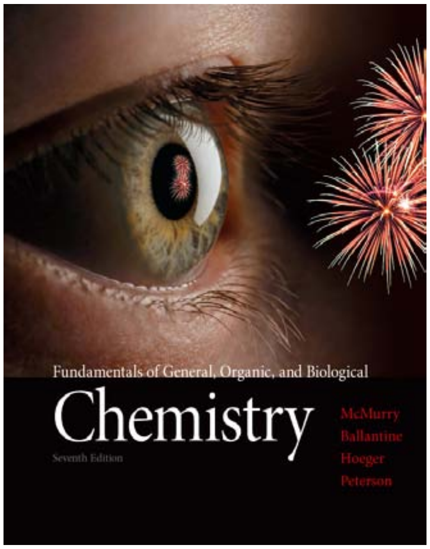
Chapter 1 Lecture



Fundamentals of General, Organic, and Biological Chemistry

7th Edition

McMurry, Ballantine, Hoeger, Peterson

Chapter One

Matter and Measurements

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ALWAYS LEARNING PEARSON

Outline

1.1 Chemistry: The Central Science
1.2 States of Matter
1.3 Classification of Matter
1.4 Chemical Elements and Symbols
1.5 Elements and the Periodic Table
1.6 Chemical Change: An Example of a Chemical Reaction
1.7 Physical Quantities
1.8 Measuring Mass, Length, and Volume
1.9 Measurement and Significant Figures
1.10 Scientific Notation
1.11 Rounding Off Numbers
1.12 Problem Solving: Unit Conversions and Estimating Answers
1.13 Temperature, Heat, and Energy
1.14 Density and Specific Gravity

Goals

1. **What is matter and how is it classified?**

   Be able to discuss the properties and states of matter, distinguish between mixtures and pure substances, elements, and compounds.

2. **How are chemical elements represented?**

   Be able to name and give the symbols of common elements.

3. **What kinds of properties does matter have?**

   Be able to distinguish between chemical and physical properties.

Goals, *Continued*

4. **What units are used to measure properties, and how can a quantity be converted from one unit to another?**

   Be able to name and use the metric and SI units for mass, length, volume, and temperature; and be able to convert quantities from one unit to another.

5. **How good are the reported measurements?**

   Be able to interpret the significant figures in a measurement and round off numbers in calculations involving measurements.

Goals, *Continued*

6. **How are large and small numbers best represented?**
   Be able to interpret prefixes and express numbers in scientific notation.
7. **What techniques are used to solve problems?**
   Be able to analyze a problem, use the factor-label method, and check the result to ensure that it makes sense.
8. **What are temperature, specific heat, density, and specific gravity?**
   Be able to define these quantities and use them in calculations.

1.1 Chemistry: The Central Science

- **Chemistry—**The study of the nature, properties and transformations of matter.
- **Matter—**Anything that has mass and occupies space, that is, things you can see, touch, taste, or smell.
- **Scientific Method—**The process of observation, hypothesis, and experimentation used to expand a body of knowledge.

1.1 Chemistry: The Central Science

Property—A characteristic useful for identifying a substance or object. These include size, color, and temperature, as well as *chemical composition* and *chemical reactivity.*

Physical Change—A change that does not affect the chemical makeup of a substance or object.

Chemical Change—A change in the chemical makeup of a substance.

1.2 States of Matter

Matter exists in three forms:

- **Solid—** A substance that has a definite shape and volume.
- **Liquid—** A substance that has a definite volume but assumes the shape of its container.
- **Gas—** A substance that has neither a definite volume nor a definite shape.

1.2 States of Matter

Many substances can exist in all three phases, and participate in **changes of state.**

(a) Ice: A solid has a definite volume and a definite shape independent of its container.

(b) Water: A liquid has a definite volume but a variable shape that depends on its container.

(c) Steam: A gas has both variable volume and shape that depend on its container.

1.3 Classification of Matter

Substances are classified as **pure substances** or **mixtures**.

