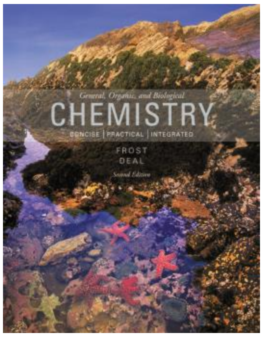


Lecture Presentation

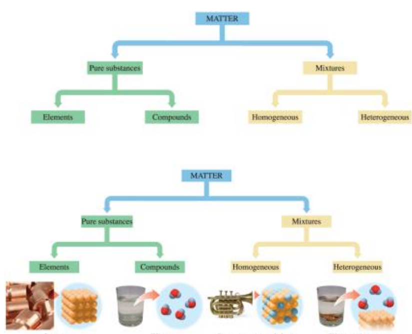
Chapter 1

Chemistry Basics: Matter and Measurement

Julie Klare
Fortis College
Smyrna, GA

Outline

1.1 Classifying Matter



1.1 Classifying Matter

- Matter can be broadly classified as a *mixture* or a *pure substance.*
- Mixtures can be classified as *homogeneous* or *heterogeneous.*
- Pure substances are classified as *elements* or *compounds*.

1.1 Classifying Matter

- A **mixture** is a combination of two or more substances.
- A mixture can be separated into its different components.
- A **homogeneous mixture** is one whose composition is the same throughout.
- A **heterogeneous mixture** is not uniform, but varies throughout.

1.1 Classifying Matter

- A **pure substance** is matter that is made up of only one substance.
- An **element** is the simplest type of matter because it is made up of only one type of atom.
- An **atom** is the smallest unit of matter that keeps its unique characteristics.
- A **compound** is a pure substance made of two or more elements chemically joined together.

1.1 Classifying Matter

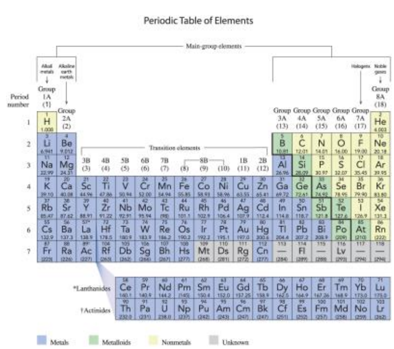
- All *elements* are listed individually on the *periodic table of the elements*.
- The rows on the periodic table are *periods*, and the columns are *groups*.
- Groups are numbered across the top of the periodic table; periods are numbered down the left side.
- Metals are on the left and nonmetals on the right, with a staircase-shaped dividing line between the two.

1.1 Classifying Matter

- Chemical formulas show the type and number of each element present in a compound.
- For example, water's chemical formula is H_2O, and it contains two hydrogen atoms and one oxygen atom.

$$H_2O$$

1.2 Elements, Compounds, and the Periodic Table



1.2 Elements, Compounds, and the Periodic Table

- The periodic table consists of many small blocks. Each has a letter or two in its center and numbers above and below.
- The letters are the **chemical symbol** and represent the name of each element.
- For many elements, the symbols are derived from the name of the element.
- Some symbols are derived from Latin:

sodium (natrium) = Na

gold (aurum) = Au

1.2 Elements, Compounds, and the Periodic Table

- A vertical column is a **group** of elements with **<u>similar chemical behaviors</u>**.
- Each group has a number and letter designation.
  - **<u>A designations are main-group elements.</u>**
  - **<u>B designations are transition elements.</u>**
- A system using numbers 1 through 18 for the columns, recommended by the International Union of Pure and Applied Chemistry (IUPAC), is also used.

1.2 Elements, Compounds, and the Periodic Table

- A horizontal row is known as a **period**.
- Periods are numbered from 1 to 7 with sections of Periods 6 and 7 set apart at the bottom of the periodic table.
- Each of the groups and periods has special characteristics.
- The staircase-shaped line, which begins at boron, **<u>separates metals from nonmetals</u>**.
- Elements bordered by the line, with the exception of aluminum (Al), are metalloids.

1.2 Elements, Compounds, and the Periodic Table

- A pure substance containing two or more chemically combined elements is a compound.
- Compounds combine elements in specific ratios.
- **Chemical formulas** show that water (H_2O) is composed of two particles of hydrogen and one particle of oxygen, and table salt (NaCl) is composed of one sodium and one chlorine.
- A chemical formula identifies both the type and number of particles of each of the elements in a compound.

