.

- A **chemical change** involves a change in the composition of a substance. A **chemical property** describes the type of chemical change. For example, the property of wood burning is a chemical property.
- A **physical change** does not involve a change in the composition of a substance but rather a change in a **physical property** such as temperature, volume, mass, pressure, or state.

Check your understanding of the new terms you have learned by doing Exercises 1-5.

EXERCISE 1 Identifying characteristics of matter

Match the following characteristics to one or more of the three states of matter: (a) has no shape of its own; (b) definite shape; (c) occupies the total volume of a container; (d) partially takes on the shape of a container; (e) does not take on the shape of a container; (f) readily compressible; (g) slightly compressible; (h) essentially noncompressible.

SOLUTION: Gas—(a), (c), (f); Liquid—(a), (d), (g); Solid—(b), (e), (h)

EXERCISE 2 Identifying characteristics of matter II

Match the term with the best identifying phrase:

Terms

1. Homogeneous mixture
2. Heterogeneous mixture
3. Mixture
4. Substance
5. Element
6. Compound

Phrases

a. Any kind of matter that is pure and has a fixed composition (Dalton)
b. Cannot be decomposed into simpler substances by chemical changes

c. A solution of uniform composition
d. Can be decomposed into simpler substances by chemical changes
e. Any kind of matter that can be separated into simpler substances by physical means
f. Nonuniform composition

SOLUTION: 1-c; 2-f; 3-e; 4-a; 5-b; 6-d

EXERCISE 3 Writing names or symbols of elements

With the help of the periodic table, write the name or the chemical symbol for each of the following elements: (a) F; (b) zinc; (c) potassium; (d) As; (e) Al; (f) iron; (g) helium; (h) barium; (i) Ne.

SOLUTION: (a) fluorine; (b) Zn; (c) K; (d) arsenic; (e) aluminum; (f) Fe; (g) He; (h) Ba; (i) neon

EXERCISE 4 Identifying changes of matter

Are the following changes physical or chemical: (a) the vaporization of solid carbon dioxide; (b) the explosion of solid TNT; (c) the aging of an egg with a resultant unpleasant smell; (d) the formation of a solid when honey is cooled?

SOLUTION: (a) A physical change. The form of carbon dioxide is changed from solid to gas. There is no change in its chemical composition. (b) A chemical and physical change. The explosion results from a change in the chemical composition of TNT and the formation of a gas. (c) A chemical change. A change in the composition of the egg results in the formation of a gas that has an unpleasant smell. (d) A physical change. The solid results from the crystallization of dissolved sugars; no change occurs in the chemical form of the sugars.

Signs of Chemical Rxn
1. change in color
2. odor
3. gas is formed (bubbles)
4. precipitate
5. temperature change

EXERCISE 5 Recognizing elements, compounds, and mixtures

Classify each of the following as an element, compound, or mixture: (a) a 100-percent lead bar; (b) wine; (c) gasoline; (d) carbon dioxide (CO_2).

SOLUTION: (a) Lead is an element and cannot be separated by either chemical or physical means into simpler substances. It is listed among the elements in Table 1.2 in the text. (You should know the symbols in Table 1.2.) (b) Wine is a mixture of alcohol, other components, and water. The fact that wines contain varying percentages of alcohol attests to their having different compositions. (c) We know gasoline must be a mixture because it is available with different compositions and properties (no lead, regular, and different brands with different additives). (d) CO_2 is a compound because the ratio of carbon and oxygen atoms is fixed and definite. The name also implies that it is a compound because we do not have such systematic names for mixtures.

Pb = Lead

PHYSICAL QUANTITIES AND UNITS

A physical property of a sample is measured by comparing it with a standard unit of that property. Measured quantities such as volume, length, mass, and temperature require a number and a reference label, called the unit of measurement. Two systems of unit measurements are shown in Table 1.1 on page 4.

The SI system of units is now the preferred one; however, you will find certain metric units still used. *You must become thoroughly familiar with the units in Table 1.1 before starting the next chapter.*

TABLE 1.1 Metric and SI Units

Physical quantity	Metric unit name	SI unit name[a]
Length	Meter (m)	Meter (m)
Volume	Cubic centimeter (cm^3)[b]	Cubic meter (m^3)
Mass	Gram (g)	Kilogram (kg)
Time	Second (s)	Second (s)
Energy	Calorie (cal)	Joule (J)
Pressure	Atmosphere (atm)	Newton per square meter (N/m^2)

[a] Systeme International d'Unites (SI) or International System of Units.

[b] Chemists commonly use the unit cubic centimeter when dealing with the volume of a solid, but they usually use the unit liter (L) when a substance is a liquid.

Another skill requiring proficiency is to change a number with a unit to one with a different unit. To do this, equivalence relationships exist between units. Tables 1.2 and 1.3 give some common equivalences that you will need in this chapter.

Prefixes are used with units to indicate decimal fractions (<1) or multiples (>1) of basic units.

- Example of a decimal fraction: The prefix centi- means 1/100 (= 0.01) of a basic unit; thus, 100 cm = 100 × 1/100 m = 1 m.
- Example of a multiple: The prefix kilo- means 10^3 (= 1000); thus, 1 km = 1 × 1000 m = 1000 m.

The commonly used prefixes that you must know are shown in Table 1.4 on page 5. *Memorize them.*

You do not need to memorize these tables.

TABLE 1.2 Equivalence Relationships between SI and Metric Units

Physical quantity	Metric unit name	SI unit name	Equivalence
Length	Meter	Meter	Same
Mass	Gram	Kilogram	1000 g = 1 kg
Time	Second	Second	Same
Energy	Calorie	Joule	1 cal = 4.184 J
Volume	Cubic centimeter	Cubic meter	1,000,000 cm^3 = 1 m^3
Volume	Liter	Cubic meter	1000 L = 1 m^3
Pressure	Atmosphere	Newton per squre meter	1 atm = 0.1754 N/m^2

Vol = L × W × h
= cm^3
Therefore, 1mL = 1cm^3

TABLE 1.3 Equivalence Relationships between Metric and English Units

Physical quantity	English unit symbol	Metric unit symbol	Equivalence
Mass	lb (= 16 oz)	g	1 lb = 453.6 g
Length	ft (= 12 in.)	m	3.272 ft = 1 m
Length	in.	cm	1 in. = 2.54 cm
Length	mi (= 5280 ft)	m	1 mi = 1609 m
Volume	qt	L	1.057 qt = 1 L

TABLE 1.4 Commonly Used Prefixes for Scientific Measurement in Chemistry

Prefix	Fraction or multiple of base unit	Abbreviation
Deci-	$10^{-1}\left(\frac{1}{10}\right)$	d
Centi-	$10^{-2}\left(\frac{1}{100}\right)$	c
Milli-	$10^{-3}\left(\frac{1}{1000}\right)$	m
Micro-	$10^{-6}\left(\frac{1}{1{,}000{,}000}\right)$	μ
Nano-	$10^{-9}\left(\frac{1}{1{,}000{,}000{,}000}\right)$	n
Pico-	$10^{-12}\left(\frac{1}{1{,}000{,}000{,}000{,}000}\right)$	p
Kilo-	10^{3} (1000)	k
Mega-	10^{6} (1,000,000)	M
Giga-	10^{9} (1,000,000,000)	G

You do not need to memorize this.

EXERCISE 6 Determining relative magnitudes of quantities

(Skip) * * * But be sure to understand negative exponents!

i.e. $10^2 = 100$ but $10^{-2} = .01$ or $\frac{1}{100}$

Which quantity of each pair is larger: **(a)** 1 nm or 1 micrometer; **(b)** 1 picogram or 1 cg; **(c)** 1 megagram or 1 milligram?

SOLUTION: Change the pairs so that each quantity is represented by either a fraction or a multiple of the same basic metric unit. Then from their relative magnitudes you can determine which is larger.

(a)

$$1 \text{ nm} = 1 \text{ nanometer} = 10^{-9} \text{ meter}$$
$$1 \text{ micrometer} = 1\ \mu\text{m} = 10^{-6} \text{ meter}$$

One micrometer is larger in value than one nanometer because the fraction $10^{-6}\left(\frac{1}{1{,}000{,}000}\right)$ is larger in magnitude than the fraction $10^{-9}\left(\frac{1}{1{,}000{,}000{,}000}\right)$.

(b)

$$1 \text{ picogram} = 1 \text{ pg} = 10^{-12} \text{ gram}$$
$$1 \text{ cg} = 1 \text{ centigram} = 10^{-2} \text{ gram}$$

One centigram is larger in value than one picogram because the fraction $10^{-2}\left(\frac{1}{100}\right)$ is larger in magnitude than the fraction $10^{-12}\left(\frac{1}{1{,}000{,}000{,}000{,}000}\right)$.

(c)

$$1 \text{ megagram} = 1 \text{ Mg} = 10^{6} \text{ gram}$$
$$1 \text{ mg} = 1 \text{ milligram} = 10^{-3} \text{ gram}$$

One megagram is larger in value than one milligram because the multiple 10^{6} (1,000,000) is larger in magnitude than the fraction $10^{-3}\left(\frac{1}{1{,}000}\right)$.

(skip)

EXERCISE 7 Recognizing units with measurements

With what types of measurements are the following units associated?

g, L, m, km, cm, Mg, pg, cm^3

SOLUTION: Mass (g, Mg, pg); volume (L, cm^3); length (m, km). Note that the prefixes such as M- and c- do not change the type of unit. However, the type of unit can be changed if it is raised to some power, as is the case for cm^3. The unit cm^3 means cm × cm × cm, which is a unit for volume ($V = 1 \times w \times h$).