- **Pure substance—**A substance that has a uniform chemical composition throughout.
- **Mixture—**A blend of two or more substances, each of which retains its chemical identity.
  - **Homogeneous mixture—**A uniform mixture that has the same composition throughout.
  - **Heterogeneous mixture—**A non-uniform mixture that has regions of different composition.

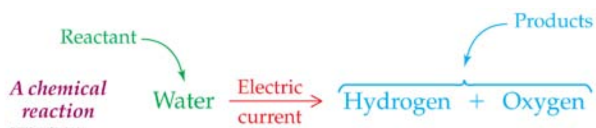
1.3 Classification of Matter

Components of a mixture can be separated without changing their identities.

Pure substances can be classified into two groups:

- **Element—**A fundamental substance that cannot be broken down chemically into any simpler substance.
- **Chemical compound—**A pure substance that can be broken down into simpler substances by chemical reactions.

1.3 Classification of Matter



Water can be chemically changed by passing an electric current through it to produce hydrogen and oxygen.

- The **reactant** (water) is written on the left.
- The **products** (hydrogen and oxygen) are written on the right.
- An arrow connects the two parts to indicate a **chemical reaction**. The conditions necessary are written above and below the arrow.

1.3 Classification of Matter

Aspirin—A Case Study

Aspirin was discovered through a combination of serendipity and the scientific method.

- Willow bark and leaves were prescribed for pain and fever by Hippocrates as early as 400 B.C.
- Salicin was isolated in 1828, but often caused stomach irritation.
- Acetylsalicylic acid was first synthesized in 1853.
- In 1971, it was discovered that aspirin suppresses the production of prostaglandins, leading to the development of new analgesic drugs.

1.4 Chemical Elements and Symbols

- 118 elements have been identified. 91 occur naturally.
- A one- or two-letter shorthand is used for each element.
  - The first letter is always capitalized.
  - The second letter, if any, is always lowercase.
- Most of the symbols are based on the elements commonly used names.
- A few symbols are based on Latin names for the elements.

1.4 Chemical Elements and Symbols

Not all elements occur with equal abundance.

TABLE 1.3 Elemental Composition of the Earth's Crust and the Human Body*

Earth's Crust		Human Body	
Oxygen	46.1%	Oxygen	61%
Silicon	28.2%	Carbon	23%
Aluminum	8.2%	Hydrogen	10%
Iron	5.6%	Nitrogen	2.6%
Calcium	4.1%	Calcium	1.4%
Sodium	2.4%	Phosphorus	1.1%
Magnesium	2.3%	Sulfur	0.20%
Potassium	2.1%	Potassium	0.20%
Titanium	0.57%	Sodium	0.14%
Hydrogen	0.14%	Chlorine	0.12%

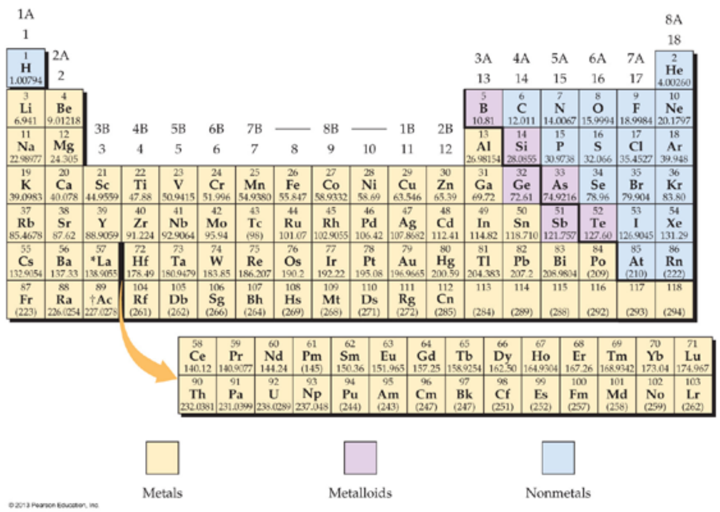
**Mass percent values are given.*

1.4 Chemical Elements and Symbols

- The symbols of elements are combined to form **chemical formulas**.
- Subscripts show how many *atoms* (the smallest fundamental units) of each element are in a given chemical compound.