1.2 Elements, Compounds, and the Periodic Table

- Knowing the units of a dose or measurement is critical.
- Pharmaceutical and scientific measurements often use the **metric system**, part of the International System of Units (SI units).

1.3 Math Counts

- The standard unit for mass is the **kilogram** (kg).
- The standard unit for volume is the **liter** (L).
- The standard unit for length is the **meter** (m).

1.3 Math Counts

- To deal efficiently with quantities that are much larger or much smaller than each another, the SI system employs a set of prefixes that can be applied to the base unit.

TABLE 1.1 Metric Prefixes

Prefix	Abbreviation	Relationship to Base Unit
giga	G	1,000,000,000 ×
mega	M	1,000,000 ×
kilo	k	1000 ×
base unit (has no prefix)		1 × (gram, liter, meter)
deci	d	÷10
centi	c	÷100
milli	m	÷1000
micro	μ	÷1,000,000
nano	n	÷1,000,000,000

1.3 Math Counts

- Quantities that can related to each other by an equal sign are called **equivalent units.**
- Such equivalencies can be used as **conversion factors** to convert one unit to another using one or more of these factors.

$$\frac{10\text{ dg}}{1\text{ g}} \text{ or } \frac{1\text{ g}}{10\text{ dg}}$$

- Conversion factors allow you to convert a quantity in one unit to the equivalent quantity in a larger or smaller unit.

1.3 Math Counts

- This use of converting units to an equivalent unit is also called **dimensional analysis**.

Step 1: Determine the units on your final answer.

Step 2: Establish the given information.

Step 3: Decide how to set up the problem. Which conversion factor should be used to leave the desired unit in the answer?

Step 4: Solve the problem.

$$\text{Given unit} \times \left(\frac{\text{desired unit}}{\text{given unit}}\right)$$

1.3 Math Counts

- **Significant Figures**
  - All measurements have some level of uncertainty.
  - Measuring matter relies on the precision of the instruments that we use to measure it.
  - It is important to report calculated answers reasonably.
  - In any measurement, the **significant figures** are the digits known with certainty plus the estimated digit.
  - Working with significant figures allows us to represent and convey the uncertainty in a given measurement.

1.3 Math Counts

- **Significant Figures—Zeros**
  - If a terminal zero (at the end or on the right) in a number is significant, put in a decimal point.
- **Exact Numbers**
  - Numbers used in conversion factors and when counting items are exact numbers with an infinite number of significant figures.

1.3 Math Counts

TABLE 1.2 Counting Significant Figures in Measurements

Rule	Measurement	Number of Significant Figures
1. A digit is significant if it is		
a. not a zero	41 g 15.3 m	2 3
b. a zero between nonzero digits	101 L 6.071 kg	3 4
c. a zero at the end of a number with a decimal point	20. g 9.800 °C	2 4
2. A zero is not significant if it is		
a. at the beginning of a number with a decimal point	0.03 L 0.00024 g	1 2
b. in a large number without a decimal point	12,000 km 3,450,000 m	2 3

1.3 Math Counts

- **Calculating Numbers and Rounding**
  - Adding, subtracting, multiplying, or dividing can result in numbers that seem more certain than they are.
  - Manipulating measurements with arithmetic cannot *increase* their certainty.
- **Rules for Significant Figures in Calculations**
  - ***Addition and Subtraction.*** Answers should be given to the least number of decimal places in the measured numbers.
  - ***Multiplication and Division.*** Answers should be given to the least number of significant digits in the measured numbers.

1.3 Math Counts

- **Rules for Rounding Numbers**
  - If the leftmost digit to be removved is 4 or less, simply remove it and the remaining digits.
  - If the leftmost digit to be dropped is 5 or greater, increase the last retained digit by 1 and remove all other digits.
  - If rounding a large number with no decimal point, zeros are substituted for numbers that are not significant.
  - When conducting multiple-step calculations, do not round answers until the end of the calculation.

1.3 Math Counts

- **Scientific Notation**
  - The general form for scientific notation is

    $$C \times 10^{n}$$

    where C is called the **coefficient** and is a number between 1 and 9 and n is the exponent telling us the number of tens places that apply.
  - A positive exponent tells us that the actual number is greater than 1.
  - A negative exponent tells us the number is less than 1.
  - In scientific notation, the coefficient shows the number of significant figures.