EXERCISE 8 Comparing English to SI System of Units

What is the advantage of the metric system in comparison to the English system?

SOLUTION: In the metric system, all quantities larger or smaller than the basic unit involve multiplication of the basic unit value by some power of 10 (for example, $10^3 = 1000$, $10^{-1} = \frac{1}{10}$, and so on). This is not true of the English system. Smaller or larger quantities of the basic unit in the English system are newly defined units. For example, 4000 qt equals 1000 gal, not 4 "kiloquarts." Many more conversion factors are required in the English unit system than in the metric unit system.

EXERCISE 9 Knowing acceptable SI volume unit

Suggest a reason for the fact that 1 μkL (microkiloliter) is not accepted as an appropriate SI unit for volume.

SOLUTION: The expression 1 μkL involves two prefixes, micro- (μ) and kilo- (k), yielding a compound prefix. This can be confusing, particularly if three or four prefixes are used. Thus, we do not use more than one prefix when expressing numbers. Instead of 1 μkL (microkiloliter), we write 1 mL (milliliter).

UNCERTAINTY IN MEASUREMENTS: SIGNIFICANT FIGURES

Quantities in chemistry are of two types:

- **Exact numbers:** These result from counting objects such as coins or occur as defined numbers such as exact conversion factors.
- **Inexact numbers:** These are obtained from measurements and require judgment. Uncertainties exist in their values.

Review sig figs

Measured quantities (inexact numbers) are reported so that the last digit is the first uncertain digit. All certain digits and the first uncertain digit are referred to as **significant figures**. For example:

- 2.86: 2 and 8 are certain and well known. The number 6 is the first that is subject to judgment and is uncertain. The first uncertain digit is assumed to have an uncertainty of ±1: 2.86 ± 0.01. The number 2.86 has three significant figures.
- 0.0020: Zeroes to the left of the first nonzero digit in a number with a decimal point are not significant. The first three zeroes are not significant because they are to the left of the 2 and also define the decimal point. The zero to the right of the 2 is significant. This number has only two significant figures.
- 100: Trailing zeroes that define a decimal point may or may not be significant. Unless stated, assume they are not significant. Therefore, 100 has one significant figure unless otherwise stated; if it is determined from counting objects, it has three significant figures.

Scientific notation removes the ambiguity of knowing how many significant figures a number possesses.

- The form of a number in scientific notation is A.BC $\times$ 10^x. If $x < 1$, the number is less than 1. If $x > 1$, the number is greater than 1.
- Only significant digits are shown. The number 0.0020 becomes 2.0×10^{-3}.

Calculated numbers must show the correct number of significant figures. The rules for doing this are:

1. Addition and Subtraction: The final answer should have the same uncertainty as the quantity in the calculation with the greatest uncertainty. In the following example the first uncertain digit in each quantity is in bold.

   32**5**.24 (uncertainty = ±0.01)
   + 2**1**.4 (uncertainty = ±0.1)
   + 14**5** (uncertainty = ±1)
   49**1**.64 (uncertainty in final answer is ±1)

If the answer is rounded off at the boldfaced digit in 491.64 to 492, the answer has the same uncertainty as for 145, the number in the calculation with the greatest uncertainty.

Review this but do not memorize

2. Multiplication and division: When multiplying or dividing numbers round off the final calculated answer so that it has the same number of significant figures as the least certain number (the one with fewest number of significant figures) in the calculation. A little caution must be used when applying this rule. For example, to divide 101 by 95, you might be tempted to report the final answer to two significant figures because 95 appears to be the least certain number. Yet 95 has almost three significant figures; there is very little difference in error between 1 in 95 and 1 in 101. Thus, in this case it makes more sense to round off the final answer to three significant figures. Use common sense in problems when a number is close in magnitude to 100, 1000, 10,000, and so on.
3. Exact or defined numbers are not used in determining the number of significant figures in a final answer. If you use the equation $A = 4\pi r^2$ to calculate the surface area of a sphere, the number 4 is considered to have an infinite number of significant figures and is therefore not used in determining the uncertainty of the calculated area.

Caution: The final answer of a calculation determined using a calculator often has more digits than any of the numbers in the calculation. You may have to round off the answer to the correct number of significant figures. The rules for rounding off numbers in a calculated answer are:

1. When the number immediately following the last digit to be retained (the first uncertain digit) is less than 5 then the last digit is retained unchanged. If 6.4362 is rounded off to four significant figures it becomes 6.436.

2. When the number immediately following the last digit to be retained is 5* or greater, then increase the last digit by 1. If 6.4366 is to be rounded off to four significant figures it becomes 6.437.

*Note: Your instructor may use an expanded approach: If there are no other numbers or only zeroes beyond the 5, then the last retained digit is increased by 1 if it is odd and left unchanged if it is even. Or if there are numbers other than zero beyond the 5, then the last digit retained is increased by 1. For example, when three significant figures are required, 2.2350 becomes 2.14 (3 is odd and thus it is increased by 1) and 2.1453 becomes 2.15.

Note: Do not round numbers until you have completed your calculation.

EXERCISE 10 Determining uncertain digits

Remembering that measured values are reported to ±1 uncertainty in the last digit, except for those values determined by counting observable objects, or unless otherwise stated, determine the first uncertain digit in each of the following numbers: **(a)** 10.03 kg; **(b)** 5 apples; **(c)** 5.02 ± 0.02 m.

SOLUTION: **(a)** The 3 in 10.03 kg is uncertain to ±1. **(b)** This is an exact measured value determined by counting. There is no uncertain digit. **(c)** The 2 in 5.02 m is uncertain to ±2.

the Atlantic / Pacific method always works!

EXERCISE 11 Determining the number of significant figures

The precision of a measurement is indicated by the number of significant figures associated with the reported value. How many significant figures does each number possess: **(a)** 225; **(b)** 10,004; **(c)** 0.0025; **(d)** 1.0025; **(e)** 0.002500; **(f)** 14,100; **(g)** 14,100.0?

SOLUTION: Try this technique: If a quantity contains a *decimal* point, draw an arrow *starting* at the *left* through all zeros up to the first nonzero digit; the digits remaining are significant. If the quantity does *not* contain a decimal point, draw an arrow *starting* at the *right* through all zeros up to the first nonzero digit; the digits remaining are significant.

(a) 225 ⟵ three significant figures (No decimal point—draw arrow to left)
(b) 10,004 ⟵ five significant figures
(c) ~~0.00~~25 two significant figures (Contains a decimal point—draw arrow to the right)
(d) ⟶ 1.0025 five significant figures
(e) ~~0.00~~2500 four significant figures
(f) 14,1~~00~~—three significant figures; however, because of our lack of knowledge about the significance of the two trailing zeros, this number also could have four or five significant figures
(g) ⟶ 14,100.0 six significant figures

Know this *

EXERCISE 12 Writing numbers in scientific notation

Write the numbers in Exercise 11 using scientific notation.

SOLUTION: Move the decimal in the appropriate direction so that it is to the right of the first nonzero digit reported in the number. If the decimal is moved to the left, multiply the resulting quantity by 10 raised to a power that equals the number of digits the decimal is moved past. If the decimal is moved to the right, the power of 10 is again the number of digits the decimal is moved past, but with a negative sign. That is, a number that is greater than 1 will appear as $A.BC \times 10^x$, while one that is less than 1 will appear as $A.BC \times 10^{-x}$. **(a)** 225 = 2.25×10^2. The decimal is moved two digits to the left, thus the power of 10 is 2. **(b)** 10,004 = 1.0004×10^4. The power of 10 is 4 because the decimal is moved four places to the left. **(c)** 0.0025 = 0002.5×10^{-3} = 2.5×10^{-3}. The power of 10 is −3 because the decimal is moved three places to the right. Note that the nonsignificant zeros are omitted. **(d)** 1.0025. We do not write 1.0025×10^0. **(e)** 0.002500 = 0002.500×10^{-3} = 2.500×10^{-3}. The zeros after the 2.5 are written because they define the number of significant figures. **(f)** 14,100 = 1.4100×10^4 = 1.41×10^4 if the zeros in 14,100 are not significant. **(g)** 14,100.0 = 1.41000×10^4.

EXERCISE 13 Rounding answers in calculations

Round the answers in the following problems to the correct number of significant figures:

(a) 12.25 + 1.32 + 1.2 = 14.770 least precise = 2 sig figs
(b) 13.7325 − 14.21 = −0.4775 Least precise = 4 sig figs
(c) 12300 + 2.11 = 12302.11 least precise = 3 sig figs
no decimal!!

In general, round to the least precise number

SOLUTION: In each problem identify the quantity with the greatest uncertainty and use this uncertainty to determine the correct number of significant figures for the answer. **(a)** The 1.2 has the greatest uncertainty, ±0.1. Therefore, the answer must be rounded to one digit to the right of the decimal point: 14.8. **(b)** 14.21 has the greatest uncertainty, ±0.01. Therefore, the answer must be rounded to two digits to the right of the decimal point: −0.48. *Note*: An answer obtained by subtraction may have fewer significant figures than either number used. **(c)** 12300 has an uncertainty of ±100. The trailing zeros are not identified as being significant; therefore, we normally assume that they are not significant. (If the trailing zeros are significant, they should have been identified as 12300., or preferably 1.2300×10^4.) The answer must be rounded at the hundreds place: 12300. Any digit to the right of the hundreds place is assigned a zero value.

