AP CHEMISTRY SUMMER ASSIGNMENT

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Students,

Frist, you have taken on a class that is worthwhile and beneficial but is also difficult and hard. You will spend many nights studying, reviewing, practicing, calculating, and memorizing. The reward on the other side of these long, arduous nights is a passing score on the AP Chemistry exam and an A in my class. The benefits don't stop there! You will be more prepared for college having already taken a rigorous course!

Second, you will have one morning a week that will be dedicated to a lab day. This day will stay constant throughout the year unless a schedule change occurs at which time, I will give you ample time to change around your schedule. These sessions will start at 7 a.m. Some weeks will have no labs. Some weeks will start later than 7 a.m. A class calendar will be handed out at the beginning of the year with lab dates and times.

Third, as we get to March and April, we will begin after school review sessions/practice tests. The lab days will stop and these will begin. You are required to attend one a week unless it is an excused absence.

Fourth, you have a summer assignment to review some general chemistry learned in your sophomore year chemistry class. Science is ever changing, so more and more material is added onto the exam. For us to not fall behind, we must finish the first three chapters in the first two weeks of school. I have given you access to a pdf of the first three chapters of the book and a pdf that is a workbook. (I would print off the workbook – but you don't have to) As you are reading through each chapter, fill out the workbook.

Fifth, watch these videos from this Bozeman website. This will assist you with completing some of the workbook problems. http://www.bozemanscience.com/chemistry

Sixth, you must email me at least **THREE SEPARATE TIMES** throughout the summer asking me questions on the material – preferably questions you don't know the answer to. This shows me that you can communicate well, you are working on the assignment, and you're ready for the rigor of AP. I would rather you work on this assignment 20 minutes a day until you finish, but I understand that we have summer vacations that get in the way. Do not wait until the last few days of summer to start this assignment.

Finally, AP Chemistry becomes a family throughout this whole process. So really decide if you're willing and ready to be part of this awesome AP Chem family.

Can't wait to receive your emails,

Have a blessed summer,

Mrs. Geradine

CHAPTER 1

Chemistry: The Study of Change

As you work through your Focus Review Guide, keep this chapter's Big Ideas in mind:

AP BIG IDEAS A LOOK AHEAD

1. The chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangement of atoms. These atoms retain their identity in chemical reactions.
2. Chemical and physical properties of materials can be explained by the structure and the arrangement of atoms, ions, or molecules and the forces between them.
3. Changes in matter involve the rearrangement and/or reorganization of atoms and/or the transfer of electrons.
4. Rates of chemical reactions are determined by details of the molecular collisions.
5. The laws of thermodynamics describe the essential role of energy and explain and predict the direction of changes in matter.
6. Any bond or intermolecular attraction that can be formed can be broken. These two processes are in a dynamic competition, sensitive to initial conditions and external perturbations.

Chapter Overview

Chemistry is the study of matter, with special focus on the changes that matter can undergo. The study of chemistry begins with observation of and categorization of change, and extends through classification of matter by type and/or phase. The Advanced Placement (AP) Chemistry course organizes the study of chemistry around six Big Ideas. These Big Ideas are introduced in Chapter 1. Chemistry is not just a collection of already known facts, but an evolving field of scientific inquiry. The AP science practices identify the quantitative analysis skills needed to study chemistry. Special emphasis is placed on the ability to represent change in the appropriate manner using SI units, chemical equations, and scientific notation.

1.1 Chemistry: A Science for the Twenty-First Century

Prerequisite Knowledge

Recall It

Chemistry is a very old science, yet chemistry has never lost relevance. Chemistry is referred to as 'the central science' because every other science can trace its lineage to chemistry. Chemistry remains an active discipline, with chemists continuing to advance our knowledge about how chemistry impacts our everyday lives.

Review It

On the next page, you will find a chart that has three branches of science studied in high school. Under each, place an item, event or phenomena that relates to the subject AND to chemistry in the twenty-first century. An example has been placed in each column to get you started.

Environmental Science	Physics	Biology
-pH testing of water samples to determine acidity levels	-semiconductors for microchips in computers	-genetic studies that rely on the DNA sequence

1.2 Introducing the AP Big Ideas

Recall It

The AP Chemistry curriculum is organized around six central ideas, called the Big Ideas. Each of the Big Ideas is subdivided into important concepts called Essential Knowledge (EK). Remember that each Big Idea will be interwoven throughout the textbook, and you are likely to encounter multiple Big Ideas in any given chapter.

Review It

In your own words, summarize each of the AP Chemistry Big Ideas.

Big Idea 1	Chemical elements are the fundamental building blocks of matter, and these building blocks have certain unchanging properties.
Big Idea 2	
Big Idea 3	
Big Idea 4	
Big Idea 5	
Big Idea 6	

1.2 Introducing the AP Big Ideas *continued*

Use It

Identify each item in the list below as a compound or an element:

water

iron

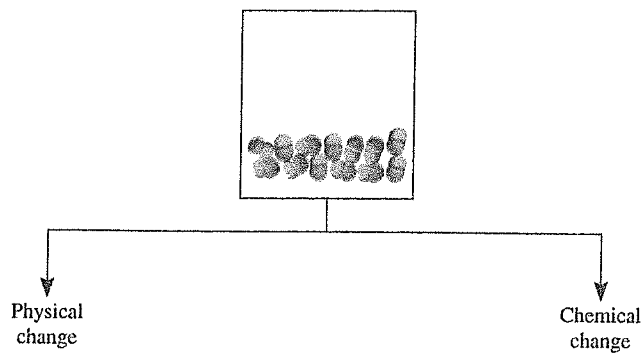
solid carbon in the form of a diamond

$CaCl_2$

gasoline

aluminum foil

Below, you will find a diagram of a compound made up of two elements, each represented using spheres of different colors. Below each, create a diagram to show each of the indicated changes:



Beside each statement, place the term solid, liquid or gas that correctly applies to the statement.

Molecules are tightly held together with little room for motion.

Molecules are close together, but can move past one another with little restriction.

Molecules are far apart and free to move in all directions.

1.3 Introducing the AP Science Practices

Recall It

Chemistry is an active and evolving scientific field. Chemists, like all scientists, possess a number of skills that allow them to propose and investigate scientific questions, make predictions, analyze data, and justify conclusions. The AP Chemistry exam will expect all AP Chemistry students to be proficient in these science practices.

1.3 Introducing the AP Science Practices *continued*

Review It

For each of the seven science practices, identify how you might use the practice in your study of chemistry:

Science Practice	Example Use
SP 1: Using Models	A chemical modeling kit to understand the structure of molecules
SP 2: Using Math to Solve Problems	
SP 3: Engaging in Scientific Questioning	
SP 4: Collecting Data	
SP 5: Analyzing Data	
SP 6: Articulating Claims	
SP 7: Connecting the Pieces	

1.4 Review: The Scientific Method

Prerequisite Knowledge

Recall It

The scientific method, which you should be familiar with from previous science classes, is the systematic manner by which research is done. You will notice that many of the steps of the scientific method correspond to the skills identified in the AP science practices. Research (and experiments) can be broadly classified into two main classes; quantitative research, which strives to generate numerical data as the product, and qualitative research, which strives to generate observable change as the product. Research begins with a hypothesis, which upon testing can become theory, which upon further testing might become law.

Review: The Scientific Method *continued*

Review It

Label the following observations as either qualitative or quantitative

The metal bar was 15 cm long.

The temperature of the metal bar was 21°C.

The surface of the metal was shiny.

The metal bar heated quickly.

It took 17 minutes for the bar to cool from 35°C to 21°C.

Describe the difference between a *hypothesis*, a *theory*, and a *law*.

Review: Measurements

Prerequisite Knowledge

Recall It

Measurement in science is done using the revised metric system known as the SI system. Each quantity has a base unit associated with it; the meter for length, the kilogram for mass, etc. There are three different yet interconvertible temperature scales. Fahrenheit, which you are likely familiar with from the daily weather report, is used rarely in science.

Review It

A student is given a sample of one of seven metals and asked to identify it based on the density of the material. A copy of the data table collected by the student is given below.

Mass of metal: 32.06g
Initial volume of the graduated cylinder: 50.40mL
Final volume of the graduated cylinder: 54.85

What is the density of the unknown metal?

Substance	Density (g/cm^3)
Aluminum	2.70
Iron	7.9
Lead	11.3
Gold	19.3
Osmium	22.6
Chromium	7.15
Tin	7.27

The student is provided a table with the densities of seven possible metals which could comprise the unknown sample. Based on the calculated density of the unknown sample, can the student identify the metal? Justify your answer.

1.5 Review: Measurements *continued*

A sample of osmium is found to have a volume of 15.95 cm^3. What would be the mass of this sample? Use the density of osmium from the table above.

1.6 Review: Handling Numbers

Prerequisite Knowledge

Recall It

Data handling is of extreme importance. Depending on the circumstance, a chemist may be handling very large or very small numbers. This becomes ungainly without the use of scientific notation. The number of digits that are included when data is recorded or manipulated is governed by the rules of significant figures, which involves specific rules for addition and subtraction, multiplication and division. Also important is the distinction between accuracy and precision.

Review It

A dart player throws three darts and all three dots land in the bull's eye. However, the dart player was aiming for the triple 20 with all three darts. Comment on the precision and accuracy of the player.

Use It

The table below shows the volumetric tolerances for a variety of chemical glassware.

Tolerances for Volumetric Ware			
Maximum Allowed Error, mL			
Volume, mL	**Volumetric Flask**	**Volumetric Pipet**	**Burette**
5	-	0.01	-
10	-	0.02	-
25	0.03	0.03	0.03
50	0.05	0.05	0.05
100	0.08	0.08	0.10
250	0.10	-	-
500	0.15	-	-
1000	0.30	-	-

*Data from J.S. Fritz and G. H. Schenk. *Quantitative Analytical Chemistry*, 3rd edition. Allyn & Bacon, Boston, 1974. p. 560.

If you filled a 25 mL volumetric pipet exactly to the line, what range of values would the volume exist between?

Which piece of equipment would best be used to measure exactly 25 mL of solution: a volumetric flask or a pipette? Explain your answer.

Perform the indicated calculations, reporting your final answer to the correct number of significant figures.

86.25 g + 14.3 g =

16.4 g/cm^3 × 1.35 cm^3 =

3 bars of gold × 25.3 kg per bar =

1.7 Review: Dimensional Analysis

Prerequisite Knowledge

Recall It

Of the myriad techniques that scientists use to handle or manipulate data, one of the most powerful is dimensional analysis, or factor-label method. This allows the interconversion of units between SI and the English system, or within SI itself. Factor label relies heavily on conversion factors, which can be readily created from any unit of measurement.

1.7 Review: Dimensional Analysis *continued*

Review It

You plan on traveling to Europe this summer and wish to convert some of your US dollars ($) to euro (€)

You are told that 1 US dollar is the equivalent to 0.73 euro.

How many euro would you get out of $450?

How much would you need in dollars to be able to convert have 1000€?

If 8 gold coins weigh 43.5 g, what would 31 gold coins weigh?

Fill in the sequence of unit conversions needed to change from fl oz to m^3.

$\boxed{\text{fl oz}} \rightarrow \boxed{} \rightarrow \boxed{} \rightarrow \boxed{m^3}$

Given that there are 0.033814 fl oz in 1 mL, use your sequence from the question above to convert 828 fl oz to m^3.

Explain how you arrived at the number of significant figures in this answer.

1.8 Review: Problem Solving – Information, Assumptions, and Simplifications

Prerequisite Knowledge

Recall It

It is often useful to estimate the solution to a problem rather than solve it discretely. One must be careful to address any uncertainty or assumption that is made in the solution of a problem, and to consider its impact on the ultimate solution.

Review It

A pile of gravel is in the approximate shape of an inverted cone, with a diameter of 18 m and a height of 34 m. The gravel has a density of 1255 kg/m^3.

The formula for the volume of a cone $V = \frac{1}{3}\pi r^2 h$, what assumptions can you make in your calculation for the volume of the pile of gravel?

Use this approximation to find the volume of the gravel pile.

Using your answer to the above question and the density of the gravel, determine the approximate mass of the gravel pile.

What assumptions were made in this calculation?

AP CHAPTER SUMMARY

Summarize It

These questions were posed in the Chemistry chapter opener (page 2). Answer them using the knowledge you've gained from this chapter.

1. What are the six AP Big Ideas?

2. What are the seven AP Science Practices?

3. How is the study of chemistry relevant to your life?

CHAPTER 2

Atoms, Molecules, and Ions

As you work through your Focus Review Guide, keep this chapter's Big Ideas in mind:

AP BIG IDEAS A LOOK AHEAD

- Atomic theory, which has been built over time by many scientists, describes the structure of the atom.
- The fundamental structure of the atom defines how atoms come together to form higher order structures.

Chapter Overview

The study of modern chemistry is said to have begun with Dalton's atomic theory, which itself was based on the theories of Democritus and others; Dalton's theory suggested that all matter was composed of small particles called *atoms*. These atoms combine in predictable ratios to form compounds, have defined structures, and can be described by their atomic number and/or their mass number. The representation of chemical compounds relies heavily on shorthand notation by way of chemical formulas, which can take on several varieties. Chemical compounds can be further subdivided into ionic and covalent compounds, which are defined by the mode by which the elements bind together. Chemical compounds are given systematic names that can be predicted from the proper chemical formula, and vice versa.

2.1 The Atomic Theory

Recall It

Although many chemists contributed to the ultimate description of the atom, it was John Dalton who is credited with the first cogent proposal of a set of hypotheses that governed the atom's behavior. Dalton's atomic theory is based on four fundamental hypotheses:

1. All elements are composed of atoms, which are very small particles.
2. All atoms of a given element are identical, and the atoms of different elements are themselves different.
3. Atoms combine in predictable ratios.
4. Atoms cannot be created or destroyed in reactions, only rearranged in some manner.

Essential Knowledge covered

1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
1.D.1: As is the case with all scientific models, any model of the atom is subject to refinement and change in response to new experimental results. In that sense, an atomic model is not regarded as an exact description of the atom, but rather a theoretical construct that fits a set of experimental data.

Review It

Nitrogen dioxide gas (NO_2) can react with itself to form a single different compound. Which of the following is the most likely formula for this new compound? Justify your choice.

a. NO
b. N_2O
c. N_2O_4
d. NO_4

Use It

The Greek philosopher Democritus named the small particles that he thought all things were made of 'atomos,' which means 'indivisible' or 'uncuttable.' Would he still choose that name if he were alive today and knew what you know? Why or why not?

2.2 The Structure of the Atom

Recall It

The atom itself is composed of three fundamental subatomic particles; the proton, neutron, and electron. The electron is the fundamental unit of negative charge and is located in the periphery of the atom. The proton is the fundamental unit of positive charge and is located in the nucleus of the atom along with the neutron, which has no charge. All of these particles have known masses.

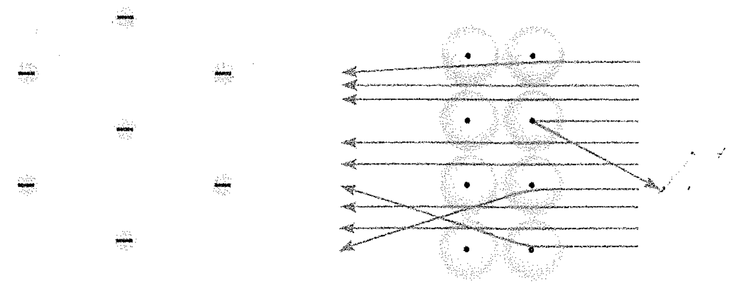
Essential Knowledge covered

1.B.1: The atom is composed of negatively charged electrons, which can leave the atom, and a positively charged nucleus that is made of protons and neutrons. The attraction of the electrons to the nucleus is the basis of the structure of the atom. Coulomb's law is qualitatively useful for understanding the structure of the atom.
1.D.1: As is the case with all scientific models, any model of the atom is subject to refinement and change in response to new experimental results. In that sense, an atomic model is not regarded as an exact description of the atom, but rather a theoretical construct that fits a set of experimental data.

Review It

Complete the following table with respect to the fundamental subatomic particles:

Particle	Mass (g)	Charge (C)	Unit of Charge
Proton			
Electron			
Neutron			



Label the regions of positive and negative charge in both Thomson's 'Plum Pudding' model and Rutherford's Gold Foil Experiment. Now, imagine that both Rutherford and Thomson exist in 2017 and that Rutherford wants to inform Thomson of his discovery using social media. In 140 characters or less, what would Rutherford say?

Use It

Why do scientists refer to Atomic Theory and not Atomic Law?

2.3 Atomic Number, Mass Number, and Isotopes

Recall It

An atom is described by its atomic number, which defines the number of protons and electrons in an uncharged atom. The atomic number can be used in conjunction with the mass number to determine how many neutrons are present. Any two atoms which have the same atomic number but different mass numbers are related as isotopes.

Essential Knowledge covered

1.D.1: As is the case with all scientific models, any model of the atom is subject to refinement and change in response to new experimental results. In that sense, an atomic model is not regarded as an exact description of the atom, but rather a theoretical construct that fits a set of experimental data.
1.D.2: An early model of the atom stated that all atoms of an element are identical. Mass spectrometry data demonstrate evidence that contradicts this early model.

2.3 Atomic Number, Mass Number, and Isotopes *continued*

Review It

On the periodic table, the atomic symbol (either one or two letters) is typically shown with one or two numbers beside it. Define what each number means and how they are calculated.

Most elements have more than one isotope, which is an atom that has the same atomic number but a different mass number. How many protons and neutrons are in the following isotope of He (helium four)?

$^{4}_{2}He$

Use It

Why do all atoms of the same element have the same atomic number but not necessarily the same mass number?

2.4 The Periodic Table

Recall It

The periodic table contains, in a lightly encoded sense, a wealth of information about the elements. In it, the elements are grouped according to the similarity of their physical and chemical properties.

Essential Knowledge covered

1.C.1: Many properties of atoms exhibit periodic trends that are reflective of the periodicity of electronic structure.

Review It

Several groups have characteristic names. Give an example of two elements in each group/family:

Alkali Metals:

Alkaline Earth Metals:

Halogens:

Noble Gases:

Using the periodic table in the front of your textbook, determine what category of elements comprises the majority of the periodic table.

How do metals differ from nonmetals?

Use It

You are building a device that uses, in its construction, a significant amount of magnesium (Mg). You notice that your supply of Mg has run out, but you have aluminum as well as beryllium available. Explain which element would most likely be the best element to use as a substitute, given that you do not want to change the device's physical properties?

2.5 Molecules and Ions

Recall It

Most of the elements exist in nature in a combined form with other elements. The manner by which these compounds form is determined by how the elements come together. Molecules are discrete collections of atoms that unite in a predictable fashion, while ions are atoms or groups of atoms that have characteristic levels of positive or negative charge.

Essential Knowledge covered

1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
1.B.1: The atom is composed of negatively charged electrons, which can leave the atom, and a positively charged nucleus that is made of protons and neutrons. The attraction of the electrons to the nucleus is the basis of the structure of the atom. Coulomb's law is qualitatively useful for understanding the structure of the atom.

Review It

Referring to the periodic table in the front of your textbook, how many hydroxide ions would need to combine with a single cation of the given element to form a neutral compound? List the number necessary for each of the common monatomic ion possible for each element, or "no reaction" if the listed ion would not combine with a hydroxide ion.

Fe:

Na:

Mn:

Sn:

N:

Ag:

F:

In the previous example, what is common between all the ions that will form neutral compounds with a hydroxide ion?

A subscripted number immediately following an atomic symbol indicates the number of that species present in a molecule. Identify the lowest whole–number ratio of hydrogen atoms to oxygen atoms in the following compounds:

H_3PO_4:

H_2O:

H_2O_2:

Use It

Turn to the periodic table in the front of your textbook. It appears that only elements coded in green tend to form cations, whereas elements in blue tend to form anions. What general trend about ionic bonds can you conclude from this?

Using the periodic table in the front of your textbook, identify the trends relating to the preferred cationic or anionic charge and group number. The principle group numbers are followed by the letter A.

2.6 Chemical Formulas

Recall It

Chemical formulas are the shorthand representations used to express the composition of molecules or ions using atomic symbols. They exist in several varieties: molecular formulas, which show the exact number of atoms of each element that are present in the smallest unit of a substance, empirical formulas, which indicate which elements are present in a substance and what the smallest whole-number ratio of their atoms is, but not the exact number of each atom, and ionic formulas, which more closely resemble empirical formulas due to the makeup of an ionic compound.

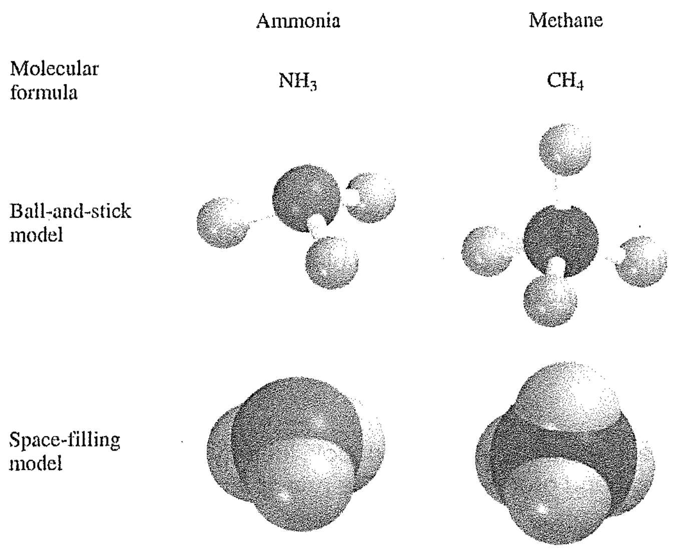
Essential Knowledge covered

1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
2.C.2: Ionic bonding results from the net attraction between oppositely charged ions, closely packed together in a crystal lattice.

2.6 Chemical Formulas *continued*

Review It

How do the empirical formulas for water (H_2O) and hydrogen peroxide (H_2O_2) compare? How do they differ?



Label all the atoms in both ball-and-stick and space filling models above according to the formula given in the molecular formula.

Ionic compounds are, by nature, electrically neutral. What are the formulas for potassium oxide and calcium oxide:

Use It

Draw the ball-and-stick models for ethane, $H_3C—CH_3$. What is ethane's empirical formula?

2.7 Naming Compounds

Prerequisite Knowledge

Recall It

There exist systematic processes for the naming of all compounds in chemistry, whether they are ionic, molecular, organic, binary, ternary, acids, bases, etc. The focus of nomenclature, the subject of naming, must be the ability to determine the class of compounds to which a target compound belongs. Once the compound's classification is determined, one must simply follow the nomenclature rules that exist for that particular class.

Review It

Complete the table, using appropriate rules for the nomenclature of ionic compounds:

Cation	Anion	Formula	Name
			magnesium bicarbonate
		$SrCl_2$	
Fe^{3+}	NO_2^-		
		$SnBr_4$	

Write formulas for the following names. In the last box, indicate the type of compound present.

Name	Formula	Compound Type
Copper(I) cyanide		
Perbromic acid		
Hydrogen sulfide		
Tin(II) fluoride		
Boron trichloride		
Tetraphosphorus decasulfide		

2.8 Introduction to Organic Compounds

Prerequisite Knowledge

Recall It

Organic chemistry is the study of chemical compounds that contain carbon; carbon is unique in its ability to form compounds which consist of carbon atoms bound to one another in a nearly limitless series. The variability of organic compounds is astounding.

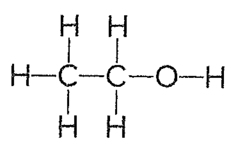
2.8 Introduction to Organic Compounds *continued*

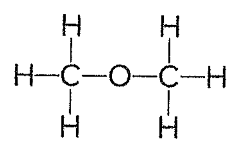
Review It

Organic chemistry assigns names based on the number of carbon atoms present. Complete the chart indicating the names of the alkanes that contain 1–10 carbon atoms.

# Carbons	Name	# Carbons	Name
1	methane	6	
2		7	
3		8	
4		9	
5		10	

Isomers are compound that have the same molecular formulas but different structural formulas. Are the compounds shown below identical or isomers? Explain your answer.





AP CHAPTER SUMMARY

Summarize It

These questions were posed in the Chemistry chapter opener (page 41). Answer them using the knowledge you've gained from this chapter.

1. How does an atom differ from a molecule?

2. What developments have changed our understanding of atoms and molecules since Dalton's proposed atomic theory?

3. What particles comprise an atom?

4. How do elements differ at the atomic level? Are all atoms of a given element identical?

5. What is the basic structure of the periodic table?

6. Can atoms form ions? How do ionic compounds maintain electrical neutrality?

CHAPTER 3

Mass Relationships in Chemical Reactions

As you work through your Focus Review Guide, keep this chapter's Big Ideas in mind:

AP BIG IDEAS A LOOK AHEAD

An understanding of atomic mass, molar mass and percent composition is important for performing calculations involving chemical reactions.

3 Determining the amounts of products that can form when given amounts of reactants are used allows for the calculation of reaction yields.

Chapter Overview

While a single atom cannot be weighed, chemists have many methods to interact with matter on a quantitative basis. The atomic mass is the average mass of all naturally occurring isotopes for a particular element, and the sum of the atomic masses of the elements in a molecule yields the molecular mass. These masses can be used to calculate the number of particles present via Avogadro's Number. Chemists quantify the number of atoms present in a sample using the mole. The mole permits the balancing of chemical equations which, in turn, allows for the determination of percent composition as well as more in depth discussions of stoichiometry.

3.1 Atomic Mass

Recall It

Initially, discussions of mass involve the amu, or the atomic mass unit, which is defined as 1/12th of the mass of one Carbon-12 atom. Due to the propensity of atoms to form more than one stable isotope, however, it is more common to use the average atomic mass, which factors in the mass and percent abundance of each isotope (i.e. a weighted average). This is the number that is reflected on the periodic table.

Essential Knowledge covered

1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.

Review It

A scientist found an element with three naturally occurring isotopes, shown in the chart below. The most abundant is an atom with an equal number of protons and neutrons.

Isotope	Mass (amu)	Abundance (%)
A	22.00	82.4
B	23.01	15.2
C	25.02	2.4

Use this information in the table above to determine the average atomic mass for the element.

3.1 Atomic Mass *continued*

Why is the average atomic mass closer to the atomic mass of isotope A than it is to isotope B or isotope C?

Use It

The element box below is that of boron. It has two naturally occurring isotopes: boron-10 and boron-11.

5
B
10.81

Fill in the table for each isotope of boron:

Isotope	Number of protons	Number of neutrons
Boron-10		
Boron-11		

Boron-10 has an abundance of 19.9%. Verify the value of the average atomic mass.

Explain any difference between your calculated value and the value in the element box.

3.2 Avogadro's Number and the Molar Mass of an Element

Recall It

The mole is one of the most fundamental concepts to master in the study of chemistry. The mole is simply a counting number, like pair or dozen, and reflects the presence of exactly 6.022×10^{23} (called Avogadro's Number) things. Indeed, the mole is the linchpin that allows conversion of the mass of an element to the number of atoms present of that element via the molar mass, which is the mass of one mole of that element. Care should be taken to accurately produce the correct conversion factors when attempting this type of conversion via factor-label method.

3.2 Avogadro's Number and the Molar Mass of an Element *continued*

Essential Knowledge covered

1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.

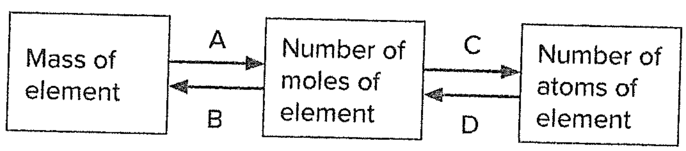
Review It

Fill in the blanks below. Do not use a calculator on this problem.

12.0 g of carbon-12 is 1 mol, so 36.0 g of the material would be ________ mol and 1.20 g would be ________ mol.

Find the periodic table in the front of your textbook. How many moles of atoms are in a 2.50 g sample of magnesium?

Use It



In the flow chart above, there are four arrows labeled A, B, C, and D. Each of these arrows represents a mathematical operation that must be done to move from value to value. What mathematical operation is needed for each numbered step?

A:

B:

C:

D:

3.2 Avogadro's Number and the Molar Mass of an Element *continued*

Use the flow chart on the previous page to determine the number of atoms in a 334.21 g sample of iron if the molar mass of iron is 55.845 g/mol.

3.3 Molecular Mass

Recall It

It is quite straightforward to move from calculations involving only one element to calculations involving a number of elements, such as in a compound. Chemists use molecular mass to describe the sum of the atomic masses in the molecule and formula mass to describe the sum of the atomic masses in one formula unit of an ionic compound.

Essential Knowledge covered

1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.

Review It

Determine the number of each type of atom in one molecule of diphosphorus pentoxide.

For each compound, determine the formula masses (in amu):

Acetylsalicylic acid, $C_9H_8O_4$, the key ingredient in many pain relievers

Aluminum chloride, $AlCl_3$, added to many antiperspirants

3.3 Molecular Mass *continued*

Use It

You are given a sample of 255.30 g of propane (C_3H_8) and are asked to determine the number of moles this sample contains, assuming the molar mass of hydrogen is 1.0079 g/mol and the molar mass of carbon is 12.01 g/mol.

Determine the number of moles in the 255.30 g sample.

3.4 The Mass Spectrometer

Recall It

A mass spectrometer is an instrument that allows for the precise experimental measurement of atomic and/or molecular masses.

Essential Knowledge covered

1.D.2: An early model of the atom stated that all atoms of an element are identical. Mass spectrometry data demonstrate evidence that contradicts this early model.

Review It

An element exists with three naturally occurring isotopes:

Isotope	Mass (amu)	Abundance (%)
A	82.00	78.2
B	85.05	17.8
C	87.10	4.0

3.4 The Mass Spectrometer *continued*

Sketch a mass spectrum for this element.

Use It

Explain how you arrived at your mass spectrum above.

3.5 Percent Composition of Compounds

Recall It

The percent composition (by mass) of a compound can be readily calculated, given both the molar mass of the element and the molar mass of the compound of interest. Additionally, since the calculation rests on having the proper ratio of elements, one must be able to accurately produce the correct chemical formula for the compound.

Essential Knowledge covered

1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
1.A.2: Chemical analysis provides a method for determining the relative number of atoms in a substance, which can be used to identify the substance or determine its purity.
1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.

Review It

Fill in the blanks for the flow chart that can be used to find the empirical formula when given percent composition by mass for each element in a compound.

mass percent			empirical formula

Identify what must be done mathematically to move between steps in the flow chart on p. 25:

Determine the percent composition for each element in the compound for sodium carbonate, Na_2CO_3.

A molecule is made up of 50.05% sulfur and 49.95% oxygen. Use the flow chart on p. 25 to determine the empirical formula of the compound.

Use It

Two large deposits of aluminum were found in two areas where a mining company was looking to open a new aluminum mine. Site A has aluminum in the form of Al_2O_3 while at Site B, the aluminum is on the form of $KAl(SO_4)_2$.

What considerations would the mining company use to make a decision as to where to open a mining operation?

Determine the percent of aluminum in each compound.

3.5 Percent Composition of Compounds *continued*

Calculate the number of grams of aluminum that would be present in 100 kg of each material.

Determine what ratio of mass deposits would be needed for both sites to be considered as sites to open a mining operation.

3.6 Experimental Determination of Empirical Formulas

Recall It

Chemists have the ability to experimentally measure the empirical, and by extension molecular, formula using appropriate instrumentation. The empirical formula, that which contains the subscripts in the smallest whole number ratio, can be converted to the molecular formula as long as molar mass is also known.

Essential Knowledge covered

1.A.2: Chemical analysis provides a method for determining the relative number of atoms in a substance, which can be used to identify the substance or determine its purity.
1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.

Review It

Identify which are possible molecular formulae for the empirical formula CH_2.

C_2H_4 C_2H_6 C_5H_{10} C_3H_8 C_6H_6 CH_4 C_3H_6

Use It

A compound containing only carbon and hydrogen was burned in an apparatus that can be used to detect the amount of carbon dioxide and water produced in the process. When 15.00 g of the material is burned, 50.78 g of carbon dioxide and 10.39 g of water are produced. The molar mass of this compound is found to be approximately 26 g.

Create flow charts to illustrate the process needed to find the molecular formula:

mass of CO_2		
		moles of H

Use this flow chart to find the molecular formula for the compound.

3.7 Chemical Reactions and Chemical Equations

Recall It

The true power of chemistry lies in the ability to convert one material into another by chemical reaction; these are traditionally represented by balanced chemical equations. By convention, reactants are listed on the left and products on the right of an arrow. It is required by the Law of Conservation of Mass that the number of each type of element in the products must be the same as in the reactants, so the overall reaction must be balanced. There are a number of methods by which equations can be balanced.

Essential Knowledge covered

1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
1.E.1: Physical and chemical processes can be depicted symbolically; when this is done, the illustration must conserve all atoms of all types.
1.E.2: Conservation of atoms makes it possible to compute the masses of substances involved in physical and chemical processes. Chemical processes result in the formation of new substances, and the amount of these depends on the number and the types and masses of elements in the reactants, as well as the efficiency of the transformation.
3.A.1: A chemical change may be represented by a molecular, ionic, or net ionic equation.
3.C.1: Production of heat or light, formation of a gas, and formation of a precipitate and/or a color change are possible evidences that a chemical change has occurred.

Review It

Balance the combustion reaction of propane, C_3H_8. In this reaction, propane combines with oxygen gas to produce carbon dioxide and water.

$$\_\_C_3H_8 + \_\_O_2 \rightarrow \_\_CO_2 + \_\_H_2O$$

Fill in the chart for a balance sheet of atoms for the reaction above:

Element	Number on reactant side	Number on product side
Carbon		
Hydrogen		
Oxygen		

Another student balanced the reaction as

$3\ C_3H_8 + 15\ O_2 \rightarrow 9\ CO_2 + 12\ H_2O$

What would you say to this student about their answer? Be specific and correct any issues that need correcting in what they did.

Use It

Balance each chemical reaction given here, and show the balance sheet for the number of each type of atom once each reaction has been balanced.

a. $Fe + Cl_2 \rightarrow FeCl_3$

Element	Number on reactant side	Number on product side
Iron		
Chlorine		

b. $Na + H_2O \rightarrow NaOH + H_2$

Element	Number on reactant side	Number on product side
Sodium		
Hydrogen		
Oxygen		

c. $C_6H_6 + O_2 \rightarrow CO_2 + H_2O$

Element	Number on reactant side	Number on product side
Carbon		
Hydrogen		
Oxygen		

3.8 Amounts of Reactants and Products

Recall It

One of the most critical questions surrounding the practice of chemical experimentation is "how much product will this reaction produce?" The answer lies in the field of stoichiometry, which is the quantitative study of reactants and products in chemical reactions. Typically, stoichiometric calculations are done using moles; indeed, the mole method allows for the interpretation of the coefficients in a balanced equation as simply the number of moles of each component that are present in the reaction.

Essential Knowledge covered

1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.
1.E.1: Physical and chemical processes can be depicted symbolically; when this is done, the illustration must conserve all atoms of all types.
1.E.2: Conservation of atoms makes it possible to compute the masses of substances involved in physical and chemical processes. Chemical processes result in the formation of new substances, and the amount of these depends on the number and the types and masses of elements in the reactants, as well as the efficiency of the transformation.
3.A.1: A chemical change may be represented by a molecular, ionic, or net ionic equation.
3.A.2: Quantitative information can be derived from stoichiometric calculations that utilize the mole ratios from the balanced chemical equations. The role of stoichiometry in real-world applications is important to note, so that it does not seem to be simply an exercise done only by chemists.

Review It

Balance the reaction $P(s) + O_2(g) \rightarrow P_2O_5(s)$

Once balanced, draw a representation of the balanced chemical reaction using solid spheres for the P atoms and open spheres for the O atoms.

Use It

Water decomposes into hydrogen gas and oxygen gas when electricity passes through the solution. If 24.50 g of water is fully decomposed, what mass of oxygen will be collected?

3.9 Limiting Reagents

Recall It

Seldom are reactants in a reaction present in the exact stoichiometric ratio that is required. As a result, one reactant is often used up before the others, at which point the reaction stops. The reactant that is completely consumed first is referred to as the limiting reagent, as it limits how much product can form. The other reactant(s) is/are referred to as being present in excess, and will be present in their unreacted form in some quantity once the reaction has stopped.

Essential Knowledge covered

1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.
1.E.1: Physical and chemical processes can be depicted symbolically; when this is done, the illustration must conserve all atoms of all types.
1.E.2: Conservation of atoms makes it possible to compute the masses of substances involved in physical and chemical processes. Chemical processes result in the formation of new substances, and the amount of these depends on the number and the types and masses of elements in the reactants, as well as the efficiency of the transformation.
3.A.1: A chemical change may be represented by a molecular, ionic, or net ionic equation.
3.A.2: Quantitative information can be derived from stoichiometric calculations that utilize the mole ratios from the balanced chemical equations. The role of stoichiometry in real-world applications is important to note, so that it does not seem to be simply an exercise done only by chemists.

Review It

In the process of assembling tricycles for small children, a worker has 16 frames and 60 tires with which to assemble frame and tire components.

Write a balanced equation for this process.

In the situation on the previous page, which would be considered the limiting reagent and which would be the excess reagent?

In the reaction of $Fe(s) + O_2(g) \rightarrow Fe_2O_3(s)$,

a. What is the exact stoichiometric ratio of iron to oxygen in the reaction (be sure to balance the reaction before answering)?

b. If 2 moles of iron and 2 moles of oxygen gas were to be used, which would be the limiting reagent? Explain your answer.

Use It

Determine the mass of hydrogen gas that will form when 22.53 g of solid nickel reacts with 65.16 g of lactic acid in the reaction:

$Ni(s) + HC_2H_5O_3(aq) \rightarrow Ni(C_2H_5O_3)_2(aq) + H_2(g)$

Show all work and be sure to balance the chemical reaction before you begin the calculation. Note that the molar mass of lactic acid is 90.08 g/mol, and the molar mass for nickel is 58.69 g/mol.

3.10 Reaction Yield

Recall It

Rarely do chemical reactions follow the theoretical balanced equation exactly. For this reason chemists speak of theoretical yields and actual yields, and these are united in percent yield.

Essential Knowledge covered

1.A.1: Molecules are composed of specific combinations of atoms; different molecules are composed of combinations of different elements and of combinations of the same elements in differing amounts and proportions.
1.A.3: The mole is the fundamental unit for counting numbers of particles on the macroscopic level and allows quantitative connections to be drawn between laboratory experiments, which occur at the macroscopic level, and chemical processes, which occur at the atomic level.
1.E.1: Physical and chemical processes can be depicted symbolically; when this is done, the illustration must conserve all atoms of all types.
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3.A.1: A chemical change may be represented by a molecular, ionic, or net ionic equation.
3.A.2: Quantitative information can be derived from stoichiometric calculations that utilize the mole ratios from the balanced chemical equations. The role of stoichiometry in real-world applications is important to note, so that it does not seem to be simply an exercise done only by chemists.

Review It

Silicon carbide, SiC(s) can be manufactured by reacting silicon dioxide with solid carbon at high temperatures. The second product in this reaction is carbon dioxide gas. The reaction can be written as

$SiO_2(s) + 3\ C(s) \rightarrow SiC(s) + 2\ CO_2(g)$

If 20.00 g of silicon dioxide is reacted with 14.02 g of carbon, 12.05 g of silicon carbide forms.

Determine which reactant is the limiting reagent in this process.

How many moles of silicon carbide will form in this process?

What is the theoretical yield in this situation?

Determine the reaction yield.

3.10 Reaction Yield *continued*

Use It

A classmate said that they read an article stating that the percent yield of a new chemical reaction is more than 100%. What would you say to this classmate in response to this claim?

AP CHAPTER SUMMARY

Summarize It

These questions were posed in the Chemistry chapter opener (page 79). Answer them using the knowledge you've gained from this chapter.

1. How do we determine atomic mass? What is a mole? How do we determine molar mass?

2. How does mass spectroscopy demonstrate the existence of isotopes?

3. What is the percent elemental composition of a substance? How do we experimentally determine composition?

4. What is conservation of mass?

5. How do we correctly represent a chemical change?

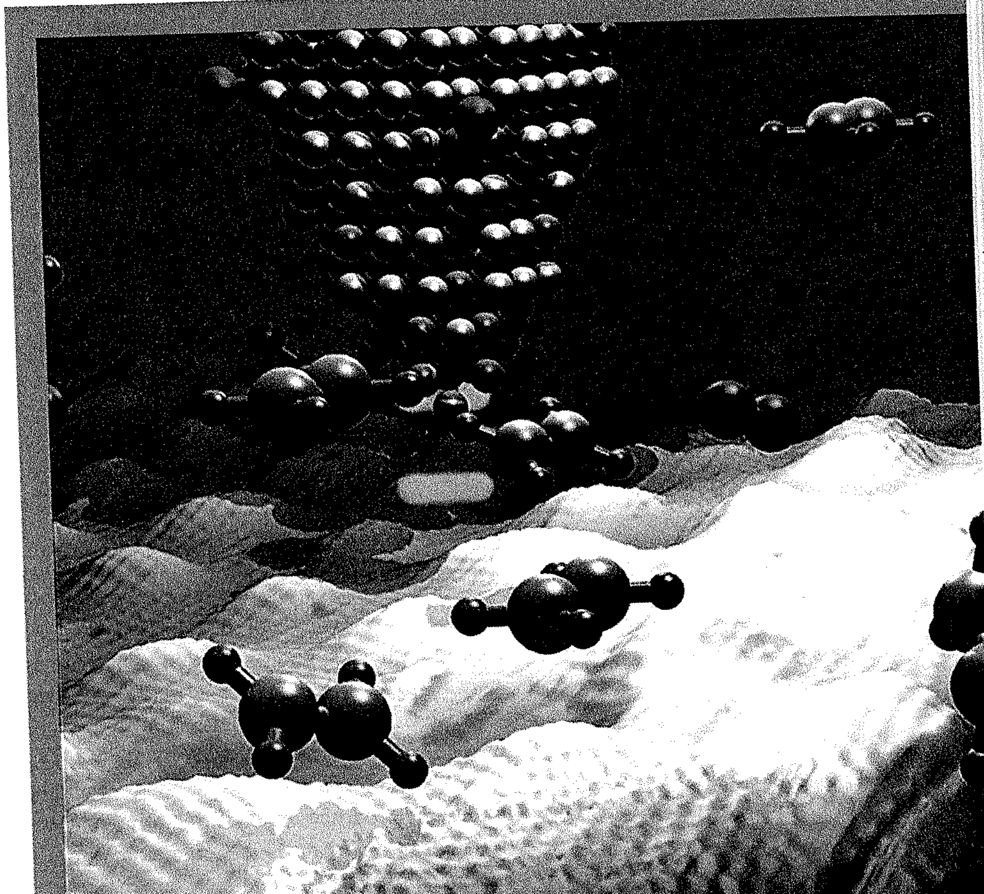
6. How do we write a properly balanced chemical reaction?

7. What evidence is used to indicate a chemical change has taken place?

8. How do we determine the amount of product formed in a chemical reaction or the amount of reactant required? How do we determine the limiting reagent? How do we determine percent yield?

CHAPTER 1

Chemistry The Study of Change



A scanning tunneling microscope probes individual small molecules when they adsorb on graphene, a single-atom thin sheet of carbon atoms.

AP CHAPTER OUTLINE

AP BIG IDEAS A LOOK AHEAD

1 2 3 4 5 6

Chemistry surrounds us. It determines the myriad of interactions needed for our bodies to function. Its laws determine the function of the food we eat and the water we drink. It is in our daily routines. Consider the car or bus ride to school. As a result of chemical interactions, a vehicle starts when the ignition is turned on and accelerates when the gas pedal is depressed. A mini explosion occurs within each cylinder and that energy is transferred to turn the wheels of the car. The tires grip the road with a prescribed air pressure within. Exhaust fumes are cleaned up by the catalytic converter. Halogen headlights, an interaction of matter and energy, show us the road in the early morning hours. The car or bus is a traveling road show of chemistry! The *challenge* of chemistry is to connect each of these visible events with the invisible particles that cause them to happen.

In Advanced Placement (AP) Chemistry, the fundamental concepts in chemistry are arranged into six Big Ideas. The study of chemistry requires not only an understanding of these Big Ideas, but also the ability to apply the seven AP Science Practices. Using the Big Ideas and Science Practices as a guide, you will better understand how your world works—from headlights to tail lights and everything in between.