H_2O	CH_4	$C_{12}H_{22}O_{11}$
2 H atoms 1 O atom	1 C atom 4 H atoms	12 C atoms 22 H atoms 11 O atoms

1.5 Elements and the Periodic Table



The periodic table, divided into metals, nonmetals, and metalloids.

1.5 Elements and the Periodic Table

Metals:

- 94 of the known elements
- Occur on the left side of the periodic table
- Solid at room temperature (except mercury)
- Usually lustrous when freshly cut
- Good conductors of heat and electricity
- Malleable rather than brittle

1.5 Elements and the Periodic Table

Nonmetals:

- 18 of the known elements
- Occur on the right side of the periodic table
- 11 are gases, 6 are solids,1 (bromine) is a liquid.
- Poor conductors of heat and electricity
- Brittle when solid

(a) (b) (c)

1.5 Elements and the Periodic Table

Metalloids:

- 6 of the known elements
- Properties are intermediate between metals and nonmetals.
- Occur in a zigzag band between metals on the left and nonmetals on the right side of the periodic table

(a) (b)

1.5 Elements and the Periodic Table

Mercury and Mercury Poisoning

- Mercury is the only metallic element that is a liquid at room temperature.
- Mercury has many uses in which it is not toxic, including the laxative calomel (Hg_2Cl_2) and dental amalgam.
- Exposure to mercury *vapor* causes adverse health effects.
- Toxicity of mercury is related to solubility. Mercury vapor accumulates in the lungs and slowly becomes soluble.
- Soluble mercury can enter the bloodstream and interfere with various biological processes.

1.6 Chemical Reactions: An Example of Chemical Change

- Nickel is a hard, shiny metal.
- Hydrogen chloride is a colorless gas that dissolves in water to form hydrochloric acid.
- When nickel is added to hydrochloric acid; the nickel is eaten away; the solution turns green; and a gas bubbles out.

(a)

1.6 Chemical Reactions: An Example of Chemical Change

A chemical reaction is taking place. This is indicated by:

- Change in color
- Dissolving of nickel
- Appearance of bubbles


(b)

1.6 Chemical Reactions: An Example of Chemical Change

The reaction of nickel with hydrochloric acid can be either written in words or a shorthand notation using symbols to represent the elements or compounds involved as reactants and products.

$$\overbrace{\text{Nickel} + \text{Hydrochloric acid}}^{\text{Reactants}} \longrightarrow \overbrace{\text{Nickel (II) chloride} + \text{Hydrogen}}^{\text{Products}}$$

$$[Ni + 2\,HCl \longrightarrow NiCl_2 + H_2]$$

1.7 Physical Quantities

Mass, volume, temperature, density, and other physical properties are called **physical quantities** and are described by both a number and a **unit**:

- **Physical quantity—**A physical property that can be measured.
- **Unit—**A defined quantity used as a standard of measurement.

1.7 Physical Quantities

Scientists have agreed on the *Système International d'Unites* (International System of Units), abbreviated *SI*.

- Mass is measured in *kilograms* (kg).
- Length is measured in *meters* (m).
- Volume is measured in *cubic meters* (m^3).
- Temperature is measured in *kelvins* (K).
- Time is measured in *seconds* (s).

1.7 Physical Quantities

TABLE 1.5 Some SI and Metric Units and Their Equivalents

Quantity	Si Unit (Symbol)	Metric Unit (Symbol)	Equivalents
Mass	Kilogram (kg)	Gram (g)	1 kg = 1000 g = 2.205 lb
Length	Meter (m)	Meter (m)	1 m = 3.280 ft
Volume	Cubic meter (m^3)	Liter (L)	1 m^3 = 1000 L = 264.2 gal
Temperature	Kelvin (K)	Celsius degree (°C)	See Section 1.13
Time	Second (s)	Second (s)	—

1.7 Physical Quantities

SI units are related to metric units, with a few differences.