EXERCISE 14 Rounding answers in calculations II

(Skip)

Round the final answer in each of the following calculations:

(a) (1.256)(2.42) = 3.03952

(b) $\frac{16.231}{2.20750} = 7.352661$

(c) $\frac{(1.1)(2.62)(13.5278)}{2.650} = 14.712121$

SOLUTION: **(a)** The least precise number in the calculation is 2.42 (three significant figures). The final answer must be rounded to three significant figures: 3.04. **(b)** The least precise number in the calculation is 16.231 (five significant figures). The final answer must be rounded to five significant figures: 7.3527. **(c)** The least precise number in the calculation is 1.1 (two significant figures). The final answer must be rounded to two significant figures: 15.

Know Celsius ⟷ Kelvin
(Skip Farenheit)

TEMPERATURE AND DENSITY

Reference Points

	Fahrenheit	Celsius	Kelvin
Freezing point of water	32°F	0°C	273.15 K
Boiling point of water	212°F	100°C	373.15 K

Intensive Properties -
Density
Boiling Point
Freezing Point

Two important concepts discussed in Chapter 1 are temperature and density.

- They are both **intensive properties** as their values are independent of the amount of substance present.
- This contrasts with **extensive properties** such as volume and mass, which depend on the amount of substance.

Temperature is a measure of the intensity of heat—the "hotness" or "coldness" of a body.

Just be able to convert between Kelvin + Celsius

- Heat is a form of energy. Heat flows from a hot object to a colder one.
- When there is no heat flow between two objects in contact, they have the same temperature.
- Three temperature scales are used: Celsius (°C), Fahrenheit (°F), and Kelvin (K). You need to know how their reference points differ and how to change between them.
- *Note that a 1-degree interval is the same on both the Celsius and Kelvin scales, but a 1°C interval equals a 1.8°F interval.* The only temperature at which both the Fahrenheit and Celsius scales are equivalent is −40° (−40°C = −40°F). This fact enables us to make conversions between the two scales using the following approach, which is different from the one given in the text.

	°F ⟶ °C	°C ⟶ °F
(a)	Add 40° to °F = (1)	Add 40° to °C = (1)
(b)	$(1) \times \frac{1°C}{1.8°F} = (2)$ **	$(1) \times \frac{1.8°F}{1°C} = (2)$ **1
(c)	Subtract 40° from (2) = °C	Subtract 40° from (2) = °F

**Notice that only step (b) is different. Step (b) converts Fahrenheit to Celsius or Celsius to Fahrenheit using the relationship that a 1-degree Celsius interval equals a 1.8-degree Fahrenheit interval. Also, do not round off until the calculation is finished.

- The following relationship is used to convert between Celsius and Kelvin temperatures.

$$K = \frac{1\ K}{1°C}(C°) + 273.15\ K$$

important *

Density (d) measures the amount of a substance (m) in a given volume (V):

- $d = \frac{\text{mass}}{\text{volume}} = \frac{m}{V}$
- Density varies with temperature because volume changes with temperature.

↑ temp will ↑ Kinetic energy of molecules and therefore ↑ volume ... since $d = m/V$ if you ↑ volume, you ↓ density

- Density can be used to change mass to volume and vice versa for the same substance.
- Chemists commonly use the following units for density: g/mL for liquids, g/cm^3 for solids, and g/L for gases.

*

EXERCISE 15 Converting temperature to a different scale

The temperature on a spring day is around 22°C. What is this temperature in degrees Fahrenheit and degrees Kelvin?

(SKIP)

SOLUTION: To change 22°C to °F, first add 40°:

$$22°C + 40° = 62°C$$

Then multiply by 1.8°F/1°C:

$$(62°C)\left(\frac{1.8°F}{1°C}\right) = 112°F$$

Finally, subtract 40° from 112°F:

$$112°F - 40° = 72°F$$

To calculate degrees Kelvin, write the relationship between the two degrees and substitute for °C:

$$K = \left(\frac{1\,K}{1°C}\right)(°C) + 273.15\,K = \left(\frac{1\,K}{1°C}\right)(22°C) + 273.15\,K = 295\,K \text{ (rounded)}$$

Note: The method presented in the text gives the same answer as follows.

$$°F = \left(\frac{1.8°F}{1°C}\right)(°C) + 32°F$$

Substituting 22°C for °C yields:

$$(°F) = \left(\frac{1.8°F}{1°C}\right)(22°C) + 32°F = 39.6°F + 32°F = 72°F \text{ (rounded)}$$

*

Understand density!

CCl_4 1.60 g/ml

H_2O 1.00 g/ml

$$\text{density} = \frac{\text{mass}}{\text{vol}} = \frac{g}{ml}$$

EXERCISE 16 Comparing densities

At 20°C, carbon tetrachloride and water have densities of 1.60 g/mL and 1.00 g/mL respectively. When water and carbon tetrachloride are poured into the same container, two layers form, one being water and the other carbon tetrachloride. Based on their densities, which one will occupy the lower layer in a container?

SOLUTION: Since carbon tetrachloride and water do not mix together permanently, the heavier substance per unit volume will fall to the bottom of the container. Carbon tetrachloride will be that substance because it has a higher mass per unit volume (density), 1.60 g/mL, than does water, 1.00 g/mL.

EXERCISE 17 Using density in a calculation

Which has the greater mass, 2.0 cm^3 of iron (d = 7.9 g/cm^3) or 1.0 cm^3 of gold (d = 19.32 g/cm^3)?

Find mass when given volume and density.

SOLUTION: The mass of a substance is related to its density by the equation $d = m/V$. Multiplying both sides of the equation by V yields $m = d \times V$. The mass of 2.0 cm^3 of iron is calculated as follows:

$$m = d \times V = \left(7.9 \frac{g}{cm^3}\right)(2.0\ cm^3) = 16\ g$$

Remember that cm^3 is the volume of a solid. (like ml is volume occupied by a liquid)

The mass of 1.0 cm^3 of gold is calculated similarly:

$$m = d \times V = \left(19.32 \frac{g}{cm^3}\right)(1.0\ cm^3) = 19\ g$$

Thus 1.0 cm^3 of gold has a greater mass than 2.0 cm^3 of iron.

DIMENSIONAL ANALYSIS

!

You need to develop the habit of including units with all measurements in calculations. *Units are handled in calculations as any algebraic symbol:*

- Numbers added or subtracted must have the same units.
- Units are multiplied as algebraic symbols: (2 L)(1 atm) = 2 L-atm
- Units are cancelled in division if they are identical. Otherwise, they are left unchanged: (3.0 m)/(2.0 mL) = 1.5 m/mL.

If you already know how to go from grams → grams then you can skip this review.

Dimensional analysis is the algebraic process of changing from one system of units to another. A fraction, called a unit conversion factor, is used to make the conversion. These fractions are obtained from an equivalence between two units. For example, consider the equality 1 in. = 2.54 cm. This equality yields two conversion factors.

$$\frac{1\ in.}{1\ in.} = \frac{2.54\ cm}{1\ in.} \text{ and } \frac{1\ in.}{2.54\ cm} = \frac{2.54\ cm}{2.54\ cm}$$

$$1 = \frac{2.54\ cm}{1\ in.} \text{ and } \frac{1\ in.}{2.54\ cm} = 1$$

Note that the two conversion factors each equal one and are the inverse of one another. They enable you to convert between units in the equality. For example, to convert from centimeters to inches or vice versa:

$$5.08\ \cancel{cm} \times \frac{1\ in.}{2.54\ \cancel{cm}} = 2.00\ in.$$

$$4.00\ \cancel{in.} \times \frac{2.54\ cm}{1\ \cancel{in.}} = 10.2\ cm$$

Note: $\text{given unit} \times \frac{\text{new unit}}{\text{given unit}} = \text{new unit}$

EXERCISE 18 Converting units of volume

Convert 10.5 L to milliliters.

SOLUTION: The required operations for converting 10.5 L to milliliters are as follows:

1. State the general relation required to convert units:

$$?\ mL = (10.5\ L)(\text{conversion factor that changes L to mL})$$

Pages 13-15 are a review of dimensional analysis + conversion units.

2. Find the conversion factor (or factors) that converts L to mL. In this case it is 1000 mL = 1 L. Using this equivalence relation, determine the appropriate ratio of units that converts L to mL. Because we want to change liters to milliliters, we will have to divide by 1 L:

$$\frac{1000\ \text{mL}}{1\ \text{L}}$$

3. Substitute 1000 mL/1 L for the conversion factor in the equation in step 1 and solve the problem.

$$?\ \text{mL} = (10.5\ \cancel{\text{L}})\left(\frac{1000\ \text{mL}}{1\ \cancel{\text{L}}}\right) = 10{,}500\ \text{mL} = 1.05 \times 10^4\ \text{mL}$$

Change the final answer to scientific notation, if necessary. It is sometimes more convenient to change all numbers to scientific notation before doing the mathematics of the problem:

$$?\ \text{mL} = (1.05 \times 10^1\ \cancel{\text{L}})\left(\frac{1 \times 10^3\ \text{mL}}{1\ \cancel{\text{L}}}\right) = 1.05 \times 10^4\ \text{mL}$$

4. Check that the units properly cancel to yield the desired unit. In this problem, if we had used the conversion factor 1 L/1000 mL instead of 1000 mL/1 L, the result would have been:

$$?\ \text{mL} \neq (1.05 \times 10^1\ \text{L})\left(\frac{1\ \text{L}}{1 \times 10^3\ \text{mL}}\right) \neq 1.05 \times 10^{-2}\,\frac{\text{L}^2}{\text{mL}}$$

The unit L^2/mL does not equal mL; therefore, we know we have used the wrong conversion factor.