1.3 Math Counts

- **Percent, ppm and ppb**
  - Percent, represented by the symbol %, means the part out of 100 total, or hundredths.
  - Percent allows us to directly compare two sets of numbers that have different total sizes.

$$\text{Percent }(\%) = \frac{\text{part}}{\text{whole}} \times 100$$

  - A fraction can be converted to a percent by dividing the numerator by the denominator, multiplying by 100, and adding a percent sign.
  - A decimal number can be converted to a percent by multiplying by 100 and adding a percent sign.

1.4 Matter: The "Stuff" of Chemistry

- **Mass**
  - Anything that takes up space can also be placed on a scale and weighed.
  - **Mass** is a measure of the amount of material in an object.
  - A common unit used to measure the mass of a substance is the gram (g).
  - The weight of an object is determined by the pull of gravity on the object, and that force changes depending on location.
  - As long as an object is weighed in roughly the same location on Earth's surface, its mass and weight will have the same measured value.

1.4 Matter: The "Stuff" of Chemistry

- **Volume**
  - **Volume** is a three-dimensional measure of the space occupied by matter.
  - In the lab, volumes are measured with a graduated cylinder or a pipet.
  - The unit typically used in the lab is the milliliter (mL).
  - In a clinical setting, volumes are often measured with calibrated syringes.
  - The typical unit in the clinical setting is the cubic centimeter (cc or cm^3).
  - One milliliter equals one cubic centimeter.

1.4 Matter: The "Stuff" of Chemistry

- **Density**
  - **Density** (*d*) is a comparison (also called a ratio) of a substance's mass (*m*) to its volume (*V*).

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

  - One gram of water has a mass of one milliliter, so the density of water is 1.00 g/mL.
  - A piece of wood will float: it is less dense than water.
  - A piece of metal will sink: it is more dense than water.
  - Because the density of a substance does not change, we can use density values as conversion factors to determine either the mass or the volume of a substance.

1.4 Matter: The "Stuff" of Chemistry

- **Specific Gravity**
  - Liquid density often is measured with respect to water.
  - The density of water is 1.00 g/mL at 4 °C.
  - The ratio of the density of a sample to the density of water is called **specific gravity** (sp gr).

$$\text{Specific gravity} = \frac{\text{density of sample}}{\text{density of water}}$$

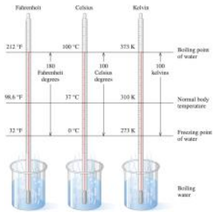
  - Specific gravity is unitless because it is a ratio.
  - The specific gravity of a liquid can be measured with a simple instrument called a *hydrometer*.

1.4 Matter: The "Stuff" of Chemistry

- **Temperature**
  - We measure the **temperature** of substance to determine its hotness or coldness.
  - This is often done using a thermometer or an electronic temperature probe.
  - In the United States, we use the Fahrenheit scale.
  - The rest of the world uses the Celsius scale.
  - Scientists use still another scale called the absolute, or Kelvin, scale, where the temperature unit is the kelvin.
  - Kelvin is the SI unit for temperature.

1.4 Matter: The "Stuff" of Chemistry

- **Temperature**
  - The most straightforward way to compare temperature scales is to compare temperatures that we are familiar with and observe their values on each scale.



1.4 Matter: The "Stuff" of Chemistry

- **Temperature**
  - The Celsius and Kelvin scales have degrees of the same size, offset by 273 degrees.

Kelvin (K) = Celsius (°C) + 273

  - 1 degree Celsius is 1.8 degrees Fahrenheit, and the "zero points" are offset by 32 degrees.

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

1.4 Matter: The "Stuff" of Chemistry

- **Body Temperature**
  - Normal body temperature is 98.6 °F or 37.0 °C.
  - Body temperature varies from person to person, changing throughout the day.
  - Human body temperature over 40 °C (104 °F) is known as *hyperthermia*: this can cause convulsions, coma, or permanent brain damage.
  - If body temperature drops below 35 °C (95 °F), *hypothermia* is present: a person in this condition feels cold, has an irregular heartbeat, and has a slow breathing rate.