- The metric unit of mass is the *gram* (g) rather than the kilogram (1g =1/1000 kg).
- The metric unit of volume is the *liter* (L) rather than the cubic meter (1L = 1/1000 m^3).
- The metric unit of temperature is the *Celsius degree* (°C) rather than the kelvin.

1.7 Physical Quantities

Derived units:

- *Speed*: *meters per second* (m/s)
- *Density: grams per cubic centimeter* (g/cm^3)

Unit sizes are often inconveniently large or small, so they can be modified using prefixes to refer to smaller or larger quantities.

1.7 Physical Quantities

TABLE 1.6 Some Prefixes for Multiples of Metric and SI Units

Prefix	Symbol	Base Unit Multiplied By*	Example
mega	M	$1{,}000{,}000 = 10^6$	1 megameter (Mm) = 10^6 m
kilo	k	$1000 = 10^3$	1 kilogram (kg) = 10^3 g
hecto	h	$100 = 10^2$	1 hectogram (hg) = 100 g
deka	da	$10 = 10^1$	1 dekaliter (daL) = 10 L
deci	d	$0.1 = 10^{-1}$	1 deciliter (dL) = 0.1 L
centi	c	$0.01 = 10^{-2}$	1 centimeter (cm) = 0.01 m
milli	m	$0.001 = 10^{-3}$	1 milligram (mg) = 0.001 g
micro	μ	$0.000\,001 = 10^{-6}$	1 micrometer (μm) = 10^{-6} m
nano	n	$0.000\,000\,001 = 10^{-9}$	1 nanogram (ng) = 10^{-9} g
pico	p	$0.000\,000\,000\,001 = 10^{-12}$	1 picogram (pg) = 10^{-12} g
femto	f	$0.000\,000\,000\,000\,001 = 10^{-15}$	1 femtogram (fg) = 10^{-15} g

**The scientific notation method of writing large and small numbers (for example, 10^6 for 1,000,000) is explained in Section 1.10.*

1.8 Measuring Mass, Length, and Volume

- **Mass**—A measure of the amount of matter in an object.
- **Weight**—A measure of the gravitational force that the earth or other large body exerts on an object.
- The *mass* of an object can be determined by comparing the *weight* of the object to the weight of a known reference standard.

1.8 Measuring Mass, Length, and Volume

TABLE 1.7 Units of Mass

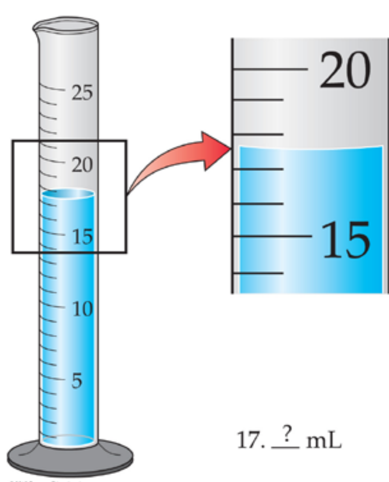
Unit	Equivalent	Unit	Equivalent
1 kilogram (kg)	= 1000 grams = 2.205 pounds	1 ton	= 2000 pounds = 907.03 kilograms
1 gram (g)	= 0.001 kilogram = 1000 milligrams = 0.035 27 ounce	1 pound (lb)	= 16 ounces = 0.454 kilogram = 454 grams
1 milligram (mg)	= 0.001 gram = 1000 micrograms	1 ounce (oz)	= 0.028 35 kilogram = 28.35 grams
1 microgram (μg)	= 0.000 001 gram = 0.001 milligram		= 28,350 milligrams

1.8 Measuring Mass, Length, and Volume

- The meter is the standard measure of length, or distance, in the SI and metric systems.
- *Volume* is the amount of space occupied by an object. The SI unit for volume—the cubic meter, m^3—is so large that the liter ($1\ L = 0.001\ m^3 = 1\ dm^3$) is much more commonly used.

1.9 Measurement and Significant Figures

- Every experimental measurement has a degree of uncertainty.
- The value recorded should use all the digits known with certainty, plus one estimated digit.
- **Significant figures—** The number of meaningful digits used to express a value.