EXERCISE 19 Converting SI units

Convert 6.23 ft^3 to the appropriate SI unit.

SOLUTION: The appropriate SI unit is meter^3 because ft^3 is based on an English unit of length, the foot. Use the same method shown in Exercise 18 to convert ft^3 to m^3.

1. ? m^3 = (6.23 ft^3) (conversion factor that changes ft^3 to m^3)
2. From Table 1.3 you find that 3.272 ft = 1 m, but there is no unit conversion given for ft^3 to m^3. What you must recognize is that if

$$1 = \frac{1\ \text{m}}{3.272\ \text{ft}}$$

then you can cube both sides of the expression

$$(1)^3 = \left(\frac{1\ \text{m}}{3.272\ \text{ft}}\right)^3 = \frac{1\ \text{m}^3}{(3.272)^3\ \text{ft}^3} = 1$$

3. Converting units yields:

$$?\ \text{m}^3 = (6.23\ \cancel{\text{ft}^3})\left(\frac{1\ \text{m}^3}{(3.272)^3\ \cancel{\text{ft}^3}}\right) = 0.178\ \text{m}^3$$

EXERCISE 20 Using multiple conversion factors

The Empire State Building in New York City was for many years the tallest building in the world. Its height is 484.7 yd to the top of the lightning rod. Convert this distance to meters.

SOLUTION: The conversion of 484.7 yd to meters is accomplished by the method used in Exercise 18:

(1) State the general relation required to convert units:

? m = (484.7 yd)(conversion factor or series of factors that changes yards to meters)

(2) Look up the required equivalences in appropriate tables. From the tables in this chapter of the *Student's Guide*, we find that 3.272 ft = 1 m. Because no conversion between yards and feet is given, you will have to go to another source or remember from your past experience that 1 yd = 3 ft. To convert yards to meters you will have to make a series of unit conversions that will look like

$$(\cancel{\text{yd}})\left(\frac{\cancel{\text{ft}}}{\cancel{\text{yd}}}\right)\left(\frac{\text{m}}{\cancel{\text{ft}}}\right) = \text{m}$$

Thus the required conversion factors are

$$\frac{3\text{ ft}}{1\text{ yd}} \text{ and } \frac{1\text{ m}}{3.272\text{ ft}}$$

(3) You could solve the problem in two steps:

$$(484.7\ \cancel{\text{yd}})\left(\frac{3\text{ ft}}{1\ \cancel{\text{yd}}}\right) = 1454\text{ ft}$$

$$(1454\ \cancel{\text{ft}})\left(\frac{1\text{ m}}{3.272\ \cancel{\text{ft}}}\right) = 444.3\text{ m}$$

However, it is usually simpler to do it all in one step:

$$?\text{ m} = (484.7\ \cancel{\text{yd}})\left(\frac{3\ \cancel{\text{ft}}}{1\ \cancel{\text{yd}}}\right)\left(\frac{1\text{ m}}{3.272\ \cancel{\text{ft}}}\right) = 444.3\text{ m}$$

(4) Check that the units properly cancel. Inspection of the previous equation shows that they do.

EXERCISE 21 Converting units of measurements having a ratio of units

Florence Griffith Joyner (USA) set a world record in the women's 100 m dash on July 16, 1988, running the distance in 10.49 s. This record has not been broken as of January 31, 2004. Assuming that the distance is exactly 100 m, what is her average speed in miles per hour?

SOLUTION: Joyner's average speed in meters per second is

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{100\text{ m}}{10.49\text{ s}} = 9.533\frac{\text{m}}{\text{s}}$$

(The answer has four significant figures because the problem says assume the distance is exact.)
This is not the units requested in the problem; the ratio m/s has to be converted to mi/hr.

(1) The conversion of units can be viewed as follows:

$$\frac{\text{m} \rightarrow \text{mi}}{\text{s} \rightarrow \text{min} \rightarrow \text{hr}}$$

Table 1.3 gives you the necessary equivalences between meters and miles and you should know the equivalences between seconds and minutes and minutes and hours:

1 mi = 1609 m
60 s = 1 min
60 min = 1 hr

(2) Using these equivalences you can now make the unit conversions by converting meters to miles in the numerator and also seconds to minutes to hours in the denominator:

$$9.533\frac{\text{m}\left(\frac{1\text{ mi}}{1609\text{ m}}\right)}{\text{s}\left(\frac{1\text{ min}}{60\text{ s}}\right)\left(\frac{1\text{ hr}}{60\text{ min}}\right)} = 21.33\frac{\text{mi}}{\text{hr}}$$

(3) An alternative way of doing the calculation is to do the conversions by first converting meters to miles and then seconds to minutes to hours as follows:

m → mi

$$\left(9.533\frac{\text{m}}{\text{s}}\right)\left(\frac{1\text{ mi}}{1609\text{ m}}\right)\left(\frac{60\text{ s}}{1\text{ min}}\right)\left(\frac{60\text{ min}}{1\text{ hr}}\right) = 21.33\frac{\text{mi}}{\text{hr}}$$

$$\text{and } \frac{1}{\text{s}} \rightarrow \frac{1}{\text{min}} \rightarrow \frac{1}{\text{hr}}$$

SELF-TEST QUESTIONS - useful to review.

Having reviewed key terms in Chapter 1, 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: (terms on next page)

1.1 Increases in value with decreasing volume.

1.2 0.01200 contains four of these.

1.3 Heat emitted from burning wood is not an example of this.

1.4 The kg unit in 1.00 kg of silver tells you about this measurement.

1.5 This property does not change with amount of material.

1.6 This property changes with amount of material.

1.7 The temperature of water at 75°C is an example.

1.8 The freezing of water is an example.

1.9 A chemical reaction is an example.

1.10 The ability of carbon to form carbon dioxide is an example.

1.11 Makes up the composition of a compound.

1.12 HF is an example.

1.13 A temperature scale with more divisions between freezing point and melting point of water than in the Kelvin scale.

Terms:

(a) Celsius
(b) Chemical change
(c) Chemical property
(d) Compound
(e) Density
(f) Elements
(g) Extensive
(h) Intensive
(i) Mass
(j) Matter
(k) Physical property
(l) Physical change
(m) Significant figures

1.14 When ice completely melts in a glass of water, there is a change from a heterogeneous *mixture* to a homogeneous one.

1.15 The *SI unit* for mass is the gram.

1.16 A *conversion factor* contains a ratio of units.

1.17 Coke is a *substance*.

1.18 −273.15°C is equivalent to 0 K.

1.19 A student measures the mass of an object three times: 3.60 g, 3.90 g, and 3.75 g. The object actually has a mass of 3.75 g. The student's measurements show good *precision*.

1.20 In problem 1.19, the average value is 3.75 g. Therefore, the student's measurements gave good *accuracy*.

1.21 A *solution* is a homogeneous mixture of two or more substances.

1.22 The basic unit of mass in the *metric system* is the same as in the SI system of units.

1.23 A *chemical reaction* involves a change in the chemical composition of substances.

1.24 A *change of state* of $N_2(l)$ to $N_2(g)$ is a chemical change.

1.25 10 g of CO_2 *gas* occupies less space than 10 g of CO_2 solid in a one liter flask.

1.26 A *solid* consists of particles closer together than in the gas state.

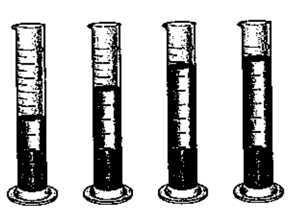
1.27 A *liquid* is slightly compressible.

1.28 According to the *law of constant composition (definite proportions)*, H_2O and H_2O_2 represent the same substance.

Problems and Short-Answer Questions

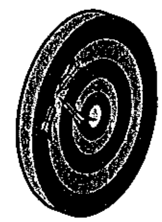
True-False Statements:

1.29 Shown below is a 10 mL graduate cylinder. It is initially filled with 5.0 mL of water. 4.0 g of a solid object with a density of 2.0 g/mL is added. Which figure below best represents the volume of water in the graduate cylinder after the solid is added?

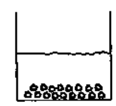


(a) (b) (c) (d)

1.30 Characterize the following dartboard in terms of precision and accuracy of the results.



1.31 Identify the type of substance described by the following figure:



1.32 You calculate your income for last year as $42,125.00. However, when you file your income tax form, you report an income of $42,000.00. You do this because you want to report a rounded-off income figure to two significant figures. Why would an IRS agent be unhappy with your use of significant figures?