As you read the chapter, consider these Essential Questions:

1. What are the six AP Big Ideas?
2. What are the seven AP Science Practices?
3. How is the study of chemistry relevant to your life?

Chemistry is an active, evolving science that has vital importance to our world, in both the realm of nature and the realm of society. Its roots are ancient, but as we will see, chemistry is every bit a modern science.

We will begin our study of AP Chemistry by introducing the AP Chemistry curriculum framework, which lays out six fundamental ideals (called Big Ideas) around which the study of chemistry is organized. In addition, we will introduce the AP Science Practices, principles of investigation relevant to all sciences. The latter sections of this chapter will serve as a review of material you have likely covered in previous courses, but which is essential for your success in AP Chemistry: the scientific method, handling numerical results and measurements, and how to solve numerical problems. In Chapter 2, we will begin to explore the microscopic world of atoms and molecules.

1.1 Chemistry: A Science for the Twenty-First Century

The Chinese characters for chemistry mean "the study of change."

Chemistry is the study of matter and the changes it undergoes. Chemistry is often called the central science, because a basic knowledge of chemistry is essential for students of biology, physics, geology, ecology, and many other subjects. Indeed, it is central to our way of life; without it, we would be living shorter lives in what we would consider primitive conditions, without automobiles, electricity, computers, CDs, and many other everyday conveniences.

Although chemistry is an ancient science, its modern foundation was laid in the nineteenth century, when intellectual and technological advances enabled scientists to break down substances into ever smaller components and consequently to explain many of their physical and chemical characteristics. The rapid development of increasingly sophisticated technology throughout the twentieth century has given us even greater means to study things that cannot be seen with the naked eye. Using computers and special microscopes, for example, chemists can analyze the structure of atoms and molecules—the fundamental units on which the study of chemistry is based—and design new substances with specific properties, such as drugs and environmentally friendly consumer products.

Compared with other subjects, chemistry is commonly believed to be more difficult, at least at the introductory level. There is some justification for this perception; for one thing, chemistry has a very specialized vocabulary. However, even if this is your first course in chemistry, you already have more familiarity with the subject than you may realize. In everyday conversations we hear words that have a chemical connection, although they may not be used in the scientifically correct sense.

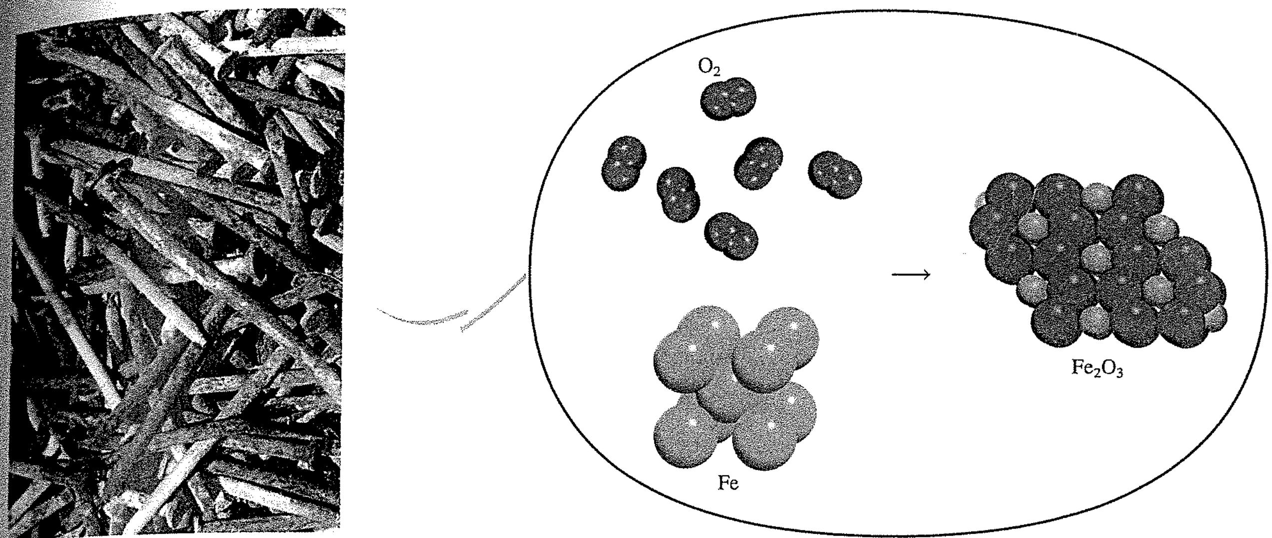


Figure 1.1 *A simplified molecular view of rust (Fe_2O_3) formation from iron (Fe) atoms and oxygen molecules (O_2). In reality, the process requires water and rust also contains water molecules.*
©B.A.E. Inc./Alamy Stock Photo

Examples are "electronic," "quantum leap," "equilibrium," "catalyst," "chain reaction," and "critical mass." Moreover, if you cook, then you are a practicing chemist! From experience gained in the kitchen, you know that oil and water do not mix and that boiling water left on the stove will evaporate. You apply chemical and physical principles when you use baking soda to leaven bread, choose a pressure cooker to shorten the time it takes to prepare soup, add meat tenderizer to a pot roast, squeeze lemon juice over sliced pears to prevent them from turning brown or over fish to minimize its odor, and add vinegar to the water in which you are going to poach eggs. Every day we observe such changes without thinking about their chemical nature. The purpose of this course is to make you think like a chemist, to look at the *macroscopic world*—the things we can see, touch, and measure directly—and visualize the particles and events of the *microscopic world* that we cannot experience without modern technology and our imaginations.

At first some students find it confusing that their chemistry instructor and textbook seem to be continually shifting back and forth between the macroscopic and microscopic worlds. Just keep in mind that the data for chemical investigations most often come from observations of large-scale phenomena, but the explanations frequently lie in the unseen and partially imagined microscopic world of atoms and molecules. In other words, chemists often see one thing (in the macroscopic world) and think another (in the microscopic world). Looking at the rusted nails in Figure 1.1, for example, a chemist might think about the basic properties of individual atoms of iron and how these units interact with other atoms and molecules to produce the observed change.

1.2 Introducing the AP Big Ideas

The Big Ideas are the central framework for AP Chemistry and provide a roadmap to understanding chemistry. They are supported by seven science practices—applications of practices, techniques and inquiry used by scientists of all disciplines. The nature of chemistry means that each Big Idea is interwoven throughout the textbook and students are likely to encounter multiple Big Ideas in a chapter.

Each of the Big Ideas is broken down in Essential Knowledge (EK) and Learning Objectives (LO). The learning objectives are the marriage between the content (the EK) and the science practices (SP) that all science students should be proficient with. Every question on the AP® Chemistry exam will be associated with a LO. The seven science practices will be discussed in more detail in section 1.3.

Big Idea 1: Chemical Elements and Matter

Big Idea 1 states that the chemical elements are fundamental building materials of matter, and all matter can be understood in terms of arrangements of atoms. These atoms retain their identity in chemical reactions.

In the third James Bond movie, *Goldfinger*, the title character was named Auric Goldfinger (hiding the Latin name for gold within his first name). In the final climatic fight scene at Fort Knox between James Bond and Oddjob, the trusted manservant of Goldfinger, James Bond throws a gold bullion bar at Oddjob, who is standing about ten feet away. Oddjob is hit squarely in the chest with the gold bar. As this is Hollywood, the gold bar bounces harmlessly to the floor and the fight continues. Gold, however, is one of the densest elements to be found on this planet. A gold bullion bar stored at Fort Knox would be 400 troy ounces (12.4 kilograms or 27.4 pounds)! If James Bond could pick it up with one hand and throw it, Oddjob certainly wouldn't be able to continue the fight.

We defined chemistry in Section 1.1 as the study of matter and the changes it undergoes. ***Matter*** is *anything that occupies space and has mass.* Matter includes things we can see and touch (such as water, earth, and trees), as well as things we cannot (such as air). Thus, everything in the universe has a "chemical" connection.

Chemists distinguish among several subcategories of matter based on composition and properties. The classifications of matter include substances, mixtures, elements, and compounds, as well as atoms and molecules, which we will consider in Chapter 2.

Substances and Mixtures

A ***substance*** is *a form of matter that has a definite (constant) composition and distinct properties.* Examples are water, ammonia, table sugar (sucrose), gold, and oxygen. Substances differ from one another in composition and can be identified by their appearance, smell, taste, and other properties.

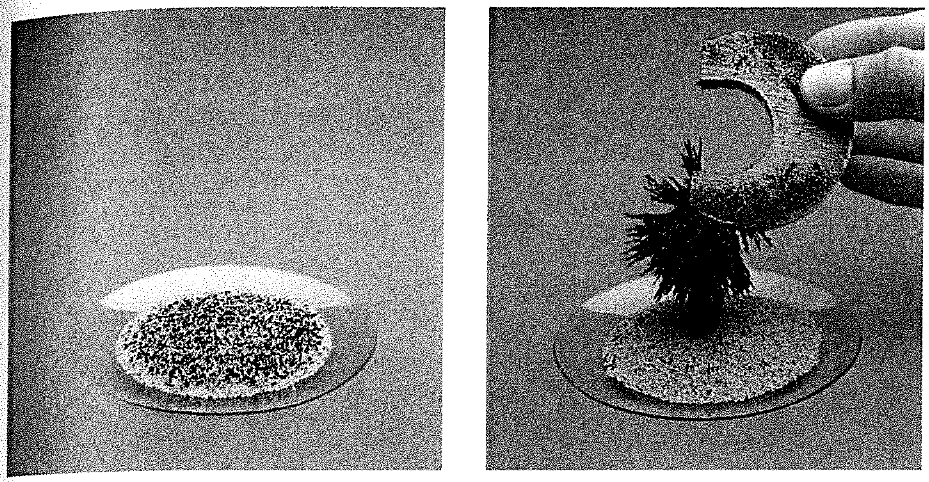
A ***mixture*** is *a combination of two or more substances in which the substances retain their distinct identities.* Some familiar examples are air, soft drinks, milk, and cement. Mixtures do not have constant composition. Therefore, samples of air collected in different cities would probably differ in composition because of differences in altitude, pollution, and so on.

Mixtures are either homogeneous or heterogeneous. When a spoonful of sugar dissolves in water we obtain a ***homogeneous mixture*** in which *the composition of the mixture is the same throughout.* If sand is mixed with iron filings, however, the sand grains and the iron filings remain separate (Figure 1.2). This type of mixture is called a ***heterogeneous mixture*** because *the composition is not uniform.*

Any mixture, whether homogeneous or heterogeneous, can be created and then separated by physical means into pure components without changing the identities of the components. Thus, sugar can be recovered from a water solution by heating the solution and evaporating it to dryness. Condensing the vapor will give us back the water component. To separate the iron-sand mixture, we can use a magnet to remove the iron filings from the sand, because sand is not attracted to the magnet [see Figure 1.2(b)]. After separation, the components of the mixture will have the same composition and properties as they did to start with.

Elements and Compounds

Substances can be either elements or compounds. An ***element*** is *a substance that cannot be separated into simpler substances by chemical means.* To date, 118 elements



(a) (b)

Figure 1.2 *(a) The mixture contains iron filings and sand. (b) A magnet separates the iron filings from the mixture. The same technique is used on a larger scale to separate iron and steel from nonmagnetic objects such as aluminum, glass, and plastics.*

have been positively identified. Most of them occur naturally on Earth. The others have been created by scientists via nuclear processes, which are the subject of Chapter 19 of this text.

For convenience, chemists use symbols of one or two letters to represent the elements. The first letter of a symbol is *always* capitalized, but any following letters are not. For example, Co is the symbol for the element cobalt, whereas CO is the formula for the carbon monoxide molecule. Table 1.1 shows the names and symbols of some of the more common elements; a complete list of the elements and their symbols appears inside the front cover of this book. The symbols of some elements are derived from their Latin names—for example, Au from *aurum* (gold), Fe from *ferrum* (iron), and Na from *natrium* (sodium)—whereas most of them come from their English names. Appendix 1 gives the origin of the names and lists the discoverers of most of the elements.

Atoms of most elements can interact with one another to form compounds. Hydrogen gas, for example, burns in oxygen gas to form water, which has properties that are distinctly different from those of the starting materials. Water is made up of two parts hydrogen and one part oxygen. This composition does not change,

Table 1.1 Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

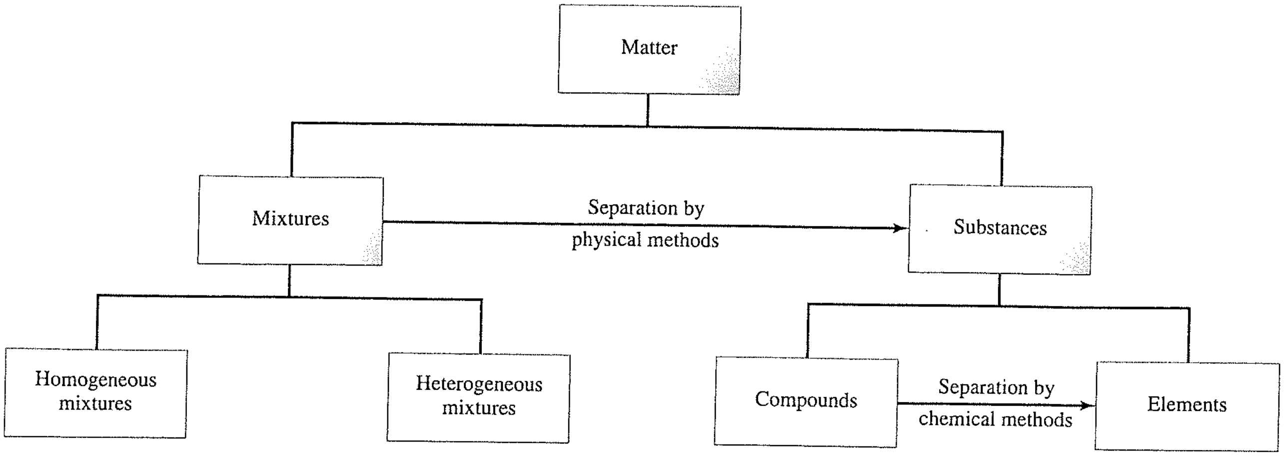


Figure 1.3 *Classification of matter.*

regardless of whether the water comes from a faucet in the United States, a lake in Outer Mongolia, or the ice caps on Mars. Thus, water is a ***compound,*** *a substance composed of atoms of two or more elements chemically united in fixed proportions.* Unlike mixtures, compounds can be separated only by chemical means into their pure components.

The relationships among elements, compounds, and other categories of matter are summarized in Figure 1.3.

Density

Two of the learning objectives for Big Idea 1 refer to using mass data to infer the composition of matter and to make a claim about the identity of a substance; one possible way to approach these learning objectives is using an objects density. The equation for density is

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

or

$$d = \frac{m}{V} \tag{1.1}$$

where d, m, and V denote density, mass, and volume, respectively. Because density is an intensive property and does not depend on the quantity of mass present, for a given substance the ratio of mass to volume always remains the same; in other words, V increases as m does. Density usually decreases with temperature.

The SI-derived unit for density is the kilogram per cubic meter (kg/m^3). This unit is awkwardly large for most chemical applications. Therefore, grams per cubic centimeter (g/cm^3) and its equivalent, grams per milliliter (g/mL), are more commonly used for solid and liquid densities. Because gas densities are often very low, we express them in units of grams per liter (g/L):

$$1\ g/cm^3 = 1\ g/mL = 1000\ kg/m^3$$

$$1\ g/L = 0.001\ g/mL$$

Table 1.2 lists the densities of several substances.

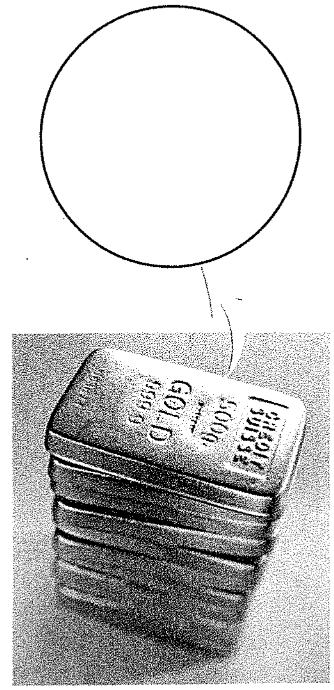
Table 1.2

Densities of Some Substances at 25°C

Substance	Density (g/cm^3)
Air*	0.001
Ethanol	0.79
Water	1.00
Graphite	2.2
Table salt	2.2
Aluminum	2.70
Diamond	3.5
Iron	7.9
Lead	11.3
Mercury	13.6
Gold	19.3
Osmium†	22.6

*Measured at 1 atmosphere
†Osmium (Os) is the densest element known

Examples 1.1 and 1.2 show density calculations.



Gold bars and the solid-state arrangement of the gold atoms.

Example 1.1

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm^3. Calculate the density of gold.

Solution We are given the mass and volume and asked to calculate the density. Therefore, from Equation (1.1), we write

$$\begin{aligned} d &= \frac{m}{V} \\ &= \frac{301 \text{ g}}{15.6 \text{ cm}^3} \\ &= 19.6 \text{ g/cm}^3 \end{aligned}$$

Practice Exercise A piece of platinum metal with a density of 21.5 g/cm^3 has a volume of 4.49 cm^3. What is its mass?

Similar problems: 1.23, 1.24

Example 1.2

The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL. Calculate the mass of 5.50 mL of the liquid.

Solution We are given the density and volume of a liquid and asked to calculate the mass of the liquid. We rearrange Equation (1.1) to give

$$\begin{aligned} m &= d \times V \\ &= 13.6 \frac{\text{g}}{\cancel{\text{mL}}} \times 5.50 \cancel{\text{mL}} \\ &= 74.8 \text{ g} \end{aligned}$$

Practice Exercise The density of sulfuric acid in a certain car battery is 1.41 g/mL. Calculate the mass of 242 mL of the liquid.

Similar problems: 1.23, 1.24

Applying Practices
For more practice determining density, go online to access the lab *Determine Density.*

Big Idea 2: Physical and Chemical Properties of Matter

Big Idea 2 states that the chemical and physical properties of materials are explained by the structure and the arrangement of the atoms that compose the material and the forces between them. This Big Idea usually breaks down into two topics: the arrangement of the atoms and the forces between them.

With the moon two days shy of its monthly rebirth, it was a dark night with a calm sea. The air temperature hovered around the freezing point. From the crow's nest three gongs rang out followed by the frantic message "iceberg right ahead". What really sunk RMS Titanic? Everyone knows that it struck an iceberg, but why did the "unsinkable" ship sink? Many theories have been proposed—faulty rivets, substandard steel used for the hull plates, a coal bunker fire that weaken the steel hull plates, a

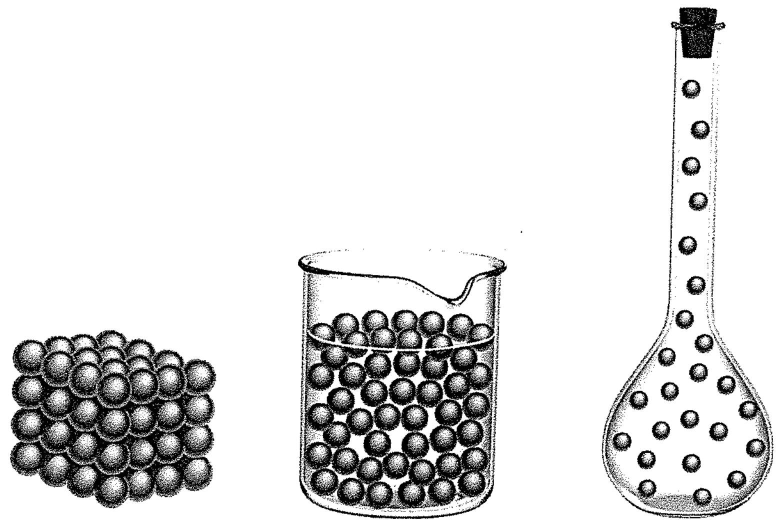


Figure 1.4 *Microscopic views of a solid, a liquid, and a gas.*

"super moon" that caused more icebergs to break off from Greenland and travel south. What is the underlying chemical reason for the sinking—the shape of the water molecule and the intermolecular forces between water molecules that result in a slight increase in volume when water freezes. The slight increase in the volume results in the density of ice to be slightly less than liquid water, hence the iceberg floated in the water and the Titanic was able to collide with it.

All substances, at least in principle, can exist in three states: solid, liquid, and gas. As Figure 1.4 shows, gases differ from liquids and solids in the distances between the molecules. In a solid, molecules are held close together in an orderly fashion with little freedom of motion. Molecules in a liquid are close together but are not held so rigidly in position and can move past one another. In a gas, the molecules are separated by distances that are large compared with the size of the molecules.

The three states of matter can be interconverted without changing the composition of the substance. Upon heating, a solid (for example, ice) will melt to form a liquid (water). (The temperature at which this transition occurs is called the melting point.) Further heating will convert the liquid into a gas. (This conversion takes place at the boiling point of the liquid.) On the other hand, cooling a gas will cause it to condense into a liquid. When the liquid is cooled further, it will freeze into the solid form.

Figure 1.5 shows the three states of water. Note that the properties of water are unique among common substances in that the molecules in the liquid state are more closely packed than those in the solid state. An example of science practice 7 (covered in section 1.3) would be to use this fact to explain why ice cubes float in your soft drink. The answer is explained using the density of liquid water and solid water (ice). Since the particles are less closely packed in the solid phase, the volume is greater, hence the density of ice is less and will float in liquid water.

Applying Practices

When liquid water freezes to ice, is it undergoing a chemical or physical change? Is it always possible to differentiate chemical from physical changes? To answer these questions, go online to access the inquiry activity *Chemical and Physical Change*.

Big Idea 3: Chemical Changes and the Reorganization of Atoms

The sentences of chemistry are balanced chemical equations. Big Idea 3 investigates chemical changes, which are represented by a balanced chemical equation and the driving forces behind chemical changes.

Boeing's new 787 Dreamliner airplanes are the most fuel-efficient passenger aircraft ever produced in the United States. They incorporate composite and lightweight materials, new flight control systems that eliminate hydraulics, and lithium-ion batteries

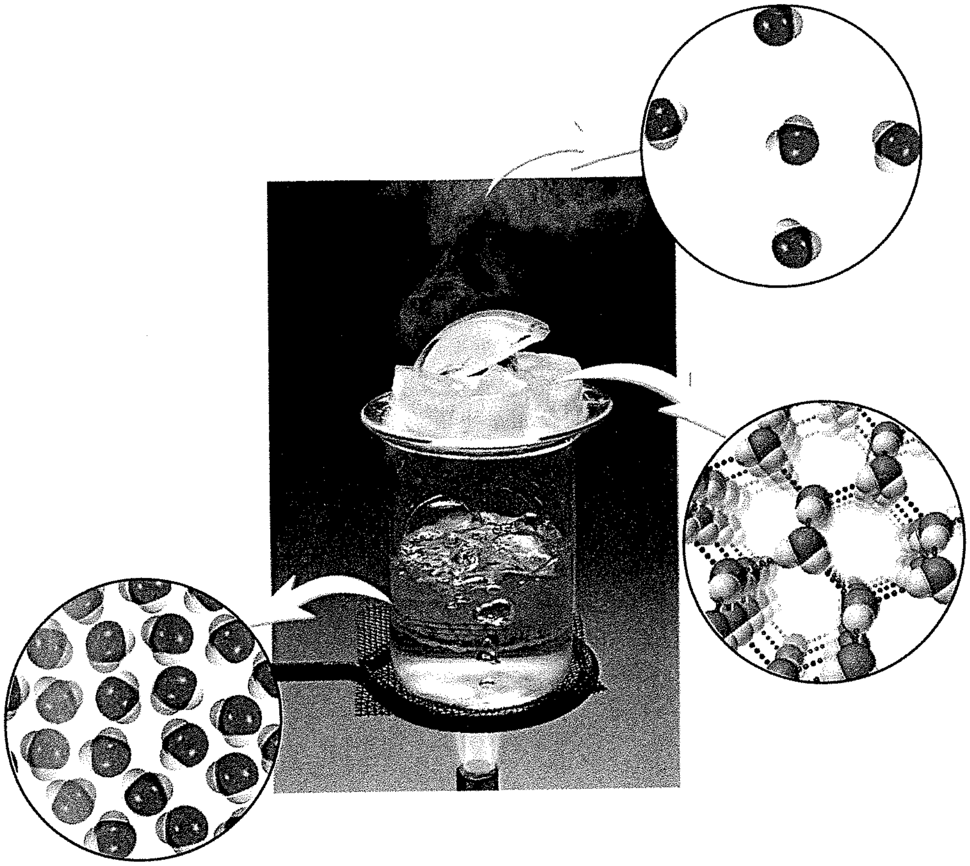


Figure 1.5 *The three states of matter. A flame changes ice into water and steam.*
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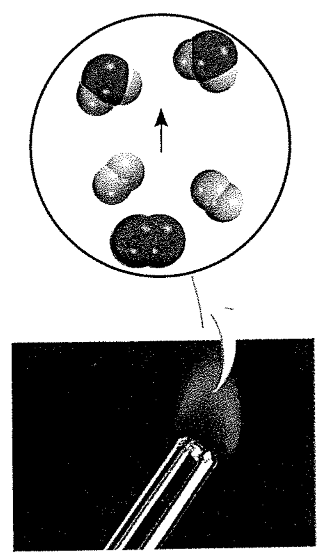
to improve fuel efficiency and range. As of March, 2017, Boeing has received orders totaling 1,211 aircraft with an average price of 265 million dollars!

On January 16, 2013, an in-flight fire occurred in the battery compartment of a Japanese airline's Dreamliner. A week before, an aircraft at Boston Logan International Airport suffered a similar fire on the ground. Fortunately, there were no major injuries with either incident. The FAA issued an emergency directive to ground all Dreamliners until the battery failures could be solved. The investigation and redesign took three months, but chemists and engineers worked together to solve the problem.

All batteries operate on the same principle—the transfer of electrons. This is the driving force behind all batteries (and a major classification of chemical reactions called oxidation-reduction reactions), which encompass a topic called electrochemistry (the electricity that results from a chemical reaction) that you will study.

Substances are identified by their properties as well as by their composition. Color, melting point, and boiling point are physical properties. A physical property can be measured and observed without changing the composition or identity of a substance. For example, we can measure the melting point of ice by heating a block of ice and recording the temperature at which the ice is converted to water. Water differs from ice only in appearance, not in composition, so this is a physical change; we can freeze the water to recover the original ice. Therefore, the melting point of a substance is a physical property. Similarly, when we say that helium gas is lighter than air, we are referring to a physical property.

On the other hand, the statement "Hydrogen gas burns in oxygen gas to form water" describes a chemical property of hydrogen, because to observe this property we must carry out a chemical change, in this case burning. After the change, the original chemical substance, the hydrogen gas, will have vanished, and all that will be left is a different chemical substance—water. We cannot recover the hydrogen from the water by means of a physical change, such as boiling or freezing.



Hydrogen burning in air to form water
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Every time we hard-boil an egg, we bring about a chemical change. When subjected to a temperature of about 100°C, the yolk and the egg white undergo changes that alter not only their physical appearance but their chemical makeup as well. When eaten, the egg is changed again, by substances in our bodies called enzymes. This digestive action is another example of a chemical change. What happens during digestion depends on the chemical properties of both the enzymes and the food.

All measurable properties of matter fall into one of two additional categories: extensive properties and intensive properties. The measured value of an extensive property depends on how much matter is being considered. Mass, which is the quantity of matter in a given sample of a substance, is an extensive property. More matter means more mass. Values of the same extensive property can be added together. For example, two copper pennies will have a combined mass that is the sum of the masses of each penny, and the length of two tennis courts is the sum of the lengths of each tennis court. Volume, defined as length cubed, is another extensive property. The value of an extensive quantity depends on the amount of matter.

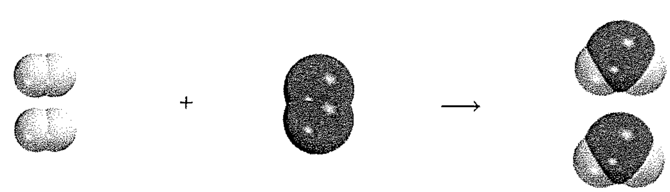
The measured value of an intensive property does not depend on how much matter is being considered. Density is an intensive property. So is temperature. Suppose that we have two beakers of water at the same temperature. If we combine them to make a single quantity of water in a larger beaker, the temperature of the larger quantity of water will be the same as it was in two separate beakers. Unlike mass, length, and volume, temperature and other intensive properties are not additive.

Big Idea 4: Rates of Chemical Reactions

The study of the rate (the change of the amount of a substance in a chemical reaction over time) at which a chemical reaction takes place is called chemical kinetics (or simply, kinetics) and is one of the major subtopics in chemistry. At times, a chemist wants to speed up the rate of a chemical reaction and other times a chemist needs to slow down the rate of a chemical reaction. It is a delicate balancing act that chemist perform.

When a juggler attempts to keep two balls in the air at the same time, the speed that he needs to move is relatively slow. As more and more balls are added, his speed has to increase to keep everything moving so he does not drop any of the balls.

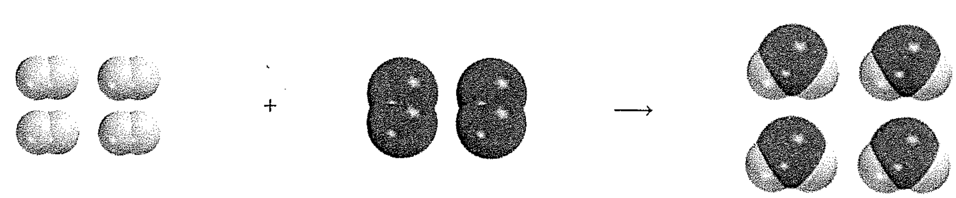
When a chemical reaction, a process in which a substance (or substances) is changed into one or more new substances, takes place, a chemical equation is written. A chemical equation can be written using chemical symbols, or can be illustrated graphical, to show what happens during a chemical reaction. One of the most important reactions for life on planet Earth is the formation of water shown below both graphically and symbolically.



Two hydrogen molecules + One oxygen molecule ⟶ Two water molecules

$$2H_2 + O_2 \longrightarrow 2H_2O$$

As you add more reactants, the rate at which products are formed will increase since there are more particles that are colliding together.



Big Idea 4 is studied in more depth in Chapter 13.

Big Idea 5: Thermodynamics and Energy in Reactions

Today's consumers have a myriad of choices to decide when purchasing a cars or truck. What color do I want, leather or cloth seats, standard or automatic are just some of the decisions to be made. Another is what type of fuel: diesel, electric, gasoline, hybrid or flex fuel.

If you choose to go with flex fuel you then have the problem of determining which fuel to use. E-85 is cheaper than gasoline; however, it has lower heat content per gallon that gasoline, which will lower your mileage, hence you need more E-85 to go the same distance as you would with pure gasoline. What does the difference in price between the E-85 and gasoline have to be in order for you to save money filling up with E-85? What combination of gasoline and E-85 should you use in order to achieve the best value? Believe it or not, there are apps to help you with this.

How much heat content is produced when a fuel is combusted is a central question of thermodynamics. It is also a central question for your health. Which food is the best food to eat? How many Calories should I have in a day? What is a Calorie? These questions are studied in chapters 6 and 17.

Thermodynamics is usually broken into two subtopics: thermochemistry (chapter 6) and thermodynamics (chapter 17). In thermochemistry, the student will study the heat energy associated with a chemical reaction. Heat energy can either be produced in a chemical reaction (the reaction gets hot) or energy can be absorbed (the reaction gets cold). In chapter 17, two main questions are investigated: to what extent will this reaction take place on its own and how can the driving force behind the reaction be manipulated to increase a desired outcome (usually to produce more product in the chemical reaction)?

Big Idea 6: Equilibrium

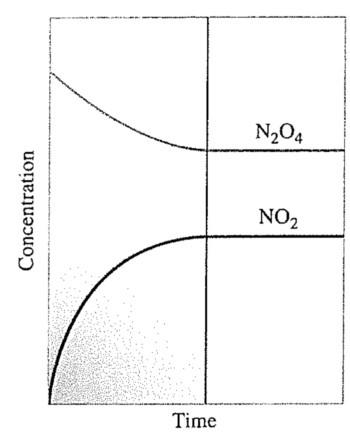
Equilibrium is, in many ways, the most important topic covered in AP® chemistry. Three whole chapters (14, 15 and 16) are devoted to the study of it. Equilibrium is the state in which there are no observable changes as time goes by. The observation is from the macroscopic level, not the microscopic level.

A typical high school in the United States starts at 8:00 am and ends at 3:00 pm. At 5:00 am the janitor may get to school, unlock the facilities and get ready for the school day. The first teacher may arrive around 6:30 am to set up a chemistry experiment. A coach may show up at 5:45 am to get ready for a 6:00 am early morning practice. Most of the students will arrive between 7:30 and 7:50, with some students (but not you) arriving late. However, the process is not one directional; people are free to come and go from the school. During the school day, students will leave to go to a doctor's appointment or go home sick. The principal may leave to attend a meeting at the administration building. Students will be arriving to school after a visit to the orthodontist; deliveries will be made for the cafeteria. Certainly between the hours of 8:30 am and 2:30 pm the number of people on the campus will remain constant. The rate of people that leave the campus and the rate of people who visit the campus are the same. This is the state of equilibrium. The macroscopic level is the number of people on the campus. The microscopic level is the actual people. At the level of the individual person the process is dynamic, not all of the same people are on campus throughout the day, but the number remains constant.

This same idea is applied to chemical reactions. To show a reversible process, a double arrow is used (⇆) to separate the reactants from the products. Consider the following reaction where gaseous dinitrogen tetroxide (N_2O_4) decomposes to form gaseous nitrogen dioxide (NO_2).

$$N_2O_4(g) \leftrightarrows 2NO_2(g)$$

Assume that you start out with a certain amount of $N_2O_4(g)$ and measure the amount of each species over time. The resulting data can be used to produce graph shown in the minor column.



Equilibrium between concentrations of dinitrogen tetraoxide (N_2O_4) and nitrogen dioxide (NO_2) over time.

At the time it takes to reach the vertical line, equilibrium has been attained. The macroscopic amounts (i.e. concentration) of N_2O_4 and NO_2 no longer changes, but at the microscopic level, the process continues—it is just that the rate at which N_2O_4 turns into NO_2 and the rate at which NO_2 forms N_2O_4 are the same.

Review of Concepts

1.2.1 The "five W's", or the "five W's and one H" (or in terms of chemistry W_5H_1) refers to questions whose answers are basic to any investigation. Who, What, Where, When, Why and How are the questions that we need to answer, whether we are investigating a chemical process or when we investigate a murder. Select three of the W_5H_1 and state how one of the Big Ideas provide the answer to your selected question.

1.3 Introducing the AP Science Practices

The seven science practices reflect the work that scientists engage in and that AP® students are expected to show competence in. They allow the students to make predictions, refine explanations and justify conclusions based on an analysis of the data and the underlying scientific principles.

Science Practice 1: Using models and representations to communicate scientific phenomena

Chemistry is a complex subject that seems to exist on two different levels, the macroscopic and the microscopic levels. The macroscopic level is what you can see with your eyes and experience with your senses. The microscopic level is what you cannot see with your eyes—it is at the particle level. The ability to use models (pictures, diagrams, symbolic representations, etc.) to explain the macroscopic properties that we see by representing the changes occurring at the microscopic level is key to understanding chemistry. Chemists uses many different forms of the microscopic representations, even for the same topic, and students need to be able to be comfortable moving between the various representations and know the correct situation for each representation.

Science Practice 2: Using math to solve problems

Mathematics is an integral part of chemistry and an essential skill for life. Students need to be able to choose appropriate mathematical processes to solve a particular problem. It is not enough to be able to "get the right answer"—rather students will need to be able to justify the selection they choose to solve a problem. Students need to be able to use estimation appropriately both in order to solve problems (get a "ballpark" answer) and to determine if the answer is a reasonable one.

Science Practice 3: Engaging in scientific questioning to guide thinking

Much to your parent's delight ("why is the sky blue?" or "what are you doing?") or apprehension ("where do babies come from?"), you have been engaging in scientific questioning about the natural world since you a baby. In order to conduct research, you must first know what you are trying to find! A meaningful laboratory experiment must first begin with a question—one that can be answered with empirical data. This leads in many cases to a hypothesis—and the beginning of the scientific method.

Science Practice 4: Collecting the correct data correctly

Now that you have your question, where do you go from here? The ability to gather data is an essential component of chemistry. Data can take many forms; personal experience or outside sources but most often in chemistry it is found by direct experimentation. The selection of the type of equipment needed and the degree of precision needed is an important skill in the chemical laboratory. In the every more complex world we live in, the skill that is increasing in importance is not to gather data, but which data to gather to answer a particular question.

Science Practice 5: Analyzing your data

Now that you have your data, what do you do with it? In many ways, the science of chemistry is the study of and the explanation for various patterns: the pattern of the periodic trends lead to the modern periodic table, the direct and inverse relationships between the various macroscopic properties of gases lead to a multitude of gas laws. Students need to be able to analyze data to identify these patterns, extrapolate behavior to new situations and to determine if there is sufficient data to answer a particular scientific question or make a valid claim about a hypothesis.

Science Practice 6: Articulating your claims

Students must be able to assembly a body of knowledge (the data) about a topic to provide the scientific foundations for their claims. Unlike other aspects of life, science requires all claims to have supporting data that has been vetted (peer reviewed) before it is accepted by the wider scientific community. However, science is an ever changing world, one where new knowledge and data is being integrated at an ever increasing pace and students need to be able to articulate why a particular set of data supports a given theory or suggests that the theory needs to be modified or replace.

Science Practice 7: Connecting the pieces

One of the hardest skills for students to develop is the ability to connect various "packets" of content knowledge across different chemistry topics. A central theme for study of chemistry is to connect atomic level interactions (particle level diagrams) to various macroscopic phenomena. A common question might be "how do the particle collisions affect the shift of a reaction at equilibrium?" Students should not have "blinders" on when they enter the chemistry classroom—bring all of your knowledge with you from your other subjects. You will have to connect the chemistry knowledge you are learning to other topics; why did the Apollo 13 astronauts develop kidney infections after they had to shut down their hydrogen fuel cells—connecting chemical reactions, electrochemistry, dehydration and its effect on human body systems.

Review of Concepts

1.3.1 Which science practice allows you to change 298×10.2 to the problem 300×10 to arrive at the answer of 3000?

1.4 Review: The Scientific Method

The material covered in the remaining sections of Chapter 1 should be a review from your previous chemistry or science courses. The material presented in these sections, however, is fundamental for your success in AP Chemistry. Refer to these sections to refresh your knowledge of the scientific method and assorted mathematical concepts.

All sciences, including the social sciences, employ variations of what is called the scientific method, a systematic approach to research. For example, a psychologist who wants to know how noise affects people's ability to learn chemistry and a chemist interested in measuring the heat given off when hydrogen gas burns in air would follow roughly the same procedure in carrying out their investigations. The first step is to carefully define the problem. The next step includes performing experiments, making careful observations, and recording information, or data, about the system—the part of the universe that is under investigation. (In the examples just discussed, the systems are the group of people the psychologist will study and a mixture of hydrogen and air.)

The data obtained in a research study may be both qualitative, consisting of general observations about the system, and quantitative, comprising numbers obtained by various measurements of the system. Chemists generally use standardized symbols and equations in recording their measurements and observations. This form of representation not only simplifies the process of keeping records, but also provides a common basis for communication with other chemists.

When the experiments have been completed and the data have been recorded, the next step in the scientific method is interpretation, meaning that the scientist attempts to explain the observed phenomenon. Based on the data that were gathered, the researcher formulates a hypothesis, a tentative explanation for a set of observations. Further experiments are devised to test the validity of the hypothesis in as many ways as possible, and the process begins anew. Figure 1.6 summarizes the main steps of the research process.

After a large amount of data has been collected, it is often desirable to summarize the information in a concise way, as a law. In science, a law is a concise verbal or mathematical statement of a relationship between phenomena that is always the same under the same conditions. For example, Sir Isaac Newton's second law of motion, which you may remember from high school science, says that force equals mass times acceleration ($F = ma$). What this law means is that an increase in the mass or in the acceleration of an object will always increase its force proportionally, and a decrease in mass or acceleration will always decrease the force.

Hypotheses that survive many experimental tests of their validity may evolve into theories. A theory is a unifying principle that explains a body of facts and/or those laws that are based on them. Theories, too, are constantly being tested. If a theory is disproved by experiment, then it must be discarded or modified so that it becomes consistent with experimental observations. Proving or disproving a theory can take years, even centuries, in part because the necessary technology may not be available. Atomic theory, which we will study in Chapter 2, is a case in point. It took more than 2000 years to work out this fundamental principle of chemistry proposed by Democritus, an ancient Greek philosopher. A more contemporary example is the search for the Higgs boson discussed on page 15.

Scientific progress is seldom, if ever, made in a rigid, step-by-step fashion. Sometimes a law precedes a theory; sometimes it is the other way around. Two scientists may start working on a project with exactly the same objective, but will end up taking drastically different approaches. Scientists are, after all, human beings, and their modes of thinking and working are very much influenced by their background, training, and personalities.

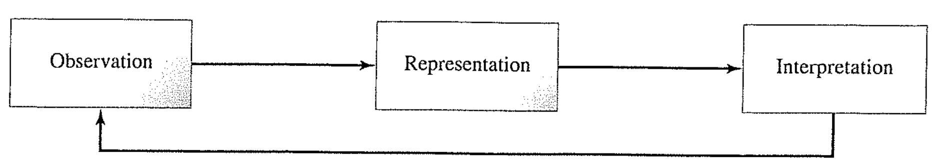


Figure 1.6 *The three levels of studying chemistry and their relationships. Observation deals with events in the macroscopic world; atoms and molecules constitute the microscopic world. Representation is a scientific shorthand for describing an experiment in symbols and chemical equations. Chemists use their knowledge of atoms and molecules to explain an observed phenomenon.*

The Search for the Higgs Boson

Have you ever wondered: Why does matter even have mass? It might seem obvious that "everything" has mass, but is that a requirement of nature? We will see later in our studies that light is composed of particles that do not have mass when at rest, and physics tells us under different circumstances the universe might not contain *anything* with mass. The search for the answer to this question illustrates nicely the process we call the scientific method.

Current theoretical models tell us that everything in the universe is based on two types of elementary particles: bosons and fermions. The building blocks of matter are constructed from fermions, while bosons are particles responsible for the force that holds the fermions together. In 1964, three different research teams independently proposed mechanisms in which a field of energy permeates the universe, and the interaction of matter with this field is due to a specific boson associated with the field. The greater the number of these bosons, the greater the interaction will be with the field. This interaction is the property we call mass, and the field and the associated boson came to be named for Peter Higgs, one of the original physicists to propose this mechanism.

This theory ignited a frantic search for the "Higgs boson," one of the most heralded quests in modern science. The Large Hadron Collider at CERN in Geneva, Switzerland (described in Chapter 19), was constructed to find evidence for the Higgs boson. In these experiments, protons are accelerated to nearly the speed of light in opposite directions in a circular 17-mile tunnel, and then allowed to collide, generating even more fundamental particles at very high energies. The data are examined for evidence of an excess of particles at an energy consistent with theoretical predictions for the Higgs boson. The ongoing process is the essence of the scientific method.

On July 4, 2012, scientists at CERN announced the discovery of the Higgs boson. It takes about 1 trillion proton–proton collisions to produce one Higgs boson event, so it requires a tremendous amount of data obtained from two independent sets of experiments to confirm the findings. In science, the quest for answers is never completely done. Our understanding can always be improved or refined, and sometimes entire tenets of accepted science are replaced by another theory that does a better job explaining the observations. For example, scientists are not sure if the Higgs boson is the only particle that confers mass to matter, or if it is only one of several such bosons predicted by other theories.