1.9 Measurement and Significant Figures

Rules for Significant Figures

Rule 1: Zeros in the middle of a number are like any other digit; they are always significant.

Rule 2: Zeros at the beginning of a number are not significant; they act only to locate the decimal point.

Rule 3: Zeros at the end of a number and *after* the decimal point are significant. It is assumed that these zeros would not be shown unless they were significant.

Rule 4: Zeros at the end of a number and *before* an implied decimal point may or may not be significant.

1.9 Measurement and Significant Figures

Some numbers are *exact* and effectively have an unlimited number of significant figures.

- A class might have *exactly* 32 students (not 31.9, 32.0, or 32.1).
- 1 foot is defined to have *exactly* 12 inches.

1.10 Scientific Notation

- **Scientific notation**—A number expressed as the product of a number between 1 and 10, times 10 raised to a power.

$$215 = 2.15 \times 100 = 2.15 \times (10 \times 10) = 2.15 \times 10^2$$

- The exponent on the 10 tells how many places the decimal point was moved to position it just after the first digit

1.10 Scientific Notation

- To express a number *smaller* than 1 in scientific notation, the decimal point is moved *to the right* until it follows the first digit. The number of places moved is the negative exponent of 10.

$$1.56 \times 10^{-8} = 0.000\,000\,015\,6$$

Negative exponent of −8, so decimal point is moved to the left eight places.

1.10 Scientific Notation

- To convert a number written in scientific notation to standard notation, the process is reversed.
  - *positive* exponent—The decimal point is moved to the *right* a number of places equal to the exponent.
  - *negative* exponent—The decimal point is moved to the *left* a number of places equal to the exponent.
- Only significant numbers are used.

1.11 Rounding Off Numbers

Calculators often display more digits than are justified by the precision of the data.

Rounding off—A procedure used for deleting nonsignificant figures

- **Rule 1:** In carrying out multiplication or division, the answer cannot have more significant figures than the original numbers.
- **Rule 2:** In carrying out addition or subtraction, the answer cannot have more digits after the decimal point than the original numbers.

1.12 Problem Solving: Unit Conversions and Estimating Answers

The simplest way to carry out calculations involving different units is to use the **factor-label method**.

- The **factor-label method** is a problem-solving procedure in which equations are set up so that unwanted units cancel and only the desired units remain.
- The **conversion factor** is an expression of the numerical relationship between two units.

1.12 Problem Solving: Unit Conversions and Estimating Answers

- Conversion factors are numerically equal to one.
- Units are treated like numbers and can thus be multiplied and divided.
- Set up an equation so that all unwanted units cancel.
- Think through a rough estimate, or *ballpark estimate*, as a check on your work.

Conversion factors between kilometers and miles

$$\frac{1\ \text{km}}{0.6214\ \text{mi}} = 1 \quad \text{or} \quad \frac{0.6214\ \text{mi}}{1\ \text{km}} = 1$$

1.12 Problem Solving: Unit Conversions and Estimating Answers

- **STEP 1:** Identify the information given, including units.
- **STEP 2:** Identify the information needed in the answer, including units.
- **STEP 3:** Find the relationship(s) between the known information and unknown answer, and plan a series of steps, for getting from one to the other.
- **STEP 4:** Solve the problem.
- **BALLPARK CHECK—**Make a ballpark estimate at the beginning and check it against your final answer to be sure the value and the units of your calculated answer are reasonable.

1.12 Problem Solving: Unit Conversions and Estimating Answers

WORKED EXAMPLE 1.11: A child is 21.5 inches long at birth. How long is this in centimeters?

BALLPARK ESTIMATE—It takes about 2.5 cm to make 1 in., so it should take 2.5 times as many centimeters to make a distance equal to approximately 20 in., or about 20 in. × 2.5 = 50 cm.

SOLUTION

STEP 1: Identify given information. Length = 21.5 inches

STEP 2: Identify answer and units. Length = ?? cm

STEP 3: Identify conversion factor. 1 in = 2.54 cm

STEP 4: Solve: 21.5 in × 2.54 cm/in = 54.6 cm

BALLPARK CHECK—54.6 cm is close to the estimate.