1.33 How many significant figures do the following numbers possess?

(a) 20.03 kg
(b) 1.90×10^3 L
(c) 120 m
(d) 10 dollar bills
(e) 0.00067 cm^3

1.34 Convert the following numbers to scientific notation form with the correct number of significant figures.

(a) 0.00067 cm^3
(b) 210.0 m
(c) 0.040 L
(d) 46900.0 g
(e) 200 rattlesnakes

1.35 Complete the following calculations and round off answers to the correct number of significant figures:

(a) $11.020 + 300.0 + 2.0030 =$
(b) $1211 + 2.205 - 1.70 =$
(c) $\dfrac{(1.425 \times 10^2)(2.61 \times 10^3)}{2.89 \times 10^5} =$
(d) $\dfrac{(0.012)(0.100)}{11.0265} =$

1.36 Make the following conversions.

(a) 1 ML to cubic centimeters.
(b) $\dfrac{1}{1000}$ g to kg
(c) 8.00 qt to milliliters
(d) 16 lb to grams

Chapter 2

Atoms, Molecules, and Ions

OVERVIEW OF THE CHAPTER

2.1, 2.2, 2.3 ATOMS

Learning Goals: You should be able to:

1. Describe the composition of an atom in terms of protons, neutrons, and electrons.
2. Give the approximate size, relative mass, and charge of an atom, proton, neutron, and electron.
3. Write the chemical symbol for an element, having been given its mass number and atomic number, and perform the reverse operation.
4. Describe the properties of the electron as seen in cathode rays. Describe the means by which J.J. Thomson determined the ratio *e/m* for the electron.
5. Describe Millikan's oil-drop experiment and indicate what property of the electron he was able to measure.
6. Cite the evidence from studies of radioactivity for the existence of subatomic particles.
7. Describe the experimental evidence for the nuclear nature of the atom.

no nuclear in AP

2.4, 2.5 PERIODIC TABLE: ATOMIC WEIGHTS AND ARRANGEMENT OF ATOMS

Learning Goals: You should be able to:

1. Use the unit of atomic mass unit (amu) in calculation of masses of atoms.
2. Define the term atomic weight and calculate the atomic weight of an element given its natural distribution of isotopes and isotopic masses.
3. Use the periodic table to determine the atomic number, atomic symbol, and atomic weight of an element.
4. Define the terms group and period and recognize the common groups of elements.
5. Use the periodic table to predict whether an element is metallic, nonmetallic, or metalloid.

2.6, 2.7 MOLECULES, MOLECULAR COMPOUNDS, IONIC COMPOUNDS, AND IONS

Learning Goals: You should be able to:

1. Define the term molecule and recognize which elements typically combine to form molecules.
2. Distinguish between empirical and molecular formulas.
3. Draw the structural and ball-and-stick formulas of a substance given its chemical formula and the linkage between atoms.

4. Use the periodic table to predict the charges of monatomic ions of non-transition elements.
5. Write the symbol and charge for an atom or ion having been given the number of protons, neutrons, electrons and perform the reverse operation.
6. Determine whether a substance is likely to be ionic or molecular. (covalent)
7. Write the simplest formula of an ionic compound having been given the charges of ions from which it is made.

NOMENCLATURE: SIMPLE INORGANIC, MOLECULAR, AND ORGANIC COMPOUNDS

Learning Goals: You should be able to:

1. Write the name of a simple inorganic compound having been given its chemical formula and perform the reverse reaction.
2. Write and name the polyatomic ions in Table 2.5.
3. Write and name acids based on anions whose names end in -ide, -ate, and -ite.
4. Write the name of simple binary molecular compounds and perform the reverse operation.
5. Define the terms hydrocarbon, alkane, and alcohol and be able to name simple alkanes and alcohols, having been given the chemical formula, and perform the reverse operation.

TOPIC SUMMARIES AND EXERCISES

ATOMS

This chapter focuses on the structure of atoms, elements in the periodic table, how atoms combine to form substances, and naming inorganic compounds. The work of many scientists, such as John Dalton, J.J. Thompson, and R. Millikan, provided experimental information for development of a modern understanding of the structure of atoms. You should know the contributions of the key historical figures discussed in this chapter: See Exercise 2.

Atoms are the smallest particles of an element that have all the same chemical properties of that element. You need to know the following information about the structure of atoms:

- The mass of a single atom is extremely small, less than 10^{-21} g.
- The center of an atom contains a small **nucleus.** A nucleus contains **protons** and **neutrons**, which make up most of the mass of an atom.
- The volume of an atom mostly consists of **electrons**.
- Protons (positively charged) and neutrons (no charge) have essentially the same mass. Electrons (negatively charged) are significantly less massive.

All atoms of the same element have the same number of protons.

- Atoms that have the same number of protons but differ in their number of neutrons are called **isotopes.**
- The general symbol for an isotope is

This. →

$^{A}_{Z}X$ — **Mass number** = number of protons + number of neutrons; Symbol of element; **Atomic number** = number of protons

Atomic Number ("Z") = # protons
* In a neutrally charged atom, #protons = # electrons

Note: The atomic number Z defines an element; it also tells you how many electrons a neutral atom of that element possesses. Thus, the number of neutrons equals A - Z.

EXERCISE 1 Writing isotopic symbols

Write the nuclear-isotope symbols for the four isotopes of sulfur with 16, 17, 18, and 20 neutrons, respectively.

SOLUTION: The atomic number of sulfur is 16, and all of its isotopes will have 16 protons. The mass number of each isotope is the sum of its number of neutrons plus its number of protons: $16 + 16 = 32$; $16 + 17 = 33$; $16 + 18 = 34$; and $16 + 20 = 36$. The nuclear-isotope symbols are $^{32}_{16}S$, $^{33}_{16}S$, $^{34}_{16}S$, and $^{36}_{16}S$. Note that the subscript before each elemental symbol (that is, the atomic number) is the same for all isotopes because the number of protons is invariant for sulfur.

atomic mass
Element
atomic number

EXERCISE 2 Developments in the history of atomic structure

The composition of the atom was determined from the experiments of many individuals. For each experiment, state the principal person associated with the experiment, the particle studied, and the property of the particle determined: **(a)** α scattering; **(b)** cathode ray; **(c)** oil drop.

SOLUTION: **(a)** E. Rutherford determined the relative volume and mass of the nucleus in the atom by observing the scattering of α particles. **(b)** J. J. Thomson determined the *e/m* ratio of the electron by using a cathode-ray tube. **(c)** R. Millikan determined the charge of an electron, by observing the motion of charged oil drops in an electric field.

Thomson - discovered electron (plum pudding)
Millikan - charge

EXERCISE 3 Using isotopic symbols to determine nuclear composition

Find the number of protons, electrons, and neutrons in the following isotopes:

(a) $^{40}_{20}Ca$; **(b)** $^{238}_{92}U$.

SOLUTION: **(a)** The number of protons is the subscript number shown in the isotopic symbol: 20. The number of electrons in a neutral atom is the same as the number of protons: 20. The number of neutrons equals the mass number minus the number of protons. This is the superscript number (mass number) minus the subscript number: $40 - 20 = 20$. **(b)** Preceding as in (a), the number of protons is 92, the number of electrons is 92, and the number of neutrons is $238 - 92 = 146$.

Isotopes vs. Ions
* Change neutrons → Change Isotope
* Change electrons → Change to an Ion

EXERCISE 4 Changing the number of electrons in an isotope

If an electron is added or removed from ^{35}Cl, what changes occur to the isotope?

SOLUTION: The isotope is still ^{35}Cl because the number of protons and the number of neutrons do not change; however, the number of electrons does. If electrons, which are negatively charged, are added to an isotope, the isotope gains negative charge. Therefore if one electron is added to ^{35}Cl, it becomes $^{35}Cl^-$. Similarly, if an electron is removed, the isotope loses negative charge and becomes positively charged because there are more protons than electrons: $^{35}Cl^+$.

PERIODIC TABLE: ATOMIC WEIGHTS AND ARRANGEMENT OF ATOMS

The masses of isotopes are measured relative to the atomic mass of $^{12}_{6}C$, which is defined to have a mass of exactly 12 **atomic mass units** (amu). A modern mass spectrometer is used to determine accurate relative masses.

- Except for a few elements, most elements have a distribution of naturally occurring isotopes. The concept of atomic weight is based on determining the average of the masses of naturally occurring isotopes weighted according to their abundances:

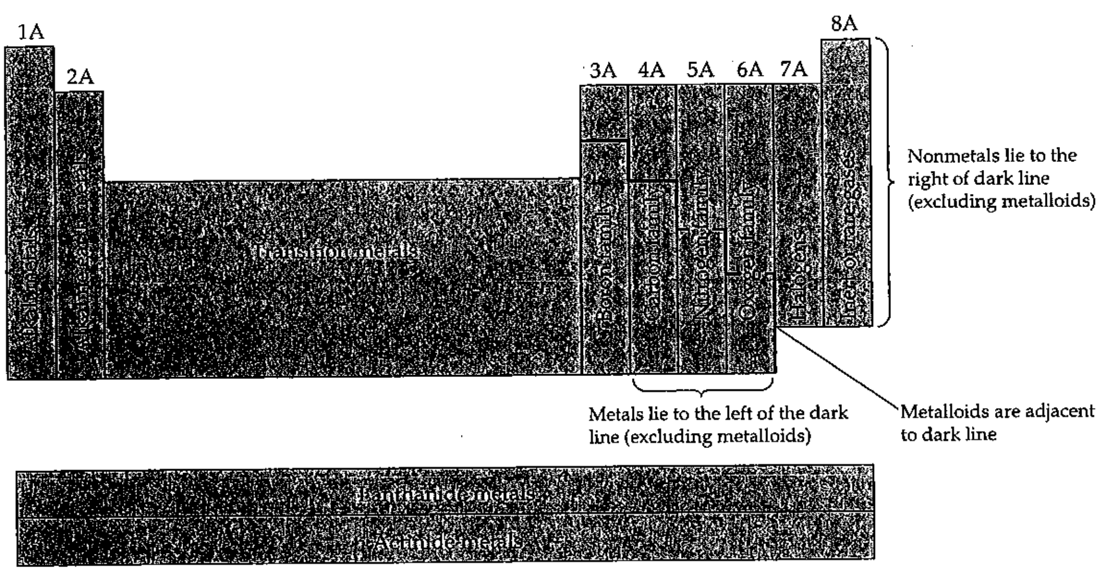
  atomic weight = % abundance × isotope 1 + % abundance × isotope 2 + ⋯⋯ + % abundance × isotope n.
- The relationship between amu and grams is 1.66054×10^{-24} g = 1 amu.