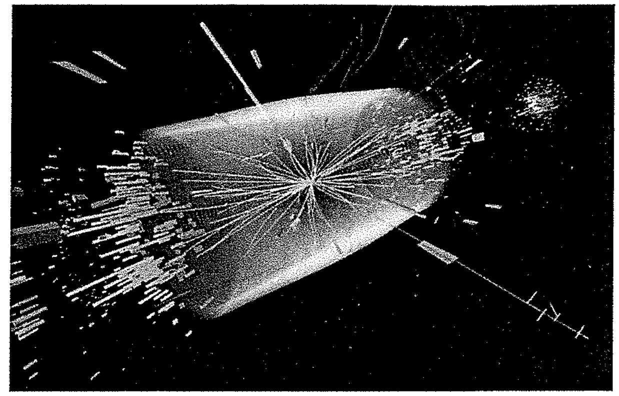


Illustration of the data obtained from decay of the Higgs boson in other particles following an 8-TeV collision at the Large Hadron Collider at CERN.
©Thomas McCauley/Lucas Taylor, CERN/Science Source

But over the long run, the scientific method has proven to be our best way of understanding the physical world. It took 50 years for experimental science to validate the existence of the Higgs boson. Using the scientific method, our understanding of our universe will continue to expand.

Thinking Critically

1. List the two types of elementary particles that make up matter in the universe and describe their specific functions.
2. How many collisions of accelerated protons are required to observe a Higgs boson event?
3. What is the physical property imparted by the Higgs boson?

The development of science has been irregular and sometimes even illogical. Great discoveries are usually the result of the cumulative contributions and experience of many workers, even though the credit for formulating a theory or a law is usually given to only one individual. There is, of course, an element of luck involved in scientific discoveries, but it has been said that "chance favors the prepared mind." It takes an alert and well-trained person to recognize the significance of an accidental discovery and to take full advantage of it. More often than not, the public learns only of spectacular scientific breakthroughs. For every success story, however, there are hundreds of cases in which scientists have spent years working on projects that ultimately led to a dead end, and in which positive achievements came only after many wrong turns and at such a slow pace that they went unheralded. Yet even the dead ends contribute something to the continually growing body of knowledge about the physical universe. It is the love of the search that keeps many scientists in the laboratory.

Review of Concepts

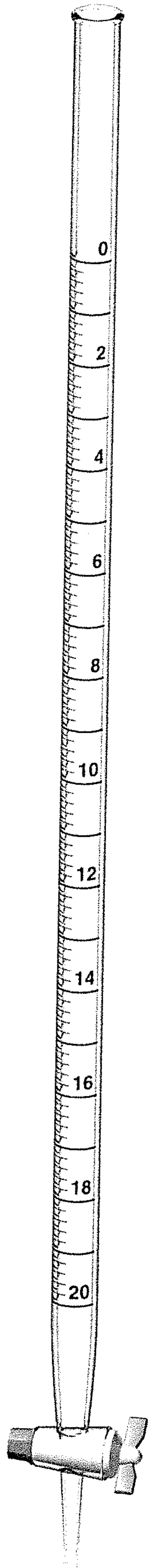
Which of the following statements is true?

(a) The hypothesis always leads to the formation of a law.
(b) The scientific method is a rigid sequence of steps in solving problems.
(c) A law summarizes a series of experimental observations; a theory provides an explanation for the observations.

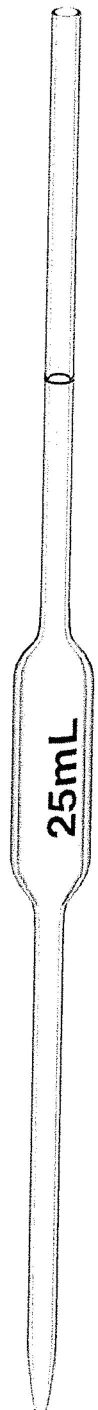
How are the seven science practices incorporated in the scientific method?

1.5 Review: Measurements

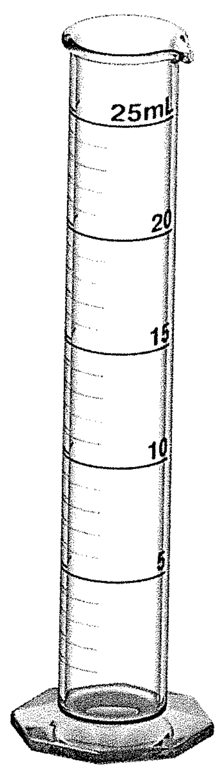
The measurements chemists make are often used in calculations to obtain other related quantities. Different instruments enable us to measure a substance's properties: The meterstick measures length or scale; the buret, the pipet, the graduated cylinder, and the volumetric flask measure volume (Figure 1.7); the balance measures mass; the thermometer measures temperature. These instruments provide measurements of macroscopic properties, which can be determined directly. Microscopic properties, on the atomic or molecular scale, must be determined by an indirect method, as we will see in Chapter 2.



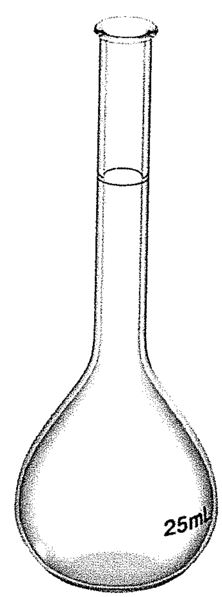
Buret



Pipet



Graduated cylinder



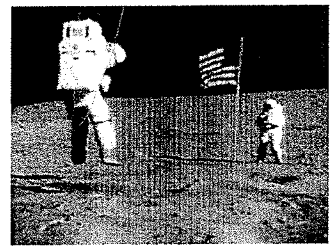
Volumetric flask

Figure 1.7 *Some common measuring devices found in a chemistry laboratory. These devices are not drawn to scale relative to one another. We will discuss the uses of these measuring devices in Chapter 4.*

Table 1.3 SI Base Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

Note that a metric prefix simply represents a number:
1 mm = 1×10^{-3} m



An astronaut jumping on the surface of the moon
Source: NASA

Table 1.4 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	μ	1/1,000,000, or 10^{-6}	1 micrometer (μm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m
femto-	f	1/1,000,000,000,000,000, or 10^{-15}	1 femtometer (fm) = 1×10^{-15} m
atto-	a	1/1,000,000,000,000,000,000, or 10^{-18}	1 attometer (am) = 1×10^{-18} m

A measured quantity is usually written as a number with an appropriate unit. To say that the distance between New York and San Francisco by car along a certain route is 5166 is meaningless. We must specify that the distance is 5166 kilometers. The same is true in chemistry; units are essential to stating measurements correctly.

SI Units

For many years, scientists recorded measurements in metric units, which are related decimally, that is, by powers of 10. In 1960, however, the General Conference of Weights and Measures, the international authority on units, proposed a revised metric system called the International System of Units (abbreviated SI, from the French Système Internationale d'Unites). Table 1.3 shows the seven SI base units. All other units of measurement can be derived from these base units. Like metric units, SI units are modified in decimal fashion by a series of prefixes, as shown in Table 1.4. We will use both metric and SI units in this book.

Measurements that we will utilize frequently in our study of chemistry include time, mass, volume, density, and temperature.

Mass and Weight

The terms "mass" and "weight" are often used interchangeably, although, strictly speaking, they are different quantities. Whereas mass is a measure of the amount of matter in an object, weight, technically speaking, is the force that gravity exerts on an

Figure 1.8 *The prototype kilogram is made of platinum-iridium alloy. It is kept in a vault at the International Bureau of Weights and Measures in Sèvres, France. In 2007 it was discovered that the alloy has mysteriously lost about 50 μg!*

object. An apple that falls from a tree is pulled downward by Earth's gravity. The mass of the apple is constant and does not depend on its location, but its weight does. For example, on the surface of the moon the apple would weigh only one-sixth what it does on Earth, because the moon's gravity is only one-sixth that of Earth. The moon's smaller gravity enabled astronauts to jump about rather freely on its surface despite their bulky suits and equipment. Chemists are interested primarily in mass, which can be determined readily with a balance; the process of measuring mass, oddly, is called weighing.

The SI unit of mass is the kilogram (kg). Unlike the units of length and time, which are based on natural processes that can be repeated by scientists anywhere, the kilogram is defined in terms of a particular object (Figure 1.8). In chemistry, however, the smaller gram (g) is more convenient:

$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^3 \text{ g}$$

Volume

The SI unit of length is the meter (m), and the SI-derived unit for volume is the cubic meter (m^3). Generally, however, chemists work with much smaller volumes, such as the cubic centimeter (cm^3) and the cubic decimeter (dm^3):

$$1 \text{ cm}^3 = (1 \times 10^{-2} \text{ m})^3 = 1 \times 10^{-6} \text{ m}^3$$
$$1 \text{ dm}^3 = (1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3$$

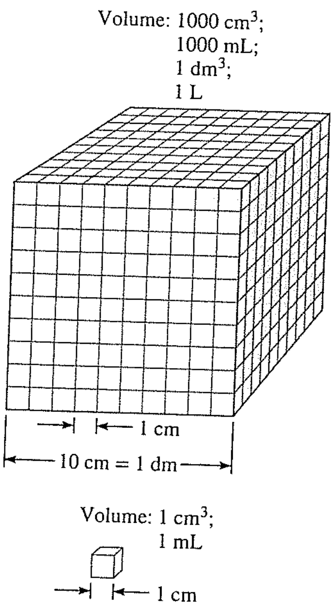


Figure 1.9 *Comparison of two volumes, 2 mL and 1000 mL.*

Another common unit of volume is the liter (L). A liter is the volume occupied by one cubic decimeter. One liter of volume is equal to 1000 milliliters (mL) or 1000 cm^3:

$$\begin{aligned} 1 \text{ L} &= 1000 \text{ mL} \\ &= 1000 \text{ cm}^3 \\ &= 1 \text{ dm}^3 \end{aligned}$$

and one milliliter is equal to one cubic centimeter:

$$1 \text{ mL} = 1 \text{ cm}^3$$

Figure 1.9 compares the relative sizes of two volumes. Even though the liter is not an SI unit, volumes are usually expressed in liters and milliliters.

Temperature Scales

Three temperature scales are currently in use. Their units are °F (degrees Fahrenheit), °C (degrees Celsius), and K (kelvin). The Fahrenheit scale, which is the most commonly used scale in the United States outside the laboratory, defines the normal freezing and boiling points of water to be exactly 32°F and 212°F, respectively. The Celsius scale divides the range between the freezing point (08C) and boiling point (100°C) of water into 100 degrees. As Table 1.2 shows, the kelvin is the SI base unit of temperature: It is the absolute temperature scale. By absolute we mean that the zero on the Kelvin scale, denoted by 0 K, is the lowest temperature that can be attained theoretically. On the other hand, 0°F and 0°C are based on the behavior of an arbitrarily chosen substance, water. Figure 1.10 compares the three temperature scales.

Note that the Kelvin scale does not have the degree sign. Also, temperatures expressed in kelvins can never be negative.

The size of a degree on the Fahrenheit scale is only 100/180, or 5/9, of a degree on the Celsius scale. To convert degrees Fahrenheit to degrees Celsius, we write

$$?°\text{C} = (°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} \tag{1.2}$$

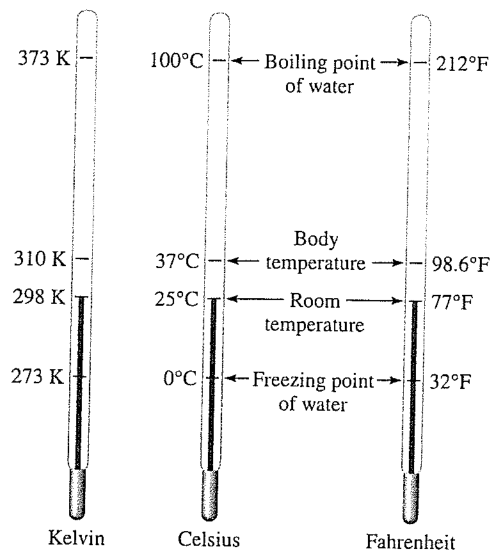


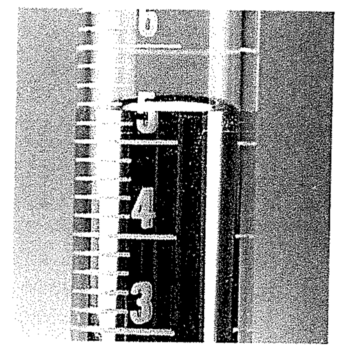
Figure 1.10 *Comparison of the three temperature scales: Celsius, and Fahrenheight, and the absolute (Kelvin) scales. Note that there are 100 divisions, or 100 degrees, between the freezing point and the boiling point of water on the Celsius scale, and there are 180 divisions, or 180 degrees, between the same two temperature limits on the Fahrenheit scale. The Celsius scale was formerly called the centrigrade scale.*

The following equation is used to convert degrees Celsius to degrees Fahrenheit:

$$?°F = \frac{9°F}{5°C} \times (°C) + 32°F \tag{1.3}$$

Both the Celsius and the Kelvin scales have units of equal magnitude; that is, one degree Celsius is equivalent to one kelvin. Experimental studies have shown that absolute zero on the Kelvin scale is equivalent to −273.15°C on the Celsius scale. Thus, we can use the following equation to convert degrees Celsius to kelvin:

$$?\ K = (°C + 273.15°C)\frac{1\ K}{1°C} \tag{1.4}$$



Mercury.
©McGraw-Hill Education/Stephen Frisch

We will frequently find it necessary to convert between degrees Celsius and degrees Fahrenheit and between degrees Celsius and kelvin. Example 1.3 illustrates these conversions.

The Chemistry in Action essay on page 20 shows why we must be careful with units in scientific work.

Example 1.3

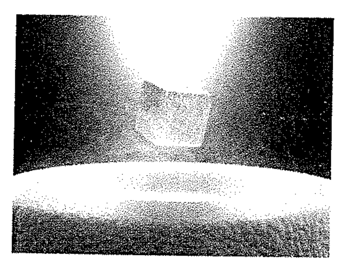
(a) Below the transition temperature of −141°C, a certain substance becomes a superconductor; that is, it can conduct electricity with no resistance. What is the temperature in degrees Fahrenheit? (b) Helium has the lowest boiling point of all the elements at −452°F. Convert this temperature to degrees Celsius. (c) Mercury, the only metal that exists as a liquid at room temperature, melts at −38.9°C. Convert its melting point to kelvins.

Solution These three parts require that we carry out temperature conversions, so we need Equations (1.2), (1.3), and (1.4). Keep in mind that the lowest temperature on the Kelvin scale is zero (0 K); therefore, it can never be negative.

(a) This is carried out by writing

$$\frac{9°F}{5°C} \times (-141°C) + 32°F = -222°F$$

(Continued)



Magnet suspended above superconductor cooled below its transition temperature by liquid nitrogen.
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CHEMISTRY *in Action*

The Importance of Units

In December 1998, NASA launched the 125-million dollar Mars Climate Orbiter, intended as Mars' first weather satellite. After a 416-million mi journey, the spacecraft was supposed to go into Mars' orbit on September 23, 1999. Instead, it entered Mars' atmosphere about 100 km (62 mi) lower than planned and was destroyed by heat. The mission controllers said the loss of the spacecraft was due to the failure to convert English measurement units into metric units in the navigation software.

Engineers at Lockheed Martin Corporation who built the spacecraft specified its thrust in pounds, which is an English unit. Scientists at NASA's Jet Propulsion Laboratory, on the other hand, had assumed that thrust data they received were expressed in metric units, as newtons. Expressed as a unit for force, 1 lb is the force due to gravitational attraction on an object of that mass. To carry out the conversion between pound and newton, we start with 1 lb 5 0.4536 kg and from Newton's second law of motion,

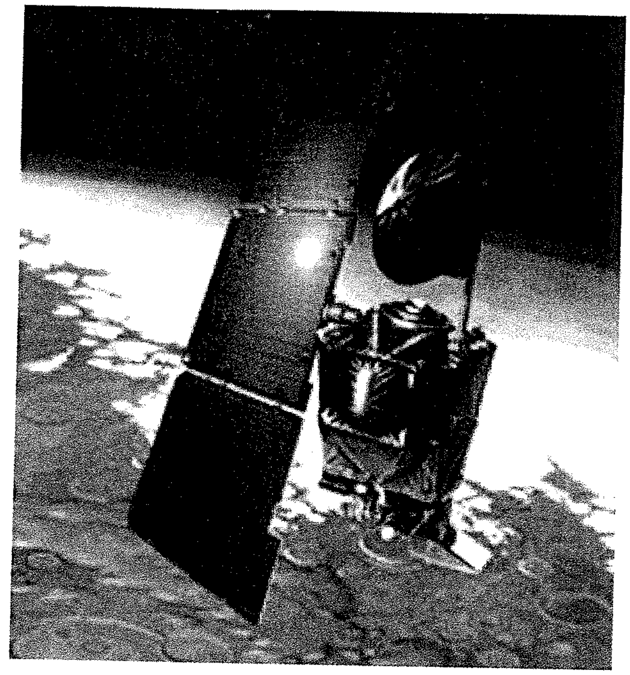
$$\begin{aligned} \text{force} &= \text{mass} \times \text{acceleration} \\ &= 0.4536 \text{ kg} \times 9.81 \text{ m/s}^2 \\ &= 4.45 \text{ kg m/s}^2 \\ &= 4.45 \text{ N} \end{aligned}$$

because 1 newton (N) = 1 kg m/s^2. Therefore, instead of converting 1 lb of force to 4.45 N, the scientists treated it as 1 N.

The considerably smaller engine thrust expressed in newtons resulted in a lower orbit and the ultimate destruction of the spacecraft. Commenting on the failure of the Mars mission, one scientist said: "This is going to be the cautionary tale that will be embedded into introduction to the metric system in elementary school, high school, and college science courses till the end of time."

Thinking Critically

1. What fundamental misunderstanding resulted in the loss of the Mars Climate Orbiter?
2. How far off was the orbiter from its intended atmospheric entry altitude and what was its fate?



Artist's conception of the Martian Climate Orbiter.
Source: NASA/JPL-Caltech

(b) Here we have

$$(-452°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}} = -269°\text{C}$$

(c) The melting point of mercury in kelvins is given by

$$(-38.9°\text{C} + 273.15°\text{C}) \times \frac{1 \text{ K}}{1°\text{C}} = 234.3 \text{ K}$$

Practice Exercise Convert (a) 327.5°C (the melting point of lead) to degrees Fahrenheit; (b) 172.9°F (the boiling point of ethanol) to degrees Celsius; and (c) 77 K. the boiling point of liquid nitrogen, to degrees Celsius.

Similar problems: 1.26, 1.27, 1.28.

Review of Concepts

1.5.1 The melting point of adamantane is 518°F. What is this melting point in kelvins?
1.5.2 What are the SI units for length? volume? mass?
1.5.3 Explain the difference between mass and weight

1.6 Review: Handling Numbers

Having surveyed some of the units used in chemistry, we now turn to techniques for handling numbers associated with measurements: scientific notation and significant figures.

Scientific Notation

Chemists often deal with numbers that are either extremely large or extremely small. For example, in 1 g of the element hydrogen there are roughly

$$602{,}200{,}000{,}000{,}000{,}000{,}000{,}000$$

hydrogen atoms. Each hydrogen atom has a mass of only

$$0.00000000000000000000000166 \text{ g}$$

These numbers are cumbersome to handle, and it is easy to make mistakes when using them in arithmetic computations. Consider the following multiplication:

$$0.0000000056 \times 0.00000000048 = 0.000000000000000002688$$

It would be easy for us to miss one zero or add one more zero after the decimal point. Consequently, when working with very large and very small numbers, we use a system called scientific notation. Regardless of their magnitude, all numbers can be expressed in the form

$$N \times 10^n$$

where N is a number between 1 and 10 and n, the exponent, is a positive or negative integer (whole number). Any number expressed in this way is said to be written in scientific notation.

Suppose that we are given a certain number and asked to express it in scientific notation. Basically, this assignment calls for us to find n. We count the number of places that the decimal point must be moved to give the number N (which is between 1 and 10). If the decimal point has to be moved to the left, then n is a positive integer; if it has to be moved to the right, n is a negative integer. The following examples illustrate the use of scientific notation:

(1) Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^2$$

Note that the decimal point is moved to the left by two places and n = 2.

(2) Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6}$$

Here the decimal point is moved to the right by six places and n = −6.

Keep in mind the following two points. First, n = 0 is used for numbers that are not expressed in scientific notation. For example, 74.6×10^0 (n = 0) is equivalent to 74.6. Second, the usual practice is to omit the superscript when n = 1. Thus, the scientific notation for 74.6 is 7.46×10 and not 7.46×10^1.

Any number raised to the power zero is equal to one.

Next, we consider how scientific notation is handled in arithmetic operations.

Addition and Subtraction

To add or subtract using scientific notation, we first write each quantity—say, N_1 and N_2—with the same exponent n. Then we combine N_1 and N_2; the exponents remain the same. Consider the following examples:

$$(7.4 \times 10^3) + (2.1 \times 10^3) = 9.5 \times 10^3$$
$$\begin{aligned}(4.31 \times 10^4) + (3.9 \times 10^3) &= (4.31 \times 10^4) + (0.39 \times 10^4)\\ &= 4.70 \times 10^4\end{aligned}$$
$$\begin{aligned}(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) &= (2.22 \times 10^{-2}) - (0.41 \times 10^{-2})\\ &= 1.81 \times 10^{-2}\end{aligned}$$

Multiplication and Division

To multiply numbers expressed in scientific notation, we multiply N_1 and N_2 in the usual way, but add the exponents together. To divide using scientific notation, we divide N_1 and N_2 as usual and subtract the exponents. The following examples show how these operations are performed:

$$\begin{aligned}(8.0 \times 10^4) \times (5.0 \times 10^2) &= (8.0 \times 5.0)(10^{4+2})\\ &= 40 \times 10^6\\ &= 4.0 \times 10^7\end{aligned}$$
$$\begin{aligned}(4.0 \times 10^{-5}) \times (7.0 \times 10^3) &= (4.0 \times 7.0)(10^{-5+3})\\ &= 28 \times 10^{-2}\\ &= 2.8 \times 10^{-1}\end{aligned}$$
$$\begin{aligned}\frac{6.9 \times 10^7}{3.0 \times 10^{-5}} &= \frac{6.9}{3.0} \times 10^{7-(-5)}\\ &= 2.3 \times 10^{12}\end{aligned}$$
$$\begin{aligned}\frac{8.5 \times 10^4}{5.0 \times 10^9} &= \frac{8.5}{5.0} \times 10^{4-9}\\ &= 1.7 \times 10^{-5}\end{aligned}$$

Significant Figures

Except when all the numbers involved are integers (for example, in counting the number of students in a class), it is often impossible to obtain the exact value of the quantity under investigation. For this reason, it is important to indicate the margin of error in a measurement by clearly indicating the number of significant figures, which are the meaningful digits in a measured or calculated quantity. When significant figures are used, the last digit is understood to be uncertain. For example, we might measure the volume of a given amount of liquid using a graduated cylinder with a scale that gives an uncertainty of 1 mL in the measurement. If the volume is found to be 6 mL, then the actual volume is in the range of 5 mL to 7 mL. We represent the volume of the liquid as (6 ± 1) mL. In this case, there is only one significant figure (the digit 6) that is uncertain by either plus or minus 1 mL. For greater accuracy, we might use a graduated cylinder that has finer divisions, so that the volume we measure is now uncertain by only 0.1 mL. If the volume of the liquid is now found to be 6.0 mL, we may express the quantity as (6.0 ± 0.1) mL, and the actual value is somewhere between 5.9 mL and 6.1 mL.

We can further improve the measuring device and obtain more significant figures, but in every case, the last digit is always uncertain; the amount of this uncertainty depends on the particular measuring device we use.

Figure 1.11 shows a modern balance. Balances such as this one are available in many general chemistry laboratories; they readily measure the mass of objects to four decimal places. Therefore, the measured mass typically will have four significant figures (for example, 0.8642 g) or more (for example, 3.9745 g). Keeping track of the number of significant figures in a measurement such as mass ensures that calculations involving the data will reflect the precision of the measurement.

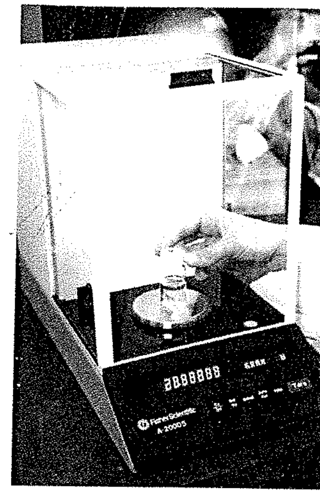


Figure 1.11 *A Fisher Scientific A-200DS Digital Recorder Precision Balance*

Student Hot Spot

Student data indicate you may struggle with significant figures. Access your eBook for additional Learning Resources on this topic.

Guidelines for Using Significant Figures

We must always be careful in scientific work to write the proper number of significant figures. In general, it is fairly easy to determine how many significant figures a number has by following these rules:

1. Any digit that is not zero is significant. Thus, 845 cm has three significant figures, 1.234 kg has four significant figures, and so on.
2. Zeros between nonzero digits are significant. Thus, 606 m contains three significant figures, 40,501 kg contains five significant figures, and so on.
3. Zeros to the left of the first nonzero digit are not significant. Their purpose is to indicate the placement of the decimal point. For example, 0.08 L contains one significant figure, 0.0000349 g contains three significant figures, and so on.
4. If a number is greater than 1, then all the zeros written to the right of the decimal point count as significant figures. Thus, 2.0 mg has two significant figures, 40.062 mL has five significant figures, and 3.040 dm has four significant figures. If a number is less than 1, then only the zeros that are at the end of the number and the zeros that are between nonzero digits are significant. This means that 0.090 kg has two significant figures, 0.3005 L has four significant figures, 0.00420 min has three significant figures, and so on.
5. For numbers that do not contain decimal points, the trailing zeros (that is, zeros after the last nonzero digit) may or may not be significant. Thus, 400 cm may have one significant figure (the digit 4), two significant figures (40), or three significant figures (400). We cannot know which is correct without more information. By using scientific notation, however, we avoid this ambiguity. In this particular case, we can express the number 400 as 4×10^2 for one significant figure, 4.0×10^2 for two significant figures, or 4.00×10^2 for three significant figures.

Example 1.4

Determine the number of significant figures in the following measurements: (a) 394 cm, (b) 5.03 g, (c) 0.714 m, (d) 0.052 kg, (e) 2.720×10^{-2} atoms, (f) 3000 mL.

Solution (a) Three, because each digit is a nonzero digit. (b) Three, because zeros between nonzero digits are significant. (c) Three, because zeros to the left of the first nonzero digit do not count as significant figures. (d) Two. Same reason as in (c). (e) Four. Because the number is greater than one, all the zeros written to the right of the decimal point count as significant figures. (f) This is an ambiguous case. The number of significant figures may be four (3.000×10^3), three (3.00×10^3), two (3.0×10^3), or one (3×10^3). This example illustrates why scientific notation must be used to show the proper number of significant figures.

Practice Exercise Determine the number of significant figures in each of the following measurements: (a) 35 mL, (b) 2008 g, (c) $0.0580\ m^3$, (d) 7.2×10^4 molecules, (e) 830 kg.

Similar problems: 1.26

A second set of rules specifies how to handle significant figures in calculations.

1. In addition and subtraction, the answer cannot have more digits to the right of the decimal point than either of the original numbers. Consider these examples:

$$\begin{array}{r l} 89.332 & \\ \underline{+\ 1.1} & \longleftarrow \text{one digit after the decimal point} \\ 90.432 & \longleftarrow \text{round off to 90.4} \end{array}$$

$$\begin{array}{r l} 2.097 & \\ \underline{-\ 0.12} & \longleftarrow \text{two digits after the decimal point} \\ 1.977 & \longleftarrow \text{round off to 1.98} \end{array}$$

The rounding-off procedure is as follows. To round off a number at a certain point we simply drop the digits that follow if the first of them is less than 5. Thus, 8.724 rounds off to 8.72 if we want only two digits after the decimal point. If the first digit following the point of rounding off is equal to or greater than 5, we add 1 to the preceding digit. Thus, 8.727 rounds off to 8.73, and 0.425 rounds off to 0.43.

2. In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the smallest number of significant figures. The following examples illustrate this rule:

$$2.8 \times 4.5039 = 12.61092 \longleftarrow \text{round off to 13}$$

$$\frac{6.85}{112.04} = 0.0611388789 \longleftarrow \text{round off to 0.0611}$$

3. Keep in mind that exact numbers obtained from definitions or by counting numbers of objects can be considered to have an infinite number of significant figures. For example, the inch is defined to be exactly 2.54 centimeters; that is,

$$1 \text{ in} = 2.54 \text{ cm}$$

Thus, the "2.54" in the equation should not be interpreted as a measured number with three significant figures. In calculations involving conversion between "in" and "cm," we treat both "1" and "2.54" as having an infinite number of significant figures. Similarly, if an object has a mass of 5.0 g, then the mass of nine such objects is

$$5.0 \text{ g} \times 9 = 45 \text{ g}$$

The answer has two significant figures because 5.0 g has two significant figures. The number 9 is exact and does not determine the number of significant figures.

Example 1.5 shows how significant figures are handled in arithmetic operations.

Student Hot Spot

Student data indicate you may struggle with handling significant figures in calculations. Access your eBook for additional Learning Resources on this topic.

Example 1.5

Carry out the following arithmetic operations to the correct number of significant figures: (a) 12,343.2 g + 0.1893 g, (b) 55.67 L − 2.386 L, (c) 7.52 m × 6.9232, (d) 0.0239 kg /46.5 mL, (e) 5.21×10^3 cm + 2.92×10^2 cm.

Solution In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the significant number of the answer is determined by the number having the smallest number of significant figures.

(a)
$$\begin{array}{r l} 12{,}343.2\ \text{g} & \\ \underline{+\qquad 0.1893\ \text{g}} & \\ 12{,}343.3893\ \text{g} & \longleftarrow \text{round off to 12,343.4 g} \end{array}$$

(Continued)

(b) $\begin{array}{r} 55.67\text{ L} \\ -\ 2.386\text{ L} \\ \hline 53.284\text{ L} \end{array}$ $\longleftarrow$ round off to 53.28 L

(c) $7.52\text{ m} \times 6.9232 = 52.06246\text{ m}$ $\longleftarrow$ round off to 52.1 m

(d) $\dfrac{0.0239\text{ kg}}{46.5\text{ mL}} = 0.0005139784946\text{ kg/mL}$ $\longleftarrow$ round off to 0.000514 kg/mL or 5.14×10^{-4} kg/mL

(e) First we change 2.92×10^2 cm to 0.292×10^3 cm and then carry out the addition $(5.21\text{ cm} + 0.292\text{ cm}) \times 10^3$. Following the procedure in (a), we find the answer is 5.50×10^3 cm.

Practice Exercise Carry out the following arithmetic operations and round off the answers to the appropriate number of significant figures: (a) 26.5862 L + 0.17 L, (b) 9.1 g − 4.682 g, (c) 7.1×10^4 dm $\times$ 2.2654×10^2 dm, (d) 6.54 g / 86.5542 mL, (e) $(7.55 \times 10^4\text{ m}) - (8.62 \times 10^3\text{ m})$.

Similar problems: 1.34, 1.50, 1.77

The preceding rounding-off procedure applies to one-step calculations. In chain calculations, that is, calculations involving more than one step, we can get a different answer depending on how we round off. Consider the following two-step calculations:

$$\begin{array}{ll} \text{First step:} & A \times B = C \\ \text{Second step:} & C \times D = E \end{array}$$

Let's suppose that A = 3.66, B = 8.45, and D = 2.11. Depending on whether we round off C to three or four significant figures, we obtain a different number for E:

Method 1	Method 2
$3.66 \times 8.45 = 30.9$	$3.66 \times 8.45 = 30.93$
$30.9 \times 2.11 = 65.2$	$30.93 \times 2.11 = 65.3$

However, if we had carried out the calculation as $3.66 \times 8.45 \times 2.11$ on a calculator without rounding off the intermediate answer, we would have obtained 65.3 as the answer for E. Although retaining an additional digit past the number of significant figures for intermediate steps helps to eliminate errors from rounding, this procedure is not necessary for most calculations because the difference between the answers is usually quite small. Therefore, for most examples and end-of-chapter problems where intermediate answers are reported, all answers, intermediate and final, will be rounded.

Accuracy and Precision

In discussing measurements and significant figures, it is useful to distinguish between accuracy and precision. Accuracy tells us how close a measurement is to the true value of the quantity that was measured. Precision refers to how closely two or more measurements of the same quantity agree with one another (Figure 1.12).

The difference between accuracy and precision is a subtle but important one. Suppose, for example, that three students are asked to determine the mass of a piece of copper wire.

The results of two successive weighings by each student are

	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	1.971 g	1.970 g	2.001 g

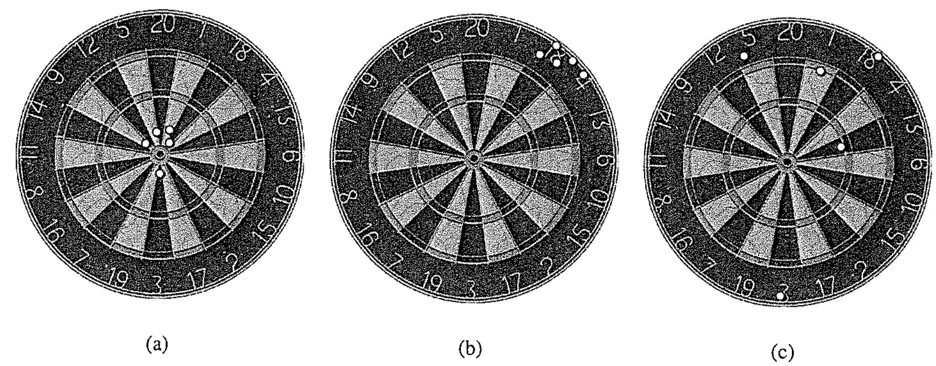
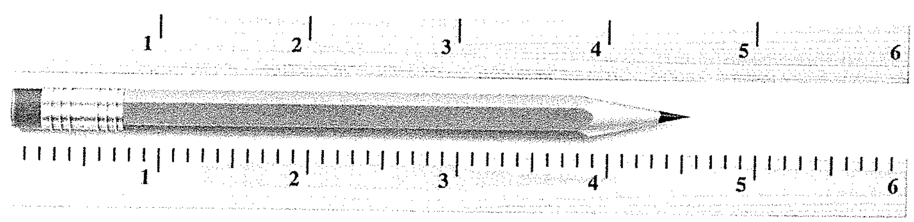


Figure 1.12 *The distribution of holes formed by darts on a dart board shows the difference between precise and accurate. (a) Good accuracy and good precision. (b) Good accuracy and poor precision. (c) Poor accuracy and poor precision.*

The true mass of the wire is 2.000 g. Therefore, Student B's results are more precise than those of Student A (1.972 g and 1.968 g deviate less from 1.970 g than 1.964 g and 1.978 g from 1.971 g), but neither set of results is very accurate. Student C's results are not only the most precise, but also the most accurate, because the average value is closest to the true value. Highly accurate measurements are usually precise too. On the other hand, highly precise measurements do not necessarily guarantee accurate results. For example, an improperly calibrated meterstick or a faulty balance may give precise readings that are in error.

Review of Concepts

1.6.1 Give the length of the pencil with proper significant figures according to which ruler you use for the measurement.



1.6.2 A student measures the density of an alloy with the following results: 10.28 g/cm^3, 9.97 g/cm^3, 10.22 g/cm^3, 10.15 g/cm^3, 9.94 g/cm^3. How should the average value for the density be reported?

1.7 Review: Dimensional Analysis in Solving Problems

Careful measurements and the proper use of significant figures, along with correct calculations, will yield accurate numerical results. But to be meaningful, the answers also must be expressed in the desired units. The procedure we use to convert between units in solving chemistry problems is called dimensional analysis (also called the factor-label method). A simple technique requiring little memorization, dimensional analysis is based on the relationship between different units that express the same physical quantity. For example, by definition 1 in = 2.54 cm (exactly). This equivalence enables us to write a conversion factor as follows:

$$\frac{1\ \text{in}}{2.54\ \text{cm}}$$

Because both the numerator and the denominator express the same length, this fraction is equal to 1. Similarly, we can write the conversion factor as

$$\frac{2.54\ \text{cm}}{1\ \text{in}}$$

which is also equal to 1. Conversion factors are useful for changing units. Thus, if we wish to convert a length expressed in inches to centimeters, we multiply the length by the appropriate conversion factor.

$$12.00\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 30.48\ \text{cm}$$

We choose the conversion factor that cancels the unit inches and produces the desired unit, centimeters. Note that the result is expressed in four significant figures because 2.54 is an exact number.

Next let us consider the conversion of 57.8 meters to centimeters. This problem can be expressed as

$$?\ \text{cm} = 57.8\ \text{m}$$

By definition,

$$1\ \text{cm} = 1 \times 10^{-2}\text{m}$$

Because we are converting "m" to "cm," we choose the conversion factor that has meters in the denominator,

$$\frac{1\ \text{cm}}{1 \times 10^{-2}\ \text{m}}$$

and write the conversion as

$$\begin{aligned} ?\ \text{cm} &= 57.8\ \cancel{\text{m}} \times \frac{1\ \text{cm}}{1 \times 10^{-2}\ \cancel{\text{m}}} \\ &= 5780\ \text{cm} \\ &= 5.78 \times 10^{3}\ \text{cm} \end{aligned}$$

Note that scientific notation is used to indicate that the answer has three significant figures. Again, the conversion factor 1 cm/1 $\times 10^{-2}$ m contains exact numbers; therefore, it does not affect the number of significant figures.

In general, to apply dimensional analysis we use the relationship

$$\text{given quantity} \times \text{conversion factor} = \text{desired quantity}$$

and the units cancel as follows:

$$\cancel{\text{given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

Remember that the unit we want appears in the numerator and the unit we want to cancel appears in the denominator.

In dimensional analysis, the units are carried through the entire sequence of calculations. Therefore, if the equation is set up correctly, then all the units will cancel except the desired one. If this is not the case, then an error must have been made somewhere, and it can usually be spotted by reviewing the solution.

A Note on Problem Solving

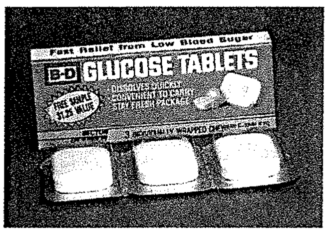
Remember that the unit we want appears in the numerator and the unit we want to cancel appears in the denominator.

At this point you have been introduced to scientific notation, significant figures, and dimensional analysis, which will help you in solving numerical problems. Chemistry is an experimental science and many of the problems are quantitative in nature. The key to success in problem solving is practice. Just as a marathon runner cannot prepare for a race by simply reading books on running and a pianist cannot give a successful concert by only memorizing the musical score, you cannot be sure of your understanding of chemistry without solving problems.

The following steps will help to improve your skill at solving numerical problems.

1. Read the question carefully. Understand the information that is given and what you are asked to solve. Frequently it is helpful to make a sketch that will help you to visualize the situation.

2. Find the appropriate equation that relates the given information and the unknown quantity. Sometimes solving a problem will involve more than one step, and you may be expected to look up quantities in tables that are not provided in the problem. Dimensional analysis is often needed to carry out conversions.
3. Check your answer for the correct sign, units, and significant figures.
4. A very important part of problem solving is being able to judge whether the answer is reasonable. It is relatively easy to spot a wrong sign or incorrect units. But if a number (say, 9) is incorrectly placed in the denominator instead of in the numerator, the answer would be too small even if the sign and units of the calculated quantity were correct.
5. One quick way to check the answer is to round off the numbers in the calculation in such a way so as to simplify the arithmetic. The answer you get will not be exact, but it will be close to the correct one.



Glucose tablets can provide diabetics with a quick method for raising their blood sugar levels.
©Leonard Lessin/Science Source

Conversion factors for some of the English system units commonly used in the United States for nonscientific measurements (for example, pounds and inches) are provided inside the back cover of this book.

Example 1.6

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Strategy The problem can be stated as

$$? \text{ mg} = 0.0833 \text{ lb}$$

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams. A metric conversion is then needed to convert grams to milligrams (1 mg = 1×10^{-3} g). Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.

Solution The sequence of conversions is

$$\text{pounds} \longrightarrow \text{grams} \longrightarrow \text{milligrams}$$

Using the following conversion factors

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3}}$$

we obtain the answer in one step:

$$? \text{ mg} = 0.0833 \cancel{\text{lb}} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = 3.78 \times 10^4 \text{ mg}$$

Check As an estimate, we note that 1 lb is roughly 500 g and that 1 g = 1000 mg. Therefore, 1 lb is roughly 5×10^5 mg. Rounding off 0.0833 lb to 0.1 lb, we get 5×10^4 mg, which is close to the preceding quantity.

Practice Exercise A roll of aluminum foil has a mass of 1.07 kg. What is its mass in pounds?

Similar problem: 1.28, 1.33

Example 1.7

A liquid helium storage tank has a volume of 275 L. What is the volume in m^3?

Strategy The problem can be stated as

$$? \text{ m}^3 = 275 \text{ L}$$

(Continued)

How many conversion factors are needed for this problem? Recall that 1 L = 1000 cm^3 and 1 cm = 1×10^{-2} m.

Solution We need two conversion factors here: one to convert liters to cm^3 and one to convert centimeters to meters:

$$\frac{1000\ \text{cm}^3}{1\ \text{L}} \quad \text{and} \quad \frac{1 \times 10^{-2}\ \text{m}}{1\ \text{cm}}$$

Because the second conversion deals with length (cm and m) and we want volume here, it must therefore be cubed to give

$$\frac{1 \times 10^{-2}\ \text{m}}{1\ \text{cm}} \times \frac{1 \times 10^{-2}\ \text{m}}{1\ \text{cm}} \times \frac{1 \times 10^{-2}\ \text{m}}{1\ \text{cm}} = \left(\frac{1 \times 10^{-2}\ \text{m}}{1\ \text{cm}}\right)^3$$

This means that 1 cm^3 = 1×10^{-6} m^3. Now we can write

$$?\ \text{m}^3 = 275\ \text{L} \times \frac{1000\ \text{cm}^3}{1\ \text{L}} \times \left(\frac{1 \times 10^{-2}\ \text{m}}{1\ \text{cm}}\right)^3 = 0.275\ \text{m}^3$$

Check From the preceding conversion factors you can show that 1 L = 1×10^{-3} m^3. Therefore, a 275-L storage tank would be equal to 275×10^{-3} m^3 or 0.275 m^3, which is the answer.

Practice Exercise The volume of a room is 1.08×10^8 dm^3. What is the volume in m^3?

Similar problem: 1.27, 1.28

Example 1.8

Liquid nitrogen is obtained from liquefied air and is used to prepare frozen goods and in low-temperature research. The density of the liquid at its boiling point (−196°C or 77 K) is 0.808 g/cm^3. Convert the density to units of kg/m^3.

Strategy The problem can be stated as

$$?\ \text{kg/m}^3 = 0.808\ \text{g/cm}^3$$

Two separate conversion are required for this problem: g $\longrightarrow$ kg and $cm^3 \longrightarrow m^3$
Recall that 1 kg = 1000 g and 1 cm = 1×10^{-2} m

Solution In Example 1.7, we saw that 1 m^3 = 1×10^6 cm^3. The conversion factors are

$$\frac{1\ \text{kg}}{1000\ \text{g}} \quad \text{and} \quad \frac{1\ \text{cm}^3}{1 \times 10^{-6}\ \text{m}^{-3}}$$

Finally,

$$?\ \text{kg/m}^3 = \frac{0.808\ \cancel{\text{g}}}{1\ \cancel{\text{cm}^3}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} \times \frac{1\ \cancel{\text{cm}^3}}{1 \times 10^{-6}\ \text{m}^3} = 808\ \text{kg/m}^3$$

Check Because 1 m^3 = 1×10^6 cm^3, we would expect much more mass in 1 m^3 than in 1 cm^3. Therefore, the answer is reasonable.

Practice Exercise The density of the lightest metal, lithium (Li) is 5.34×10^2 kg/m^3. Convert the density to g/cm^3.

Review of Concepts

1.7.1 What is the volume in L of a 1.24-kg sample of benzene (density = 0.876 g/mL)?

1.7.2 The heat of combustion of TNT is 14.5 MJ/kg. What is this heat of combustion in J/g?

1.8 Review: Problem Solving: Information, Assumptions, and Simplifications

In chemistry, as in other scientific disciplines, it is not always possible to solve a numerical problem exactly. There are many reasons why this is the case. For example, our understanding of a situation is not complete or data are not fully available. In these cases, we must learn to make an intelligent guess. This approach is sometimes called "ball-park estimates," which are simple, quick calculations that can be done on the "back of an envelope." As you can imagine, in many cases the answers are only order-of-magnitude estimates.†

In most of the example problems that you have seen so far, as well as the questions given at the end of this and subsequent chapters, the necessary information is provided; however, in order to solve important real-world problems such as those related to medicine, energy, and agriculture, you must be able to determine what information is needed and where to find it. Much of the information you might need can be found in the various tables located throughout the text, and a list of tables and important figures is given on the inside back cover. In many cases, however, you will need to go to outside sources to find the information you need. Although the Internet is a fast way to find information, you must take care that the source is reliable and well referenced. One excellent source is the National Institute of Standards and Technology (NIST).