1.13 Temperature, Heat, and Energy

- **Energy**—The capacity to do work or supply heat.
- **Temperature**—The measure of the amount of heat energy in an object. Commonly reported either in Fahrenheit (°F) or Celsius (°C) units, but the SI unit for temperature is the *kelvin* (K). (Note that we say only "kelvin," not "degrees kelvin.")

1.13 Temperature, Heat, and Energy

- The kelvin and the Celsius degree are the same size.
- Thus, a change in temperature of 1 °C is equal to a change of 1 K.
- The Celsius scale assigns a value of 0 °C to the freezing point of water.
- The Kelvin scale assigns a value of 0 K to the coldest possible temperature, *absolute zero*, which is equal to - 273.15 °C. Thus, 0 K = -273.15 °C, and +273.15 K = 0 °C.

1.13 Temperature, Heat, and Energy

- The Fahrenheit scale defines the freezing point of water as 32 °F and the boiling point of water as 212 °F.
- It takes 180 Fahrenheit degrees to cover the same range encompassed by 100 Celsius degrees.
- A change in temperature of 1.0 °C is equal to a change of 1.8 °F.

1.13 Temperature, Heat, and Energy

Temperature conversions:

$$K = {}^\circ C + 273.15$$

$$^\circ C = K - 273.15$$

$$^\circ F = \left(\frac{9\ ^\circ F}{5\ ^\circ C} \times {}^\circ C\right) + 32\ ^\circ F$$

$$^\circ C = \frac{5\ ^\circ C}{9\ ^\circ F} \times \left({}^\circ F - 32\ ^\circ F\right)$$

1.13 Temperature, Heat, and Energy

Temperature–Sensitive Materials

- Thermochromic materials change color as their temperature changes.
- These "liquid crystals" can be incorporated into plastics or paints, and can be used to monitor the temperature of the products or packages in which they are incorporated.
- Hospitals and other medical facilities now routinely use temperature strips that change color to indicate body temperature.
- In the future, we may see road signs that change color to warn of icy road conditions.

1.13 Temperature, Heat and Energy

- Energy is represented in SI units by the unit *joule* (J), but the metric unit *calorie* (cal) is still widely used.
- One calorie is the amount of heat necessary to raise the temperature of 1 g of water by 1 °C.
- A *kilocalorie* (kcal), called a *large calorie (Cal)* or *food calorie* by nutritionists, equals 1000 cal.
- One calorie raises the temperature of 1 g of water by 1 °C, but raises the temperature of 1 g of iron by 10 °C. The amount of heat needed to raise the temperature of 1 g of a substance by 1 °C is called the **specific heat.**

1.13 Temperature, Heat, and Energy

Knowing the mass and specific heat of a substance makes it possible to calculate how much heat must be added or removed to accomplish a given temperature change.

TABLE 1.10 Specific Heats of Some Common Substances

Substance	Specific Heat [cal/g °C]; [J/g °C]	
Ethanol	0.59;	2.5
Gold	0.031;	0.13
Iron	0.106;	0.444
Mercury	0.033;	0.14
Sodium	0.293;	1.23
Water	1.00;	4.18

1.14 Density and Specific Gravity

- **Density** is the physical property that relates the mass of an object to its volume; mass per unit volume.
- Most substances contract when cooled, and expand when heated.
- Water expands when it freezes, so ice floats on liquid water.

TABLE 1.11 Densities of Some Common Materials at 25 °C

Substance	Density*	Substance	Density*
Gases		Solids	
Helium	0.000 194	Ice (0 °C)	0.917
Air	0.001 185	Gold	19.3
		Human fat	0.94
Liquids		Cork	0.22–0.26
Water (3.98 °C)	1.0000	Table sugar	1.59
Urine	1.003–1.030	Balsa wood	0.12
Blood plasma	1.027	Earth	5.54

*Densities are in g/cm^3 for solids and g/mL for liquids and gases.