The periodic table lists elements in order of increasing atomic number. It is very useful in helping us to organize trends in the chemical and physical properties of elements. The manner in which elements are grouped in vertical columns(families) and horizontal rows(periods) helps you to remember that

- *Metals* are on the left side and middle (except hydrogen).
- *Nonmetals* are on the right side.
- *Metalloids* have properties of both metals and nonmetals. They are identified by the dark diagonal line toward the right side of the periodic table.
- Elements in vertical columns, called *families* or *groups*, exhibit similar chemical and physical properties, whereas elements in a horizontal row (*period*) exhibit different properties. The names of groups in the periodic table and the general location of metals, metalloids, and nonmetals are shown in Figure 2.1.

EXERCISE 5 Understanding the concept of average atomic weight

(a) What is the sum of percentages of naturally occurring isotopes for an element? (b) The following problem shows the concept of atomic weights applied to a different situation. What is the average weight of a class of students given the following information about the class: (1) The average woman weighs 122 pounds; (2) the average man weighs 165 pounds; and (3) the percentage of men in the class is 45.0%.



▲ **FIGURE 2.1** Families in the periodic table. Note that hydrogen, the first element in family 1A, is not an alkali metal.

SOLUTION: (**a**) The sum of percentages must add to 100 %; however, in experimental situations it rarely adds exactly to this value because of experimental error. (**b**) The average atomic weight of the class is calculate from the relation:

$$\text{average weight} = \%\text{ men} \times \text{average mass of men} + \%\text{ women} \times \text{average mass of women}$$

We are given the percentage of men but not that of women. We can calculate the percentage of women from relation:

$$100.0\% = \%\text{ men} + \%\text{ women} = 45.0\% + \%\text{ women}$$

$$\%\text{ women} = 100\% - 45.0\% = 55.0\%$$

Therefore, the average weight of the class is:

$$\text{average weight} = (0.450)(165\text{ pounds}) + (0.550)(122\text{ pounds}) = 141\text{ pounds}$$

EXERCISE 6 Calculating the atomic weight of an element

What is the atomic weight of antimony if it has only two naturally occurring isotopes, Sb-121 with an isotopic mass of 120.904 amu and an abundance of 57.21% and Sb-123 with an isotopic mass of 122.904 amu and an abundance of 42.79%?

The average atomic weight of an element (listed on periodic table) is the sum of the weighted averages of all the naturally occurring isotopes of that element.

SOLUTION: We can use the concept of average atomic mass discussed in the previous exercise to solve for the atomic weight of antimony:

$$\begin{aligned}\text{Atomic weight of Sb} &= 120.904\text{ amu}(0.5721) + 122.904\text{ amu}(0.4279)\\ &= 69.17\text{ amu} + 52.59\text{ amu}\\ &= 121.76\text{ amu}\end{aligned}$$

EXERCISE 7 Using a periodic table

The position of an element in the periodic table is a function of its atomic composition. Elements are arranged in order of increasing atomic number. The atomic number of an atom equals its number of protons. By referring to a periodic table, determine which *neutral* atom: (**a**) contains 50 protons; (**b**) contains 17 electrons; (**c**) has an atomic number of 56.

SOLUTION: (**a**) Tin (Sn). Its atomic number, 50, corresponds to 50 protons. (**b**) Chlorine (Cl). Its atomic number corresponds to 17 protons. This is also the number of electrons in the neutral atom. (**c**) Barium (Ba). Its atomic number is 56.

EXERCISE 8 Identifying the family of an element in the periodic table

What is the name of the family in which the following elements occur: I, Na, S, and Ca?

SOLUTION: By referring to Figure 2.2 and a periodic table we find: I, halogen; Na, alkali metal; S, oxygen family; and Ca, alkaline-earth metal.

MOLECULES, MOLECULAR COMPOUNDS, IONIC COMPOUNDS, AND IONS

Molecular compounds consist of small particles called molecules.

- A **molecule** is a small particle consisting of two or more atoms combined together in a discrete unit. Molecules have their own chemical and physical properties, which differ from the properties of the elements forming them.
- Molecules typically consist of nonmetallic elements.
- You should know the seven elements that form homonuclear diatomic molecules: H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2.
- Subscripts in a **molecular formula** tell you how many atoms are actually present

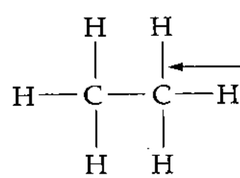
C_2H_4

2 atoms of carbon — 4 atoms of hydrogen

- An **empirical formula** shows only the simplest whole number ratio of atoms. For example, C_2H_4 is a molecular formula; its empirical formula, CH_2, is obtained by dividing the subscripts in the molecular formula by 2:

$$CH_2 = C_{2/2}H_{4/2}$$

- A **structural formula** shows "how atoms are joined together"—that is, the relative orientation and position of bonded atoms. For example, the structural formula of C_2H_6 shows the following arrangement of atoms:



A line between two atoms represents a chemical bond

An **ionic compound** consists of positively charged ions (cations) and negatively charged ions (anions).

- Ionic compounds typically contain a metal and a nonmetal.
- **Cations** are positively charged ions of metallic elements formed by the loss of electrons; e.g., Na^+ is formed by Na losing one electron.
- **Anions** are negatively charged ions of nonmetallic elements formed by the addition of electrons; e.g., S^{2-} is formed by S gaining two electrons.
- Ionic compounds do not contain unique identifiable pairs of ions. Therefore, empirical formulas are used to describe their composition.
- Ionic compounds may contain polyatomic ions such as NO_3^- as in $NaNO_3$. Polyatomic ions are like molecules except that they carry a charge.
- You should know the most common charge carried by elements in the following families:

Memorize ~~Beware~~ the formulas + charges of the common polyatomic IONS!

Know this!

Alkali metals	Alkaline-earth metals	Transition metals	Halogens	Oxygen family
1+	2+	2+, 3+	1−	2−

EXERCISE 9 Interpreting chemical formulas

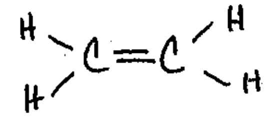
Interpret the following chemical formulas: (a) N_2; (b) NH_3; (c) NH_4^+; (d) H_2SO_4.

SOLUTION: (a) N_2 is a homonuclear (having atoms of the same type) molecule composed of two nitrogen atoms; this is the elemental form of nitrogen. (b) NH_3 is a heteronuclear (having more than one type of atom) molecule composed of one nitrogen atom and three hydrogen atoms. (c) NH_4^+ is an ion that has a 1+ charge and is composed of one nitrogen atom and four hydrogen atoms. (d) H_2SO_4 is a heteronuclear molecule composed of two hydrogen atoms, one sulfur atom, and four oxygen atoms.

EXERCISE 10 Writing empirical and chemical formulas

(a) Do C_2H_4 and $2\ CH_2$ represent the same substance? (b) What are the empirical formulas for B_2H_6, H_2O, and $C_2H_2O_4$?

ethene
C_2H_4 ("ethylene")



SOLUTION: (a) The formulas do not represent the same substance. C_2H_4 represents the molecular formula for the substance acetylene, a gas used in welding. $2\ CH_2$ represents two empirical formulas of acetylene. (b) To determine an empirical formula divide each subscript of a molecular formula by the largest whole number that goes into each subscript (the greatest common denominator)

$$B_{2/2}H_{6/2} = BH_3$$

There is no way to further reduce the subscripts in H_2O—this is the molecular and empirical formula for water.

$$C_{2/2}H_{2/2}O_{4/2} = CHO_2$$

EXERCISE 11 Writing chemical formulas of ionic substances

Write the formula for: (a) a neutral polyatomic compound consisting of one barium ion with a 2+ charge (written Ba^{2+}) and chlorine ions with a 1− charge (written Cl^-). (b) A polyatomic ion that has a 1− charge and consists of one boron ion with a 3^+ charge (written B^{3+}) and fluorine ions with a 1− charge (written F^-).

Remember/Criss Cross Method
i.e. "barium chloride"
Ba^{+2} Cl^{-1}
$BaCl_2$
(or)
Na^+ SO_4^{-2}
"sodium sulfate" Na_2SO_4
or
Ca^{+2} OH^-
"Calcium hydroxide" $Ca(OH)_2$

SOLUTION: (a) The sum of positive and negative charges for the ions must equal zero in a neutral compound. Two Cl^- ions are required to balance the 2+ charge of Ba^{2+}: $BaCl_2$. (b) The sum of positive and negative charges for the ions must equal the 1− charge of the compound. Four F^- ions are required so that, when they are combined with B^{3+}, the sum of the charges is 1−: BF_4^- $[+3 + 4(-1) = -1]$.

EXERCISE 12 Determining the number of electrons possessed by a molecule or ion

Determine the total number of electrons in each of the following: (a) NH_3; (b) NH_4^+; (c) NH_2^-.

SOLUTION: Determine the number of electrons for each atom in each chemical formula, sum these numbers, and then adjust the number for the charge shown. For each negative charge, one electron is added; for each positive charge, one electron is subtracted. (a) The number of electrons for each atom is the same as the number of protons. Each hydrogen atom has one proton and each nitrogen atom has 7 protons. Thus, for NH_3 the number of electrons is $(1 \times 7) + (3 \times 1) = 10$ electrons. The charge of the compound is zero, thus no adjustment for charge is necessary. (b) The number of electrons for NH_4^+ with no charge is $(1 \times 7) + (4 \times 1) = 11$ electrons. However, the species has a positive charge, thus one electron must be removed so that there are more protons than electrons: $11 - 1 = 10$ electrons for NH_4^+. (c) The number of electrons for NH_2^- with no charge is $(1 \times 7) = (2 \times 1) = 9$ electrons. The species has a charge of 1−, thus one electron must be added so that there are more electrons than protons: $9 + 1 = 10$ electrons.