In order to know what information you need, you will first have to formulate a plan for solving the problem. In addition to the limitations of the theories used in science, typically assumptions are made in setting up and solving the problems based on those theories. These assumptions come at a price, however, as the accuracy of the answer is reduced with increasing simplifications of the problem, as illustrated in Example 1.9.

Considering Example 1.9, even if the dimensions of the pencil lead were measured with greater precision, the accuracy of the final answer would be limited by the assumptions made in modeling this problem. The pencil lead is actually a mixture of graphite and clay, where the relative amounts of the two materials determine the soft-ness of the lead, so the density of the material is likely to be different than 2.2 g/cm^3. You could probably find a better value for the density of the mixture used to make No. 2 pencils, but it is not worth the effort in this case.

Example 1.9

A modern pencil "lead" is actually composed primarily of graphite, a form of carbon. Estimate the mass of the graphite core in a standard No. 2 pencil before it is sharpened.

Strategy Assume that the pencil lead can be approximated as a cylinder. Measurement of a typical unsharpened pencil gives a length of about 18 cm (subtracting the length of the eraser head) and a diameter of roughly 2 mm for the lead. The volume of a cylinder V is given by $V = \pi r^2 l$, where r is the radius and l is the length. Assuming that the lead is pure graphite, you can calculate the mass of the lead from the volume using the density of graphite given in Table 1.2.

Solution Converting the diameter of the lead to units of cm gives

$$2 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 0.2 \text{ cm}$$

(Continued)

†An order of magnitude is a factor of 10.

which, along with the length of the lead, gives

$$V = \pi\left(\frac{0.2\text{ cm}}{2}\right)^2 \times 18\text{ cm}$$
$$= 0.57\text{ cm}^3$$

Rearranging Equation (1.1) gives

$$m = d \times V$$
$$= 2.2\frac{\text{g}}{\text{cm}^3} \times -.57\text{cm}^3$$
$$= 1\text{ g}$$

Check Rounding off the values used to calculate the volume of the lead gives $3 \times (0.1\text{ cm})^2 \times 20\text{ cm} = 0.6\text{ cm}^3$. Multiplying that volume by roughly 2 g/cm^3 gives around 1 g, which agrees with the value just calculated.

Practice Exercise Estimate the mass of air in a ping pong ball.

Review of Concepts

1.8.1 What is the correct answer with the appropriate number of significant figures and units for the following energy conversion calculation?

$$3.54 \times 10^{12}\text{ erg} \times \frac{1.000\text{ J}}{1.000 \times 10^7\text{ erg}} \times \frac{1.000\text{ cal}}{4.184\text{ J}} \times \frac{1\text{ kcal}}{1000\text{ cal}} =$$

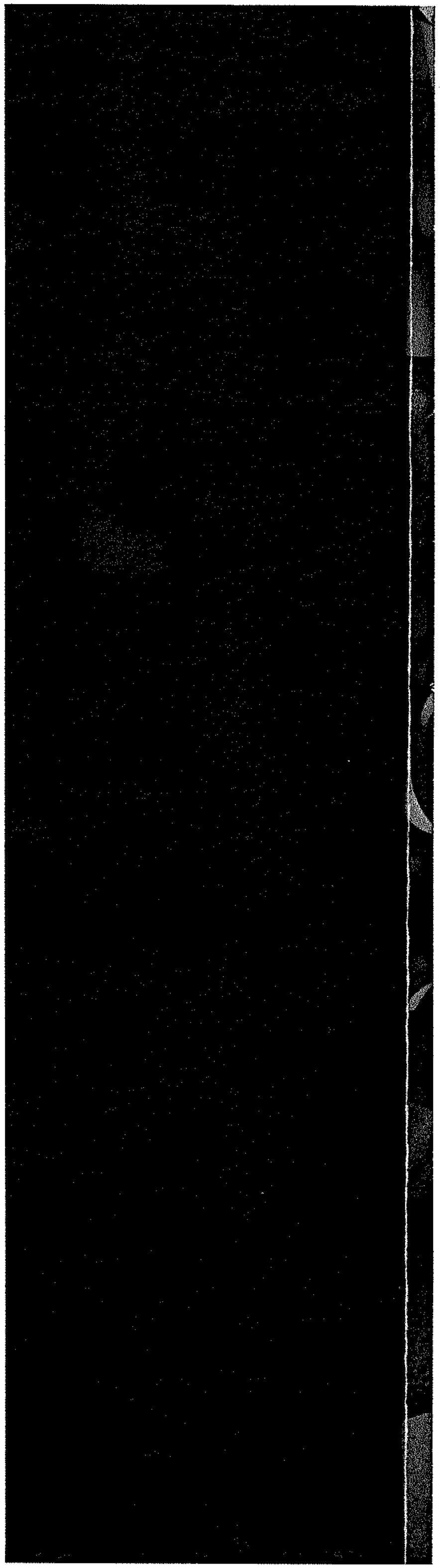
AP A Look Back at the Essential Knowledge

Chapter 1 covers the introduction to AP Chemistry: the six Big Ideas and the seven science practices. The Big Ideas are the central framework for AP Chemistry and provide a roadmap to understanding chemistry. They are supported by seven science practices—applications of practices, techniques and inquiry used by scientists of all disciplines. The science practices parallel the scientific method—the tools that scientists use to observe and understand our world. Chapter 1 also reviews information covered in previous science courses: measurement, handling numbers, and problem solving.

FOCUS REVIEW GUIDE
Complete the activities in Chapter 1 of your *Focus Review Guide* to review content essential for your AP exam.

Key Equations

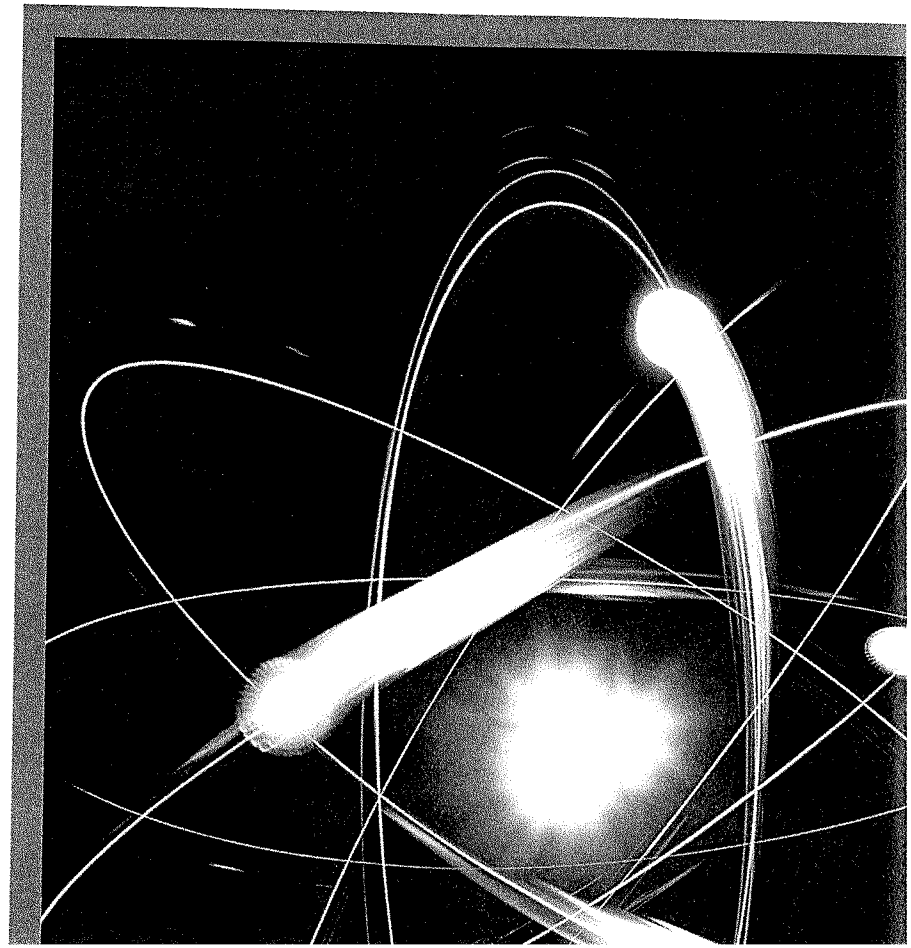
$d = \frac{m}{V}$ (1.1)	Equation for density
$?°\text{C} = (°\text{F} - 32°\text{F}) \times \frac{5°\text{C}}{9°\text{F}}$ (1.2)	Converting °F to °C
$?°\text{F} = \frac{9°\text{F}}{5°\text{C}} \times (°\text{C}) + 32°\text{F}$ (1.3)	Converting °C to °F
$?\text{ K} = (°\text{C} + 273.15°\text{C})\frac{1\text{ K}}{1°\text{C}}$ (1.4)	Converting °C to K



CHAPTER

2

Atoms, Molecules, and Ions



The nuclear model of the atom with a nucleus and orbiting electrons was devised from the work of Ernest Rutherford.

AP CHAPTER OUTLINE

AP BIG IDEAS A LOOK AHEAD

The behavior of the metallic element, sodium, is dramatic. It interacts vigorously when placed in water. Yet when a single electron is removed from an atom of sodium, it becomes a positively charged particle that is present in ordinary table salt. And it dissolves readily—*and unspectacularly*—in water. Its behavior to some degree is copied by the elements of lithium and potassium, found directly above and below sodium on the periodic table. Yet fluorine and calcium, which are one atomic number lower and higher than sodium, have very different chemical behaviors compared to sodium.

To account for chemical behavior we turn to the structure of the atom. In this respect, the number and placement of electrons is especially crucial. Although every element has a unique behavior, there are patterns among them.

The Periodic Table helps us see and use these patterns (EK.1.C.1). Finally, a system by which we name and symbolize particles exists so that sodium the atom (Na) will not be confused with sodium the ion (Na_+) or carbon monoxide (CO)—a lethal gas—is not mistaken for carbon dioxide (CO_2), the gas we exhale when breathing and swallow when drinking carbonated beverages. As you read the chapter, think about these Essential Questions:

1. How does an atom differ from a molecule? 1.A.1
2. What developments have changed our understanding of atoms and molecules since Dalton's proposed atomic theory? 1.D.1
3. What particles comprise an atom? 1.B.1
4. How do elements differ at the atomic level? Are all atoms of a given element identical? 1.D.2
5. What is the basic structure of the periodic table? 1.C.1
6. Can atoms form ions? How do ionic compounds maintain electrical neutrality? 2.C.2

Since ancient times humans have pondered the nature of matter. Our modern ideas of the structure of matter began to take shape in the early nineteenth century with Dalton's atomic theory. We now know that all matter is made of atoms, molecules, and ions. All of chemistry is concerned in one way or another with these species.

2.1 The Atomic Theory

In the fifth century B.C. the Greek philosopher Democritus expressed the belief that all matter consists of very small, indivisible particles, which he named *atomos* (meaning "uncuttable or indivisible"). Although Democritus' idea was not accepted by many of his contemporaries (notably Plato and Aristotle), somehow it endured. Experimental evidence from early scientific investigations provided support for the notion of "atomism" and gradually gave rise to the modern definitions of elements and compounds. In 1808 an English scientist and schoolteacher, John Dalton,† formulated a precise definition of the indivisible building blocks of matter that we call atoms.

Dalton's work marked the beginning of the modern era of chemistry. According to Dalton's atomic theory, the hypotheses about the nature of matter can be summarized as follows:

1. Elements are composed of extremely small particles called atoms.
2. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.
3. Compounds are composed of atoms of more than one element. For any given compound, the atoms present are always in the same ratio.
4. A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in the creation or destruction of atoms.

†John Dalton (1766–1844). English chemist, mathematician, and philosopher. In addition to the atomic theory, he also formulated several gas laws and gave the first detailed description of color blindness, from which he suffered. Dalton was described as an indifferent experimenter, and singularly wanting in the language and power of illustration. His only recreation was lawn bowling on Thursday afternoons. Perhaps it was the sight of those wooden balls that provided him with the idea of the atomic theory.

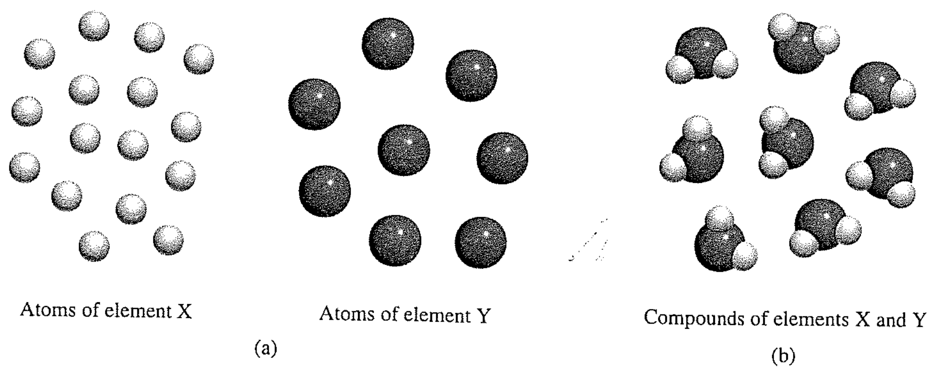


Figure 2.1 *(a) According to Dalton's atomic theory, atoms of the same element are identical, but atoms of one element are different from atoms of other elements. (b) Compound formed from atoms of elements X and Y. In this case, the ratio of the atoms of element X to the atoms of element Y is 2:1. Note that a chemical reaction results only in the rearrangement of atoms, not in their destruction or creation.*

Figure 2.1 is a schematic representation of the last three hypotheses.

Dalton's concept of an atom was far more detailed and specific than Democritus' concept. The second hypothesis states that atoms of one element are different from atoms of all other elements. Dalton made no attempt to describe the structure or composition of atoms—he had no idea what an atom is really like. But he did realize that the different properties shown by elements such as hydrogen and oxygen can be explained by assuming that hydrogen atoms are not the same as oxygen atoms.

Applying Practices

Our understanding of the world around us grows as we develop new technology. Go online to access the inquiry activity *Different Things Found in the Same Place* to explore how our understanding of the structure of the atom has changed.

The third hypothesis suggests that, to form a certain compound, we need not only atoms of the right kinds of elements, but specific numbers of these atoms as well. This idea is an extension of a law published in 1799 by Joseph Proust,† a French chemist. Proust's ***law of definite proportions*** states that *different samples of the same compound always contain its constituent elements in the same proportion by mass.* Thus, if we were to analyze samples of carbon dioxide gas obtained from different sources, we would find in each sample the same ratio by mass of carbon to oxygen. It stands to reason, then, that if the ratio of the masses of different elements in a given compound is fixed, the ratio of the atoms of these elements in the compound also must be constant.

Dalton's third hypothesis supports another important law, the ***law of multiple proportions.*** According to the law, *if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element are in ratios of small whole numbers.* Dalton's theory explains the law of multiple proportions quite simply: Different compounds made up of the same elements differ in the number of atoms of each kind that combine. For example, carbon forms two stable compounds with oxygen—namely, carbon monoxide and carbon dioxide. Modern measurement techniques indicate that one atom of carbon combines with one atom of oxygen in carbon monoxide and with two atoms of oxygen in carbon dioxide. Thus, the ratio of oxygen in carbon monoxide to oxygen in carbon dioxide is 1:2. This result is consistent with the law of multiple proportions (Figure 2.2).

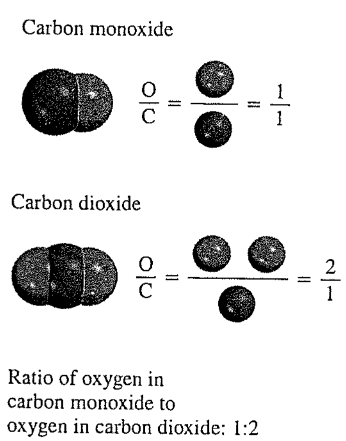


Figure 2.2 *An illustration of the law of multiple proportions.*

Dalton's fourth hypothesis is another way of stating the ***law of conservation of mass,***‡ which is that *matter can be neither created nor destroyed.* Because matter is made of atoms that are unchanged in a chemical reaction, it follows that mass must be conserved as well. Dalton's brilliant insight into the nature of matter was the main stimulus for the rapid progress of chemistry during the nineteenth century.

†Joseph Louis Proust (1754–1826). French chemist. Proust was the first person to isolate sugar from grapes.

Review of Concepts & Facts

2.1.1 The atoms of elements A (blue) and B (orange) form two compounds shown here. Do these compounds obey the law of multiple proportions?

2.2 The Structure of the Atom

On the basis of Dalton's atomic theory, we can define an ***atom*** as *the basic unit of an element that can enter into chemical combination.* Dalton imagined an atom that was both extremely small and indivisible. However, a series of investigations that began in the 1850s and extended into the twentieth century clearly demonstrated that atoms actually possess internal structure; that is, they are made up of even smaller particles, which are called *subatomic particles.* This research led to the discovery of three such particles—electrons, protons, and neutrons.

The Electron

In the 1890s, many scientists became caught up in the study of ***radiation,*** *the emission and transmission of energy through space in the form of waves.* Information gained from this research contributed greatly to our understanding of atomic structure. One device used to investigate this phenomenon was a cathode ray tube, the forerunner of the television tube (Figure 2.3). It is a glass tube from which most of the air has been

Video
Cathode Ray Tube

‡According to Albert Einstein, mass and energy are alternate aspects of a single entity called *mass-energy.* Chemical reactions usually involve a gain or loss of heat and other forms of energy. Thus, when energy is lost in a reaction, for example, mass is also lost. Except for nuclear reactions (see Chapter 19), however, changes of mass in chemical reactions are too small to detect. Therefore, for all practical purposes mass is conserved.

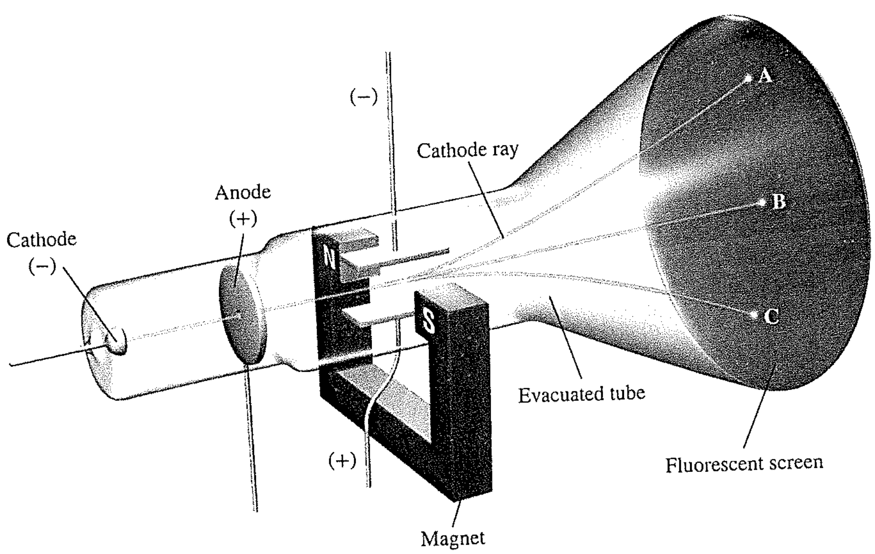
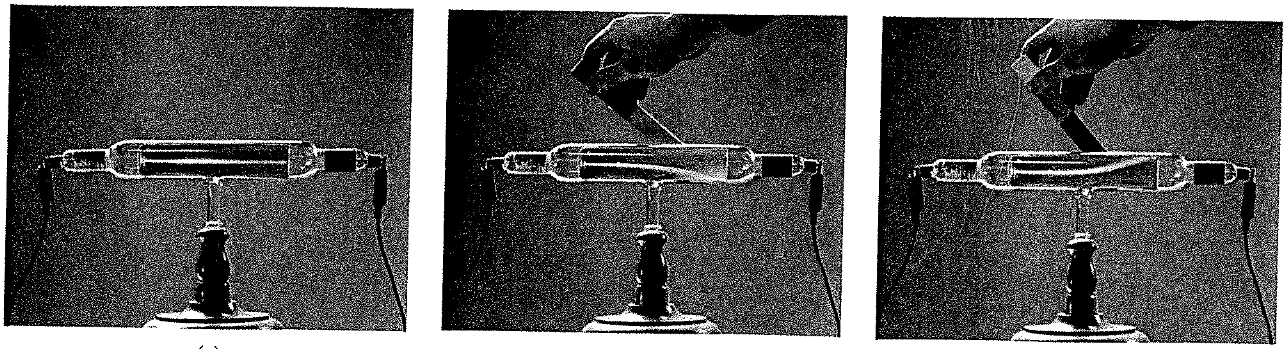


Figure 2.3 *A cathode ray tube with an electric field perpendicular to the direction of the cathode rays and an external magnetic field. The symbols N and S denote the north and south poles of the magnet. The cathode rays will strike the end of the tube at A in the presence of a magnetic field, at C in the presence of an electric field, and at B when there are no external fields present or when the effects of the electric field and magnetic field cancel each other.*



(a) (b) (c)

Figure 2.4 *(a) A cathode ray produced in a discharge tube traveling from the cathode (left) to the anode (right). The ray itself is invisible, but the fluorescence of a zinc sulfide coating on the glass causes it to appear green. (b) The cathode ray is bent downward when a bar magnet is brought toward it. (c) When the polarity of the magnet is reversed, the ray bends in the opposite direction.*

evacuated. When the two metal plates are connected to a high-voltage source, the negatively charged plate, called the *cathode,* emits an invisible ray. The cathode ray is drawn to the positively charged plate, called the *anode,* where it passes through a hole and continues traveling to the other end of the tube. When the ray strikes the specially coated surface, it produces a strong fluorescence, or bright light.

Electrons are normally associated with atoms. However, they can also be studied individually.

In some experiments, two electrically charged plates and a magnet were added to the *outside* of the cathode ray tube (see Figure 2.3). When the magnetic field is on and the electric field is off, the cathode ray strikes point A. When only the electric field is on, the ray strikes point C. When both the magnetic and the electric fields are off or when they are both on but balanced so that they cancel each other's influence, the ray strikes point B. According to electromagnetic theory, a moving charged body behaves like a magnet and can interact with electric and magnetic fields through which it passes. Because the cathode ray is attracted by the plate bearing positive charges and repelled by the plate bearing negative charges, it must consist of negatively charged particles. We know these *negatively charged particles* as ***electrons.*** Figure 2.4 shows the effect of a bar magnet on the cathode ray.

Video
Millikan Oil Drop

An English physicist, J. J. Thomson,† used a cathode ray tube and his knowledge of electromagnetic theory to determine the ratio of electric charge to the mass of an individual electron. The number he came up with was -1.76×10^{8} C/g, where C stands for *coulomb,* which is the unit of electric charge. Thereafter, in a series of experiments carried out between 1908 and 1917, R. A. Millikan‡ succeeded in measuring the charge of the electron with great precision. His work proved that the charge on each electron was exactly the same. In his experiment, Millikan examined the motion of single tiny drops of oil that picked up static charge from ions in the air. He suspended the charged drops in air by applying an electric field and followed their motions through a microscope (Figure 2.5). Using his knowledge of electrostatics, Millikan found the charge of an electron to be -1.6022×10^{-19} C. From these data he calculated the mass of an electron:

$$\begin{aligned}\text{mass of an electron} &= \frac{\text{charge}}{\text{charge/mass}}\\ &= \frac{-1.6022 \times 10^{-19}\ \text{C}}{-1.76 \times 10^{8}\ \text{C/g}}\\ &= 9.10 \times 10^{-28}\ \text{g}\end{aligned}$$

This is an exceedingly small mass.

†Joseph John Thomson (1856–1940). British physicist who received the Nobel Prize in Physics in 1906 for discovering the electron.

‡Robert Andrews Millikan (1868–1953). American physicist who was awarded the Nobel Prize in Physics in 1923 for determining the charge of the electron.

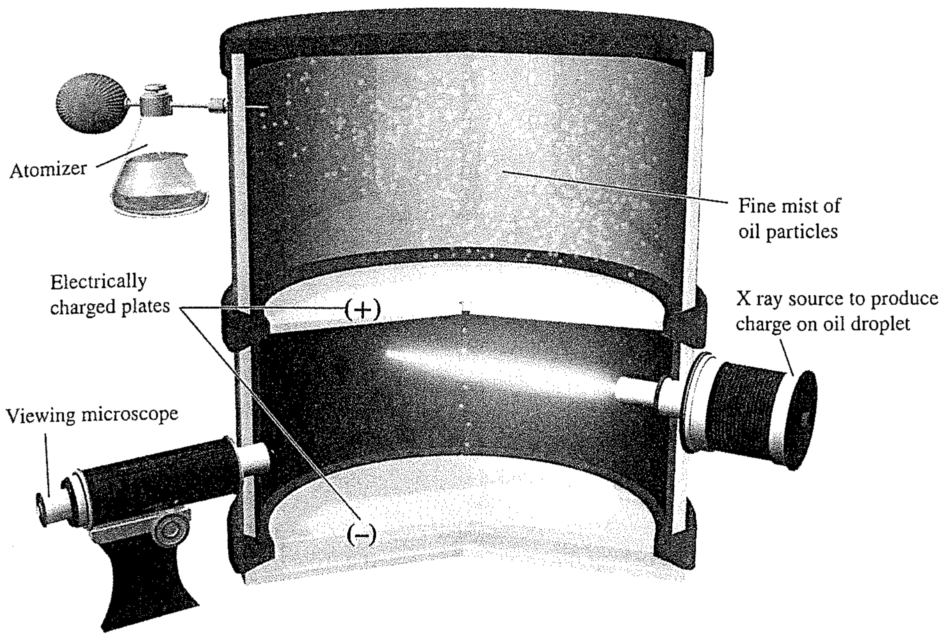


Figure 2.5 *Schematic diagram of Millikan's oil drop experiment.*

Radioactivity

In 1895 the German physicist Wilhelm Röntgen[§] noticed that cathode rays caused glass and metals to emit very unusual rays. This highly energetic radiation penetrated matter, darkened covered photographic plates, and caused a variety of substances to fluoresce. Because these rays could not be deflected by a magnet, they could not contain charged particles as cathode rays do. Röntgen called them X rays because their nature was not known.

Not long after Röntgen's discovery, Antoine Becquerel,[†] a professor of physics in Paris, began to study the fluorescent properties of substances. Purely by accident, he found that exposing thickly wrapped photographic plates to a certain uranium compound caused them to darken, even without the stimulation of cathode rays. Like X rays, the rays from the uranium compound were highly energetic and could not be deflected by a magnet, but they differed from X rays because they arose spontaneously. One of Becquerel's students, Marie Curie,[‡] suggested the name ***radioactivity*** to describe this *spontaneous emission of particles and/or radiation.* Since then, any element that spontaneously emits radiation is said to be *radioactive.*

Video
Alpha, Beta, and Gamma Rays

Three types of rays are produced by the *decay,* or breakdown, of radioactive substances such as uranium. Two of the three are deflected by oppositely charged metal plates (Figure 2.6). ***Alpha*** (α) ***rays*** consist of *positively charged particles,* called α ***particles,*** and therefore are deflected by the positively charged plate. ***Beta*** (β) ***rays,*** or β ***particles,*** are electrons and are deflected by the negatively charged plate. The third type of radioactive radiation consists of high-energy rays called ***gamma*** (γ) ***rays.*** Like X rays, γ rays have no charge and are not affected by an external field.

[§]Wilhelm Konrad Röntgen (1845–1923). German physicist who received the Nobel Prize in Physics in 1901 for the discovery of X rays.

[†]Antoine Henri Becquerel (1852–1908). French physicist who was awarded the Nobel Prize in Physics in 1903 for discovering radioactivity in uranium.

[‡]Marie (Marya Sklodowska) Curie (1867–1934). Polish-born chemist and physicist. In 1903 she and her French husband, Pierre Curie, were awarded the Nobel Prize in Physics for their work on radioactivity. In 1911, she again received the Nobel prize, this time in chemistry, for her work on the radioactive elements radium and polonium. She is one of only three people to have received two Nobel prizes in science. Despite her great contribution to science, her nomination to the French Academy of Sciences in 1911 was rejected by one vote because she was a woman! Her daughter Irene, and son-in-law Frederic Joliot-Curie, shared the Nobel Prize in Chemistry in 1935.

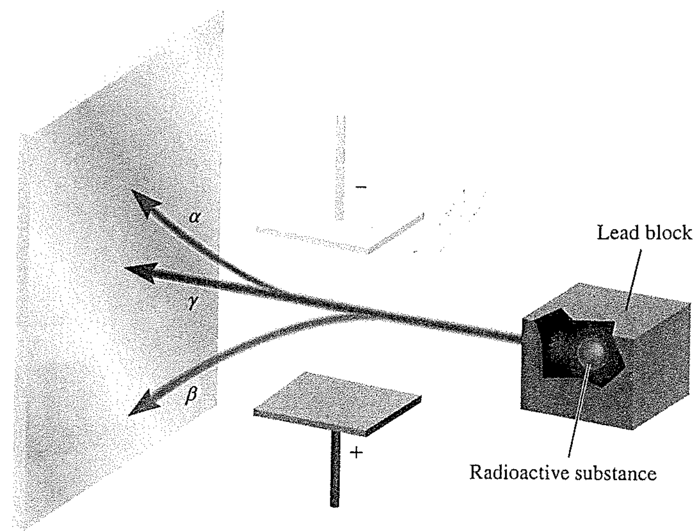


Figure 2.6 *Three types of rays emitted by radioactive elements. β rays consist of negatively charged particles (electrons) and are therefore attracted by the positively charged plate. The opposite holds true for α rays—they are positively charged and are drawn to the negatively charged plate. Because γ rays have no charges, their path is unaffected by an external electric field.*

The Proton and the Nucleus

By the early 1900s, two features of atoms had become clear: They contain electrons, and they are electrically neutral. To maintain electric neutrality, an atom must contain an equal number of positive and negative charges. Therefore, Thomson proposed that an atom could be thought of as a uniform, positive sphere of matter in which electrons are embedded like raisins in a cake (Figure 2.7). This so-called "plum-pudding" model was the accepted theory for a number of years.

Video
α-Particle Scattering

Video
Rutherford's Experiment

In 1910 the New Zealand physicist Ernest Rutherford,† who had studied with Thomson at Cambridge University, decided to use α particles to probe the structure of atoms. Together with his associate Hans Geiger‡ and an undergraduate named Ernest Marsden,§ Rutherford carried out a series of experiments using very thin foils of gold and other metals as targets for α particles from a radioactive source (Figure 2.8). They observed that the majority of particles penetrated the foil either undeflected or with only a slight deflection. But every now and then an α particle was scattered (or deflected) at a large angle. In some instances, an α particle actually bounced back in the direction from which it had come! This was a most surprising finding, for in Thomson's model the positive charge of the atom was so diffuse that the positive α particles should have passed through the foil with very little deflection. To quote Rutherford's initial reaction when told of this discovery: "It was as incredible as if you had fired a 15-inch shell at a piece of tissue paper and it came back and hit you."

Rutherford explained the results of the α-scattering experiment by introducing a new model for the atom. According to Rutherford, most of the atom must be empty space. This explains why the majority of α particles passed through the gold foil with little or no deflection. The atom's positive charges, Rutherford proposed, are all concentrated in the ***nucleus,*** which is *a dense central core within the atom.* Whenever an α particle came close to a nucleus in the scattering experiment, it experienced a large repulsive force and therefore a large deflection. Moreover, an α particle traveling directly toward a nucleus would be completely repelled and its direction would be reversed.

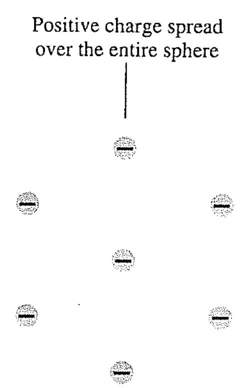


Figure 2.7 *Thomson's model of the atom, sometimes described as the "plum-pudding" model, after a traditional English dessert containing raisins. The electrons are embedded in a uniform, positively charged sphere.*

†Ernest Rutherford (1871–1937). New Zealand physicist. Rutherford did most of his work in England (Manchester and Cambridge Universities). He received the Nobel Prize in Chemistry in 1908 for his investigations into the structure of the atomic nucleus. His often-quoted comment to his students was that "all science is either physics or stamp-collecting."

‡Johannes Hans Wilhelm Geiger (1882–1945). German physicist. Geiger's work focused on the structure of the atomic nucleus and on radioactivity. He invented a device for measuring radiation that is now commonly called the Geiger counter.

§Ernest Marsden (1889–1970). English physicist. It is gratifying to know that at times an undergraduate can assist in winning a Nobel prize. Marsden went on to contribute significantly to the development of science in New Zealand.

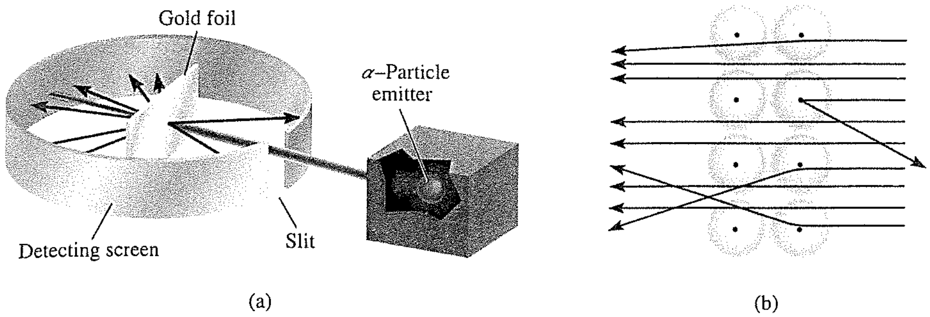


Figure 2.8 *(a) Rutherford's experimental design for measuring the scattering of α particles by a piece of gold foil. Most of the α particles passed through the gold foil with little or no deflection. A few were deflected at wide angles. Occasionally an α particle was turned back. (b) Magnified view of α particles passing through and being deflected by nuclei.*

The positively charged particles in the nucleus are called ***protons***. In separate experiments, it was found that each proton carries the same *quantity* of charge as an electron and has a mass of 1.67262×10^{-24} g—about 1840 times the mass of the oppositely charged electron.

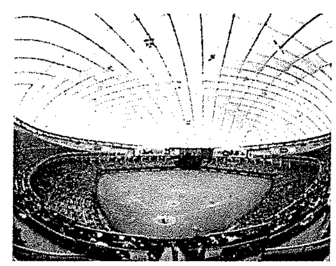
At this stage of investigation, scientists perceived the atom as follows: The mass of a nucleus constitutes most of the mass of the entire atom, but the nucleus occupies only about $1/10^{13}$ of the volume of the atom. We express atomic (and molecular) dimensions in terms of the SI unit called the *picometer (pm),* where

A common non-SI unit for atomic length is the angstrom (Å; 1 Å = 100 pm).

$$1 \text{ pm} = 1 \times 10^{-12} \text{ m}$$

A typical atomic radius is about 100 pm, whereas the radius of an atomic nucleus is only about 5×10^{-3} pm. You can appreciate the relative sizes of an atom and its nucleus by imagining that if an atom were the size of a sports stadium, the volume of its nucleus would be comparable to that of a small marble. Although the protons are confined to the nucleus of the atom, the electrons are conceived of as being spread out about the nucleus at some distance from it.

The concept of atomic radius is useful experimentally, but we should not infer that atoms have well-defined boundaries or surfaces. We will learn later that the outer regions of atoms are relatively "fuzzy."



If the size of an atom were expanded to that of this sports stadium, the size of the nucleus would be that of a marble.

The Neutron

Rutherford's model of atomic structure left one major problem unsolved. It was known that hydrogen, the simplest atom, contains only one proton and that the helium atom contains two protons. Therefore, the ratio of the mass of a helium atom to that of a hydrogen atom should be 2:1. (Because electrons are much lighter than protons, their contribution to atomic mass can be ignored.) In reality, however, the ratio is 4:1. Rutherford and others postulated that there must be another type of subatomic particle in the atomic nucleus; the proof was provided by another English physicist, James Chadwick,† in 1932. When Chadwick bombarded a thin sheet of beryllium with α particles, a very high-energy radiation similar to γ rays was emitted by the metal. Later experiments showed that the rays actually consisted of a third type of subatomic particles. Chadwick named these subatomic particles ***neutrons,*** because they proved to be *electrically neutral particles with a mass slightly greater than that of protons.* The mystery of the mass ratio could now be explained. In the helium nucleus there are two protons and two neutrons, but in the hydrogen nucleus there is only one proton and no neutrons; therefore, the ratio is 4:1.

Applying Practices

Conduct your own simulation of Rutherford's experiment by accessing the lab *Simulation of Rutherford's Gold Foil Experiment* available online.

Figure 2.9 shows the location of the elementary particles (protons, neutrons, and electrons) in an atom. There are other subatomic particles, but the electron, the proton, and the neutron are the three fundamental components of the atom that are important in chemistry. Table 2.1 shows the masses and charges of these three elementary particles.

†James Chadwick (1891–1972). British physicist. In 1935 he received the Nobel Prize in Physics for proving the existence of neutrons.

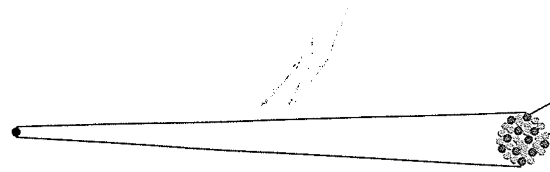


Figure 2.9 *The protons and neutrons of an atom are packed in an extremely small nucleus. Electrons are shown as "clouds" around the nucleus.*

Table 2.1 Mass and Charge of Subatomic Particles

Particle	Mass (g)	Charge: Coulomb	Charge: Charge Unit
Electron*	9.10938×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	$+1$
Neutron	1.67493×10^{-24}	0	0

*More refined measurements have given us a more accurate value of an electron's mass than Millikan's.

2.3 Atomic Number, Mass Number, and Isotopes

All atoms can be identified by the number of protons and neutrons they contain. The ***atomic number (Z)*** is *the number of protons in the nucleus of each atom of an element.* In a neutral atom the number of protons is equal to the number of electrons, so the atomic number also indicates the number of electrons present in the atom. The chemical identity of an atom can be determined solely from its atomic number. For example, the atomic number of fluorine is 9. This means that each fluorine atom has 9 protons and 9 electrons. Or, viewed another way, every atom in the universe that contains 9 protons is correctly named "fluorine."

Protons and neutrons are collectively called *nucleons*.

The ***mass number (A)*** is *the total number of neutrons and protons present in the nucleus of an atom of an element.* Except for the most common form of hydrogen, which has one proton and no neutrons, all atomic nuclei contain both protons and neutrons. In general, the mass number is given by

$$\begin{aligned}\text{mass number} &= \text{number of protons} + \text{number of neutrons} \\ &= \text{atomic number} + \text{number of neutrons}\end{aligned} \tag{2.1}$$

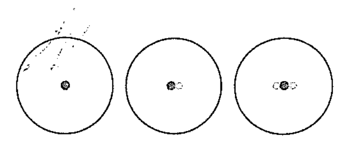
The number of neutrons in an atom is equal to the difference between the mass number and the atomic number, or $(A - Z)$. For example, if the mass number of a particular boron atom is 12 and the atomic number is 5 (indicating 5 protons in the nucleus), then the number of neutrons is $12 - 5 = 7$. Note that all three quantities (atomic number, number of neutrons, and mass number) must be positive integers, or whole numbers.

Atoms of a given element do not all have the same mass despite what the first hypothesis of Dalton's atomic theory states. Most elements have two or more ***isotopes,*** *atoms that have the same atomic number (Z) but different mass numbers (A).* For example, there are three isotopes of hydrogen. One, simply known as hydrogen, has one proton and no neutrons. The *deuterium* isotope contains one proton and one neutron, and *tritium* has one proton and two neutrons. The accepted way to denote the atomic number and mass number of an atom of an element (X) is as follows:

$$\text{mass number} \rightarrow {}^{A}_{Z}\text{X} \leftarrow \text{atomic number}$$

Thus, for the isotopes of hydrogen, we write

$$\underset{\text{hydrogen}}{{}^{1}_{1}\text{H}} \quad \underset{\text{deuterium}}{{}^{2}_{1}\text{H}} \quad \underset{\text{tritium}}{{}^{3}_{1}\text{H}}$$



As another example, consider two common isotopes of uranium with mass numbers of 235 and 238, respectively:

$${}^{235}_{92}\text{U} \quad {}^{238}_{92}\text{U}$$

The first isotope is used in nuclear reactors and atomic bombs, whereas the second isotope lacks the properties necessary for these applications. With the exception of hydrogen, which has different names for each of its isotopes, isotopes of elements are identified by their mass numbers. Thus, the preceding two isotopes are called uranium-235 (pronounced "uranium two thirty-five") and uranium-238 (pronounced "uranium two thirty-eight").

The chemical properties of an element are determined primarily by the protons and electrons in its atoms; neutrons do not take part in chemical changes under normal conditions. Therefore, isotopes of the same element have similar chemistries, forming the same types of compounds and displaying similar reactivities.

Example 2.1 shows how to calculate the number of protons, neutrons, and electrons using atomic numbers and mass numbers.

Example 2.1

Give the number of protons, neutrons, and electrons in each of the following species: (a) $^{20}_{11}$Na, (b) $^{22}_{11}$Na, (c) ^{17}O, and (d) carbon-14.

Strategy Recall that the superscript denotes the mass number (A) and the subscript denotes the atomic number (Z). Mass number is always greater than atomic number. (The only exception is $^{1}_{1}$H, where the mass number is equal to the atomic number.) In a case where no subscript is shown, as in parts (c) and (d), the atomic number can be deduced from the element symbol or name. To determine the number of electrons, remember that because atoms are electrically neutral, the number of electrons is equal to the number of protons.

Solution

(a) The atomic number is 11, so there are 11 protons. The mass number is 20, so the number of neutrons is $20 - 11 = 9$. The number of electrons is the same as the number of protons; that is, 11.

(b) The atomic number is the same as that in (a), or 11. The mass number is 22, so the number of neutrons is $22 - 11 = 11$. The number of electrons is 11. Note that the species in (a) and (b) are chemically similar isotopes of sodium.

(c) The atomic number of O (oxygen) is 8, so there are 8 protons. The mass number is 17, so there are $17 - 8 = 9$ neutrons. There are 8 electrons.

(d) Carbon-14 can also be represented as ^{14}C. The atomic number of carbon is 6, so there are $14 - 6 = 8$ neutrons. The number of electrons is 6.

Practice Exercise

How many protons, neutrons, and electrons are in the following isotope of copper: ^{63}Cu?

Similar problems: 2.15, 2.16.

Review of Concepts & Facts

2.3.1 What is the atomic number of an element if one of its isotopes has 117 neutrons and a mass number of 195?

2.3.2 How many neutrons are in an atom of ^{114}Cd?

2.3.3 Which of the following two symbols provides more information: ^{17}O or $_8O$?

2.4 The Periodic Table

More than half of the elements known today were discovered between 1800 and 1900. During this period, chemists noted that many elements show strong similarities to one another. Recognition of periodic regularities in physical and chemical behavior and the need to organize the large volume of available information about the structure and properties of elemental substances led to the development of the ***periodic table,*** *a chart in which elements having similar chemical and physical properties are grouped together.* Figure 2.10 shows the modern periodic table in which the elements are arranged by atomic number (Z) (shown above the element symbol) in *horizontal rows* called ***periods*** and in *vertical columns* known as ***groups*** or ***families,*** according to similarities in their chemical properties.