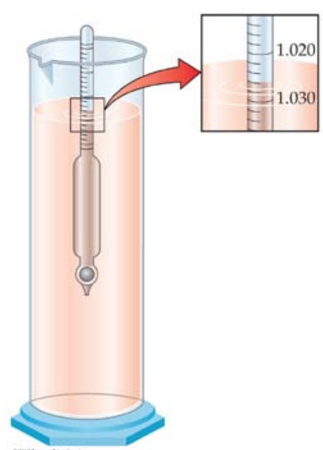
1.14 Density and Specific Gravity

Obesity and Body Fat

- Obesity is defined by reference to *body mass index* (BMI)—Mass in kilograms divided by the square of height in meters.
- 25 or above is overweight, and 30 or above is obese. By these standards, approximately 61% of the U.S. population is overweight.
- Body fat is most easily measured by the skinfold thickness method. The thickness of the fat layer beneath the skin is measured with calipers. Comparing the results to those in a standard table gives an estimation of percentage body fat.
- A more accurate assessment of body fat can be made by underwater immersion. The higher the percentage of body fat, the more buoyant the person and the greater the difference between land weight and underwater body weight. Checking observed buoyancy on a standard table gives an estimation of body fat percentage.

1.14 Density and Specific Gravity

- **Specific gravity** is the density of a substance divided by the density of water at the same temperature.
- At normal temperatures, the density of water is very close to 1 g/mL.
- At normal temperatures, the specific gravity of a substance is numerically equal to its density.
- The specific gravity of a liquid can be measured using an instrument called a *hydrometer*.
- In medicine, a hydrometer called a *urinometer* is used to indicate the amount of solids dissolved in urine.



Chapter Summary

1. What is matter and how is it classified?

Matter has mass and occupies volume.

Matter can be classified as *solid*, *liquid*, or *gas*.

- A solid has a definite volume and shape.
- A liquid has a definite volume, but indefinite shape.
- A gas has neither a definite volume nor shape.

Chapter Summary, *Continued*

1. What is matter and how is it classified?

Matter can be classified by composition as being either *pure* or a *mixture*.

- Every pure substance is either an *element* or a *chemical compound*.
  - Elements are fundamental substances that cannot be chemically changed into anything simpler.
  - A chemical compound can be broken down by chemical change into simpler substances.
- Mixtures are composed of two or more pure substances and can be separated by physical means.

Chapter Summary, *Continued*

2. How are chemical elements represented?

Elements are represented by one- or two-letter symbols.

All the known elements are commonly organized into a form called the *periodic table*.

Most elements are *metals*, 18 are *nonmetals*, and 6 are *metalloids*.

3. What kinds of properties does matter have?

Physical properties can be seen without changing the identity of the substance

Chemical properties can only be seen or measured when the substance undergoes a *chemical change*.

Chapter Summary, *Continued*

4. What units are used to measure properties, and how can a quantity be converted from one unit to another?

A *physical quantity* is described by a number and a *unit*.

Units are those of the International System of Units (*SI units*) or the *metric system*.

- Mass is measured in *kilograms* (kg) or *grams* (g).
- Length is measured in *meters* (m).
- Volume is measured in *cubic meters* in the SI system and in *liters* (L) or *milliliters* (mL) in the metric system.
- Temperature is measured in *kelvins* (K) in the SI system and in *degrees Celsius* (°C) in the metric system.

A measurement in one unit can be converted to another unit by multiplying by a *conversion factor* that expresses the exact relationship between the units.

Chapter Summary, *Continued*

5. **How good are the reported measurements?**

   The exactness of a measurement is indicated by *rounding off* the final answer using the correct number of *significant figures*.

6. **How are large and small numbers best represented?**

   Small and large quantities are usually written in *scientific notation.*

7. **What techniques are used to solve problems?**

   Problems are solved with the *factor-label method*, in which units are also multiplied and divided.

8. **What are temperature, specific heat, density, and specific gravity?**

   Temperature is a measure of how hot or cold an object is.

   Specific heat is the amount of heat necessary to raise the temperature of 1 g of the substance by 1 °C.

   Density relates mass to volume.

   Specific gravity of a liquid is the density of the liquid divided by the density of water at the same temperature.