1.4 Matter: The "Stuff" of Chemistry

- **Energy**
  - **Energy** is the ability to do work.
  - Stored energy is **potential energy**.
  - The energy of motion is **kinetic energy**.
  - Energy takes various forms, but it is never created or destroyed.
  - This is the law of **conservation of energy**.
  - The SI unit for energy is the **joule** (J).
  - A **calorie** is the amount of energy that raises the temperature of one gram of water one degree Celsius.
  - A nutritional **Calorie** (Cal) is 1000 times larger than a calorie.

1 cal = 4.184 J

1 Cal = 1000 cal

1.4 Matter: The "Stuff" of Chemistry

- **Heat and Specific Heat**
  - **Heat** is kinetic energy flowing from a warmer body to a colder one.
  - Every substance has the ability to absorb or lose heat as the temperature changes.
  - The **specific heat capacity**, or specific heat of a substance, is the amount of heat needed to raise the temperature of one gram of a substance by 1 °C.
  - Metals have low specific heat values.
  - Water has a very high specific heat.

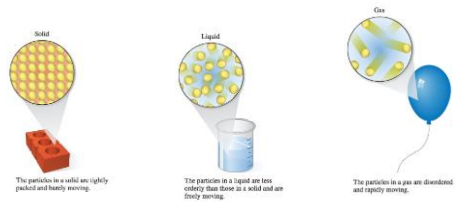
1.4 Matter: The "Stuff" of Chemistry

TABLE 1.4 Specific Heats of Various Substances

Substance	Specific Heat (cal/g °C)
Water (liquid)	1.00
Human body	0.83
Paraffin wax	0.60
Wood, soft	0.34
Wood, hard	0.29
Air	0.24
Aluminum	0.21
Table salt	0.21
Brick	0.20
Stainless steel	0.12
Iron	0.11
Copper	0.092
Silver	0.056
Gold	0.031

1.4 Matter: The "Stuff" of Chemistry

- **States of Matter**
  - A **state of matter** is the physical form in which the matter exists. The three most common states of matter are solid, liquid, and gas.



1.4 Matter: The "Stuff" of Chemistry

- **States of Matter**
  - The particles in a solid are tightly packed together and moving only slightly.
  - **Solids** have a definite shape and volume.
  - The particles in a liquid are less orderly and moving freely.
  - A **liquid** has a definite volume, but takes the shape of its container.
  - The particles in a gas have no arrangement, are far apart from each other, move at high rates of speed, and often collide with each other and with the walls of their container.
  - A **gas** has no definite shape or volume.

1.4 Matter: The "Stuff" of Chemistry

TABLE 1.5 Properties of Solids, Liquids, and Gases

Property of a Substance	Solid	Liquid	Gas
Shape	Definite shape	Adopts shape of container	Adopts shape of container
Volume	Definite volume	Definite volume	Fills volume of container
Kinetic energy	Lowest of the three states	More than solid, less than gas	Highest of the three states
Positioning of particles	Closely packed and fixed	Loosely packed, but random	Far apart and random
Attraction between particles	Very strong	Strong	Practically none

1.5 Measuring Matter

- **Accuracy and Precision**
  - Accurate measurements are close to the actual or true value.
  - Precise measurements are similar in value, but may not be close to the actual value.
  - In taking measurements, it is best to measure with both **accuracy** and **precision**.
  - This can be accomplished by taking measurements several times and averaging their values.

1.5 Measuring Matter

- **Units and Dosing**
  - Health care professionals use SI or metric units, but must also be familiar with the U.S. customary system of measurement.

TABLE 1.6 Equivalent Units in SI and U.S. Customary Systems

Property	U.S Customary Unit	Metric Equivalent	U.S. Customary Equivalent Unit
Mass	Pound (lb)	2.205 lb = 1 kg	1 lb = 16 oz
Volume	Quart (qt)	1.06 qt = 1 L	1 qt = 4 cups
	Fluid ounce (fl oz)	1 oz = 29.6 mL	1 cup = 8 oz
	Teaspoon (tsp)	1 tsp = 5 mL	1 fl oz = 6 tsp
Length	Mile (mi)	1 mi = 0.62 km	1 mi = 5,280 ft
	Inch (in.)	1 in. = 2.54 cm	1 ft = 12 in.