The elements can be divided into three categories—metals, nonmetals, and metalloids. A ***metal*** is *a good conductor of heat and electricity,* whereas a ***nonmetal*** is usually *a poor conductor of heat and electricity.* A **metalloid** *has properties that are intermediate between those of metals and nonmetals.* Figure 2.10 shows that the majority of

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1 **H**																	2 **He**
3 **Li**	4 **Be**											5 **B**	6 **C**	7 **N**	8 **O**	9 **F**	10 **Ne**
11 **Na**	12 **Mg**											13 **Al**	14 **Si**	15 **P**	16 **S**	17 **Cl**	18 **Ar**
19 **K**	20 **Ca**	21 **Sc**	22 **Ti**	23 **V**	24 **Cr**	25 **Mn**	26 **Fe**	27 **Co**	28 **Ni**	29 **Cu**	30 **Zn**	31 **Ga**	32 **Ge**	33 **As**	34 **Se**	35 **Br**	36 **Kr**
37 **Rb**	38 **Sr**	39 **Y**	40 **Zr**	41 **Nb**	42 **Mo**	43 **Tc**	44 **Ru**	45 **Rh**	46 **Pd**	47 **Ag**	48 **Cd**	49 **In**	50 **Sn**	51 **Sb**	52 **Te**	53 **I**	54 **Xe**
55 **Cs**	56 **Ba**	57 **La**	72 **Hf**	73 **Ta**	74 **W**	75 **Re**	76 **Os**	77 **Ir**	78 **Pt**	79 **Au**	80 **Hg**	81 **Tl**	82 **Pb**	83 **Bi**	84 **Po**	85 **At**	86 **Rn**
87 **Fr**	88 **Ra**	89 **Ac**	104 **Rf**	105 **Db**	106 **Sg**	107 **Bh**	108 **Hs**	109 **Mt**	110 **Ds**	111 **Rg**	112 **Cn**	113 **Nh**	114 **Fl**	115 **Mc**	116 **Lv**	117 **Ts**	118 **Og**

58 **Ce**	59 **Pr**	60 **Nd**	61 **Pm**	62 **Sm**	63 **Eu**	64 **Gd**	65 **Tb**	66 **Dy**	67 **Ho**	68 **Er**	69 **Tm**	70 **Yb**	71 **Lu**
90 **Th**	91 **Pa**	92 **U**	93 **Np**	94 **Pu**	95 **Am**	96 **Cm**	97 **Bk**	98 **Cf**	99 **Es**	100 **Fm**	101 **Md**	102 **No**	103 **Lr**

Metals

Metalloids

Nonmetals

Figure 2.10 *The modern periodic table. The elements are arranged according to the atomic numbers above their symbols. With the exception of hydrogen (H), nonmetals appear at the far right of the table. The two rows of metals beneath the main body of the table are conventionally set apart to keep the table from being too wide. Actually, cerium (Ce) should follow lanthanum (La), and thorium (Th) should come right after actinium (Ac). The 1–18 group designation has been recommended by the International Union of Pure and Applied Chemistry (IUPAC) but is not yet in wide use. In this text, we use the standard U.S. notation for group numbers (1A–8A and 1B–8B).*

CHEMISTRY *in Action*

Distribution of Elements on Earth and in Living Systems

The majority of elements are naturally occurring. How are these elements distributed on Earth, and which are essential to living systems?

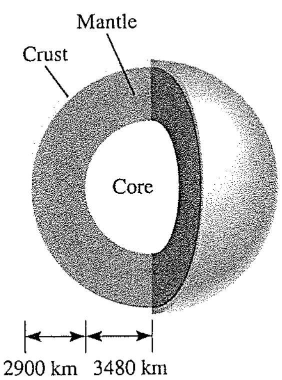
Earth's crust extends from the surface to a depth of about 40 km (about 25 mi). There is a solid core consisting mostly of iron at the center of Earth. Surrounding the core is a layer called the *mantle,* which consists of hot fluid containing iron, carbon, silicon, and sulfur.

Of the 83 elements that are found in nature, 12 make up 99.7 percent of Earth's crust by mass. They are, in decreasing order of natural abundance, oxygen (O), silicon (Si), aluminum (Al), iron (Fe), calcium (Ca), magnesium (Mg), sodium (Na), potassium (K), titanium (Ti), hydrogen (H), phosphorus (P), and manganese (Mn). In discussing the natural abundance of the elements, we should keep in mind that (1) the elements are not evenly distributed throughout Earth's crust, and (2) most elements occur in combined forms.

The accompanying table lists the essential elements in the human body. Of special interest are the *trace elements,* such as iron (Fe), copper (Cu), zinc (Zn), iodine (I), and cobalt (Co), which together make up about 0.1 percent of the body's mass. These elements are necessary for biological functions such as growth, transport of oxygen for metabolism, and defense against disease. There is a delicate balance in the amounts of these elements in the human body. Too much or too little over an extended period of time can lead to serious illness, retardation, or even death.

Thinking Critically

1. Describe the delicate balance maintained by trace elements in the human body.
2. List the most abundant elements found in the *mantle* layer of the Earth
3. Given a homogenous, 200-gram sample from the Earth's crust, approximately how much calcium (in grams) would you expect to find?

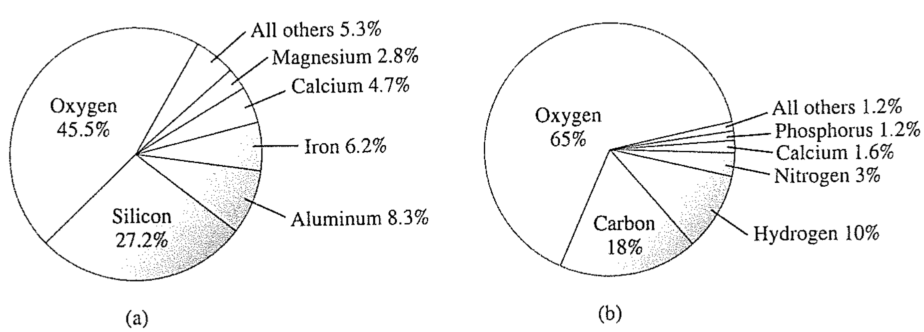


Structure of Earth's interior.

Essential Elements in the Human Body

Element	Percent by Mass*	Element	Percent by Mass*
Oxygen	65	Sodium	0.1
Carbon	18	Magnesium	0.05
Hydrogen	10	Iron	<0.05
Nitrogen	3	Cobalt	<0.05
Calcium	1.6	Copper	<0.05
Phosphorus	1.2	Zinc	<0.05
Potassium	0.2	Iodine	<0.05
Sulfur	0.2	Selenium	<0.01
Chlorine	0.2	Fluorine	<0.01

*Percent by mass *gives the mass of the element in grams present in a 100-g sample.*



(a) Natural abundance of the elements in percent by mass. For example, oxygen's abundance is 45.5 percent. This means that in a 100-g sample of Earth's crust there are, on the average, 45.5 g of the element oxygen. (b) Abundance of elements in the human body in percent by mass.

known elements are metals; only 17 elements are nonmetals, and 8 elements are metalloids. From left to right across any period, the physical and chemical properties of the elements change gradually from metallic to nonmetallic.

Elements are often referred to collectively by their periodic table group number (Group 1A, Group 2A, and so on). However, for convenience, some element groups have been given special names. *The Group 1A elements (Li, Na, K, Rb, Cs, and Fr)* are called ***alkali metals,*** and *the Group 2A elements (Be, Mg, Ca, Sr, Ba, and Ra)* are called ***alkaline earth metals.*** *Elements in Group 7A (F, Cl, Br, I, and At)* are known as ***halogens,*** and *elements in Group 8A (He, Ne, Ar, Kr, Xe, and Rn)* are called ***noble gases,*** or *rare gases.*

The periodic table is a handy tool that correlates the properties of the elements in a systematic way and helps us to make predictions about chemical behavior. We will take a closer look at this keystone of chemistry in Chapter 8.

The Chemistry in Action essay "Distribution of Elements on Earth and in Living Systems" describes the distribution of the elements on Earth and in the human body.

Review of Concepts & Facts

2.4.1 In viewing the periodic table, do chemical properties change more markedly across a period or down a group?

2.4.2 Identify the following as a metal, metalloid, or nonmetal: (a) K, (b) Se, (c) Sb, (d) W.

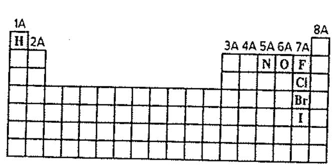
2.5 Molecules and Ions

Of all the elements, only the six noble gases in Group 8A of the periodic table (He, Ne, Ar, Kr, Xe, and Rn) exist in nature as single atoms. For this reason, they are called *monatomic* (meaning "one atom") gases. Most matter is composed of molecules or ions formed by atoms.

Molecules

We will discuss the nature of chemical bonds in Chapters 9 and 10.

A ***molecule*** is an *aggregate of at least two atoms in a definite arrangement held together by chemical forces* (also called *chemical bonds*). A molecule may contain atoms of the same element or atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions stated in Section 2.1. Thus, a molecule is not necessarily a compound, which, by definition, is made up of two or more elements (see Section 1.4). Hydrogen gas, for example, is a pure element, but it consists of molecules made up of two H atoms each. Water, on the other hand, is a molecular compound that contains hydrogen and oxygen in a ratio of two H atoms and one O atom. Like atoms, molecules are electrically neutral.



Elements that exist as diatomic molecules.

The hydrogen molecule, symbolized as H_2, is called a ***diatomic molecule,*** because it *contains only two atoms.* Other elements that normally exist as diatomic molecules are nitrogen (N_2) and oxygen (O_2), as well as the Group 7A elements—fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). Of course, a diatomic molecule can contain atoms of different elements. Examples are hydrogen chloride (HCl) and carbon monoxide (CO).

The vast majority of molecules contain more than two atoms. They can be atoms of the same element, as in ozone (O_3), which is made up of three atoms of oxygen, or they can be combinations of two or more different elements. *Molecules containing more than two atoms* are called ***polyatomic molecules.*** Like ozone, water (H_2O) and ammonia (NH_3) are polyatomic molecules.

Ions

An ***ion*** is *an atom or a group of atoms that has a net positive or negative charge.* The number of positively charged protons in the nucleus of an atom remains the same during ordinary chemical changes (called chemical reactions), but negatively charged electrons may be lost or gained. The loss of one or more electrons from a neutral atom results in a ***cation,*** *an ion with a net positive charge.* For example, a sodium atom (Na) can readily lose an electron to become a sodium cation, which is represented by Na^+:

In Chapter 8 we will see why atoms of different elements gain (or lose) a specific number of electrons.

Na Atom	Na^+ Ion
11 protons	11 protons
11 electrons	10 electrons

On the other hand, an ***anion*** is *an ion whose net charge is negative* due to an increase in the number of electrons. A chlorine atom (Cl), for instance, can gain an electron to become the chloride ion Cl^-:

Cl Atom	Cl^- Ion
17 protons	17 protons
17 electrons	18 electrons

Sodium chloride (NaCl), ordinary table salt, is called an ***ionic compound,*** because it is *formed from cations and anions.*

An atom can lose or gain more than one electron. Examples of ions formed by the loss or gain of more than one electron are Mg^{2+}, Fe^{3+}, S^{2-}, and N^{3-}. These ions, as well as Na^+ and Cl^-, are called ***monatomic ions,*** because they *contain only one atom.* Figure 2.11 shows the charges of a number of monatomic ions. With very few exceptions, metals tend to form cations and nonmetals form anions.

In addition, two or more atoms can combine to form an ion that has a net positive or net negative charge. ***Polyatomic ions*** such as OH^- (hydroxide ion), CN^- (cyanide ion), and NH_4^+ (ammonium ion) are *ions containing more than one atom.*

Review of Concepts & Facts

2.5.1 What does S_8 signify? How does it differ from 8S?

2.5.2 Determine the number of protons and electrons for the following ions: (a) P^{3-}, (b) Ti^{4+}, (c) Ge^{2+}.

2.5.3 What is the symbol and charge for an ion containing 24 protons and 22 electrons?

1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
Li^+													C^{4-}	N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}				Cr^{2+} Cr^{3+}	Mn^{2+} Mn^{3+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+} Ni^{3+}	Cu^+ Cu^{2+}	Zn^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}									Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}		Te^{2-}	I^-	
Cs^+	Ba^{2+}									Au^+ Au^{3+}	Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}				

Figure 2.11 *Common monatomic ions arranged according to their positions in the periodic table. Note that the* Hg_2^{2+} *ion contains two atoms.*

2.6 Chemical Formulas

Chemists use ***chemical formulas*** to *express the composition of molecules and ionic compounds in terms of chemical symbols.* By composition we mean not only the elements present but also the ratios in which the atoms are combined. Here we are concerned with two types of formulas: molecular formulas and empirical formulas.

Molecular Formulas

A ***molecular formula*** *shows the exact number of atoms of each element in the smallest unit of a substance.* In our discussion of molecules, each example was given with its molecular formula in parentheses. Thus, H_2 is the molecular formula for hydrogen, O_2 is oxygen, O_3 is ozone, and H_2O is water. The subscript numeral indicates the number of atoms of an element present. There is no subscript for O in H_2O because there is only one atom of oxygen in a molecule of water, and so the number "one" is omitted from the formula. Note that oxygen (O_2) and ozone (O_3) are allotropes of oxygen. An ***allotrope*** is *one of two or more distinct forms of an element.* Two allotropic forms of the element carbon—diamond and graphite—are dramatically different not only in properties but also in their relative cost.

Molecular Models

See back endpaper for color codes for atoms.

Molecules are too small for us to observe directly. An effective means of visualizing them is by the use of molecular models. Two standard types of molecular models are currently in use: *ball-and-stick* models and *space-filling* models (Figure 2.12). In ball-and-stick model kits, the atoms are wooden or plastic balls with holes in them. Sticks or springs are used to represent chemical bonds. The angles they form between atoms approximate the bond angles in actual molecules. With the exception of the H atom, the balls are all the same size and each type of atom is represented by a specific color. In space-filling models, atoms are represented by truncated balls held together by snap fasteners, so that the bonds are not visible. The balls are proportional in size to atoms.

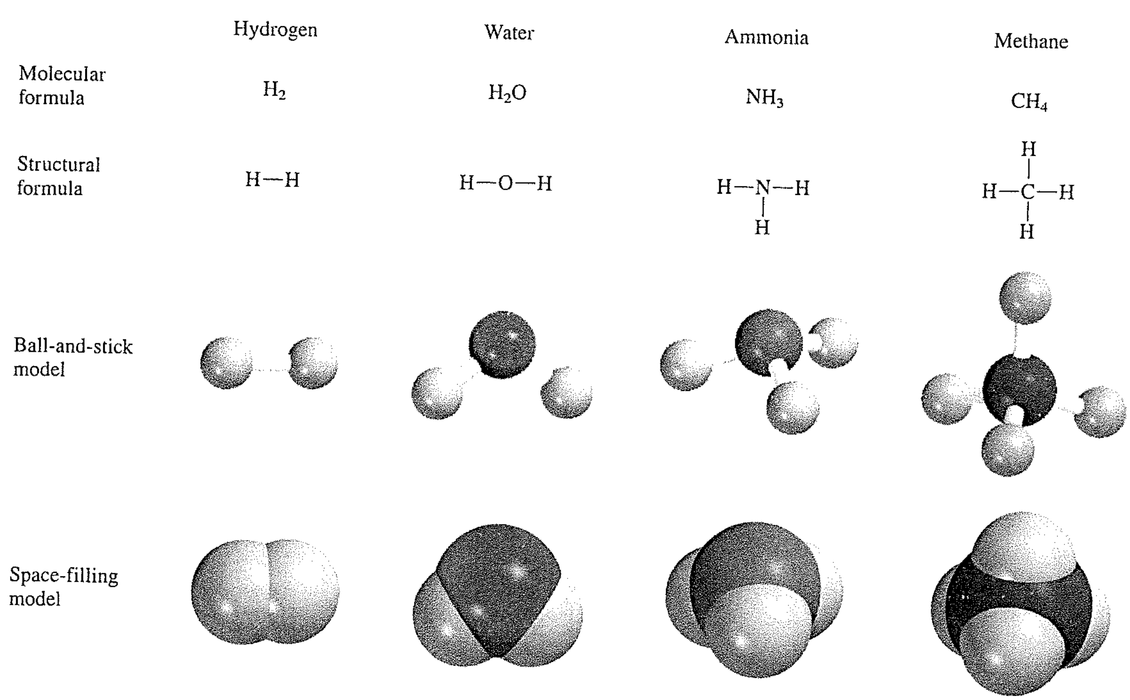


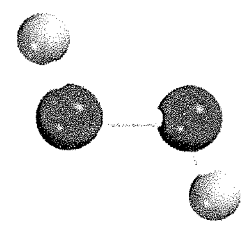
Figure 2.12 *Molecular and structural formulas and molecular models of four common molecules.*

The first step toward building a molecular model is writing the ***structural formula,*** which *shows how atoms are bonded to one another in a molecule.* For example, it is known that each of the two H atoms is bonded to an O atom in the water molecule. Therefore, the structural formula of water is H—O—H. A line connecting the two atomic symbols represents a chemical bond.

Ball-and-stick models show the three-dimensional arrangement of atoms clearly, and they are fairly easy to construct. However, the balls are not proportional to the size of atoms. Furthermore, the sticks greatly exaggerate the space between atoms in a molecule. Space-filling models are more accurate because they show the variation in atomic size. Their drawbacks are that they are time consuming to put together and they do not show the three-dimensional positions of atoms very well. Molecular modeling software can also be used to create ball-and-stick and space-filling models. We will use both models extensively in this text.

Empirical Formulas

The molecular formula of hydrogen peroxide, a substance used as an antiseptic and as a bleaching agent for textiles and hair, is H_2O_2. This formula indicates that each hydrogen peroxide molecule consists of two hydrogen atoms and two oxygen atoms. The ratio of hydrogen to oxygen atoms in this molecule is 2:2 or 1:1. The empirical formula of hydrogen peroxide is HO. Thus, the ***empirical formula*** *tells us which elements are present and the simplest whole-number ratio of their atoms.* However, empirical formulas do not necessarily tell us the actual number of atoms in a given molecule. As another example, consider the compound hydrazine (N_2H_4), which is used as a rocket fuel. The empirical formula of hydrazine is NH_2. Although the ratio of nitrogen to hydrogen is 1:2 in both the molecular formula (N_2H_4) and the empirical formula (NH_2), only the molecular formula tells us the actual number of N atoms (two) and H atoms (four) present in a hydrazine molecule.



H_2O_2

Empirical formulas are the *simplest* chemical formulas; they are written by reducing the subscripts in the molecular formulas to the smallest possible whole numbers. Molecular formulas are the *true* formulas of molecules. If we know the molecular formula, we also know the empirical formula, but the reverse is not true. Why, then, do chemists bother with empirical formulas? As we will see in Chapter 3, when chemists analyze an unknown compound, the first step is usually the determination of the compound's empirical formula. With additional information, it is possible to deduce the molecular formula.

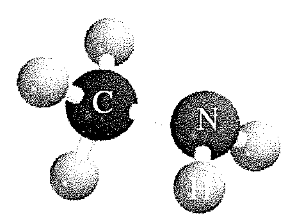
The word *empirical* means "derived from experiment." As we will see in Chapter 3, empirical formulas are determined experimentally.

For many molecules, the molecular formula and the empirical formula are one and the same. Some examples are water (H_2O), ammonia (NH_3), carbon dioxide (CO_2), and methane (CH_4).

Examples 2.2 and 2.3 deal with writing molecular formulas from molecular models and writing empirical formulas from molecular formulas.

Example 2.2

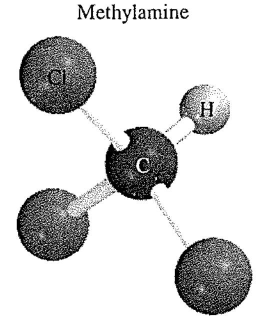
Write the molecular formula of methylamine, a colorless gas used in the production of pharmaceuticals and pesticides, from its ball-and-stick model, shown here.



Methylamine

Solution Refer to the labels (also see end papers). There are five H atoms, one C atom, and one N atom. Therefore, the molecular formula is CH_5N. However, the standard way of writing the molecular formula for methylamine is CH_3NH_2 because it shows how the atoms are joined in the molecule.

Practice Exercise Write the molecular formula of chloroform, which is used as a solvent and a cleaning agent. The ball-and-stick model of chloroform is shown here.



Chloroform

Similar problems: 2.47, 2.48.

Example 2.3

Write the empirical formulas for the following molecules: (a) diborane (B_2H_6), used in rocket propellants; (b) dimethyl fumarate ($C_8H_{12}O_4$), a substance used to treat psoriasis, a skin disease; and (c) vanillin ($C_8H_8O_3$), a flavoring agent used in foods and beverages.

Strategy Recall that to write the empirical formula, the subscripts in the molecular formula must be converted to the smallest possible whole numbers.

Solution

(a) There are two boron atoms and six hydrogen atoms in diborane. Dividing the subscripts by 2, we obtain the empirical formula BH_3.

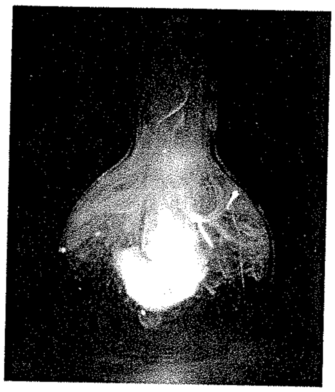
(b) In dimethyl fumarate there are 8 carbon atoms, 12 hydrogen atoms, and 4 oxygen atoms. Dividing the subscripts by 4, we obtain the empirical formula C_2H_3O. Note that if we had divided the subscripts by 2, we would have obtained the formula $C_4H_6O_2$. Although the ratio of carbon to hydrogen to oxygen atoms in $C_4H_6O_2$ is the same as that in C_2H_3O (2:3:1), $C_4H_6O_2$ is not the simplest formula because its subscripts are not in the smallest whole-number ratio.

(c) Because the subscripts in $C_8H_8O_3$ are already the smallest possible whole numbers, the empirical formula for vanillin is the same as its molecular formula.

Practice Exercise

Write the empirical formula for caffeine ($C_8H_{10}N_4O_2$), a stimulant found in tea and coffee.

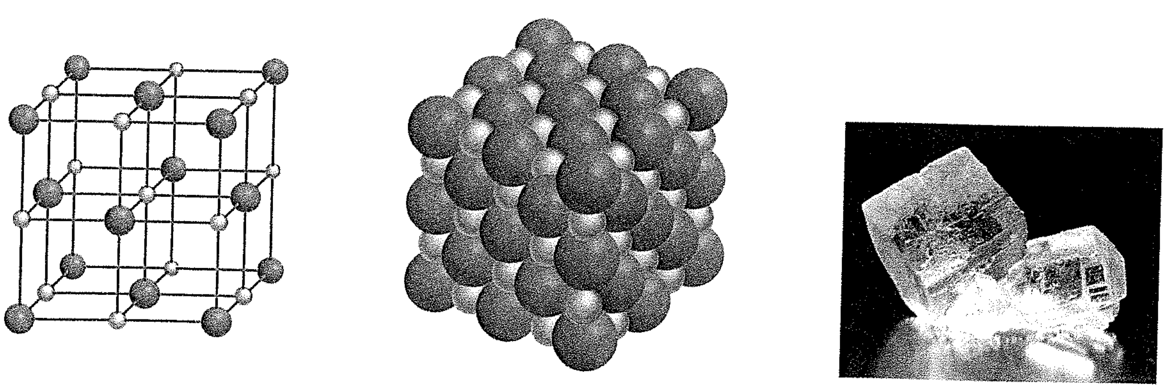
Similar problems: 2.45, 2.46.



Sodium metal reacting with chlorine gas to form sodium chloride.

Formula of Ionic Compounds

The formulas of ionic compounds are usually the same as their empirical formulas, because ionic compounds do not consist of discrete molecular units. For example, a solid sample of sodium chloride (NaCl) consists of equal numbers of Na^+ and Cl^- ions arranged in a three-dimensional network called a ***lattice*** (Figure 2.13). In such a compound there is a 1:1 ratio of cations to anions so that the compound is electrically neutral. As you can see in Figure 2.13, no Na^+ ion in NaCl is associated with just one particular Cl^- ion. In fact, each Na^+ ion is equally held by six surrounding Cl^- ions and vice versa. Thus, NaCl is the empirical formula for sodium chloride. In other ionic compounds, the actual structure may be different, but the arrangement of cations and anions is always such that the compounds are all electrically neutral. Note that the charges on the cation and anion are not shown in the formula for an ionic compound.



(a) (b) (c)

Figure 2.13 *(a) Structure of solid NaCl. (b) In reality, the cations are in contact with the anions. In both (a) and (b), the smaller spheres represent Na^+ ions and the larger spheres, Cl^- ions. (c) Crystals of NaCl.*

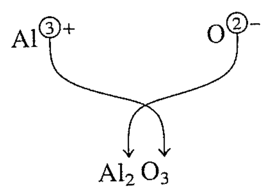
For ionic compounds to be electrically neutral, the sum of the charges on the cation and anion in each formula unit must be zero. If the charges on the cation and anion are numerically different, we apply the following rule to make the formula electrically neutral: *The subscript of the cation is numerically equal to the charge on the anion, and the subscript of the anion is numerically equal to the charge on the cation.* If the charges are numerically equal, then no subscripts are necessary. This rule follows from the fact that because the formulas of ionic compounds are usually empirical formulas, the subscripts must always be reduced to the smallest ratios. Let us consider some examples.

Refer to Figure 2.11 for charges of cations and anions.

- **Potassium Bromide.** The potassium cation K^+ and the bromine anion Br^- combine to form the ionic compound potassium bromide. The sum of the charges is $+1 + (-1) = 0$, so no subscripts are necessary. The formula is KBr.
- **Zinc Iodide.** The zinc cation Zn^{2+} and the iodine anion I^- combine to form zinc iodide. The sum of the charges of one Zn^{2+} ion and one I^- ion is $+2 + (-1) = +1$. To make the charges add up to zero we multiply the -1 charge of the anion by 2 and add the subscript "2" to the symbol for iodine. Therefore the formula for zinc iodide is ZnI_2.
- **Aluminum Oxide.** The cation is Al^{3+} and the oxygen anion is O^{2-}. The following diagram helps us determine the subscripts for the compound formed by the cation and the anion:

Student Hot Spot

Student data indicate you may struggle with formulas of ionic compounds. Access your eBook for additional Learning Resources on this topic.

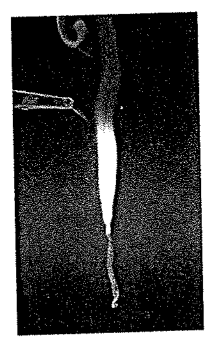


The sum of the charges is $2(+3) + 3(-2) = 0$. Thus, the formula for aluminum oxide is Al_2O_3.

Note that in each of the three examples, the subscripts are in the smallest ratios.

Example 2.4

Magnesium nitride is used to prepare Borazon, a very hard compound used in cutting tools and machine parts. Write the formula of magnesium nitride, containing the Mg^{2+} and N^{3-} ions.



When magnesium burns in air, it forms both magnesium oxide and magnesium nitride.

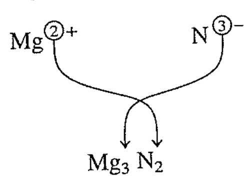
©Charles D. Winters/Science Source

Strategy Our guide for writing formulas for ionic compounds is electrical neutrality; that is, the total charge on the cation(s) must be equal to the total charge on the anion(s). Because the charges on the Mg^{2+} and N^{3-} ions are not equal, we know the formula cannot be MgN. Instead, we write the formula as Mg_xN_y, where x and y are subscripts to be determined.

Solution To satisfy electrical neutrality, the following relationship must hold:

$$(+2)x + (-3)y = 0$$

Solving, we obtain $x/y = 3/2$. Setting $x = 3$ and $y = 2$, we write



Check The subscripts are reduced to the smallest whole-number ratio of the atoms because the chemical formula of an ionic compound is usually its empirical formula.

Practice Exercise Write the formulas of the following ionic compounds: (a) chromium sulfate (containing the Cr^{3+} and SO_4^{2-} ions) and (b) titanium oxide (containing the Ti^{4+} and O^{2-} ions).

Similar problems: 2.43, 2.44.

Review of Concepts & Facts

2.6.1 Match each of the diagrams shown here with the following ionic compounds: Al_2O_3, LiH, Na_2S, $Mg(NO_3)_2$. (Green spheres represent cations and red spheres represent anions.)

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

2.6.2 Write the formulas for the following ionic compounds: (a) calcium iodide (containing the Ca^{2+} and I^- ions), (b) gallium sulfide (containing the Ga^{3+} and S^{2-} ions).

2.7 Naming Compounds

When chemistry was a young science and the number of known compounds was small, it was possible to memorize their names. Many of the names were derived from their physical appearance, properties, origin, or application—for example, milk of magnesia, laughing gas, limestone, caustic soda, lye, washing soda, and baking soda.

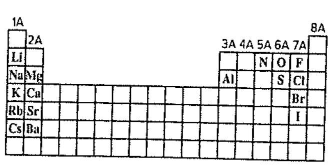
Today the number of known compounds is well over 66 million. Fortunately, it is not necessary to memorize their names. Over the years chemists have devised a clear system for naming chemical substances. The rules are accepted worldwide, facilitating communication among chemists and providing a useful way of labeling an overwhelming variety of substances. Mastering these rules now will prove beneficial almost immediately as we proceed with our study of chemistry.

To begin our discussion of chemical *nomenclature,* the naming of chemical compounds, we must first distinguish between inorganic and organic compounds. ***Organic compounds*** *contain carbon, usually in combination with elements such as hydrogen, oxygen, nitrogen, and sulfur.* All other compounds are classified as ***inorganic compounds.*** For convenience, some carbon-containing compounds, such as carbon monoxide (CO), carbon dioxide (CO_2), carbon disulfide (CS_2), compounds containing the cyanide group (CN^-), and carbonate (CO_3^{2-}) and bicarbonate (HCO_3^-) groups are considered to be inorganic compounds. Section 2.8 gives a brief introduction to organic compounds.

For names and symbols of the elements, see the list of The Elements with Their Symbols and Atomic Masses.

To organize and simplify our venture into naming compounds, we can divide inorganic compounds into four categories: ionic compounds, molecular compounds, acids and bases, and hydrates.

Ionic Compounds



The most reactive metals (green) and the most reactive nonmetals (blue) combine to form ionic compounds.

In Section 2.5 we learned that ionic compounds are made up of cations (positive ions) and anions (negative ions). With the important exception of the ammonium ion, NH_4^+, all cations of interest to us are derived from metal atoms. Metal cations take their names from the elements. For example,

Element			**Name of Cation**
Na	sodium	Na^+	sodium ion (or sodium cation)
K	potassium	K^+	potassium ion (or potassium cation)
Mg	magnesium	Mg^{2+}	magnesium ion (or magnesium cation)
Al	aluminum	Al^{3+}	aluminum ion (or aluminum cation)

Video
Formation of an Ionic Compound

Many ionic compounds are ***binary compounds,*** or *compounds formed from just two elements.* For binary compounds, the first element named is the metal cation, followed by the nonmetallic anion. Thus, NaCl is sodium chloride. The anion is named by taking the first part of the element name (chlorine) and adding "-ide." Potassium

Table 2.2 **The "-ide" Nomenclature of Some Common Monatomic Anions According to Their Positions in the Periodic Table**

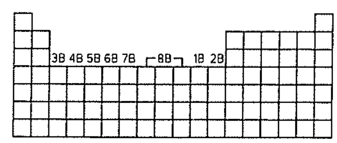
Group 4A	Group 5A	Group 6A	Group 7A
C carbide (C^{4-})*	N nitride (N^{3-})	O oxide (O^{2-})	F fluoride (F^-)
Si silicide (Si^{4-})	P phosphide (P^{3-})	S sulfide (S^{2-})	Cl chloride (Cl^-)
		Se selenide (Se^{2-})	Br bromide (Br^-)
		Te telluride (Te^{2-})	I iodide (I^-)

*The word "carbide" is also used for the anion C_2^{2-}.

bromide (KBr), zinc iodide (ZnI_2), and aluminum oxide (Al_2O_3) are also binary compounds. Table 2.2 shows the "-ide" nomenclature of some common monatomic anions according to their positions in the periodic table.

The "-ide" ending is also used for certain anion groups containing different elements, such as hydroxide (OH^-) and cyanide (CN^-). Thus, the compounds LiOH and KCN are named lithium hydroxide and potassium cyanide, respectively. These and a number of other such ionic substances are called ***ternary compounds,*** meaning *compounds consisting of three elements.* Table 2.3 lists alphabetically the names of a number of common cations and anions.

Certain metals, especially the *transition metals,* can form more than one type of cation. Take iron as an example. Iron can form two cations: Fe^{2+} and Fe^{3+}. An older nomenclature system that is still in limited use assigns the ending "-ous" to the cation with fewer positive charges and the ending "-ic" to the cation with more positive charges:



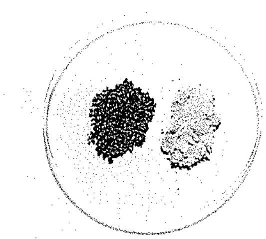
The transition metals are the elements in Groups 1B and 3B–8B (see Figure 2.10).

Fe^{2+}	ferrous ion
Fe^{3+}	ferric ion

The names of the compounds that these iron ions form with chlorine would thus be

$FeCl_2$	ferrous chloride
$FeCl_3$	ferric chloride

This method of naming ions has some distinct limitations. First, the "-ous" and "-ic" suffixes do not provide information regarding the actual charges of the two cations involved. Thus, the ferric ion is Fe^{3+}, but the cation of copper named cupric has the formula Cu^{2+}. In addition, the "-ous" and "-ic" designations provide names for only two different elemental cations.



$FeCl_2$ (left) and $FeCl_3$ (right).
©McGraw-Hill Education/Ken Karp

Some metals can form three or more ions with different positive charges in compounds. Therefore, it has become increasingly common to designate different cations with Roman numerals. This is called the Stock† system. In this system, the Roman numeral I indicates one positive charge on the metal cation, II means two positive charges on the metal cation, and so on. For example, manganese (Mn) atoms can assume several different positive charges:

Mn^{2+}: MnO	manganese(II) oxide
Mn^{3+}: Mn_2O_3	manganese(III) oxide
Mn^{4+}: MnO_2	manganese(IV) oxide

These names are pronounced "manganese-two oxide," "manganese-three oxide," and "manganese-four oxide." Using the Stock system, we denote the ferrous ion and the ferric ion as iron(II) and iron(III), respectively; ferrous chloride becomes iron(II)

Nontransition metals such as tin (Sn) and lead (Pb) can also form more than one type of cations.

†Alfred E. Stock (1876–1946). German chemist. Stock did most of his research in the synthesis and characterization of boron, beryllium, and silicon compounds. He was the first scientist to explore the dangers of mercury poisoning.

Table 2.3 Names and Formulas of Some Common Inorganic Cations and Anions

Cation	Anion
aluminum (Al^{3+})	bromide (Br^-)
ammonium (NH_4^+)	carbonate (CO_3^{2-})
barium (Ba^{2+})	chlorate (ClO_3^-)
cadmium (Cd^{2+})	chloride (Cl^-)
calcium (Ca^{2+})	chromate (CrO_4^{2-})
cesium (Cs^+)	cyanide (CN^-)
chromium(III) or chromic (Cr^{3+})	dichromate ($Cr_2O_7^{2-}$)
cobalt(II) or cobaltous (Co^{2+})	dihydrogen phosphate ($H_2PO_4^-$)
copper(I) or cuprous (Cu^+)	fluoride (F^-)
copper(II) or cupric (Cu^{2+})	hydride (H^-)
hydrogen (H^+)	hydrogen carbonate or bicarbonate (HCO_3^-)
iron(II) or ferrous (Fe^{2+})	hydrogen phosphate (HPO_4^{2-})
iron(III) or ferric (Fe^{3+})	hydrogen sulfate or bisulfate (HSO_4^-)
lead(II) or plumbous (Pb^{2+})	hydroxide (OH^-)
lithium (Li^+)	iodide (I^-)
magnesium (Mg^{2+})	nitrate (NO_3^-)
manganese(II) or manganous (Mn^{2+})	nitride (N^{3-})
mercury(I) or mercurous (Hg_2^{2+})*	nitrite (NO_2^-)
mercury(II) or mercuric (Hg^{2+})	oxide (O^{2-})
potassium (K^+)	permanganate (MnO_4^-)
rubidium (Rb^+)	peroxide (O_2^{2-})
silver (Ag^+)	phosphate (PO_4^{3-})
sodium (Na^+)	sulfate (SO_4^{2-})
strontium (Sr^{2+})	sulfide (S^{2-})
tin(II) or stannous (Sn^{2+})	sulfite (SO_3^{2-})
zinc (Zn^{2+})	thiocyanate (SCN^-)

*Mercury(I) exists as a pair as shown.

Student Hot Spot

Student data indicate you may struggle with naming an ionic compound from its formula. Access your eBook for additional Learning Resources on this topic.

chloride, and ferric chloride is called iron(III) chloride. In keeping with modern practice, we will favor the Stock system of naming compounds in this textbook.

Examples 2.5 and 2.6 illustrate how to name ionic compounds and write formulas for ionic compounds based on the information given in Figure 2.11 and Tables 2.2 and 2.3.

Example 2.5

Name the following compounds: (a) $Fe(NO_3)_2$, (b) Na_2HPO_4, and (c) $(NH_4)_2SO_3$.

Strategy Our reference for the names of cations and anions is Table 2.3. Keep in mind that if a metal can form cations of different charges (see Figure 2.11), we need to use the Stock system.

Solution

(a) The nitrate ion (NO_3^-) bears one negative charge, so the iron ion must have two positive charges. Because iron forms both Fe^{2+} and Fe^{3+} ions, we need to use the Stock system and call the compound iron(II) nitrate.

(Continued)

(b) The cation is Na^+ and the anion is HPO_4^{2-} (hydrogen phosphate). Because sodium only forms one type of ion (Na^+), there is no need to use sodium(I) in the name. The compound is sodium hydrogen phosphate.

(c) The cation is NH_4^+ (ammonium ion) and the anion is SO_3^{2-} (sulfite ion). The compound is ammonium sulfite.

Practice Exercise Name the following compounds: (a) PbO and (b) $LiClO_3$.

Similar problems: 2.57(b), (e), (f).

Example 2.6

Student Hot Spot

Student data indicate you may struggle with naming an ionic compound from its formula. Access your eBook for additional Learning Resources on this topic.

Write chemical formulas for the following compounds: (a) mercury(I) nitrate, (b) cesium oxide, and (c) strontium nitride.

Strategy We refer to Table 2.3 for the formulas of cations and anions. Recall that the Roman numerals in the Stock system provide useful information about the charges of the cation.

Solution

(a) The Roman numeral shows that the mercury ion bears a +1 charge. According to Table 2.3, however, the mercury(I) ion is diatomic (that is, Hg_2^{2+}) and the nitrate ion is NO_3^-. Therefore, the formula is $Hg_2(NO_3)_2$.

Note that the subscripts of this ionic compound are not reduced to the smallest ratio because the Hg(I) ion exists as a pair or dimer.

(b) Each oxide ion bears two negative charges, and each cesium ion bears one positive charge (cesium is in Group 1A, as is sodium). Therefore, the formula is Cs_2O.

(c) Each strontium ion (Sr^{2+}) bears two positive charges, and each nitride ion (N^{3-}) bears three negative charges. To make the sum of the charges equal zero, we must adjust the numbers of cations and anions:

$$3(+2) + 2(-3) = 0$$

Thus, the formula is Sr_3N_2.

Practice Exercise Write formulas for the following ionic compounds: (a) rubidium sulfate and (b) barium hydride.

Similar problems: 2.59(a), (b), (d), (h), (i).

Molecular Compounds

Unlike ionic compounds, molecular compounds contain individual molecular units. They are usually composed of nonmetallic elements (see Figure 2.10). Many molecular compounds are binary compounds. Naming binary molecular compounds is similar to naming binary ionic compounds. We place the name of the first element in the formula first, and the second element is named by adding "-ide" to the root of the element name. Some examples are

HCl	hydrogen chloride
HBr	hydrogen bromide
SiC	silicon carbide

It is quite common for one pair of elements to form several different compounds. In these cases, confusion in naming the compounds is avoided by the use of Greek prefixes to denote the number of atoms of each element present (Table 2.4). Consider the following examples:

CO	carbon monoxide
CO_2	carbon dioxide
SO_2	sulfur dioxide
SO_3	sulfur trioxide
NO_2	nitrogen dioxide
N_2O_4	dinitrogen tetroxide

Table 2.4

Greek Prefixes Used in Naming Molecular Compounds

Prefix	Meaning
mono-	1
di-	2
tri-	3
tetra-	4
penta-	5
hexa-	6
hepta-	7
octa-	8
nona-	9
deca-	10

The following guidelines are helpful in naming compounds with prefixes:

- The prefix "mono-" may be omitted for the first element. For example, PCl_3 is named phosphorus trichloride, not monophosphorus trichloride. Thus, the absence of a prefix for the first element usually means there is only one atom of that element present in the molecule.
- For oxides, the ending "a" in the prefix is sometimes omitted. For example, N_2O_4 may be called dinitrogen tetroxide rather than dinitrogen tetraoxide.

Exceptions to the use of Greek prefixes are molecular compounds containing hydrogen. Traditionally, many of these compounds are called either by their common, nonsystematic names or by names that do not specifically indicate the number of H atoms present:

B_2H_6	diborane
CH_4	methane
SiH_4	silane
NH_3	ammonia
PH_3	phosphine
H_2O	water
H_2S	hydrogen sulfide

Binary compounds containing carbon and hydrogen are organic compounds; they do not follow the same naming conventions. We will discuss the naming of organic compounds in Chapter 24.

Note that even the order of writing the elements in the formulas for hydrogen compounds is irregular. In water and hydrogen sulfide, H is written first, whereas it appears last in the other compounds.

Writing formulas for molecular compounds is usually straightforward. Thus, the name arsenic trifluoride means that there are three F atoms and one As atom in each molecule, and the molecular formula is AsF_3. Note that the order of elements in the formula is the same as in its name.

Student Hot Spot

Student data indicate you may struggle with naming molecular compounds. Access your eBook for additional Learning Resources on this topic.

Example 2.7

Name the following molecular compounds: (a) PBr_5 and (b) As_2O_5.

Strategy We refer to Table 2.4 for the prefixes used in naming molecular compounds.

Solution

(a) Because there are five bromine atoms present, the compound is phosphorus pentabromide.

(b) There are two arsenic atoms and five oxygen atoms present, so the compound is diarsenic pentoxide. Note that the "a" is omitted in "penta."

Practice Exercise Name the following molecular compounds: (a) NF_3 and (b) Cl_2O_7.

Similar problems: 2.57(c), (i), (j).

Example 2.8

Write chemical formulas for the following molecular compounds: (a) bromine trifluoride and (b) diboron trioxide.

Strategy We refer to Table 2.4 for the prefixes used in naming molecular compounds.

Solution

(a) Because there are three fluorine atoms and one bromine atom present, the formula is BrF_3.

(b) There are two boron atoms and three oxygen atoms present, so the formula is B_2O_3.

Practice Exercise Write chemical formulas for the following molecular compounds: (a) sulfur tetrafluoride and (b) dinitrogen pentoxide.

Similar problems: 2.59(g), (j).

Figure 2.14 summarizes the steps for naming ionic and binary molecular compounds.