EXERCISE 13 Writing binary chemical formulas

Write the expected formula when the following elements combine to form compounds: (a) Al and O; (b) B and Cl; (c) Ca and F; (d) Na and I; (e) S and O; and (f) H and Ca.

SOLUTION: (a) Al forms the 3+ state in salts. Oxygen forms the oxide ion, O^{2-}. Al_2O_3. (b) Boron is a member of the family containing aluminum; thus, you should expect it to behave in a similar manner to aluminum. BCl_3. (c) Calcium forms Ca^{2+} and flourine forms F^-. CaF_2. (d) Na forms Na^+ and iodine in the presence of metals forms I^-. NaI. (e) Sulfur and oxygen, both nonmetals, form nonmetal oxides: For example, $SO_2(g)$ and $SO_3(g)$. (f) Ca is an active metal. In the presence of an active metal hydrogen forms the hydride ion, H^-: CaH_2.

NOMENCLATURE: SIMPLE INORGANIC, MOLECULAR, AND ORGANIC COMPOUNDS

Chemists not only write chemical formulas for compounds but also give them reasonably systematic names. Rules for naming simple inorganic and binary molecular compounds are found in Section 2.8 in the text. *These rules should be committed to memory; and you should then practice them by naming anions, cations, ionic compounds, oxyanions, acids, and simple binary molecular compounds.* Attempt Exercises 13-19 *after* you have reviewed nomenclature rules. The solutions to many of the exercises provide explanations which reinforce what you have learned.

In Section 2.9 the concept of organic chemistry and hydrocarbons and the names of simple hydrocarbons are introduced. Hydrocarbons are compounds containing hydrogen and carbon atoms.

- One class of hydrocarbons is called an *alkane*, in which each carbon atom is bonded to four other atoms; all bonds are single. You need to learn the prefixes for alkanes in Table 2.6 in the text.
- A *functional group* is a nonhydrogen atom or group of atoms that replaces a hydrogen atom in a hydrocarbon. *Alcohols* contain the – OH functional group and their names end in – ol.

EXERCISE 14 Correcting chemical formulas when they are written incorrectly

The following chemical formulas are written incorrectly. Correct them so that they are in the proper form: (a) ClNa; (b) CaClCl; (c) $NH_4NH_4SO_4$; (d) H^2S; (e) FXeF.

SOLUTION: (a) NaCl. The cation (Na^+) is always placed before the anion (Cl^-). (b) $CaCl_2$. A subscript is used to indicate the number of Cl^- ions present. (c) $(NH_4)_2SO_4$. If there is more than one type of polyatomic ion present, its formula is enclosed in parentheses, and subscripts are used as necessary. (d) H_2S. The 2 must be a subscript when it indicates the number of the same type of atoms or ions present. (e) XeF_2.

EXERCISE 15 Naming cations

Name the following cations: (a) H^+; (b) Na^+; (c) Be^{2+}; (d) Al^{3+}; (e) Co^{2+}; (f) Co^{3+}.

SOLUTION: Monatomic cations have the same names as the elements from which they are formed. Some monatomic cations commonly exist in two common ion states; for example, Fe^{2+} and Fe^{3+}. When naming such ions, indicate the charge of the ion in Roman numerals enclosed within parentheses after the name of the element. Sometimes an older method is used, in which the suffix *-ic* indicates the higher charge and *-ous* the lower one. This method is further complicated by the fact that the Latin name of the element is often used. Thus Fe^{2+} is iron(II) or ferrous, and Fe^{3+} is iron(III) or ferric. When in doubt, use the newer method. (a) Hydrogen. (b) Sodium. (c) Beryllium. (d) Aluminum. (e) Cobalt(II) or cobaltous. (f) Cobalt(III) or cobaltic.

Stock System
(Roman Numerals)
Iron II chloride
is Fe^{+2} and Cl^{-1}
$FeCl_2$

EXERCISE 16 Naming anions

Name the following anions: (a) F^-; (b) S^{2-}; (c) OH^-; (d) CN^-; (e) PO_4^{3-}; (f) PO_3^{3-}; (g) IO^-; (h) IO_2^-; (i) IO_3^-; (j) IO_4^-; (k) HS^-; (l) $H_2PO_4^-$.

SOLUTION: (a) Fluoride. Monatomic anions have *-ide* added to the stem of the element's name. (b) Sulfide. (c) Hydroxide. Treated as a monatomic anion. (d) Cyanide. Treated as a monatomic anion. (e) Phosphate. When two oxyanions exist, as in the cases of PO_4^{3-} and PO_3^{3-}, the one with more oxygen atoms ends in *-ate*. (f) Phosphite. The one with fewer oxygen atoms ends in *-ite*. (g) Hypoiodite. The prefix *hypo-* indicates one less oxygen atom than in IO_2^-. (h) Iodite. (i) Iodate. (j) Periodate. The prefix *per-* indicates one more oxygen atom than in IO_3^-. (k) Hydrogen sulfide. *Hydrogen* is added to the name of the anion to indicate that there is one hydrogen atom present. (l) Dihydrogen phosphate. *Dihydrogen* is added to phosphate to indicate that there are two hydrogen atoms present.

Know -ide vs -ate
ie. hydrogen sulfide (H_2S)
hydrogen sulfate (H_2SO_4)

EXERCISE 17 Naming ionic and molecular compounds

Name the following compounds: (a) ZnS; (b) $(NH_4)_2SO_4$; (c) $FeCl_2$; (d) $KClO_4$; (e) $SnCl_2$; (f) PCl_3; (g) SF_6; (h) CO.

sodium nitride (Na_3N)
sodium nitrate (Na_3NO_3)

SOLUTION: Remember: When naming ionic compounds cations are named before anions; in molecular compounds the atoms are named in the order found in the molecular formula. For binary molecular compounds, the first atom is given its elemental name, and the second is named as if it were a negative ion. (a) Zinc sulfide. Zinc exists only in the 2+ state. (b) Ammonium sulfate. The use of *di-* before *ammonium* to indicate two NH_4^+ ions is not necessary because the −2 charge of SO_4^{2-} automatically requires two NH_4^+ ions. (c) Iron(II) chloride or, using the older method, ferrous chloride. (d) Potassium perchlorate. (e) Tin(II) chloride or stannous chloride. (f) Phosphorus trichloride. A compound formed from two nonmetallic elements is named as if it were an ionic compound, with the appropriate prefixes added to indicate how many atoms are present. (g) Sulfur hexafluoride. The prefix *hexa-* indicates six fluorine atoms. (h) Carbon monoxide.

EXERCISE 18 Naming acids

Name the following acids in water: (a) HCN; (b) H_2S; (c) H_2CO_3; (d) H_3PO_4.

SOLUTION: (a) Hydrocyanic acid. For hydrogen acids formed from monatomic anions, *hydro-* is added to the name of the anion—in this case, *cyanide*—and the *-ide* ending of the anion is changed to an *-ic* ending. (b) Hydrosulfuric acid. Again, the prefix *hydro-* indicates a hydrogen acid, and *-ic* is added to the end of sulfur. (c) Carbonic acid. When an acid is derived from a polyatomic anion whose name ends in *-ate*—*carbonate* (CO_3^{2-}) in this case—the ending *-ic* is added to the name of the central atom of the polyatomic anion, and no prefix is used. (d) Phosphoric acid. Its name is derived from the polyatomic anion *phosphate* (PO_4^{3-}). The

-ate ending is dropped and is changed to *-oric*. The *-ic* ending for acids is associated with acids formed from anions ending in *-ate*.

EXERCISE 19 Naming molecular compounds

Name the following substances: (a) SF_6; (b) NO; (c) N_2O; (d) P_2O_5.

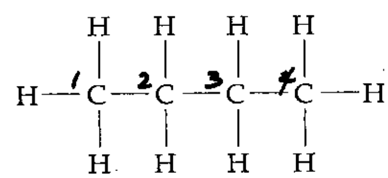
SOLUTION: We will use the prefixes in Table 2.6 of the text to name the substances. (a) sulfur hexaflouride; (b) nitrogen oxide; (c) dinitrogen oxide; (d) diphosphorus pentaoxide.

Organic Chemistry is NOT on the AP exam. We will review it for purposes of recognizing organic molecules which do appear on the exam.

EXERCISE 20 Writing names and chemical formulas of alkanes

Consider the alkane called butane. (a) Assuming the carbon atoms are in a straight chain, write a structural formula for butane. (b) What is its molecular formula? (c) If one of the hydrogen atoms attached to a carbon atom at the end of the chain is replaced with – OH, what is its name?

SOLUTION: (a) The ending *-ane* in propane tells you that it is an alkane. Referring to Table 2.6 in the text we read that the prefix *but-* means four carbon atoms. We can write



Butane

Skip Nuclear - NOT on the AP exam at all.

(b) With the structural formula written we can determine the molecular formula by counting the number of carbon and hydrogen atoms: C_4H_{10}. (c) If one of the hydrogen atoms at the end of the carbon chain is replaced by – OH, an alcohol functional group, the following alcohol is produced: Butanol.

SELF-TEST QUESTIONS

Key Terms

Having reviewed key terms in Chapter 2, 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:

2.1 A general name for a positive ion.

2.2 A sulfide ion is an example of this species.

2.3 The least massive particle in an atom as discussed in the chapter.

2.4 A particle possessing no charge in the nucleus.

2.5 Helium atom with a +2 charge.

2.6 $^{85}_{37}Rb$ and $^{87}_{37}Rb$ are examples.