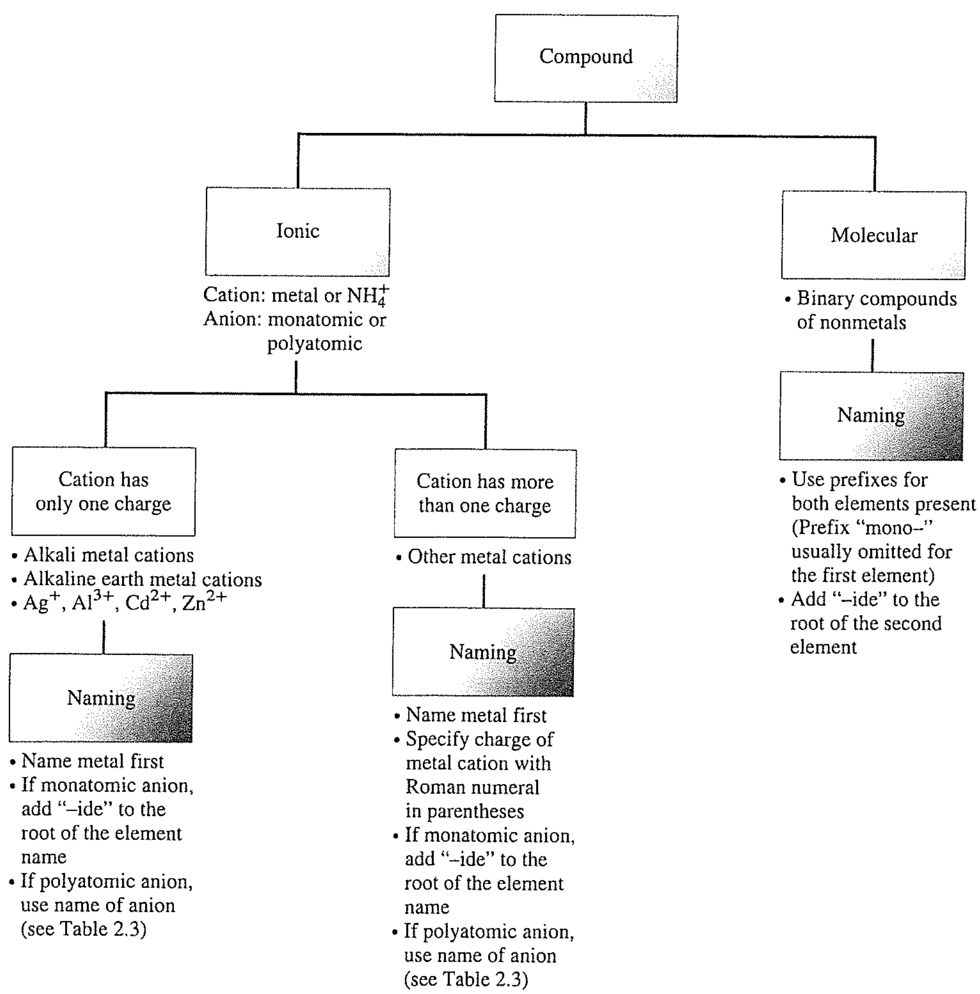
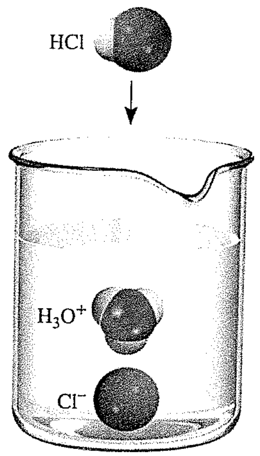


Figure 2.14 *Steps for naming ionic and binary molecular compounds.*



When dissolved in water, the HCl molecule is converted to the H^+ and Cl^- ions. The H^+ ion is associated with one or more water molecules, and is usually represented as H_3O^+.

Acids and Bases

Naming Acids

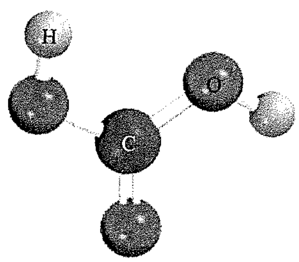
An ***acid*** can be described as *a substance that yields hydrogen ions (H^+) when dissolved in water.* (H^+ is equivalent to one proton, and is often referred to that way.) Formulas for acids contain one or more hydrogen atoms as well as an anionic group. Anions whose names end in "-ide" form acids with a "hydro-" prefix and an "-ic" ending, as shown in Table 2.5. In some cases two different names seem to be assigned to the same chemical formula.

HCl	hydrogen chloride
HCl	hydrochloric acid

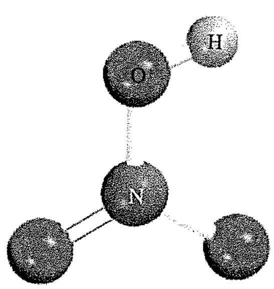
The name assigned to the compound depends on its physical state. In the gaseous or pure liquid state, HCl is a molecular compound called hydrogen chloride. When it is dissolved in water, the molecules break up into H^+ and Cl^- ions; in this state, the substance is called hydrochloric acid.

Oxoacids are acids that *contain hydrogen, oxygen, and another element (the central element).* The formulas of oxoacids are usually written with the H first, followed by the central element and then O. We use the following five common acids as our references in naming oxoacids:

H_2CO_3	carbonic acid
$HClO_3$	chloric acid
HNO_3	nitric acid
H_3PO_4	phosphoric acid
H_2SO_4	sulfuric acid



H_2CO_3



HNO_3

Often two or more oxoacids have the same central atom but a different number of O atoms. Starting with our reference oxoacids whose names all end with "-ic," we use the following rules to name these compounds.

1. Addition of one O atom to the "-ic" acid: The acid is called "per . . .-ic" acid. Thus, adding an O atom to $HClO_3$ changes chloric acid to perchloric acid, $HClO_4$.
2. Removal of one O atom from the "-ic" acid: The acid is called "-ous" acid. Thus, nitric acid, HNO_3, becomes nitrous acid, HNO_2.
3. Removal of two O atoms from the "-ic" acid: The acid is called "hypo . . .-ous" acid. Thus, when $HBrO_3$ is converted to HBrO, the acid is called hypobromous acid.

Table 2.5 Some Simple Acids

Molecular Compound	Acid	Corresponding Anion
HF (hydrogen fluoride)	HF (hydrofluoric acid)	F^- (fluoride)
HCl (hydrogen chloride)	HCl (hydrochloric acid)	Cl^- (chloride)
HBr (hydrogen bromide)	HBr (hydrobromic acid)	Br^- (bromide)
HI (hydrogen iodide)	HI (hydroiodic acid)	I^- (iodide)
HCN (hydrogen cyanide)	HCN (hydrocyanic acid)	CN^- (cyanide)
H_2S (hydrogen sulfide)	H_2S (hydrosulfuric acid)	S^{2-} (sulfide)

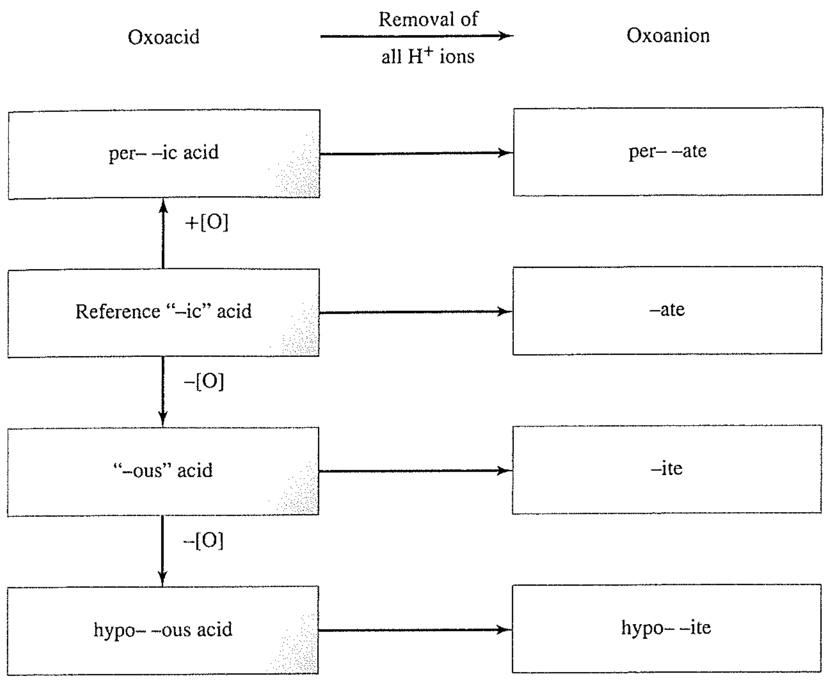
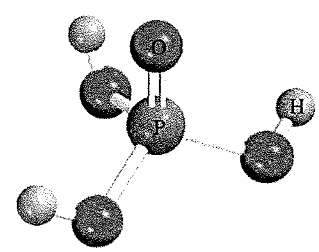


Figure 2.15 *Naming oxoacids and oxoanions.*

The rules for naming ***oxoanions,*** *anions of oxoacids,* are as follows:

1. When all the H ions are removed from the "-ic" acid, the anion's name ends with "-ate." For example, the anion CO_3^{2-} derived from H_2CO_3 is called carbonate.
2. When all the H ions are removed from the "-ous" acid, the anion's name ends with "-ite." Thus, the anion ClO_2^- derived from $HClO_2$ is called chlorite.
3. The names of anions in which one or more but not all the hydrogen ions have been removed must indicate the number of H ions present. For example, consider the anions derived from phosphoric acid:

H_3PO_4	phosphoric acid
$H_2PO_4^-$	dihydrogen phosphate
HPO_4^{2-}	hydrogen phosphate
PO_4^{3-}	phosphate



H_3PO_4

Note that we usually omit the prefix "mono-" when there is only one H in the anion. Figure 2.15 summarizes the nomenclature for the oxoacids and oxoanions, and Table 2.6 gives the names of the oxoacids and oxoanions that contain chlorine.

Example 2.9 deals with the nomenclature for an oxoacid and an oxoanion.

Table 2.6 **Names of Oxoacids and Oxoanions That Contain Chlorine**

Acid	Corresponding Anion
$HClO_4$ (perchloric acid)	ClO_4^- (perchlorate)
$HClO_3$ (chloric acid)	ClO_3^- (chlorate)
$HClO_2$ (chlorous acid)	ClO_2^- (chlorite)
$HClO$ (hypochlorous acid)	ClO^- (hypochlorite)

Student Hot Spot

Student data indicate you may struggle with naming oxoacids and compounds with oxoanions. Access the SmartBook to view additional Learning Resources on this topic.

Example 2.9

Name the following oxoacid and oxoanions: (a) H_2SO_3, a very unstable acid formed when $SO_2(g)$ reacts with water, (b) $H_2AsO_4^-$, once used to control ticks and lice on livestock, and (c) SeO_3^{2-}, used to manufacture colorless glass. H_3AsO_4 is arsenic acid, and H_2SeO_4 is selenic acid.

Strategy We refer to Figure 2.15 and Table 2.6 for the conventions used in naming oxoacids and oxoanions.

Solution

(a) We start with our reference acid, sulfuric acid (H_2SO_4). Because H_2SO_3 has one fewer O atom, it is called sulfurous acid.

(b) Because H_3AsO_4 is arsenic acid, the AsO_4^{3-} ion is named arsenate. The $H_2AsO_4^-$ anion is formed by adding two H^+ ions to AsO_4^{3-}, so $H_2AsO_4^-$ is called dihydrogen arsenate.

(c) The parent acid is H_2SeO_3. Because the acid has one fewer O atom than selenic acid (H_2SeO_4), it is called selenous acid. Therefore, the SeO_3^{2-} anion derived from H_2SeO_3 is called selenite.

Practice Exercise Name the following oxoacid and oxoanion: (a) HBrO and (b) HSO_4^-.

Similar problems: 2.58(f).

Naming Bases

A ***base*** can be described as *a substance that yields hydroxide ions (OH^-) when dissolved in water.* Some examples are

NaOH	sodium hydroxide
KOH	potassium hydroxide
$Ba(OH)_2$	barium hydroxide

Ammonia (NH_3), a molecular compound in the gaseous or pure liquid state, is also classified as a common base. At first glance this may seem to be an exception to the definition of a base. But note that as long as a substance *yields* hydroxide ions when dissolved in water, it need not contain hydroxide ions in its structure to be considered a base. In fact, when ammonia dissolves in water, NH_3 reacts partially with water to yield NH_4^+ and OH^- ions. Thus, it is properly classified as a base.

Hydrates

Hydrates are *compounds that have a specific number of water molecules attached to them.* For example, in its normal state, each unit of copper(II) sulfate has five water molecules associated with it. The systematic name for this compound is copper(II) sulfate pentahydrate, and its formula is written as $CuSO_4 \cdot 5H_2O$. The water molecules can be driven off by heating. When this occurs, the resulting compound is $CuSO_4$, which is sometimes called *anhydrous* copper(II) sulfate; "anhydrous" means that the compound no longer has water molecules associated with it (Figure 2.16). Some other hydrates are

$BaCl_2 \cdot 2H_2O$	barium chloride dihydrate
$LiCl \cdot H_2O$	lithium chloride monohydrate
$MgSO_4 \cdot 7H_2O$	magnesium sulfate heptahydrate
$Sr(NO_3)_2 \cdot 4H_2O$	strontium nitrate tetrahydrate

Figure 2.16 *$CuSO_4 \cdot 5H_2O$ (left) is blue; $CuSO_4$ (right) is white.*

Table 2.7 Common and Systematic Names of Some Compounds

Formula	Common Name	Systematic Name
H_2O	Water	Dihydrogen monoxide
NH_3	Ammonia	Trihydrogen nitride
CO_2	Dry ice	Solid carbon dioxide
NaCl	Table salt	Sodium chloride
N_2O	Laughing gas	Dinitrogen monoxide
$CaCO_3$	Marble, chalk, limestone	Calcium carbonate
CaO	Quicklime	Calcium oxide
$Ca(OH)_2$	Slaked lime	Calcium hydroxide
$NaHCO_3$	Baking soda	Sodium hydrogen carbonate
$Na_2CO_3 \cdot 10H_2O$	Washing soda	Sodium carbonate decahydrate
$MgSO_4 \cdot 7H_2O$	Epsom salt	Magnesium sulfate heptahydrate
$Mg(OH)_2$	Milk of magnesia	Magnesium hydroxide
$CaSO_4 \cdot 2H_2O$	Gypsum	Calcium sulfate dihydrate

Familiar Inorganic Compounds

Some compounds are better known by their common names than by their systematic chemical names. Familiar examples are listed in Table 2.7.

Review of Concepts & Facts

2.7.1 Why is it that the name for $SeCl_2$, selenium dichloride, contains a prefix, but the name for $SrCl_2$, strontium chloride, does not?

2.7.2 Why is the following question ambiguous: What is the name of HF? What additional information is needed to answer the question?

2.7.3 Name the following compounds: (a) Cs_2SO_3, (b) $Cu(NO_2)_2$.

2.7.4 Write formulas for the following compounds: (a) cobalt(II) hydrogen sulfate, (b) chromium(III) cyanide.

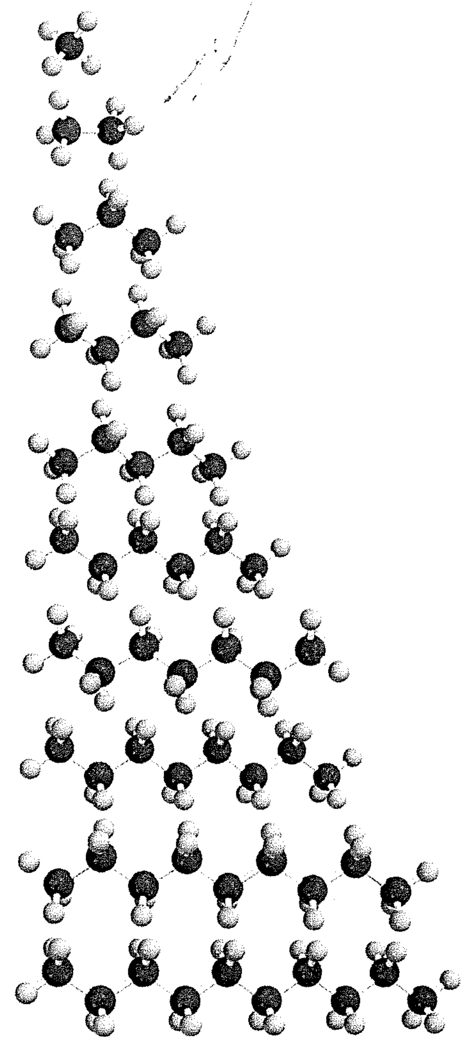
2.7.5 What are the correct names for the following compounds? (a) BrF_5, (b) HIO_4.

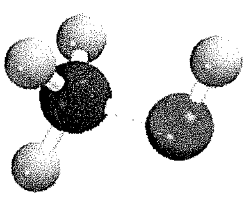
2.8 Introduction to Organic Compounds

The simplest type of organic compounds is the *hydrocarbons,* which contain only carbon and hydrogen atoms. The hydrocarbons are used as fuels for domestic and industrial heating, for generating electricity and powering internal combustion engines, and as starting materials for the chemical industry. One class of hydrocarbons is called the *alkanes.* Table 2.8 shows the names, formulas, and molecular models of the first 10 *straight-chain* alkanes, in which the carbon chains have no branches. The name of an

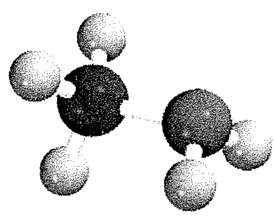
Table 2.8 **The First 10 Straight-Chain Alkanes**

Name	Formula	Molecular Model
Methane	CH_4	
Ethane	C_2H_6	
Propane	C_3H_8	
Butane	C_4H_{10}	
Pentane	C_5H_{12}	
Hexane	C_6H_{14}	
Heptane	C_7H_{16}	
Octane	C_8H_{18}	
Nonane	C_9H_{20}	
Decane	$C_{10}H_{22}$	

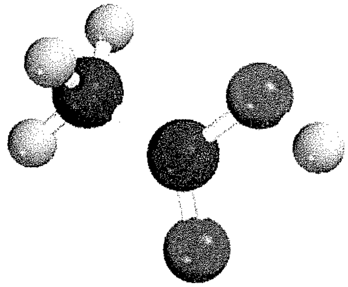




CH_3OH



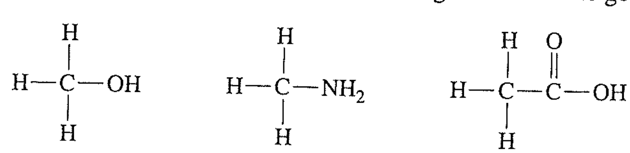
CH_3NH_2



CH_3COOH

alkane depends on the number of carbon atoms in the molecule and all the names end with "-ane." Starting with C_5H_{12}, we use the Greek prefixes in Table 2.4 to indicate the number of carbon atoms present.

The chemistry of organic compounds is largely determined by the *functional groups,* which consist of one or a few atoms bonded in a specific way. For example, when an H atom in methane is replaced by a hydroxyl group (—OH), an amino group (—NH_2), and a carboxyl group (—COOH), the following molecules are generated:



Methanol Methylamine Acetic acid

The chemical properties of these molecules can be predicted based on the reactivity of the functional groups. Although the nomenclature of the major classes of organic compounds and their properties in terms of the functional groups will not be discussed until Chapter 24, we will frequently use organic compounds as examples to illustrate chemical bonding, acid-base reactions, and other properties throughout the book.

Review of Concepts & Facts

2.8.1 How many different molecules can you generate by replacing one H atom with a hydroxyl group (—OH) in butane (see Table 2.8)?

AP A Look Back at the Essential Knowledge

The desire to understand the composition of matter can be traced to when the Ancient Greeks first postulated the concept of an atom. John Dalton provided the first modern atomic theory, postulating that all matter is composed of atoms, that atoms are indivisible, and all atoms of a given element are identical. In addition, he proposed that atoms could combine together in small whole number ratios to form compounds, and that atoms were neither created nor destroyed during chemical reactions.

A series of experiments determined that contrary to Dalton's atomic theory atoms are divisible and consist of several specifically organized particles. Atoms consist of a small dense nucleus containing positively charged protons and neutrons which have no charge. Smaller, negatively charged electrons are in motion at a significantly large distance about the nucleus. Atoms of different elements differ in the number of protons they contain. The number of protons also determines an element's atomic number. All atoms of a given elements contain the same number of protons, but they may differ in the number of neutrons. Atoms of the same element with a different number of neutrons are called isotopes, and a given element may be composed of several different isotopes. The mass number of an isotope is the number of protons and neutrons. Atoms of two or more elements can combine in more than one whole number ratio, resulting in more than one compound from the same set of elements. This is the law of multiple proportions.

All known elements are arranged in a systematic fashion known as the periodic table. Elements are arranged by groups and periods and classified as metals, non-metals, and metalloids based on physical properties such as conductivity. Molecular formulas consist of appropriate elemental symbols and the smallest whole number subscripts indicating the number of each element in the compound. Another type of chemical formula, the empirical formula, is used to show the simplest whole number ratio of elements in a compound. Elements and molecules can gain or lose electrons resulting in positively or negatively charges ion. These ionic units can be combined together in small, whole number units to create a neutral formula unit.

FOCUS REVIEW GUIDE

Complete the activities in Chapter 2 of your *Focus Review Guide* to review content essential for your AP exam.

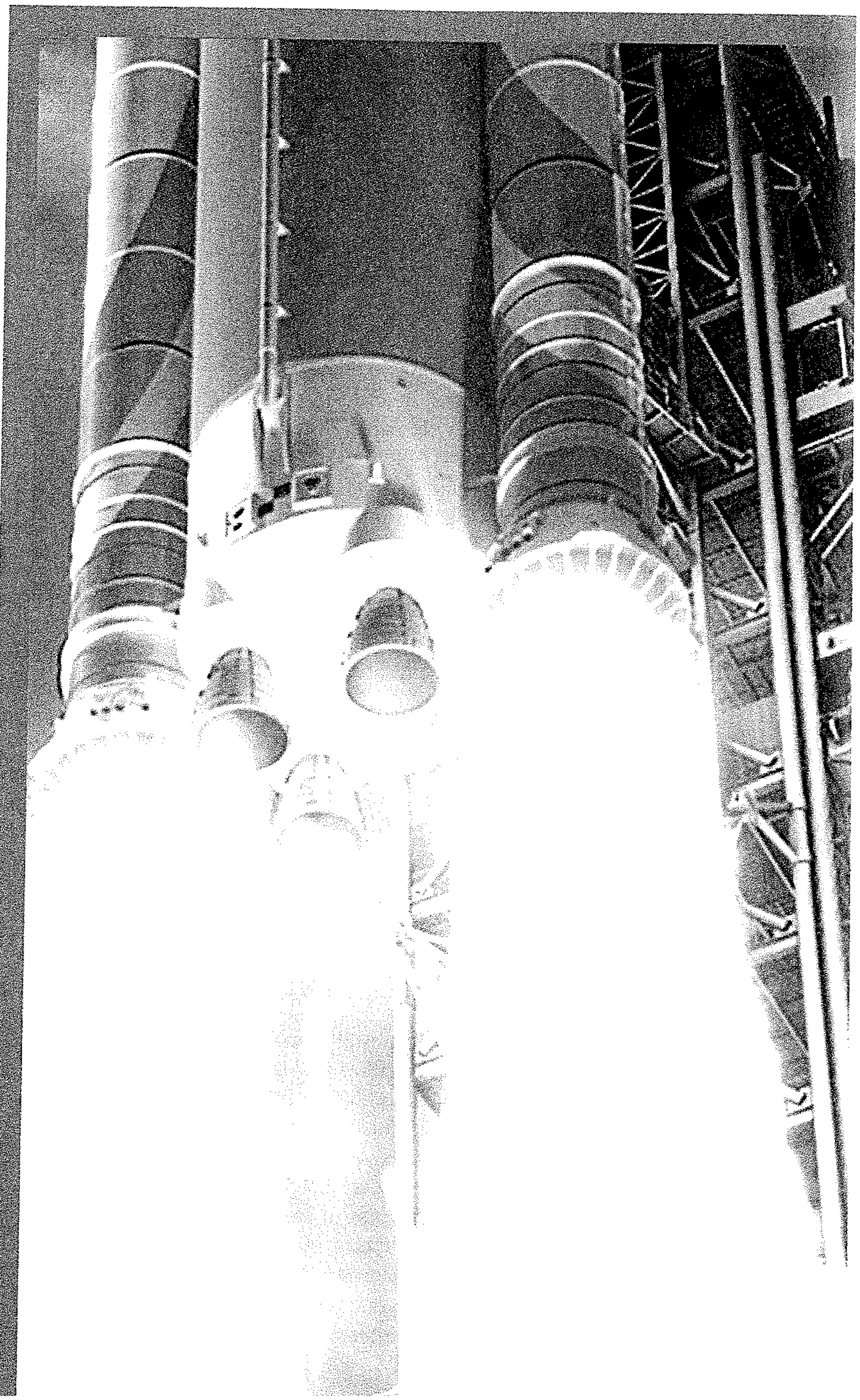
Learning Objectives

- Outline Dalton's hypotheses about the nature of matter. (Section 2.1)
- Understand the concept of the atom and the nature of an element. (Section 2.2)
- Assess the importance of experiments conducted by Thomson, Millikan, Röntgen, and Rutherford, and how they influenced our understanding of the nature and structure of atoms. (Section 2.2)
- Summarize the different types of radiation that radioactive substances can produce. (Section 2.2)
- Describe the location and physical properties of electrons, protons, and neutrons. (Section 2.2)
- Explain the nature and importance of isotopes. (Section 2.3)
- Calculate the mass number of an isotope. (Section 2.3)
- Utilize the mass number of an isotope to solve for the number of electrons, protons, or neutrons, given other relevant information. (Section 2.3)
- Recognize the general organization of the periodic table with respect to metals, metalloids, nonmetals, groups, and periods. (Section 2.4)

CHAPTER

3

Mass Relationships in Chemical Reactions



The reaction between aluminum and ammonium perchlorate fuels the solid rocket boosters to send this rocket into space.
NASA

AP CHAPTER OUTLINE

AP BIG IDEAS A LOOK AHEAD 3

Chemists and cooks have much in common. They perform procedures, which call for a set of ingredients in prescribed amounts. The amounts can be varied but only proportionally. Their starting materials are changed—often with the help of heat—producing new substances with new properties but containing all of the original elements in their original quantities. Whether it's baking chocolate chip cookies or creating a fireworks display, it is crucial that the cook and the chemist know exactly what starting materials they are using and in what amounts they are to be combined. For this to happen, the cook follows a recipe. The chemist uses a chemical equation to determine starting materials and quantities.

To use an equation, the chemist must know how to count particles. Counting particles is challenging since atomic sizes are very small and numerous in any visible sample. To help, a large count unit called the "mole" is used. The mole unit is linked to the mass of a substance, allowing the chemist to go from a macroscopic measurement of mass to the microscopic level of particles (EK.1.A.3). While a cook uses measuring cups and teaspoons, a chemist uses masses, moles, and a chemical equation to determine the amounts of reactant needed to produce a desired amount of product. And from the actual amount of product formed, the efficiency of the reaction can be determined (EK.1.E.2).As you read the chapter, think about these Essential Questions:

1. How do we determine atomic mass? What is a mole? How do we determine molar mass? **1.A.3**
2. How does mass spectroscopy demonstrate the existence of isotopes? **1.D.2**
3. What is the percent elemental composition of a substance? How do we experimentally determine composition? **1.A.2**
4. What is conservation of mass? **1.E.2**
5. How do we correctly represent a chemical change? **1.E.2**
6. How do we write a properly balanced chemical reaction? **3.A.1**
7. What evidence is used to indicate a chemical change has taken place? **3.C.1**
8. How do we determine the amount of product formed in a chemical reaction or the amount of reactant required? How do we determine the limiting reagent? How do we determine percent yield? **3.A.2**

In this chapter, we will consider the masses of atoms and molecules and what happens to them when chemical changes occur. Our guide for this discussion will be the law of conservation of mass.

3.1 Atomic Mass

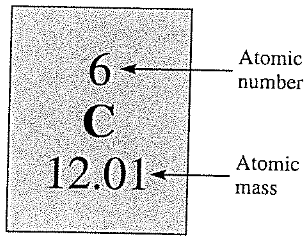
In this chapter, we will use what we have learned about chemical structure and formulas in studying the mass relationships of atoms and molecules. These relationships in turn will help us to explain the composition of compounds and the ways in which composition changes.

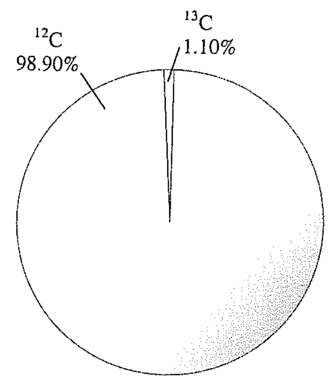
The mass of an atom depends on the number of electrons, protons, and neutrons it contains. Knowledge of an atom's mass is important in laboratory work. But atoms are extremely small particles—even the smallest speck of dust that our unaided eyes can detect contains as many as 1×10^{16} atoms! Clearly we cannot weigh a single atom, but it is possible to determine the mass of one atom *relative* to another experimentally. The first step is to assign a value to the mass of one atom of a given element so that it can be used as a standard.

Section 3.4 describes a method for determining atomic mass.

By international agreement, ***atomic mass*** (sometimes called *atomic weight*) is *the mass of the atom in atomic mass units (amu)*. One ***atomic mass unit*** is defined as *a mass exactly equal to one-twelfth the mass of one carbon-12 atom.* Carbon-12 is the carbon isotope that has six protons and six neutrons. Setting the atomic mass of carbon-12 at 12 amu provides the standard for measuring the atomic mass of the other elements. For example, experiments have shown that, on average, a hydrogen atom is only 8.400 percent as massive as the carbon-12 atom. Thus, if the mass of one carbon-12 atom is exactly 12 amu, the atomic mass of hydrogen must be 0.08400×12 amu or 1.008 amu. Similar calculations show that the atomic mass of oxygen is 16.00 amu and that of iron is 55.85 amu. Thus, although we do not know just how much an average iron atom's mass is, we know that it is approximately 56 times as massive as a hydrogen atom.

One atomic mass unit is also called one dalton.





Natural abundances of C-12 and C-13 isotopes.

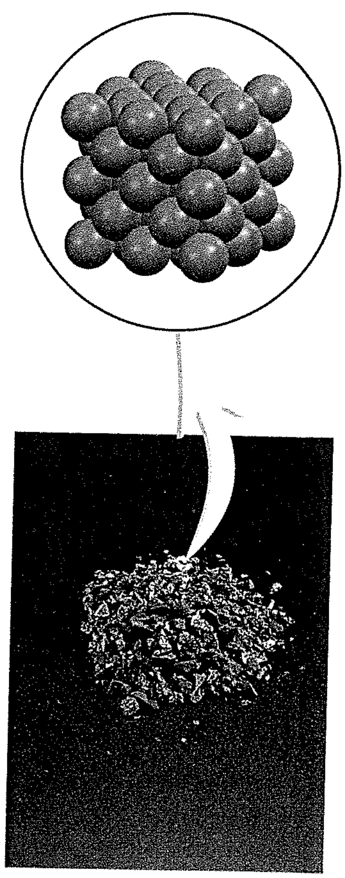
Average Atomic Mass

When you look up the atomic mass of carbon in a periodic table, you will find that i value is not 12.00 amu but 12.01 amu. The reason for the difference is that most natu rally occurring elements (including carbon) have more than one isotope. This mear that when we measure the atomic mass of an element, we must generally settle for th *average* mass of the naturally occurring mixture of isotopes. For example, the natur abundances of carbon-12 and carbon-13 are 98.90 percent and 1.10 percent, respec tively. The atomic mass of carbon-13 has been determined to be 13.00335 amu. Thu the average atomic mass of carbon can be calculated as follows:

$$\begin{aligned}\text{average atomic mass}&\\ \text{of natural carbon} &= (0.9890)\,(12\text{ amu}) + (0.0110)\,(13.00335\text{ amu})\\ &= 12.01\text{ amu}\end{aligned}$$

Note that in calculations involving percentages, we need to convert percentages to frac tions. For example, 98.90 percent becomes 98.90/100, or 0.9890. Because there ar many more carbon-12 atoms than carbon-13 atoms in naturally occurring carbon, th average atomic mass is much closer to 12 amu than to 13 amu.

It is important to understand that when we say that the atomic mass of carbon i 12.01 amu, we are referring to the *average* value. If carbon atoms could be examine individually, we would find either an atom of atomic mass exactly 12 amu or one o 13.00335 amu, but never one of 12.01 amu. Example 3.1 shows how to calculate th average atomic mass of an element.



Boron and the solid-state structure of boron.
©McGraw-Hill Education/Stephen Frisch

Example 3.1

Boron is used in the manufacture of ceramics and polymers such as fiberglass. The atomic masses of its two stable isotopes, $^{10}_{5}B$ (19.80 percent) and $^{11}_{5}B$ (80.20 percent), are 10.0129 amu and 11.0093 amu, respectively. The boron-10 isotope is also important as a neutron-capturing agent in nuclear reactors. Calculate the average atomic mass of boron.

Strategy Each isotope contributes to the average atomic mass based on its relative abundance. Multiplying the mass of an isotope by its fractional abundance (not percent) will give the contribution to the average atomic mass of that particular isotope.

Solution First the percent abundances are converted to fractions: 19.80 percent to 19.80/100 or 0.1980 and 80.20 percent to 80.20/100 or 0.8020. We find the contribution to the average atomic mass for each isotope, and then add the contributions together to obtain the average atomic mass.

$$(0.1980)(10.0129\text{ amu}) + (0.8020)(11.0093\text{ amu}) = 10.8129\text{ amu}$$

Check The average atomic mass should be between the two isotopic masses; therefore, the answer is reasonable. Note that because there are more $^{11}_{5}B$ isotopes than $^{10}_{5}B$ isotopes, the average atomic mass is closer to 11.0093 amu than to 10.0129 amu.

Practice Exercise The atomic masses of the two stable isotopes of copper, $^{63}_{29}Cu$ (69.17 percent) and $^{65}_{29}Cu$ (30.83 percent), are 62.9296 amu and 64.9278 amu, respectively. Calculate the average atomic mass of copper.

Similar problems: 3.5, 3.6.

The atomic masses of many elements have been accurately determined to five or six significant figures. However, for our purposes we will normally use atomic masses accurate only to four significant figures (see List of the Elements with Their Symbols and Atomic Masses). For simplicity, we will omit the word "average" when we discuss the atomic masses of the elements.

Review of Concepts & Facts

3.1.1 There are two stable isotopes of iridium: ^{191}Ir (190.96 amu) and ^{193}Ir (192.96 amu). If you were to randomly pick an iridium atom from a large collection of iridium atoms, which isotope are you more likely to select?

3.1.2 The hypothetical element Q occurs as a mixture of 37.50 percent ^{47}Q (47.054 amu) and 62.50 percent ^{51}Q (50.924 amu). Calculate the average atomic mass of Q.

3.2 Avogadro's Number and the Molar Mass of an Element

Atomic mass units provide a relative scale for the masses of the elements. But because atoms have such small masses, no usable scale can be devised to weigh them in calibrated units of atomic mass units. In any real situation, we deal with macroscopic samples containing enormous numbers of atoms. Therefore, it is convenient to have a special unit to describe a very large number of atoms. The idea of a unit to denote a particular number of objects is not new. For example, the pair (2 items), the dozen (12 items), and the gross (144 items) are all familiar units. Chemists measure atoms and molecules in moles.

In the SI system the ***mole (mol)*** is *the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 g (or 0.012 kg) of the carbon-12 isotope.* The actual number of atoms in 12 g of carbon-12 is determined experimentally. This number is called ***Avogadro's number*** (N_A), in honor of the Italian scientist Amedeo Avogadro.† The currently accepted value is

The adjective formed from the noun "mole" is "molar."

$$N_A = 6.0221413 \times 10^{23}$$

Generally, we round Avogadro's number to 6.022×10^{23}. Thus, just as 1 dozen oranges contains 12 oranges, 1 mole of hydrogen atoms contains 6.022×10^{23} H atoms. Figure 3.1 shows samples containing 1 mole each of several common elements.

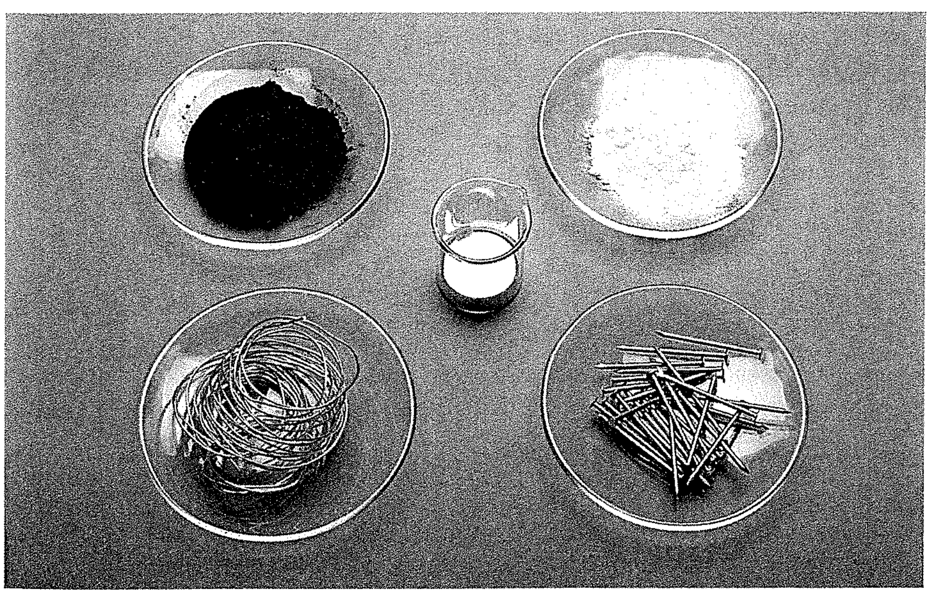


Figure 3.1 *One mole each of several common elements. Carbon (black charcoal powder), sulfur (yellow powder), iron (as nails), copper wires, and mercury (shiny liquid metal).*
©McGraw-Hill Education/Stephen Frisch

†Lorenzo Romano Amedeo Carlo Avogadro di Quaregua e di Cerreto (1776–1856). Italian mathematical physicist. He practiced law for many years before he became interested in science. His most famous work, now known as Avogadro's law (see Chapter 5), was largely ignored during his lifetime, although it became the basis for determining atomic masses in the late nineteenth century.

The enormity of Avogadro's number is difficult to imagine. For example, spreading 6.022×10^{23} oranges over the entire surface of Earth would produce a layer 9 mi into space! Because atoms (and molecules) are so tiny, we need a huge number to study them in manageable quantities.

In calculations, the units of molar mass are g/mol or kg/mol.

We have seen that 1 mole of carbon-12 atoms has a mass of exactly 12 g and contains 6.022×10^{23} atoms. This mass of carbon-12 is its ***molar mass*** (***$\mathcal{M}$***), defined as *the mass (in grams or kilograms) of 1 mole of units* (such as atoms or molecules) *of a substance*. Note that the molar mass of carbon-12 (in grams) is numerically equal to its atomic mass in amu. Likewise, the atomic mass of sodium (Na) is 22.99 amu and its molar mass is 22.99 g; the atomic mass of phosphorus is 30.97 amu and its molar mass is 30.97 g; and so on. If we know the atomic mass of an element, we also know its molar mass.

The molar masses of the elements are given on a periodic table.

Knowing the molar mass and Avogadro's number, we can calculate the mass of a single atom in grams. For example, we know the molar mass of carbon-12 is 12 g and there are 6.022×10^{23} carbon-12 atoms in 1 mole of the substance; therefore, the mass of one carbon-12 atom is given by

$$\frac{12 \text{ g carbon-12 atoms}}{6.022 \times 10^{23} \text{ carbon-12 atoms}} = 1.993 \times 10^{-23} \text{ g}$$

We can use the preceding result to determine the relationship between atomic mass units and grams. Because the mass of every carbon-12 atom is exactly 12 amu, the number of atomic mass units equivalent to 1 gram is

$$\frac{\text{amu}}{\text{gram}} = \frac{12 \text{ amu}}{1 \cancel{\text{carbon-12 atom}}} \times \frac{1 \cancel{\text{carbon-12 atom}}}{1.993 \times 10^{-23} \text{ g}}$$
$$= 6.022 \times 10^{23} \text{ amu/g}$$

Thus,

$$1 \text{ g} = 6.022 \times 10^{-24} \text{ g}$$

and

$$1 \text{ amu} = 1.661 \times 10^{-24} \text{ g}$$

After some practice, you can use the equations in Figure 3.2 in calculations: $n = m/\mathcal{M}$ and $N = nN_A$.

This example shows that Avogadro's number can be used to convert from the atomic mass units to mass in grams and vice versa.

The notions of Avogadro's number and molar mass enable us to carry out conversions between mass and moles of atoms and between moles and number of atoms (Figure 3.2). We will employ the following conversion factors in the calculations:

Student Hot Spot

Student data indicate you may struggle with converting mass of an element into moles and atoms. Access your eBook for additional Learning Resources on this topic.

$$\frac{1 \text{ mol X}}{\text{molar mass of X}} \quad \text{and} \quad \frac{1 \text{ mol X}}{6.022 \times 10^{23} \text{ X atoms}}$$

where X represents the symbol of an element. Using the proper conversion factors we can convert one quantity to another, as Examples 3.2–3.4 show.

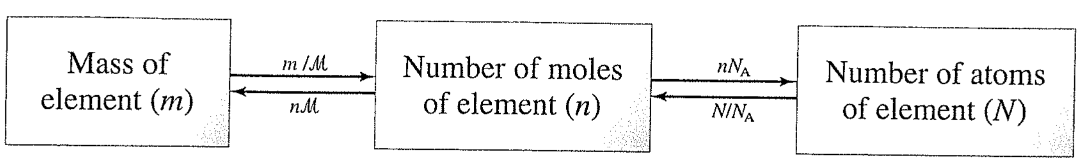


Figure 3.2 *The relationships between mass (m in grams) of an element and number of moles of an element (n) and between number of moles of an element and number of atoms (N) of an element. $\mathcal{M}$ is the molar mass (g/mol) of the element and N_A is Avogadro's number.*

Example 3.2

Helium (He) is a valuable gas used in industry, low-temperature research, deep-sea diving tanks, and balloons. How many moles of He atoms are in 6.46 g of He?

Strategy We are given grams of helium and asked to solve for moles of helium. What conversion factor do we need to convert between grams and moles? Arrange the appropriate conversion factor so that grams cancel and the unit moles is obtained for your answer.

Solution The conversion factor needed to convert between grams and moles is the molar mass. In the periodic table, we see that the molar mass of He is 4.003 g. This can be expressed as

$$1 \text{ mol He} = 4.003 \text{ g He}$$

From this equality, we can write two conversion factors

$$\frac{1 \text{ mol He}}{4.003 \text{ g He}} \quad \text{and} \quad \frac{4.003 \text{ g He}}{1 \text{ mol He}}$$

The conversion factor on the left is the correct one. Grams will cancel, leaving the unit mol for the answer—that is,

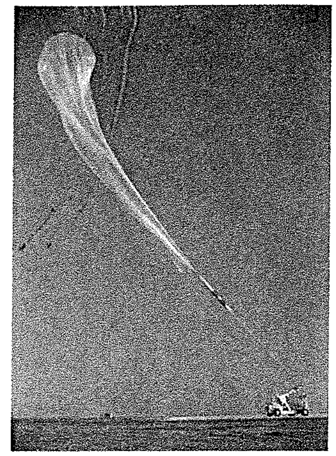
$$6.46 \cancel{\text{g He}} \times \frac{1 \text{ mol He}}{4.003 \cancel{\text{g He}}} = 1.61 \text{ mol He}$$

Thus, there are 1.61 moles of He atoms in 6.46 g of He.

Check Because the given mass (6.46 g) is larger than the molar mass of He, we expect to have more than 1 mole of He.

Practice Exercise How many moles of magnesium (Mg) are there in 87.3 g of Mg?

Similar problem: 3.15.



A scientific research helium balloon.

National Scientific Balloon Facility/Palestine, Texas

Example 3.3

Zinc (Zn) is a silvery metal that is used in making brass (with copper) and in plating iron to prevent corrosion. How many grams of Zn are in 0.356 mole of Zn?

Strategy We are trying to solve for grams of zinc. What conversion factor do we need to convert between moles and grams? Arrange the appropriate conversion factor so that moles cancel and the unit grams are obtained for your answer.

Solution The conversion factor needed to convert between moles and grams is the molar mass. In the periodic table, we see the molar mass of Zn is 65.39 g. This can be expressed as

$$1 \text{ mol Zn} = 65.39 \text{ g Zn}$$

From this equality, we can write two conversion factors

$$\frac{1 \text{ mol Zn}}{65.39 \text{ g Zn}} \quad \text{and} \quad \frac{65.39 \text{ g Zn}}{1 \text{ mol Zn}}$$

The conversion factor on the right is the correct one. Moles will cancel, leaving unit of grams for the answer. The number of grams of Zn is

$$0.356 \cancel{\text{mol Zn}} \times \frac{65.39 \text{ g Zn}}{1 \cancel{\text{mol Zn}}} = 23.3 \text{ g Zn}$$

Thus, there are 23.3 g of Zn in 0.356 mole of Zn.

Zinc.

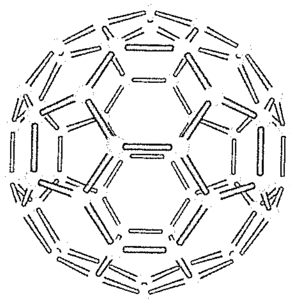
©Charles D. Winters/Science Source

(Continued)

Check Does a mass of 23.3 g for 0.356 mole of Zn seem reasonable? What is the mass of 1 mole of Zn?

Practice Exercise Calculate the number of grams of lead (Pb) in 12.4 moles of lead.

Similar problem: 3.16.



Buckminsterfullerene (C_{60}) or "buckyball."

Example 3.4

The C_{60} molecule is called buckminsterfullerene because its shape resembles the geodesic domes designed by the visionary architect R. Buckminster Fuller. What is the mass (in grams) of one C_{60} molecule?

Strategy The question asks for the mass of one C_{60} molecule. Determine the moles of C atoms in one C_{60} molecule, and then use the molar mass of C to calculate the mass of one molecule in grams.

Solution Because one C_{60} molecule contains 60 C atoms, and 1 mole of C contains 6.022×10^{23} C atoms and has a mass of 12.011 g, we can calculate the mass of one C_{60} molecule as follows:

$$1\ \cancel{C_{60}\ \text{molecule}} \times \frac{60\ \cancel{\text{C atoms}}}{1\ \cancel{C_{60}\ \text{molecule}}} \times \frac{1\ \cancel{\text{mol C}}}{6.022 \times 10^{23}\ \cancel{\text{C atoms}}} \times \frac{12.01\ \text{g}}{1\ \cancel{\text{mol C}}} = 1.197 \times 10^{-21}\ \text{g}$$

Check Because 6.022×10^{23} atoms of C have a mass 12.01 g, a molecule containing only 60 carbon atoms should have a significantly smaller mass.

Practice Exercise Gold atoms form small clusters containing a fixed number of atoms. What is the mass (in grams) of one Au_{31} cluster?

Similar problems: 3.20, 3.21.

Applying Practices

Conduct your own experimental determination of Avogadro's number by accessing the online lab, *Determining Avogadro's Number.*

Review of Concepts & Facts

3.2.1 Determine which of the following contains the largest number of atoms: (a) 7.68 g of He, (b) 112 g of Fe, (c) 389 g of Hg.

3.2.2 How many moles of rubidium (Rb) are there in 3.75×10^{24} Rb atoms?

3.2.3 What is the mass in grams of 1.68 moles of vanadium (V)?

3.3 Molecular Mass

If we know the atomic masses of the component atoms, we can calculate the mass of a molecule. The ***molecular mass*** (sometimes called *molecular weight*) is *the mass (in amu) of one molecule.* It is determined by the sum of the atomic masses of each atom in a molecule. For example, the molecular mass of H_2O is

$$2(\text{atomic mass of H}) + \text{atomic mass of O}$$

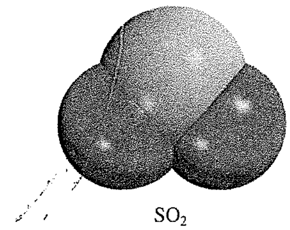
or

$$2(1.008\ \text{amu}) + 16.00\ \text{amu} = 18.02\ \text{amu}$$

In general, we need to multiply the atomic mass of each element by the number of atoms of that element present in the molecule and sum over all the elements. Example 3.5 illustrates this approach.

Example 3.5

Calculate the molecular masses (in amu) of the following compounds: (a) sulfur dioxide (SO_2), a gas that is responsible for acid rain, and (b) caffeine ($C_8H_{10}N_4O_2$), a stimulant present in tea, coffee, and cola beverages.



SO_2

Strategy How do atomic masses of different elements combine to give the molecular mass of a compound?

Solution To calculate molecular mass, we need to sum all the atomic masses in the molecule. For each element, we multiply the atomic mass of the element by the number of atoms of that element in the molecule. We find atomic masses in the periodic table.

(a) There are two O atoms and one S atom in SO_2, so that

$$\begin{aligned}\text{molecular mass of } SO_2 &= 32.07 \text{ amu} + 2(16.00 \text{ amu})\\ &= 64.07 \text{ amu}\end{aligned}$$

(b) There are eight C atoms, ten H atoms, four N atoms, and two O atoms in caffeine, so the molecular mass of $C_8H_{10}N_4O_2$ is given by

$$8(12.01 \text{ amu}) + 10(1.008 \text{ amu}) + 4(14.01 \text{ amu}) + 2(16.00 \text{ amu}) = 194.20 \text{ amu}$$

Practice Exercise What is the molecular mass of methanol (CH_4O)?

Similar problems: 3.23, 3.24.

From the molecular mass we can determine the molar mass of a molecule or compound. The molar mass of a compound (in grams) is numerically equal to its molecular mass (in amu). For example, the molecular mass of water is 18.02 amu, so its molar mass is 18.02 g. Note that 1 mole of water weighs 18.02 g and contains 6.022×10^{23} H_2O *molecules,* just as 1 mole of elemental carbon contains 6.022×10^{23} carbon *atoms.*

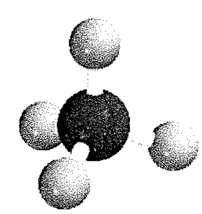
Student Hot Spot

Student data indicate you may struggle with determining moles from a given mass. Access the SmartBook to view additional Learning Resources on this topic.

As Examples 3.6 and 3.7 show, a knowledge of the molar mass enables us to calculate the numbers of moles and individual atoms in a given quantity of a compound.

Example 3.6

Methane (CH_4) is the principal component of natural gas. How many moles of CH_4 are present in 6.07 g of CH_4?

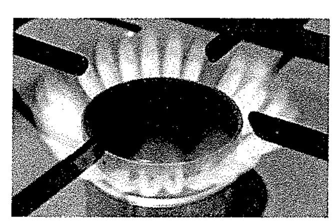


CH_4

Strategy We are given grams of CH_4 and asked to solve for moles of CH_4. What conversion factor do we need to convert between grams and moles? Arrange the appropriate conversion factor so that grams cancel and the unit moles are obtained for your answer.

Solution The conversion factor needed to convert between grams and moles is the molar mass. First we need to calculate the molar mass of CH_4, following the procedure in Example 3.5:

$$\begin{aligned}\text{molar mass of } CH_4 &= 12.01 \text{ g} + 4(1.008 \text{ g})\\ &= 16.04 \text{ g}\end{aligned}$$



Methane gas burning on a cooking range.

©Steve Allen/Getty Images

Because

$$1 \text{ mol } CH_4 = 16.04 \text{ g } CH_4$$

the conversion factor we need should have grams in the denominator so that the unit g will cancel, leaving the unit mol in the numerator:

$$\frac{1 \text{ mol } CH_4}{16.04 \text{ g } CH_4}$$

(Continued)

We now write

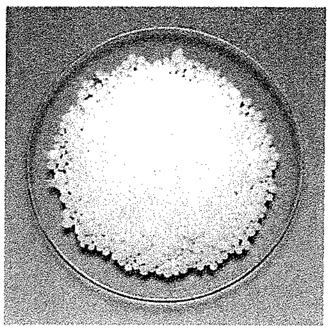
$$6.07\ \cancel{g\ CH_4} \times \frac{1\ mol\ CH_4}{16.04\ \cancel{g\ CH_4}} = 0.378\ mol\ CH_4$$

Thus, there is 0.378 mole of CH_4 in 6.07 g of CH_4.

Check Should 6.07 g of CH_4 equal less than 1 mole of CH_4? What is the mass of 1 mole of CH_4?

Practice Exercise Calculate the number of moles of chloroform ($CHCl_3$) in 198 g of chloroform.

Similar problem: 3.26.



Urea.
©McGraw-Hill Education/Ken Karp

Example 3.7

How many hydrogen atoms are present in 25.6 g of urea [$(NH_2)_2CO$], which is used as a fertilizer, in animal feed, and in the manufacture of polymers? The molar mass of urea is 60.06 g.

Strategy We are asked to solve for atoms of hydrogen in 25.6 g of urea. We cannot convert directly from grams of urea to atoms of hydrogen. How should molar mass and Avogadro's number be used in this calculation? How many moles of H are in 1 mole of urea?

Solution To calculate the number of H atoms, we first must convert grams of urea to moles of urea using the molar mass of urea. This part is similar to Example 3.2. The molecular formula of urea shows there are four moles of H atoms in one mole of urea molecule, so the mole ratio is 4:1. Finally, knowing the number of moles of H atoms, we can calculate the number of H atoms using Avogadro's number. We need two conversion factors: molar mass and Avogadro's number. We can combine these conversions

$$\text{grams of urea} \longrightarrow \text{moles of urea} \longrightarrow \text{moles of H} \longrightarrow \text{atoms of H}$$

into one step:

$$25.6\ \cancel{g\ (NH_2)_2CO} \times \frac{1\ \cancel{mol\ (NH_2)_2CO}}{60.06\ \cancel{g\ (NH_2)_2CO}} \times \frac{4\ \cancel{mol\ H}}{1\ \cancel{mol\ (NH_2)_2CO}} \times \frac{6.022 \times 10^{23}\ H\ atoms}{1\ \cancel{mol\ H}}$$
$$= 1.03 \times 10^{24}\ H\ atoms$$

Check Does the answer look reasonable? How many atoms of H would 60.06 g of urea contain?

Practice Exercise How many H atoms are in 72.5 g of isopropanol (rubbing alcohol), C_3H_8O?

Similar problems: 3.27, 3.28.

Note that the combined mass of a Na^+ ion and a Cl^- ion is equal to the combined mass of a Na atom and a Cl atom.

Finally, note that for ionic compounds like NaCl and MgO that do not contain discrete molecular units, we use the term *formula mass* instead. The formula unit of NaCl consists of one Na^+ ion and one Cl^- ion. Thus, the formula mass of NaCl is the mass of one formula unit:

$$\text{formula mass of NaCl} = 22.99\ \text{amu} + 35.45\ \text{amu}$$
$$= 58.44\ \text{amu}$$

and its molar mass is 58.44 g.

Review of Concepts & Facts

3.3.1 Determine the molecular mass and the molar mass of citric acid, $H_3C_6H_5O_7$.
3.3.2 What is the mass in grams of 0.382 moles of caffeine, $C_8H_{10}O_2N_4$?
3.3.3 How many oxygen atoms are in 124 g of calcium phosphate, $Ca_3(PO_4)_2$ (molar mass = 310.2 g)?

3.4 The Mass Spectrometer

The most direct and accurate method for determining atomic and molecular masses is mass spectrometry, depicted in Figure 3.3. In one type of a *mass spectrometer,* a gaseous sample is bombarded by a stream of high-energy electrons. Collisions between the electrons and the gaseous atoms (or molecules) produce positive ions by dislodging an electron from each atom or molecule. These positive ions (of mass m and charge e) are accelerated by two oppositely charged plates as they pass through the plates. The emerging ions are deflected into a circular path by a magnet. The radius of the path depends on the charge-to-mass ratio (that is, e/m). Ions of smaller e/m ratio trace a wider curve than those having a larger e/m ratio, so that ions with equal charges but different masses are separated from one another. The mass of each ion (and hence its parent atom or molecule) is determined from the magnitude of its deflection. Eventually the ions arrive at the detector, which registers a current for each type of ion. The amount of current generated is directly proportional to the number of ions, so it enables us to determine the relative abundance of isotopes.

The first mass spectrometer, developed in the 1920s by the English physicist F. W. Aston,† was crude by today's standards. Nevertheless, it provided indisputable evidence of the existence of isotopes—neon-20 (atomic mass 19.9924 amu and natural abundance 90.92 percent) and neon-22 (atomic mass 21.9914 amu and natural abundance 8.82 percent). When more sophisticated and sensitive mass spectrometers became available, scientists were surprised to discover that neon has a third stable isotope with an atomic mass of 20.9940 amu and natural abundance 0.257 percent (Figure 3.4). This

Note that it is possible to determine the molar mass of a compound without knowing its chemical formula.

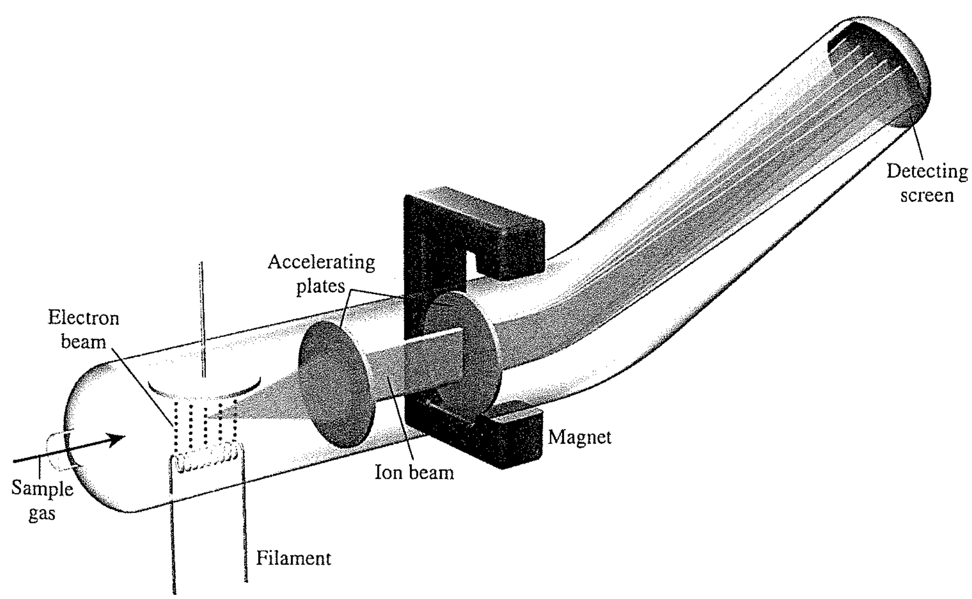


Figure 3.3 *Schematic diagram of one type of mass spectrometer.*

†Francis William Aston (1877–1945). English chemist and physicist. He was awarded the Nobel Prize in Chemistry in 1922 for developing the mass spectrometer.

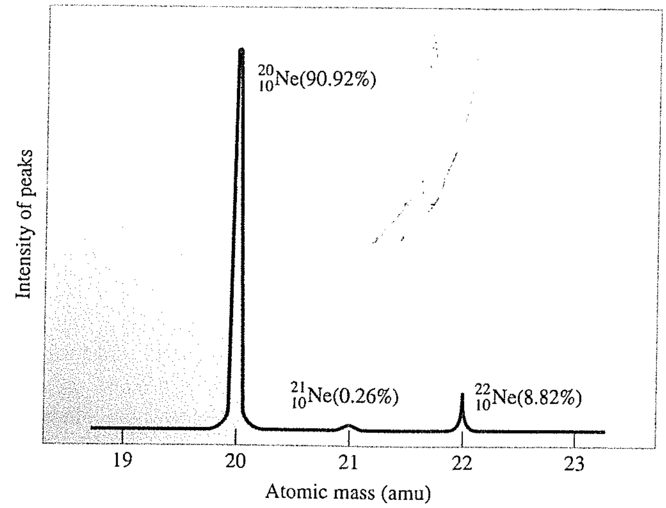


Figure 3.4 *The mass spectrum of the three isotopes of neon.*

Applying Practices
Mass spectrometry is a powerful tool. To learn more about how scientists use this tool, go online to access the inquiry activity *Distant Moons*.

example illustrates how very important experimental accuracy is to a quantitative science like chemistry. Early experiments failed to detect neon-21 because its natural abundance is just 0.257 percent. In other words, only 26 in 10,000 Ne atoms are neon-21. The masses of molecules can be determined in a similar manner by the mass spectrometer.

Review of Concepts & Facts

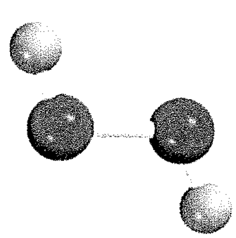
3.4.1 Explain how the mass spectrometer enables chemists to determine the average atomic mass of chlorine, which has two stable isotopes (^{35}Cl and ^{37}Cl).

3.5 Percent Composition of Compounds

As we have seen, the formula of a compound tells us the numbers of atoms of each element in a unit of the compound. However, suppose we needed to verify the purity of a compound for use in a laboratory experiment. From the formula we could calculate what percent of the total mass of the compound is contributed by each element. Then, by comparing the result to the percent composition obtained experimentally for our sample, we could determine the purity of the sample.

The ***percent composition by mass*** is the *percent by mass of each element in a compound.* Percent composition is obtained by dividing the mass of each element in 1 mole of the compound by the molar mass of the compound and multiplying by 100 percent. Mathematically, the percent composition of an element in a compound is expressed as

$$\text{percent composition of an element} = \frac{n \times \text{molar mass of element}}{\text{molar mass of compound}} \times 100\% \qquad (3.1)$$



H_2O_2

where n is the number of moles of the element in 1 mole of the compound. For example, in 1 mole of hydrogen peroxide (H_2O_2) there are 2 moles of H atoms and 2 moles of O atoms. The molar masses of H_2O_2, H, and O are 34.02 g, 1.008 g, and 16.00 g, respectively. Therefore, the percent composition of H_2O_2 is calculated as follows:

$$\%H = \frac{2 \times 1.008 \text{ g H}}{34.02 \text{ g } H_2O_2} \times 100\% = 5.926\%$$

$$\%O = \frac{2 \times 16.00 \text{ g O}}{34.02 \text{ g } H_2O_2} \times 100\% = 94.06\%$$

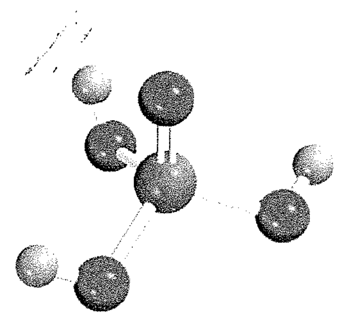
The sum of the percentages is 5.926% + 94.06% = 99.99%. The small discrepancy from 100 percent is due to the way we rounded off the molar masses of the elements. If we had used the empirical formula HO for the calculation, we would have obtained the same percentages. This is so because both the molecular formula and empirical formula tell us the percent composition by mass of the compound.

Student Hot Spot

Student data indicate you may struggle with percent composition by mass. Access your eBook for additional Learning Resources on this topic.

Example 3.8

Phosphoric acid (H_3PO_4) is a colorless, syrupy liquid used in detergents, fertilizers, toothpastes, and in carbonated beverages for a "tangy" flavor. Calculate the percent composition by mass of H, P, and O in this compound.



H_3PO_4

Strategy Recall the procedure for calculating a percentage. Assume that we have 1 mole of H_3PO_4. The percent by mass of each element (H, P, and O) is given by the combined molar mass of the atoms of the element in 1 mole of H_3PO_4 divided by the molar mass of H_3PO_4, then multiplied by 100 percent.

Solution The molar mass of H_3PO_4 is 97.99 g. The percent by mass of each of the elements in H_3PO_4 is calculated as follows:

$$\%\text{H} = \frac{3(1.008\text{ g})\text{ H}}{97.99\text{ g H}_3\text{PO}_4} \times 100\% = 3.086\%$$

$$\%\text{P} = \frac{30.97\text{ g P}}{97.99\text{ g H}_3\text{PO}_4} \times 100\% = 31.61\%$$

$$\%\text{O} = \frac{4(16.00\text{ g})\text{ O}}{97.99\text{ g H}_3\text{PO}_4} \times 100\% = 65.31\%$$

Check Do the percentages add to 100 percent? The sum of the percentages is (3.086% + 31.61% + 65.31%) = 100.01%. The small discrepancy from 100 percent is due to the way we rounded off.

Practice Exercise Calculate the percent composition by mass of each of the elements in sulfuric acid (H_2SO_4).

Similar problem: 3.38.

Student Hot Spot

Student data indicate you may struggle with determining empirical formulas. Access your eBook for additional Learning Resources on this topic.

The procedure used in the example can be reversed if necessary. Given the percent composition by mass of a compound, we can determine the empirical formula of the compound (Figure 3.5). Because we are dealing with percentages and the sum of all the percentages is 100 percent, it is convenient to assume that we started with 100 g of a compound, as Example 3.9 shows.

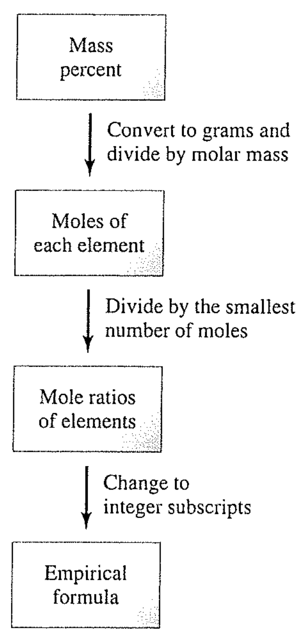


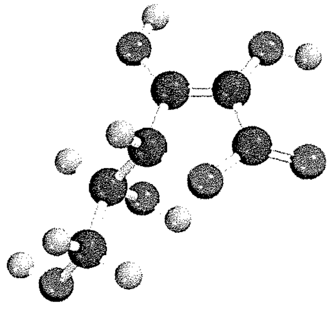
Figure 3.5 *Procedure for calculating the empirical formula of a compound from its percent compositions.*

Example 3.9

Ascorbic acid (vitamin C) cures scurvy. It is composed of 40.92 percent carbon (C), 4.58 percent hydrogen (H), and 54.50 percent oxygen (O) by mass. Determine its empirical formula.

Strategy In a chemical formula, the subscripts represent the ratio of the number of moles of each element that combine to form one mole of the compound. How can we convert from mass percent to moles? If we assume an exactly 100-g sample of the compound, do we know the mass of each element in the compound? How do we then convert from grams to moles?

(Continued)



The molecular formula of ascorbic acid is $C_6H_8O_6$.

Solution If we have 100 g of ascorbic acid, then each percentage can be converted directly to grams. In this sample, there will be 40.92 g of C, 4.58 g of H, and 54.50 g of O. Because the subscripts in the formula represent a mole ratio, we need to convert the grams of each element to moles. The conversion factor needed is the molar mass of each element. Let n represent the number of moles of each element so that

$$n_C = 40.92\ \cancel{g\ C} \times \frac{1\ \text{mol C}}{12.01\ \cancel{g\ C}} = 3.407\ \text{mol C}$$

$$n_H = 4.58\ \cancel{g\ H} \times \frac{1\ \text{mol H}}{1.008\ \cancel{g\ H}} = 4.54\ \text{mol H}$$

$$n_O = 54.50\ \cancel{g\ O} \times \frac{1\ \text{mol O}}{16.00\ \cancel{g\ O}} = 3.406\ \text{mol O}$$

Thus, we arrive at the formula $C_{3.407}H_{4.54}O_{3.406}$, which gives the identity and the mole ratios of atoms present. However, chemical formulas are written with whole numbers. Try to convert to whole numbers by dividing all the subscripts by the smallest subscript (3.406):

$$\text{C:}\ \frac{3.407}{3.406} \approx 1 \qquad \text{H:}\ \frac{4.54}{3.406} = 1.33 \qquad \text{O:}\ \frac{3.406}{3.406} = 1$$

where the ≈ sign means "approximately equal to." This gives $CH_{1.33}O$ as the formula for ascorbic acid. Next, we need to convert 1.33, the subscript for H, into an integer. This can be done by a trial-and-error procedure:

$$1.33 \times 1 = 1.33$$
$$1.33 \times 2 = 2.66$$
$$1.33 \times 3 = 3.99 \approx 4$$

Because 1.33×3 gives us an integer (4), we multiply all the subscripts by 3 and obtain $C_3H_4O_3$ as the empirical formula for ascorbic acid.

Check Are the subscripts in $C_3H_4O_3$ reduced to the smallest whole numbers?

Practice Exercise Determine the empirical formula of a compound having the following percent composition by mass: K: 24.75 percent; Mn: 34.77 percent; O: 40.51 percent.

Similar problems: 3.43, 3.44.

Chemists often want to know the actual mass of an element in a certain mass of a compound. For example, in the mining industry, this information will tell the scientists about the quality of the ore. Because the percent composition by mass of the elements in the substance can be readily calculated, such a problem can be solved in a rather direct way.

Chalcopyrite.

Example 3.10

Chalcopyrite ($CuFeS_2$) is a principal mineral of copper. Calculate the number of kilograms of Cu in 3.71×10^3 kg of chalcopyrite.

Strategy Chalcopyrite is composed of Cu, Fe, and S. The mass due to Cu is based on its percentage by mass in the compound. How do we calculate mass percent of an element?

Solution The molar masses of Cu and $CuFeS_2$ are 63.55 g and 183.5 g, respectively. The mass percent of Cu is therefore

$$\%\text{Cu} = \frac{\text{molar mass of Cu}}{\text{molar mass of } CuFeS_2} \times 100\%$$
$$= \frac{63.55\ \text{g}}{183.5\ \text{g}} \times 100\% = 34.63\%$$

(Continued)

To calculate the mass of Cu in a 3.71×10^3 kg sample of $CuFeS_2$, we need to convert the percentage to a fraction (that is, convert 34.63 percent to 34.63/100, or 0.3463) and write

$$\text{mass of Cu in } CuFeS_2 = 0.3463 \times (3.71 \times 10^3 \text{ kg}) = 1.28 \times 10^3 \text{ kg}$$

Check As a ball-park estimate, note that the mass percent of Cu is roughly 33 percent, so that a third of the mass should be Cu—that is, $\frac{1}{3} \times 3.71 \times 10^3 \text{ kg} \approx 1.24 \times 10^3$ kg. This quantity is quite close to the answer.

Practice Exercise Calculate the number of grams of Al in 371 g of Al_2O_3.

Similar problem: 3.41.

Review of Concepts & Facts

3.5.1 Without doing detailed calculations, estimate whether the percent composition by mass of Sr is greater than or smaller than that of O in strontium nitrate [$Sr(NO_3)_2$].

3.5.2 What is the percent composition by mass of each of the elements present in fumaric acid, $C_4H_4O_4$?

3.5.3 Determine the empirical formula of a compound having the following percent composition by mass: P: 43.64 percent; O: 56.36 percent.

3.6 Experimental Determination of Empirical Formulas

The fact that we can determine the empirical formula of a compound if we know the percent composition enables us to identify compounds experimentally. The procedure is as follows. First, chemical analysis tells us the number of grams of each element present in a given amount of a compound. Then, we convert the quantities in grams to number of moles of each element. Finally, using the method given in Example 3.9, we find the empirical formula of the compound.

As a specific example, let us consider the compound ethanol. When ethanol is burned in an apparatus such as that shown in Figure 3.6, carbon dioxide (CO_2) and water (H_2O) are given off. Because neither carbon nor hydrogen was in the inlet gas, we can conclude that both carbon (C) and hydrogen (H) were present in ethanol and that oxygen (O) may also be present. (Molecular oxygen was added in the combustion process, but some of the oxygen may also have come from the original ethanol sample.)

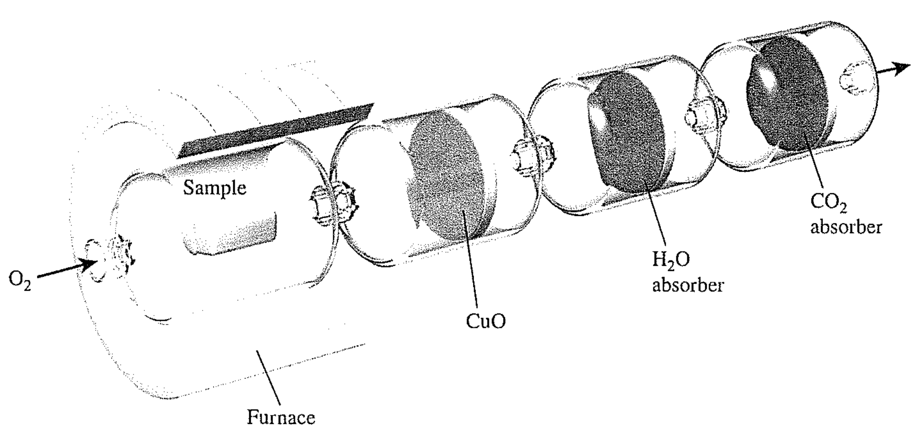


Figure 3.6 *Apparatus for determining the empirical formula of ethanol. The absorbers are substances that can retain water and carbon dioxide, respectively. CuO is used to ensure complete combustion of all carbon to CO_2.*

The masses of CO_2 and of H_2O produced can be determined by measuring the ir crease in mass of the CO_2 and H_2O absorbers, respectively. Suppose that in one exper ment the combustion of 11.5 g of ethanol produced 22.0 g of CO_2 and 13.5 g of H_2(We can calculate the mass of carbon and hydrogen in the original 11.5-g sample c ethanol as follows:

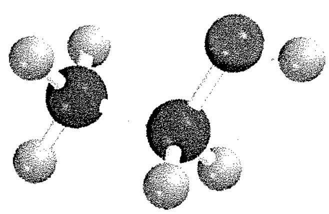
$$\begin{aligned}\text{mass of C} &= 22.0\ \text{g}\ \cancel{CO_2} \times \frac{1\ \cancel{\text{mol}\ CO_2}}{44.01\ \cancel{\text{g}\ CO_2}} \times \frac{1\ \cancel{\text{mol C}}}{1\ \cancel{\text{mol}\ CO_2}} \times \frac{12.01\ \text{g C}}{1\ \cancel{\text{mol C}}}\\ &= 6.00\ \text{g C}\\ \text{mass of H} &= 13.5\ \cancel{\text{g}\ H_2O} \times \frac{1\ \text{mol}\ \cancel{H_2O}}{18.02\ \cancel{\text{g}\ H_2O}} \times \frac{2\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol}\ H_2O}} \times \frac{1.008\ \text{g H}}{1\ \cancel{\text{mol H}}}\\ &= 1.51\ \text{g H}\end{aligned}$$

Thus, 11.5 g of ethanol contains 6.00 g of carbon and 1.51 g of hydrogen. The remain der must be oxygen, whose mass is

$$\begin{aligned}\text{mass of O} &= \text{mass of sample} - (\text{mass of C} + \text{mass of H})\\ &= 11.5\ \text{g} - (6.00\ \text{g} + 1.51\ \text{g})\\ &= 4.0\ \text{g}\end{aligned}$$

The number of moles of each element present in 11.5 g of ethanol is

$$\begin{aligned}\text{moles of C} &= 6.00\ \text{g}\ \cancel{\text{C}} \times \frac{1\ \text{mol C}}{12.01\ \text{g}\ \cancel{\text{C}}} = 0.500\ \text{mol C}\\ \text{moles of H} &= 1.51\ \text{g}\ \cancel{\text{H}} \times \frac{1\ \text{mol H}}{1.008\ \text{g}\ \cancel{\text{H}}} = 1.50\ \text{mol H}\\ \text{moles of O} &= 4.00\ \text{g}\ \cancel{\text{O}} \times \frac{1\ \text{mol O}}{16.00\ \text{g}\ \cancel{\text{O}}} = 0.25\ \text{mol O}\end{aligned}$$



It happens that the molecular formula of ethanol is the same as its empirical formula.

The formula of ethanol is therefore $C_{0.50}H_{1.5}O_{0.25}$ (we round off the number of moles tc two significant figures). Because the number of atoms must be an integer, we divide the subscripts by 0.25, the smallest subscript, and obtain for the empirical formula C_2H_6O.

Now we can better understand the word "empirical," which literally means "basec only on observation and measurement." The empirical formula of ethanol is determinec from analysis of the compound in terms of its component elements. No knowledge oi how the atoms are linked together in the compound is required.

Determination of Molecular Formulas

The formula calculated from percent composition by mass is always the empirical formula, because the subscripts in the formula are always reduced to the smallest whole numbers. To calculate the actual, molecular formula we must know the *approximate* molar mass of the compound in addition to its empirical formula. Knowing that the molar mass of a compound must be an integral multiple of the molar mass of its empirical formula, we can use the molar mass to find the molecular formula, as Example 3.11 demonstrates.

Student Hot Spot

Student data indicate you may struggle with combustion analysis and determining molecular formulas. Access your eBook for additional Learning Resources on this topic.

Example 3.11

A sample of a compound contains 30.46 percent nitrogen and 69.54 percent oxygen by mass, as determined by a mass spectrometer. In a separate experiment, the molar mass of the compound is found to be between 90 g and 95 g. Determine the molecular formula and the accurate molar mass of the compound.

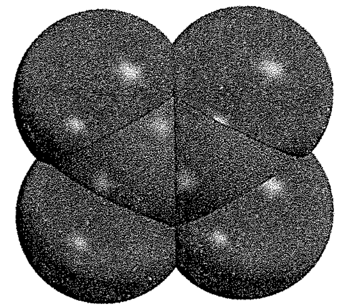
(Continued)

Strategy To determine the molecular formula, we first need to determine the empirical formula. Comparing the empirical molar mass to the experimentally determined molar mass will reveal the relationship between the empirical formula and molecular formula.

Solution We start by assuming that there are 100 g of the compound. Then each percentage can be converted directly to grams—that is, 30.46 g of N and 69.54 g of O. Let n represent the number of moles of each element so that

$$n_N = 30.46\ \cancel{g\ N} \times \frac{1\ \text{mol N}}{14.01\ \cancel{g\ N}} = 2.174\ \text{mol N}$$

$$n_O = 69.54\ \cancel{g\ O} \times \frac{1\ \text{mol O}}{16.00\ \cancel{g\ O}} = 4.346\ \text{mol O}$$



N_2O_4

Thus, we arrive at the formula $N_{2.174}O_{4.346}$, which gives the identity and the ratios of atoms present. However, chemical formulas are written with whole numbers. Try to convert to whole numbers by dividing the subscripts by the smaller subscript (2.174). After rounding off, we obtain NO_2 as the empirical formula.

The molecular formula might be the same as the empirical formula or some integral multiple of it (for example, two, three, four, or more times the empirical formula). Comparing the ratio of the molar mass to the molar mass of the empirical formula will show the integral relationship between the empirical and molecular formulas. The molar mass of the empirical formula NO_2 is

$$\text{empirical molar mass} = 14.01\ \text{g} + 2(16.0\ \text{g}) = 46.01\ \text{g}$$

Next, we determine the ratio between the molar mass and the empirical molar mass

$$\frac{\text{molar mass}}{\text{empirical molar mass}} = \frac{90\ \text{g}}{46.01\ \text{g}} \approx 2$$

The molar mass is twice the empirical molar mass. This means that there are two NO_2 units in each molecule of the compound, and the molecular formula is $(NO_2)_2$ or N_2O_4.

The actual molar mass of the compound is two times the empirical molar mass—that is, 2(46.01 g) or 92.02 g—which is between 90 g and 95 g.

Check Note that in determining the molecular formula from the empirical formula, we need only know the *approximate* molar mass of the compound. The reason is that the true molar mass is an integral multiple (1×, 2×, 3×, . . .) of the empirical molar mass. Therefore, the ratio (molar mass/empirical molar mass) will always be close to an integer.

Practice Exercise A sample of a compound containing boron (B) and hydrogen (H) contains 6.444 g of B and 1.803 g of H. The molar mass of the compound is about 30 g. What is its molecular formula?

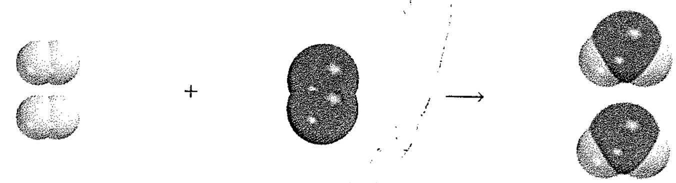
Similar problems: 3.48, 3.49, 3.52.

Review of Concepts & Facts

3.6.1 What is the molecular formula of a compound containing only carbon and hydrogen if combustion of 1.05 g of the compound produces 3.30 g CO_2 and 1.35 g H_2O and its molar mass is about 70 g?

3.7 Chemical Reactions and Chemical Equations

Having discussed the masses of atoms and molecules, we turn next to what happens to atoms and molecules in a ***chemical reaction,*** *a process in which a substance (or substances) is changed into one or more new substances.* To communicate with one another about chemical reactions, chemists have devised a standard way to represent them



Two hydrogen molecules	+	One oxygen molecule	$\longrightarrow$	Two water molecules
$2H_2$	+	O_2	$\longrightarrow$	$2H_2O$
2 moles H_2	+	1 mole O_2	$\longrightarrow$	2 moles H_2O
$2(2.02\ g) = 4.04\ g\ H_2$	+	$32.00\ g\ O_2$	$\longrightarrow$	$2(18.02\ g) = 36.04\ g\ H_2O$
36.04 g reactants			$\longrightarrow$	36.04 g product

Figure 3.7 *Three ways of representing the combustion of hydrogen. In accordance with the law of conservation of mass, the number of each type of atom must be the same on both sides of the equatic*

using chemical equations. A ***chemical equation*** *uses chemical symbols to show wl happens during a chemical reaction.* In this section, we will learn how to write cher cal equations and balance them.

Writing Chemical Equations

Consider what happens when hydrogen gas (H_2) burns in air (which contains oxyge O_2) to form water (H_2O). This reaction can be represented by the chemical equation

$$H_2 + O_2 \longrightarrow H_2O \quad (3.$$

where the "plus" sign means "reacts with" and the arrow means "to yield." Thus, tl symbolic expression can be read: "Molecular hydrogen reacts with molecular oxygen yield water." The reaction is assumed to proceed from left to right as the arrow indicate

We use the law of conservation of mass as our guide in balancing chemical equations.

Equation (3.2) is not complete, however, because there are twice as many oxyg atoms on the left side of the arrow (two) as on the right side (one). To conform with tl law of conservation of mass, there must be the same number of each type of atom both sides of the arrow—that is, we must have as many atoms after the reaction ends we did before it started. We can *balance* Equation (3.2) by placing the appropriate coe ficient (2 in this case) in front of H_2 and H_2O:

When the coefficient is 1, as in the case of O_2, it is not shown.

$$2H_2 + O_2 \longrightarrow 2H_2O$$

This *balanced chemical equation* shows that "two hydrogen molecules can combine react with one oxygen molecule to form two water molecules" (Figure 3.7). Because tl ratio of the number of molecules is equal to the ratio of the number of moles, the equ tion can also be read as "2 moles of hydrogen molecules react with 1 mole of oxyge molecules to produce 2 moles of water molecules." We know the mass of a mole each of these substances, so we can also interpret the equation as "4.04 g of H_2 rea with 32.00 g of O_2 to give 36.04 g of H_2O." These three ways of reading the equatio are summarized in Figure 3.7.

We refer to H_2 and O_2 in Equation (3.2) as ***reactants,*** which are *the starting mater als in a chemical reaction.* Water is the ***product,*** which is *the substance formed as result of a chemical reaction.* A chemical equation, then, is just the chemist's shorthan description of a reaction. In a chemical equation, the reactants are conventionally wri ten on the left and the products on the right of the arrow:

$$\text{reactants} \longrightarrow \text{products}$$

To provide additional information, chemists often indicate the physical states of th reactants and products by using the letters *g, l,* and *s* to denote gas, liquid, and solic respectively. For example,

$$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$$
$$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$$

To represent what happens when sodium chloride (NaCl) is added to water, we write

$$NaCl(s) \xrightarrow{H_2O} NaCl(aq)$$

where *aq* denotes the aqueous (that is, water) environment. Writing H_2O above the arrow symbolizes the physical process of dissolving a substance in water, although it is sometimes left out for simplicity.

Knowing the states of the reactants and products is especially useful in the laboratory. For example, when potassium bromide (KBr) and silver nitrate ($AgNO_3$) react in an aqueous environment, a solid, silver bromide (AgBr), is formed. This reaction can be represented by the equation

$$KBr(aq) + AgNO_3(aq) \longrightarrow KNO_3(aq) + AgBr(s)$$

If the physical states of reactants and products are not given, an uninformed person might try to bring about the reaction by mixing solid KBr with solid $AgNO_3$. These solids would react very slowly or not at all. Imagining the process on the microscopic level, we can understand that for a product like silver bromide to form, the Ag^+ and Br^- ions would have to come in contact with each other. However, these ions are locked in place in their solid compounds and have little mobility. (Here is an example of how we explain a phenomenon by thinking about what happens at the molecular level, as discussed in Section 1.2.)

Student Hot Spot

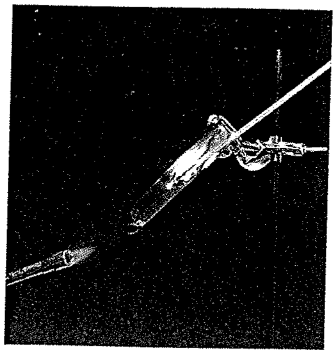
Student data indicate you may struggle with interpreting chemical equations. Access your eBook for additional Learning Resources on this topic.

Balancing Chemical Equations

Suppose we want to write an equation to describe a chemical reaction that we have just carried out in the laboratory. How should we go about doing it? Because we know the identities of the reactants, we can write their chemical formulas. The identities of products are more difficult to establish. For simple reactions it is often possible to guess the product(s). For more complicated reactions involving three or more products, chemists may need to perform further tests to establish the presence of specific compounds.

Once we have identified all the reactants and products and have written the correct formulas for them, we assemble them in the conventional sequence—reactants on the left separated by an arrow from products on the right. The equation written at this point is likely to be *unbalanced;* that is, the number of each type of atom on one side of the arrow differs from the number on the other side. In general, we can balance a chemical equation by the following steps:

1. Identify all reactants and products and write their correct formulas on the left side and right side of the equation, respectively.
2. Begin balancing the equation by trying different coefficients to make the number of atoms of each element the same on both sides of the equation. We can change the coefficients (the numbers preceding the formulas) but not the subscripts (the numbers within formulas). Changing the subscripts would change the identity of the substance. For example, $2NO_2$ means "two molecules of nitrogen dioxide," but if we double the subscripts, we have N_2O_4, which is the formula of dinitrogen tetroxide, a completely different compound.
3. First, look for elements that appear only once on each side of the equation with the same number of atoms on each side: The formulas containing these elements must have the same coefficient. Therefore, there is no need to adjust the coefficients of these elements at this point. Next, look for elements that appear only once on each side of the equation but in unequal numbers of atoms. Balance these elements. Finally, balance elements that appear in two or more formulas on the same side of the equation.
4. Check your balanced equation to be sure that you have the same total number of each type of atoms on both sides of the equation arrow.



Heating potassium chlorate produces oxygen, which supports the combustion of a wood splint.

Let's consider a specific example. In the laboratory, small amounts of oxyge can be prepared by heating potassium chlorate ($KClO_3$). The products are oxyge (O_2) and potassium chloride (KCl). From this information, we write

$$KClO_3 \longrightarrow KCl + O_2$$

(For simplicity, we omit the physical states of reactants and products.) All three ments (K, Cl, and O) appear only once on each side of the equation, but only for I Cl do we have equal numbers of atoms on both sides. Thus, $KClO_3$ and KCl must the same coefficient. The next step is to make the number of O atoms the same on sides of the equation. Because there are three O atoms on the left and two O aton the right of the equation, we can balance the O atoms by placing a 2 in front of K and a 3 in front of O_2.

$$2KClO_3 \longrightarrow KCl + 3O_2$$

Finally, we balance the K and Cl atoms by placing a 2 in front of KCl:

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

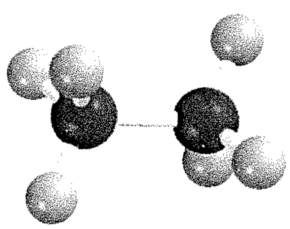
As a final check, we can draw up a balance sheet for the reactants and products w the number in parentheses indicates the number of atoms of each element:

Reactants	Products
K (2)	K (2)
Cl (2)	Cl (2)
O (6)	O (6)

Note that this equation could also be balanced with coefficients that are multiples (for $KClO_3$), 2 (for KCl), and 3 (for O_2); for example,

$$4KClO_3 \longrightarrow 4KCl + 6O_2$$

However, it is common practice to use the *simplest* possible set of whole-number c ficients to balance the equation. Equation (3.3) conforms to this convention.