2.7 The number 85 in $^{85}_{37}Rb$.

2.8 The number 37 in $^{85}_{37}Rb$.

2.9 HCl is an example of this species.

2.10 Occupies a small volume of an atom.

2.11 NO_3^- is an anion and also called this type of particle.

2.12 Shows smallest whole-number ratio of atoms.

2.13 Shows arrangement of atoms and bonds.

2.14 Name of any type of charged atom.

2.15 H_2O_2 is an example of this type of chemical formula.

2.16 Contains a beam of electrons.

2.17 Contains two atoms bonded together which act as a unit.

2.18 Emits gamma rays.

2.19 Particles of an element having the same atomic number.

Terms:

(a) alpha particle
(b) anion
(c) atomic number
(d) atoms
(e) cation
(f) cathode ray
(g) diatomic molecule
(h) electron
(i) empirical formula
(j) ion
(k) isotopes
(l) mass number
(m) molecule
(n) molecular formula
(o) neutron
(p) nucleus
(q) polyatomic
(r) radioactivity
(s) structural formula

AP Chem

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 DETERMINATION OF 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.

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 EXERCISES

CHEMICAL EQUATIONS: BALANCING AND PREDICTING PRODUCTS OF REACTIONS

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

- 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:

Know combustion Reactions!

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

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:

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

(Calcium oxide) (Calcium carbonate)

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

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

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, $5O_2$, are needed on the reactant side to equal the 10 oxygen atoms on the product side. The balanced chemical equation is

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

(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 4HF. 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 $2H_2O$ on the reactant side. The balancing of the hydrogen atoms also causes the oxygen atoms to be balanced. The balanced equation is

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

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.

Know Combustion Reactions!! **
> one reactant is an organic molecule
> one reactant must be oxygen
> products are Always CO_2 and H_2O

EXERCISE 2 Writing and balancing a combustion reaction

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

SOLUTION: *Analyze:* We are given butane, C_4H_{10}, and are asked to write the balance 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 as a reactant and recalling that carbon dioxide and water are the typical products when a hydrocarbon is combusted. 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 ten hydrogen atoms in butane, therefore, place a 5 in front of H_2O:

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

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:

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

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:

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

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

Predict the products!

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) $3CO_2(g) + 4K(s) \rightarrow 2K_2CO_3(s) + C(s)$: single replacement, form a carbonate
(2) $CaH_2(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + 2H_2(g)$: combine H_2O, form hydroxide, liberate H_2 gas
(3) $2SbCl_3(s) + 3H_2S(g) \rightarrow Sb_2S_3(s) + 6HCl(g)$: double replacement
(4) $2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$: single replacement by anions, liberate SO_2

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

Familiarize yourself with what is likely to happen.

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

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

$\frac{\text{part}}{\text{whole}} \times 100 = \%$ comp

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%.
- $\%$ 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**.
- 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.
- 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

diatomic < 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)

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

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

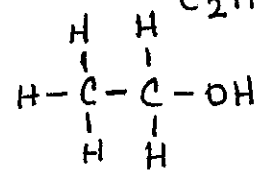
$$\text{O: (2 atoms O)}\left(\frac{16.00\text{ amu}}{1\text{ atom O}}\right) = \underline{32.00\text{ amu}}$$

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

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

Ethanol : C_2H_6O
or C_2H_5OH



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: (GFM)

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

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

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

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

Use your molar mass to make Conversions

Dimensional Analysis!

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: We should recall that the mass of one mole of a substance is its molecular weight expressed in grams.

* moles → grams

$$\#\text{moles}\left|\frac{\text{grams}}{1\text{ mole}}\right. = \#\text{moles}$$

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:

* grams → moles

$$\text{grams}\left|\frac{1\text{ mole}}{\text{grams}}\right. = \#\text{grams}$$

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

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.

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

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:

molecules =

grams	1 mole	# molecules
	gms	1 mole

↑ start with ↑ molar mass ↑ Avogadro's #

(use dimensional analysis & Conversion factors)

$$46.08 \text{ g ethanol} = 1 \text{ mol ethanol} \qquad \text{[Molar Mass]}$$

$$1 \text{ mol ethanol} = 6.022 \times 10^{23} \text{ molecules ethanol} \qquad \text{[Avogadro's number]}$$

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.

(They make this harder to show you the specifics)

C: 12 amu

in ethanol (C_2H_5OH)

$$\frac{2\times 12}{(2\times 12)+(6\times 1)+(16)} = \frac{24}{46} = .5217$$

$$.523 \times 100 = 52\%$$

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:

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

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.

$Na_2B_4O_7$: "Borax" - sodium borate (detergent)
boron + oxygen → borate

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:

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

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

Start
what do we have? .3478 g Na

Grams Na ⟶ *moles* Na ⟶ *moles* $Na_2B_4O_7$ ⟶ *grams* $Na_2B_4O_7$

END: What unit do I need to end up with?

The required equivalences are

1 mol $Na_2B_4O_7$ = 2 mol Na [from formula of $Na_2B_4O_7$]*

1 mol $Na_2B_4O_7$ = 201.24 g $Na_2B_4O_7$ [from molar mass of $Na_2B_4O_7$]

1 mol Na = 23.00 g Na [from molar mass of Na]

Gm Na	1 mol Na	mol $Na_2B_4O_7$	gm $Na_2B_4O_7$
	g Na	mol Na	mol $Na_2B_4O_7$
	molar mass Na	molar ratio	GFM

(START) molar mass

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

molar ratio GFM

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

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

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.

✻ ✻ DETERMINATION OF EMPIRICAL AND MOLECULAR FORMULAS

(A) find empirical formula

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 experimental 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 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.

1. Assume 100 gm sample (% → gm)
2. Determine moles element using molar mass
3. Divide by smallest # moles
4. Get to Whole numbers by multiplying

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

Find Molecular Formula

B) 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?

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.

① Assume a 100 gram sample

∴ 53% = 53.0gm Al

47% = 47.0gm O

Solve:

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

② Determine moles of each element

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

$$\frac{53\text{gm Al}}{} \Big| \frac{1\text{mol Al}}{27\text{g Al}} = 1.96 \text{ mol}$$

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

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

$$\frac{47\text{gm O}}{} \Big| \frac{1\text{mol O}}{16 \text{ g O}} = 2.94\text{mol}$$

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

③ Divide by smallest # moles

Al : 1.96 ÷ 1.96 = 1

O : 2.94 ÷ 1.96 = 1.5

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:

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

(4) Multiply subscripts to make whole numbers
Al : 1×2 = 2
O : 1.5×2 = 3
Al_2O_3

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 provides quantitative mass-mole relationships among the reactants and products. The branch of chemistry that deals with these quantitative relationships is termed **stoichiometry**. When we talk about the "stoichiometric relationships in a chemical reaction," we are referring to the mass-mole relationships given by the balanced chemical equation that describes the reaction of interest.

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

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

Understand that coefficients give us the molar ratios of reactants + products in a chemical rxn.

- 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 in all cases is equivalent to a mole:

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

- This last interpretation of the balanced chemical equation is the one that enables us to derive mass-mole relationships among reactants and products. The numerical coefficients in front of the reactants and products give the ratios of moles in which the chemical substances react. For example, since 2 mol of O_2 is required to react with 1 mol of CH_4, then 2 × 2, or 4 moles of O_2, are required to react with 2 × 1, or 2 moles of CH_4—that is, O_2 always reacts with CH_4 in a 2:1 mole ratio. You can represent this stoichiometrically equivalent ratio 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 stoichiometrically equivalent ratio for the formation or products as

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

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

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

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

- 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 percent 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

They give you .5mol Al to start with and you need to end up with moles of Al_2O_3

$\boxed{\text{moles} \rightarrow \text{moles}}$

(use molar ratio)

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

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

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:

$$\textit{Moles}\ \text{Al} \rightarrow \textit{moles}\ Al_2O_3$$

$$.5\,\text{mol Al} \times \frac{2\,\text{mol}\ Al_2O_3}{4\,\text{mol Al}} = 0.25\ \text{mol}\ Al_2O_3$$

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

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

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:

Need to use molar ratio AND molar mass conversion.

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

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

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

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

molar ratio

molar mass

dimensional analysis

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

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

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:

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

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:

note that since we are using Cl_2, you need the molar mass of Chlorine as diatomic

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

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

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

molar mass Cl_2

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

molar ratio

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

molar mass (GFM) 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

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

(b) Calculate the percent yield of the reaction

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

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

10g C_2H_6 (a) $2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$
50g O_2 (Find limiting reagent)

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

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

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

$$2\ \text{mol}\ C_2H_6 \approx 7\ \text{mol}\ O_2$$
$$30.08\ \text{g}\ C_2H_6 = 1\ \text{mol}\ C_2H_6$$
$$32.00\ \text{g}\ O_2 = 1\ \text{mol}\ O_2$$

* Calculate grams O_2 needed:

$$\frac{10g\,C_2H_6}{} \left| \frac{1mol\,C_2H_6}{30g\,C_2H_6} \right| \frac{7mol\,O_2}{2mol\,C_2H_6} \left| \frac{32g\,O_2}{1mol\,O_2} \right. = \frac{2240}{60} = 37.3 \text{ grams}$$

O_2 needed

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:

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

grams → molar mass → Molar Ratio → molar mass → grams

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

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

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

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

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

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

$Al(OH)_3$ 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

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

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 HNO_3.

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

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