C_2H_6

Now let us consider the combustion (that is, burning) of the natural gas compor ethane (C_2H_6) in oxygen or air, which yields carbon dioxide (CO_2) and water. The balanced equation is

$$C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$$

We see that the number of atoms is not the same on both sides of the equation for of the elements (C, H, and O). In addition, C and H appear only once on each side of equation; O appears in two compounds on the right side (CO_2 and H_2O). To balance C atoms, we place a 2 in front of CO_2:

$$C_2H_6 + O_2 \longrightarrow 2CO_2 + H_2O$$

To balance the H atoms, we place a 3 in front of H_2O:

$$C_2H_6 + O_2 \longrightarrow 2CO_2 + 3H_2O$$

At this stage, the C and H atoms are balanced, but the O atoms are not because there seven O atoms on the right-hand side and only two O atoms on the left-hand side of t equation. This inequality of O atoms can be eliminated by writing $\frac{7}{2}$ in front of the O_2 on the left-hand side:

$$C_2H_6 + \tfrac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$$

The "logic" for using $\frac{7}{2}$ as a coefficient is that there were seven oxygen atoms on the right-hand side of the equation, but only a pair of oxygen atoms (O_2) on the left. To balance them we ask how many *pairs* of oxygen atoms are needed to equal seven oxygen atoms. Just as 3.5 pairs of shoes equal seven shoes, $\frac{7}{2}O_2$ molecules equal seven O atoms. As the following tally shows, the equation is now balanced.

Reactants	**Products**
C (2)	C (2)
H (6)	H (6)
O (7)	O (7)

However, we normally prefer to express the coefficients as whole numbers rather than as fractions. Therefore, we multiply the entire equation by 2 to convert $\frac{7}{2}$ to 7:

$$2C_2H_6 + 7O_2 \longrightarrow 4CO_2 + 6H_2O$$

The final tally is

Reactants	**Products**
C (4)	C (4)
H (12)	H (12)
O (14)	O (14)

Note that the coefficients used in balancing the last equation are the smallest possible set of whole numbers.

In Example 3.12 we will continue to practice our equation-balancing skills.

Example 3.12

When aluminum metal is exposed to air, a protective layer of aluminum oxide (Al_2O_3) forms on its surface. This layer prevents further reaction between aluminum and oxygen, and it is the reason that aluminum beverage cans do not corrode. [In the case of iron, the rust, or iron(III) oxide, that forms is too porous to protect the iron metal underneath, so rusting continues.] Write a balanced equation for the formation of Al_2O_3.

Strategy Remember that the formula of an element or compound cannot be changed when balancing a chemical equation. The equation is balanced by placing the appropriate coefficients in front of the formulas. Follow the procedure for balancing equations described earlier.

Solution The unbalanced equation is

$$Al + O_2 \longrightarrow Al_2O_3$$

In a balanced equation, the number and types of atoms on each side of the equation must be the same. We see that there is one Al atom on the reactants side and there are two Al atoms on the product side. We can balance the Al atoms by placing a coefficient of 2 in front of Al on the reactants side.

$$2Al + O_2 \longrightarrow Al_2O_3$$

(Continued)

There are two O atoms on the reactants side, and three O atoms on the product side of the equation. We can balance the O atoms by placing a coefficient of $\frac{3}{2}$ in front of O_2 on the reactants side.

$$2Al + \tfrac{3}{2}O_2 \longrightarrow Al_2O_3$$

This is a balanced equation. However, equations are normally balanced with the smallest set of *whole*-number coefficients. Multiplying both sides of the equation by 2 gives whole-number coefficients.

$$2(2Al + \tfrac{3}{2}O_2 \longrightarrow Al_2O_3)$$

or

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

Check For an equation to be balanced, the number and types of atoms on each side of the equation must be the same. The final tally is

Reactants	Products
Al (4)	Al (4)
O (6)	O (6)

The equation is balanced. Also, the coefficients are reduced to the simplest set of whole numbers.

Practice Exercise Balance the equation representing the reaction between iron(III) oxide (Fe_2O_3) and carbon monoxide (CO) to yield iron (Fe) and carbon dioxide (CO_2).

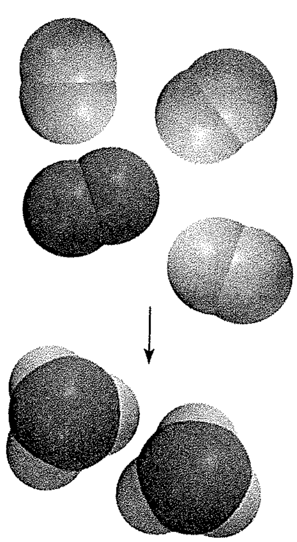
Similar problems: 3.55, 3.56.

Review of Concepts & Facts

3.7.1 Which parts of the equation shown here are essential for a balanced equation and which parts are helpful if we want to carry out the reaction in the laboratory?

$$BaH_2(s) + 2H_2O(l) \longrightarrow Ba(OH)_2(aq) + 2H_2(g)$$

3.7.2 What is the balanced equation representing the reaction between potassium thiosulfate ($K_2S_2O_3$) and iodine (I_2) to produce potassium tetrathionate ($K_2S_4O_6$) and potassium iodide (KI)?



The synthesis of NH_3 from H_2 and N_2.

3.8 Amounts of Reactants and Products

A basic question raised in the chemical laboratory is, "How much product will be formed from specific amounts of starting materials (reactants)?" Or in some cases, we might ask the reverse question, "How much starting material must be used to obtain a specific amount of product?" To interpret a reaction quantitatively, we need to apply our knowledge of molar masses and the mole concept. ***Stoichiometry*** is *the quantitative study of reactants and products in a chemical reaction.*

Whether the units given for reactants (or products) are moles, grams, liters (for gases), or some other units, we use moles to calculate the amount of product formed in a reaction. This approach is called the ***mole method,*** which means simply that *the stoichiometric coefficients in a chemical equation can be interpreted as the number of moles of each substance.* For example, industrially ammonia is synthesized from hydrogen and nitrogen as follows:

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

The stoichiometric coefficients show that one molecule of N_2 reacts with three molecules of H_2 to form two molecules of NH_3. It follows that the relative numbers of moles are the same as the relative number of molecules:

$N_2(g)$	+	$3H_2(g)$	$\longrightarrow$	$2NH_3(g)$
1 molecule		3 molecules		2 molecules
6.022×10^{23} molecules		$3(6.022 \times 10^{23}$ molecules)		$2(6.022 \times 10^{23}$ molecules)
1 mol		3 mol		2 mol

Thus, this equation can also be read as "1 mole of N_2 gas combines with 3 moles of H_2 gas to form 2 moles of NH_3 gas." In stoichiometric calculations, we say that three moles of H_2 are equivalent to two moles of NH_3, that is,

$$3 \text{ mol } H_2 \simeq 2 \text{ mol } NH_3$$

where the symbol $\simeq$ means "stoichiometrically equivalent to" or simply "equivalent to." This relationship enables us to write the conversion factors

$$\frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3} \quad \text{and} \quad \frac{2 \text{ mol } NH_3}{3 \text{ mol } H_2}$$

Similarly, we have 1 mol $N_2 \simeq 2$ mol NH_3 and 1 mol $N_2 \simeq 3$ mol H_2.

Let's consider a simple example in which 6.0 moles of H_2 react completely with N_2 to form NH_3. To calculate the amount of NH_3 produced in moles, we use the conversion factor that has H_2 in the denominator and write

$$\begin{aligned}\text{moles of } NH_3 \text{ produced} &= 6.0 \cancel{\text{mol } H_2} \times \frac{2 \text{ mol } NH_3}{3 \cancel{\text{mol } H_2}} \\ &= 4.0 \text{ mol } NH_3\end{aligned}$$

Now suppose 16.0 g of H_2 react completely with N_2 to form NH_3. How many grams of NH_3 will be formed? To do this calculation, we note that the link between H_2 and NH_3 is the mole ratio from the balanced equation. So we need to first convert grams of H_2 to moles of H_2, then to moles of NH_3, and finally to grams of NH_3. The conversion steps are

$$\text{grams of } H_2 \longrightarrow \text{moles of } H_2 \longrightarrow \text{moles of } NH_3 \longrightarrow \text{grams of } NH_3$$

First, we convert 16.0 g of H_2 to number of moles of H_2, using the molar mass of H_2 as the conversion factor.

$$\begin{aligned}\text{moles of } H_2 &= 16.0 \cancel{\text{g } H_2} \times \frac{1 \text{ mol } H_2}{2.016 \cancel{\text{g } H_2}} \\ &= 7.94 \text{ mol } H_2\end{aligned}$$

Next, we calculate the number of moles of NH_3 produced.

$$\begin{aligned}\text{moles of } NH_3 &= 7.94 \cancel{\text{mol } H_2} \times \frac{2 \text{ mol } NH_3}{3 \cancel{\text{mol } H_2}} \\ &= 5.29 \text{ mol } NH_3\end{aligned}$$

Finally, we calculate the mass of NH_3 produced in grams using the molar mass of NH_3 as the conversion factor.

$$\begin{aligned}\text{grams of } NH_3 &= 5.29 \cancel{\text{mol } NH_3} \times \frac{17.03 \text{ g } NH_3}{1 \cancel{\text{mol } NH_3}} \\ &= 90.1 \text{ g } NH_3\end{aligned}$$

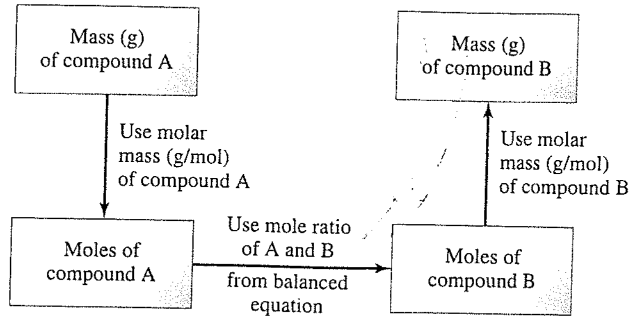


Figure 3.8 *The procedure for calculating the amounts of reactants or products in a reaction using the mole method.*

These three separate calculations can be combined in a single step as follows:

$$\text{grams of } NH_3 = 16.0\ \cancel{g\ H_2} \times \frac{1\ \cancel{mol\ H_2}}{2.016\ \cancel{g\ H_2}} \times \frac{2\ \cancel{mol\ NH_3}}{3\ \cancel{mol\ H_2}} \times \frac{17.03\ g\ NH_3}{1\ \cancel{mol\ NH_3}}$$
$$= 90.1\ g\ NH_3$$

Similarly, we can calculate the mass in grams of N_2 consumed in this reaction. The conversion steps are

$$\text{grams of } H_2 \longrightarrow \text{moles of } H_2 \longrightarrow \text{moles of } N_2 \longrightarrow \text{grams of } N_2$$

By using the relationship 1 mol $N_2 \simeq 3$ mol H_2, we write

$$\text{grams of } N_2 = 16.0\ \cancel{g\ H_2} \times \frac{1\ \cancel{mol\ H_2}}{2.016\ \cancel{g\ H_2}} \times \frac{1\ \cancel{mol\ N_2}}{3\ \cancel{mol\ H_2}} \times \frac{28.02\ g\ N_2}{1\ \cancel{mol\ N_2}}$$
$$= 74.1\ g\ N_2$$

The general approach for solving stoichiometry problems is summarized next.

1. Write a balanced equation for the reaction.
2. Convert the given amount of the reactant (in grams or other units) to number of moles.
3. Use the mole ratio from the balanced equation to calculate the number of moles of product formed.
4. Convert the moles of product to grams (or other units) of product.

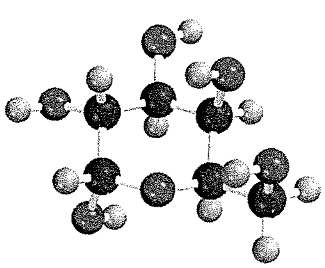
Student Hot Spot

Student data indicate you may struggle with stoichiometry conversions. Access your eBook for additional Learning Resources on this topic.

Figure 3.8 shows these steps. Sometimes we may be asked to calculate the amount of a reactant needed to form a specific amount of product. In those cases, we can reverse the steps shown in Figure 3.8.

Examples 3.13 and 3.14 illustrate the application of this approach.

Example 3.13



$C_6H_{12}O_6$

The food we eat is degraded, or broken down, in our bodies to provide energy for growth and function. A general overall equation for this very complex process represents the degradation of glucose ($C_6H_{12}O_6$) to carbon dioxide (CO_2) and water (H_2O):

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

If 856 g of $C_6H_{12}O_6$ is consumed by a person over a certain period, what is the mass of CO_2 produced?

(Continued)

Strategy Looking at the balanced equation, how do we compare the amounts of $C_6H_{12}O_6$ and CO_2? We can compare them based on the *mole ratio* from the balanced equation. Starting with grams of $C_6H_{12}O_6$, how do we convert to moles of $C_6H_{12}O_6$? Once moles of CO_2 are determined using the mole ratio from the balanced equation, how do we convert to grams of CO_2?

Solution We follow the preceding steps and Figure 3.8.

Step 1: The balanced equation is given in the problem.

Step 2: To convert grams of $C_6H_{12}O_6$ to moles of $C_6H_{12}O_6$, we write

$$856\ \cancel{g\ C_6H_{12}O_6} \times \frac{1\ mol\ C_6H_{12}O_6}{180.2\ \cancel{g\ C_6H_{12}O_6}} = 4.750\ mol\ C_6H_{12}O_6$$

Step 3: From the mole ratio, we see that 1 mol $C_6H_{12}O_6 \simeq 6$ mol CO_2. Therefore, the number of moles of CO_2 formed is

$$4.750\ \cancel{mol\ C_6H_{12}O_6} \times \frac{6\ mol\ CO_2}{1\ \cancel{mol\ C_6H_{12}O_6}} = 28.50\ mol\ CO_2$$

Step 4: Finally, the number of grams of CO_2 formed is given by

$$28.50\ \cancel{mol\ CO_2} \times \frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}} = 1.25 \times 10^3\ g\ CO_2$$

After some practice, we can combine the conversion steps

$$\text{grams of } C_6H_{12}O_6 \longrightarrow \text{moles of } C_6H_{12}O_6 \longrightarrow \text{moles of } CO_2 \longrightarrow \text{grams of } CO_2$$

into one equation:

$$\begin{aligned}\text{mass of } CO_2 &= 856\ \cancel{g\ C_6H_{12}O_6} \times \frac{1\ \cancel{mol\ C_6H_{12}O_6}}{180.2\ \cancel{g\ C_6H_{12}O_6}} \times \frac{6\ \cancel{mol\ CO_2}}{1\ \cancel{mol\ C_6H_{12}O_6}} \times \frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}} \\ &= 1.25 \times 10^3\ g\ CO_2\end{aligned}$$

Check Does the answer seem reasonable? Should the mass of CO_2 produced be larger than the mass of $C_6H_{12}O_6$ reacted, even though the molar mass of CO_2 is considerably less than the molar mass of $C_6H_{12}O_6$? What is the mole ratio between CO_2 and $C_6H_{12}O_6$?

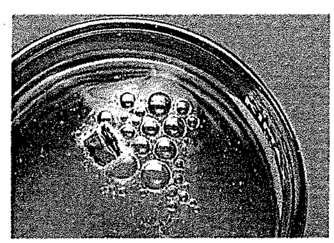
Practice Exercise Methanol (CH_3OH) burns in air according to the equation

$$2CH_3OH + 3O_2 \longrightarrow 2CO_2 + 4H_2O$$

If 209 g of methanol are used up in a combustion process, what is the mass of H_2O produced?

Similar problem: 3.72.

Example 3.14



Lithium reacting with water to produce hydrogen gas.

All alkali metals react with water to produce hydrogen gas and the corresponding alkali metal hydroxide. A typical reaction is that between lithium and water:

$$2Li(s) + 2H_2O(l) \longrightarrow 2LiOH(aq) + H_2(g)$$

How many grams of Li are needed to produce 9.89 g of H_2?

(Continued)

Strategy The question asks for number of grams of reactant (Li) to form a specific amount of product (H_2). Therefore, we need to reverse the steps shown in Figure 3.8. From the equation we see that 2 mol Li $\simeq$ 1 mol H_2.

Solution The conversion steps are

$$\text{grams of } H_2 \longrightarrow \text{moles of } H_2 \longrightarrow \text{moles of Li} \longrightarrow \text{grams of Li}$$

Combining these steps into one equation, we write

$$9.89\ \cancel{g\ H_2} \times \frac{1\ \cancel{mol\ H_2}}{2.016\ \cancel{g\ H_2}} \times \frac{2\ \cancel{mol\ Li}}{1\ \cancel{mol\ H_2}} \times \frac{6.941\ g\ Li}{1\ \cancel{mol\ Li}} = 68.1\ g\ Li$$

Check There are roughly 5 moles of H_2 in 9.89 g H_2, so we need 10 moles of Li. From the approximate molar mass of Li (7 g), does the answer seem reasonable?

Practice Exercise The reaction between nitric oxide (NO) and oxygen to form nitrogen dioxide (NO_2) is a key step in photochemical smog formation:

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

How many grams of O_2 are needed to produce 2.21 g of NO_2?

Similar problem: 3.64.

Review of Concepts & Facts

3.8.1 Which of the following statements is correct for the equation shown here?

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$

(a) 6 g of H_2O are produced for every 4 g of NH_3 reacted.
(b) 1 mole of NO is produced per mole of NH_3 reacted.
(c) 2 moles of NO are produced for every 3 moles of O_2 reacted.

3.8.2 Silicon reacts with chromium(III) oxide according to the equation

$$3Si(s) + 2Cr_2O_3(s) \longrightarrow 3SiO_2(s) + 4Cr(s)$$

If 59.4 g of silicon is consumed in the reaction, what is the mass of SiO_2 produced?

3.8.3 At high temperatures, magnesium reacts with nitrogen gas according to the equation

$$3Mg(s) + N_2(g) \longrightarrow Mg_3N_2(s)$$

How many grams of magnesium are needed to produce 25.0 g of Mg_3N_2?

3.9 Limiting Reactants

Video
Limiting Reagent

When a chemist carries out a reaction, the reactants are usually not present in exact ***stoichiometric amounts,*** that is, *in the proportions indicated by the balanced equation.* Because the goal of a reaction is to produce the maximum quantity of a useful compound from the starting materials, frequently a large excess of one reactant is supplied to ensure that the more expensive reactant is completely converted to the desired product. Consequently, some reactant will be left over at the end of the reaction. *The reactant used up first in a reaction* is called the ***limiting reactant,*** because the maximum amount of product formed depends on how much of this reactant was originally present. When this reactant is used up, no more product can be formed. ***Excess reactants*** are the *reactants present in quantities greater than necessary to react with the quantity of the limiting reagent.*

The concept of the limiting reactant is analogous to the relationship between men and women in a dance contest at a club. If there are 14 men and only 9 women, then only 9 female/male pairs can compete. Five men will be left without partners. The number of women thus *limits* the number of men that can dance in the contest, and there is an *excess* of men.

Consider the industrial synthesis of methanol (CH_3OH) from carbon monoxide and hydrogen at high temperatures:

$$CO(g) + 2H_2(g) \longrightarrow CH_3OH(g)$$

Suppose initially we have 4 moles of CO and 6 moles of H_2 (Figure 3.9). One way to determine which of two reactants is the limiting reagent is to calculate the number of moles of CH_3OH obtained based on the initial quantities of CO and H_2. From the preceding definition, we see that only the limiting reactant will yield the *smaller* amount of the product. Starting with 4 moles of CO, we find the number of moles of CH_3OH produced is

$$4 \cancel{\text{mol CO}} \times \frac{1 \text{ mol } CH_3OH}{1 \cancel{\text{mol CO}}} = 4 \text{ mol } CH_3OH$$

and starting with 6 moles of H_2, the number of moles of CH_3OH formed is

$$6 \cancel{\text{mol } H_2} \times \frac{1 \text{ mol } CH_3OH}{2 \cancel{\text{mol } H_2}} = 3 \text{ mol } CH_3OH$$

Because H_2 results in a smaller amount of CH_3OH, it must be the limiting reagent. Therefore, CO is the excess reagent.

In stoichiometric calculations involving limiting reactants, the first step is to decide which reactant is the limiting reactant. After the limiting reactant has been identified, the rest of the problem can be solved as outlined in Section 3.8. Example 3.15 illustrates this approach.

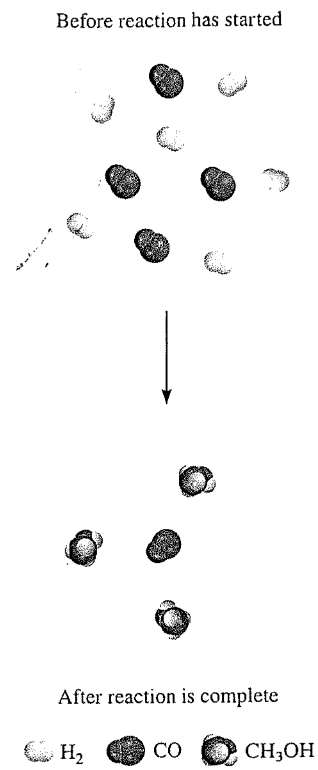


Figure 3.9 *At the start of the reaction, there were six H_2 molecules and four CO molecules. At the end, all the H_2 molecules are gone and only one CO molecule is left. Therefore, H_2 molecule is the limiting reactant and CO is the excess reagent. Each molecule can also be treated as one mole of the substance in this reaction.*

Example 3.15

The synthesis of urea, $(NH_2)_2CO$, is considered to be the first recognized example of preparing a biological compound from nonbiological reactants, challenging the notion that biological processes involved a "vital force" present only in living systems. Today urea is produced industrially by reacting ammonia with carbon dioxide:

$$2NH_3(g) + CO_2(g) \longrightarrow (NH_2)_2CO(aq) + H_2O(l)$$

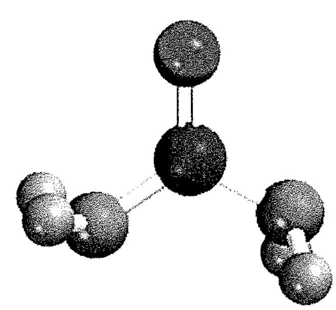
In one process, 637.2 g of NH_3 are treated with 1142 g of CO_2. (a) Which of the two reactants is the limiting reactant? (b) Calculate the mass of $(NH_2)_2CO$ formed. (c) How much excess reagent (in grams) is left at the end of the reaction?

(a) Strategy The reactant that produces fewer moles of product is the limiting reactant because it limits the amount of product that can be formed. How do we convert from the amount of reactant to the amount of product? Perform this calculation for each reactant, then compare the moles of product, $(NH_2)_2CO$, formed by the given amounts of NH_3 and CO_2 to determine which reactant is the limiting reagent.

(Continued)

Student Hot Spot

Student data indicate you may struggle with finding the limiting reactant. Access your eBook for additional Learning Resources on this topic.



$(NH_2)_2CO$

Solution We carry out two separate calculations. First, starting with 637.2 g of NH_3, we calculate the number of moles of $(NH_2)_2CO$ that could be produced if all the NH_3 reacted according to the following conversions.

$$\text{grams of } NH_3 \longrightarrow \text{moles of } NH_3 \longrightarrow \text{moles of } (NH_2)_2CO$$

Combining these conversions in one step, we write

$$\begin{aligned}\text{moles of } (NH_2)_2CO &= 637.2 \ \cancel{\text{g } NH_3} \times \frac{1 \ \cancel{\text{mol } NH_3}}{17.03 \ \cancel{\text{g } NH_3}} \times \frac{1 \text{ mol } (NH_2)_2CO}{2 \ \cancel{\text{mol } NH_3}}\\ &= 18.71 \text{ mol } (NH_2)_2CO\end{aligned}$$

Second, for 1142 g of CO_2, the conversions are

$$\text{grams of } CO_2 \longrightarrow \text{moles of } CO_2 \longrightarrow \text{moles of } (NH_2)_2CO$$

The number of moles of $(NH_2)_2CO$ that could be produced if all the CO_2 reacted is

$$\begin{aligned}\text{moles of } (NH_2)_2CO &= 1142 \ \cancel{\text{g } CO_2} \times \frac{1 \ \cancel{\text{mol } CO_2}}{44.01 \ \cancel{\text{g } CO_2}} \times \frac{1 \text{ mol } (NH_2)_2CO}{1 \ \cancel{\text{mol } CO_2}}\\ &= 25.95 \text{ mol } (NH_2)_2CO\end{aligned}$$

It follows, therefore, that NH_3 must be the limiting reactant because it produces a smaller amount of $(NH_2)_2CO$.

(b) Strategy We determined the moles of $(NH_2)_2CO$ produced in part (a), using NH_3 as the limiting reactant. How do we convert from moles to grams?

Solution The molar mass of $(NH_2)_2CO$ is 60.06 g. We use this as a conversion factor to convert from moles of $(NH_2)_2CO$ to grams of $(NH_2)_2CO$:

$$\begin{aligned}\text{mass of } (NH_2)_2CO &= 18.71 \ \cancel{\text{mol } (NH_2)_2CO} \times \frac{60.06 \text{ g } (NH_2)_2CO}{1 \ \cancel{\text{mol } (NH_2)_2CO}}\\ &= 1124 \text{ g } (NH_2)_2CO\end{aligned}$$

Check Does your answer seem reasonable? Note that 18.71 moles of product are formed. What is the mass of 1 mole of $(NH_2)_2CO$?

(c) Strategy Working backward, we can determine the amount of CO_2 that reacted to produce 18.71 moles of $(NH_2)_2CO$. The amount of CO_2 left over is the difference between the initial amount and the amount reacted.

Solution Starting with 18.71 moles of $(NH_2)_2CO$, we can determine the mass of CO_2 that reacted using the mole ratio from the balanced equation and the molar mass of CO_2. The conversion steps are

$$\text{grams of } CO_2 \longrightarrow \text{moles of } CO_2 \longrightarrow \text{moles of } (NH_2)_2CO$$

so that

$$\begin{aligned}\text{mass of } CO_2 \text{ reacted} &= 18.71 \ \cancel{\text{mol } (NH_2)_2CO} \times \frac{1 \ \cancel{\text{mol } CO_2}}{1 \ \cancel{\text{mol } (NH_2)_2CO}} \times \frac{44.01 \text{ g } CO_2}{1 \text{ mol } CO_2}\\ &= 823.4 \text{ g } CO_2\end{aligned}$$

The amount of CO_2 remaining (in excess) is the difference between the initial amount (1142 g) and the amount reacted (823.4 g):

$$\text{mass of } CO_2 \text{ remaining} = 1142 \text{ g} - 823.4 \text{ g} = 319 \text{ g}$$

(Continued)

Practice Exercise The reaction between aluminum and iron(III) oxide can generate temperatures approaching 3000°C and is used in welding metals:

$$2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$$

In one process, 124 g of Al are reacted with 601 g of Fe_2O_3. (a) Calculate the mass (in grams) of Al_2O_3 formed. (b) How much of the excess reactant is left at the end of the reaction?

Similar problem: 3.86.

Example 3.15 brings out an important point. In practice, chemists usually choose the more expensive chemical as the limiting reactant so that all or most of it will be converted to products in the reaction. In the synthesis of urea, NH_3 is invariably the limiting reactant because it is more expensive than CO_2. At other times, an excess of one reactant is used to help drive the reaction to completion, or to compensate for a side reaction that consumes that reactant. Synthetic chemists often have to calculate the amount of reactant to use based on this need to have one or more components in excess, as Example 3.16 shows.

Example 3.16

The reaction between alcohols and halogen compounds to form ethers is important in organic chemistry, as illustrated here for the reaction between methanol (CH_3OH) and methyl bromide (CH_3Br) to form dimethylether (CH_3OCH_3), which is a useful precursor to other organic compounds and an aerosol propellant.

$$CH_3OH + CH_3Br + LiC_4H_9 \longrightarrow CH_3OCH_3 + LiBr + C_4H_{10}$$

This reaction is carried out in a dry (water-free) organic solvent, and the butyl lithium (LiC_4H_9) serves to remove a hydrogen ion from CH_3OH. Butyl lithium will also react with any residual water in the solvent, so the reaction is typically carried out with 2.5 molar equivalents of that reagent. How many grams of CH_3Br and LiC_4H_9 will be needed to carry out the preceding reaction with 10.0 g of CH_3OH?

Solution We start with the knowledge that CH_3OH and CH_3Br are present in stoichiometric amounts and that LiC_4H_9 is the excess reactant. To calculate the quantities of CH_3Br and LiC_4H_9 needed, we proceed as shown in Example 3.14.

$$\text{grams of } CH_3Br = 10.0\ \cancel{g\ CH_3OH} \times \frac{1\ \cancel{mol\ CH_3OH}}{32.04\ \cancel{g\ CH_3OH}} \times \frac{1\ \cancel{mol\ CH_3Br}}{1\ \cancel{mol\ CH_3OH}} \times \frac{94.93\ g\ CH_3Br}{1\ \cancel{mol\ CH_3Br}}$$

$$= 29.6\ g\ CH_3Br$$

$$\text{grams of } LiC_4H_9 = 10.0\ \cancel{g\ CH_3OH} \times \frac{1\ \cancel{mol\ CH_3OH}}{32.04\ \cancel{g\ CH_3OH}} \times \frac{2.5\ \cancel{mol\ LiC_4H_9}}{1\ \cancel{mol\ CH_3OH}} \times \frac{64.05\ g\ LiC_4H_9}{1\ \cancel{mol\ LiC_4H_9}}$$

$$= 50.0\ g\ LiC_4H_9$$

Practice Exercise The reaction between benzoic acid (C_6H_5COOH) and octanol ($C_8H_{17}OH$) to yield octyl benzoate ($C_6H_5COOC_8H_{17}$) and water

$$C_6H_5COOH + C_8H_{17}OH \longrightarrow C_6H_5COOC_8H_{17} + H_2O$$

is carried out with an excess of $C_8H_{17}OH$ to help drive the reaction to completion and maximize the yield of product. If an organic chemist wants to use 1.5 molar equivalents of $C_8H_{17}OH$, how many grams of $C_8H_{17}OH$ would be required to carry out the reaction with 15.7 g of C_6H_5COOH?

Similar problems: 3.137, 3.138.

Review of Concepts & Facts

3.9.1 Aluminum and bromine vigorously react according to the equation

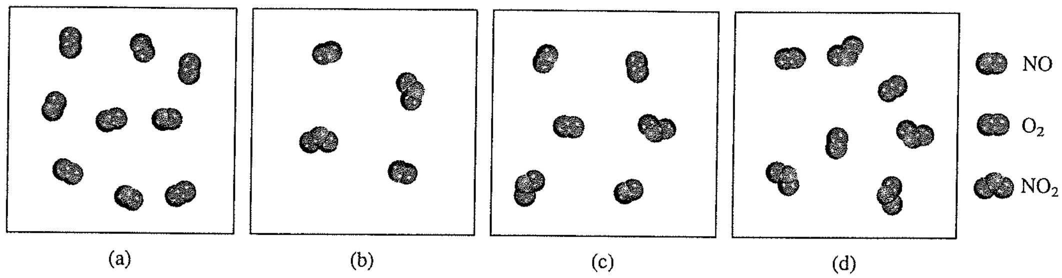
$$2Al(s) + 3Br_2(l) \longrightarrow 2AlBr_3(s)$$

(a) If 5.00 g of aluminum and 22.2 g of bromine react, what mass of $AlBr_3$ is produced? (b) What mass of the excess reactant remains at the end of the reaction?

3.9.2 Consider the following reaction:

$$2NO(g) + O_2(g) \longrightarrow 2NO_2(g)$$

Starting with the reactants shown in (a), which of the diagrams shown in (b)–(d) best represents the situation in which the limiting reactant has completely reacted?



3.10 Reaction Yield

Keep in mind that the theoretical yield is the yield that you calculate using the balanced equation. The actual yield is the yield obtained by carrying out the reaction.

The amount of limiting reactant present at the start of a reaction determines the ***theoretical yield*** of the reaction—that is, *the amount of product that would result if all the limiting reactant reacted.* The theoretical yield, then, is the *maximum* obtainable yield, predicted by the balanced equation. In practice, the ***actual yield,*** or *the amount of product actually obtained from a reaction,* is almost always less than the theoretical yield. There are many reasons for the difference between actual and theoretical yields. For instance, many reactions are reversible, and so they do not proceed 100 percent from left to right. Even when a reaction is 100 percent complete, it may be difficult to recover all of the product from the reaction medium (say, from an aqueous solution). Some reactions are complex in the sense that the products formed may react further among themselves or with the reactants to form still other products. These additional reactions will reduce the yield of the first reaction.

To determine how efficient a given reaction is, chemists often figure the ***percent yield***, which describes *the proportion of the actual yield to the theoretical yield.* It is calculated as follows:

$$\%\text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{3.4}$$

Percent yields may range from a fraction of 1 percent to 100 percent. Chemists strive to maximize the percent yield in a reaction. Factors that can affect the percent yield include temperature and pressure. We will study these effects later.

In Example 3.17 we will calculate the yield of an industrial process.

Example 3.17

Titanium is a strong, lightweight, corrosion-resistant metal that is used in rockets, aircraft, jet engines, and bicycle frames. It is prepared by the reaction of titanium(IV) chloride with molten magnesium between 950°C and 1150°C:

$$TiCl_4(g) + 2Mg(l) \longrightarrow Ti(s) + 2MgCl_2(l)$$

(Continued)

In a certain industrial operation 3.54×10^7 g of $TiCl_4$ are reacted with 1.13×10^7 g of Mg. (a) Calculate the theoretical yield of Ti in grams. (b) Calculate the percent yield if 7.91×10^6 g of Ti are actually obtained.

(a) Strategy Because there are two reactants, this is likely to be a limiting reactant problem. The reactant that produces fewer moles of product is the limiting reactant. How do we convert from amount of reactant to amount of product? Perform this calculation for each reactant, then compare the moles of product, Ti, formed.

Solution Carry out two separate calculations to see which of the two reactants is the limiting reactant. First, starting with 3.54×10^7 g of $TiCl_4$, calculate the number of moles of Ti that could be produced if all the $TiCl_4$ reacted. The conversions are

$$\text{grams of TiCl}_4 \longrightarrow \text{moles of TiCl}_4 \longrightarrow \text{moles of Ti}$$

so that

$$\begin{aligned}\text{moles of Ti} &= 3.54 \times 10^7 \cancel{\text{g TiCl}_4} \times \frac{1 \cancel{\text{mol TiCl}_4}}{189.7 \cancel{\text{g TiCl}_4}} \times \frac{1 \text{ mol Ti}}{1 \cancel{\text{mol TiCl}_4}} \\ &= 1.87 \times 10^5 \text{ mol Ti}\end{aligned}$$

Next, we calculate the number of moles of Ti formed from 1.13×10^7 g of Mg. The conversion steps are

$$\text{grams of Mg} \longrightarrow \text{moles of Mg} \longrightarrow \text{moles of Ti}$$

and we write

$$\begin{aligned}\text{moles of Ti} &= 1.13 \times 10^7 \cancel{\text{g Mg}} \times \frac{1 \cancel{\text{mol Mg}}}{24.31 \cancel{\text{g Mg}}} \times \frac{1 \text{ mol Ti}}{2 \cancel{\text{mol Mg}}} \\ &= 2.32 \times 10^5 \text{ mol Ti}\end{aligned}$$

Therefore, $TiCl_4$ is the limiting reactant because it produces a smaller amount of Ti. The mass of Ti formed is

$$1.87 \times 10^5 \cancel{\text{mol Ti}} \times \frac{47.88 \text{ g Ti}}{1 \cancel{\text{mol Ti}}} = 8.95 \times 10^6 \text{ g Ti}$$

(b) Strategy The mass of Ti determined in part (a) is the theoretical yield. The amount given in part (b) is the actual yield of the reaction.

Solution The percent yield is given by

$$\begin{aligned}\%\text{yield} &= \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \\ &= \frac{7.91 \times 10^6 \text{ g}}{8.95 \times 10^6 \text{ g}} \times 100\% \\ &= 88.4\%\end{aligned}$$

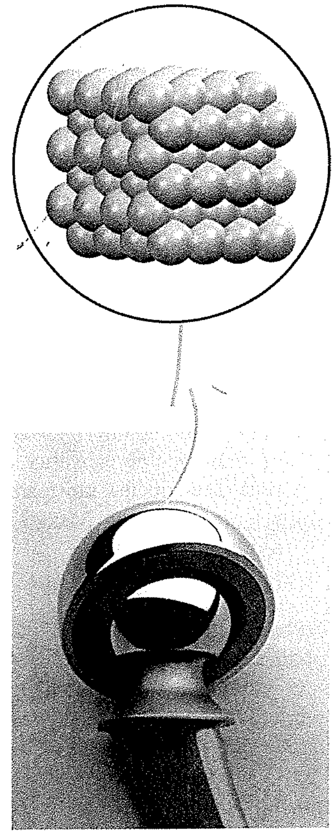
Check Should the percent yield be less than 100 percent?

Practice Exercise Industrially, vanadium metal, which is used in steel alloys, can be obtained by reacting vanadium(V) oxide with calcium at high temperatures:

$$5\text{Ca} + \text{V}_2\text{O}_5 \longrightarrow 5\text{CaO} + 2\text{V}$$

In one process, 1.54×10^3 g of V_2O_5 react with 1.96×10^3 g of Ca. (a) Calculate the theoretical yield of V. (b) Calculate the percent yield if 803 g of V are obtained.

Similar problems: 3.89, 3.90.



An artificial hip joint made of titanium and the structure of solid titanium.

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CHEMISTRY *in Action*

Chemical Fertilizers

A growing world population requires an increase in agricultural production. Farmers increase crop quality and yield by adding chemical fertilizer to their fields. N, P, K, Ca, S, and Mg are the required elements for increased growth. The preparation and properties of several nitrogen- and phosphorus-containing fertilizers illustrate some of the principles introduced in this chapter.

Nitrogen fertilizers contain nitrate (NO_3^-) salts, ammonium (NH_4^+) salts, and other compounds. Plants can absorb nitrogen in the form of nitrate directly, but ammonium salts and ammonia (NH_3) must first be converted to nitrates by the action of soil bacteria. The principal raw material of nitrogen fertilizers is ammonia, prepared by the reaction between hydrogen and nitrogen:

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

(This reaction will be discussed in detail in Chapters 13 and 14.) In its liquid form, ammonia can be injected directly into the soil.

Alternatively, ammonia can be converted to ammonium nitrate (NH_4NO_3), ammonium sulfate [$(NH_4)_2SO_4$], or ammonium hydrogen phosphate [$(NH_4)_2HPO_4$].

One method of preparing ammonium sulfate requires two steps:

$$2NH_3(aq) + CO_2(aq) + H_2O(l) \longrightarrow (NH_4)_2CO_3(aq) \quad (1)$$

$$(NH_4)_2CO_3(aq) + CaSO_4(aq) \longrightarrow (NH_4)_2SO_4(aq) + CaCO_3(s) \quad (2)$$

This approach is desirable because the starting materials—carbon dioxide and calcium sulfate—are less costly than sulfuric acid. To increase the yield, ammonia is made the limiting reagent in Reaction (1) and ammonium carbonate is made the limiting reagent in Reaction (2).

Liquid ammonia being applied to the soil before planting.

The table lists the percent composition by mass of nitrogen in some common fertilizers. The preparation of urea was discussed in Example 3.15.

Percent Composition by Mass of Nitrogen in Five Common Fertilizers

Fertilizer	% N by Mass
NH_3	82.4
NH_4NO_3	35.0
$(NH_4)_2SO_4$	21.2
$(NH_4)_2HPO_4$	21.2
$(NH_2)_2CO$	46.7

Several factors influence the choice of one fertilizer over another: (1) cost of the raw materials needed to prepare the fertilizer; (2) ease of storage, transportation, and utilization; (3) percent composition by mass of the desired element; and (4) suitability of the compound, that is, whether the compound is soluble in water and whether it can be readily taken up by plants. Considering all these factors together, we find that NH_4NO_3 is the most important nitrogen-containing fertilizer in the world, even though ammonia has the highest percentage by mass of nitrogen.

Phosphorus fertilizers are derived from phosphate rock, called *fluorapatite*, $Ca_5(PO_4)_3F$. Fluorapatite is insoluble in water, so it must first be converted to water-soluble calcium dihydrogen phosphate [$Ca(H_2PO_4)_2$]:

$$2Ca_5(PO_4)_3F(s) + 7H_2SO_4(aq) \longrightarrow 3Ca(H_2PO_4)_2(aq) + 7CaSO_4(aq) + 2HF(g)$$

For maximum yield, fluorapatite is made the limiting reagent in this reaction.

The reactions we have discussed for the preparation of fertilizers all appear relatively simple, yet much effort has been expended to improve the yields by changing conditions such as temperature, pressure, and so on. Industrial chemists usually run promising reactions first in the laboratory and then test them in a pilot facility before putting them into mass production.

Thinking Critically

1. Describe the benefits of ammonia-based fertilizers compared to nitrate-based fertilizers.
2. How does setting ammonia as the limiting reagent in Eq. (1) ensure maximum yield of ammonium sulfate?
3. Calculate the percent by mass of phosphorous in calcium dihydrogen phosphate.

Industrial processes usually involve huge quantities (thousands to millions of tons) of products. Thus, even a slight improvement in the yield can significantly reduce the cost of production. A case in point is the manufacture of chemical fertilizers, discussed in the Chemistry in Action essay "Chemical Fertilizers."

Review of Concepts & Facts

3.10.1 Can the percent yield ever exceed the theoretical yield of a reaction?

3.10.2 Sulfur trioxide (SO_3) is prepared from the oxidation of sulfur dioxide (SO_2) according to the equation

$$2SO_2(g) + O_2(g) \longrightarrow 2SO_3(g)$$

If 16.4 g of SO_2 produces 18.1 g of SO_3, what is the percent yield of the reaction?

AP A Look Back at the Essential Knowledge

Atomic mass is a relative unit and is based on a mass of exactly 12 for the carbon-12 isotope. If we have Avogadro's number (6.022×10^{23}) of C-12 isotopes, the mass would be exactly 12 grams. The mass of all elements is based on the mass of the C-12 isotope. The molar mass can be determined by summing the atomic mass of all constituent atoms of the molecule or for ionic compounds the simplest formula unit. The mass of an element can be determined using a mass spectrometer. Mass spectrometer experiments demonstrated that a given element could have different masses which lead to the discovery of isotopes.

The atomic mass listed on the periodic table is a weighted average of isotope masses for each element. The mass percent composition of a known chemical formula is determined by converting the mass fraction of each element into percentage. Percentage composition can also be used to determine the simplest, whole number ratio of elements in a substance; the empirical formula. It is also possible to determine the molecular formula from an empirical formula of the experimental molar mass of the compound is known. Mass percent composition can be experimentally determined through combustion analysis.

We can represent chemical change be writing a chemical equation with reactants on the left and products on the right separated by a reaction arrow. Atoms are not destroyed during a chemical reaction, thus mass is conserved. As a result, the same number of atoms of each element in the reactants must appear in the products. This requires us to balance chemical reactions. Using balanced chemical reactions and the conservation of mass we can determine the mass of products formed from a given mass of a single reactant or the mass of reactant required to produce a given mass of product. This process is known as stoichiometry.

Since the mass of a substance does not directly indicate the number of particles of the substance and a balanced chemical equation indicates the number of particles reacting and produced, we must be able to convert mass to the number of particles. The mole concept allows us to convert mass to moles or moles to mass since the mole indicates the number of particles. If the mass of more than one reactant is provided you may have a limiting reagent problem. In this case it is necessary to determine which reactant is the limiting reagent, the reagent which is in the smallest stoichiometric amount. This reactant will determine the mass of product that can be formed in a reaction. Many factors influence the outcome of a reaction. These factors can reduce the actual mass of products formed, the experimental yield of the reaction, compared to the theoretical or stoichiometric yield. The ratio of the actual yield to the theoretical yield multiplied by 100 is the percent yield.

FOCUS REVIEW GUIDE

Complete the activities in Chapter 3 of your *Focus Review Guide* to review content essential for your AP exam.

Learning Objectives

- Discuss the nature of the atomic mass scale. (Section 3.1)
- Determine the average atomic mass of an element from isotopic mass and relative abundance information. (Section 3.1)
- Describe and apply the concept of a mole. (Section 3.2)
- Calculate the molecular and molar mass of compounds. (Section 3.3)
- Interconvert between mass, moles, and number of atoms or molecules. (Section 3.3)
- Deduce the percent composition by mass for elements in a compound. (Section 3.4)
- Solve for the empirical formula of a compound from percent composition or from combustion analysis. (Section 3.6)