1.5 Measuring Matter

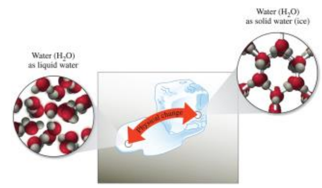
- **Reading Lab Reports**
  - Results are given along with normal limits. If the result is out of range, it is highlighted.
  - Values are listed as greater than one, but this makes the units on the numbers vary widely.
  - Most of the units are metric.
  - Mmole (millimole) is equivalent to 1/1000 moles. A mole is a unit used to count the particles in matter. A similar unit used for electrolytes is the milliequivalent (mEq).
  - In the United States, body weight is usually measured in pounds, but pharmaceuticals are often dispensed by body weight in kilograms.

1.5 Measuring Matter

- **Percents in Health**
  - *Percent Active Ingredient*: Because of the high potency of many medicines, binders are often added to increase the size of a pill.
  - *Percent of an Adult Dose*: Because children weigh less than adults, they are often administered a percent of the adult dose.
  - *Percent in Nutrition Labeling*: The amount of carbohydrate, protein, and fat and percent of the recommended daily allowance (RDA) for vitamins present in a serving.

1.6 How Matter Changes

- **Physical Change**
  - A change in the state of matter represents a physical change.
  - In a **physical change**, the form of the matter is changed, but its identity remains the same.



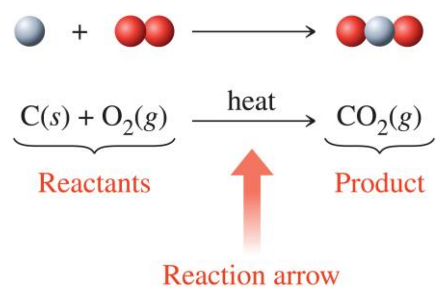
1.6 How Matter Changes

- **Chemical Change**
  - A chemical change results in a change in the chemical identity of a substance.
  - When a substance undergoes such a change, it is referred to as a **chemical reaction**.

TABLE 1.7 Sample Blood Chemistry Lab Results

Patient Name: Jane Patient		Age: 40	
Test	**Result**	**Nrml-Range**	**Units**
Blood Chemistry			
Sodium	137	135–145	mmole/L or mEq/L
Potassium	3.9	3.5–5.2	mmole/L or mEq/L
Chloride	100	97–108	mmole/L or mEq/L
Glucose	88	65–99	mg/dL
Urea Nitrogen	14	5–26	mg/dL
Calcium	9.1	8.5–10.6	mg/dL
Phosphorus	3.5	2.5–4.5	mg/dL
Protein, Total	7.1	6.0–8.5	g/dL
Iron	113	35–155	μg/dL

1.6 How Matter Changes



1.6 How Matter Changes

- **Chemical Equations**
  - A chemical equation is a way of writing a sentence about what happens in a chemical reaction.
  - Carbon and oxygen are the **reactants,** and carbon dioxide is the **product**. The reaction arrow means "react to form."
  - Special reaction conditions are often written above the reaction arrow
  - The labels in parentheses after each substance indicate its physical state—(*s*)olid, (*l*)iquid, or (*g*)as.

1.6 How Matter Changes

- **Balancing Chemical Equations**
  - The number of each reactant element equals the number of each element in the products.
  - This illustrates the **law of conservation of mass.**
  - For any chemical equation, the number of each element or compound must be the same on both sides of the equation.
  - We can balance chemical equations when necessary by adding a number, called a **coefficient**, in front of the chemical formula for a substance in the chemical equation.

1.6 How Matter Changes

- **Balancing Chemical Equations**

Step 1: Examine the original equation. Is it balanced? If not, proceed to step 2.

Step 2: Balance the equation one element at a time by adding coefficients.

Step 3: Check to see if the equation is balanced.

Chapter One Summary

- **1.1 Classifying Matter**
  - Besides classifying matter as solid, liquid, or gas, chemists use a broader classification of mixture or pure substance.
  - Mixtures can be separated into their component parts and are further classified as homogeneous (evenly mixed throughout) or heterogeneous (unevenly mixed).
  - Pure substances are made up of a single component and are classified as elements or compounds.
  - Compounds contain more than one element while elements contain a single type of atom.
  - Atoms are the smallest unit of matter with unique chemical characteristics.

Chapter One Summary (continued)

- **1.2 Elements, Compounds, and the Periodic Table**
  - The periodic table of the elements is a useful catalog of all of the elements.
  - Each block on the table contains the symbol of an element along with other useful information about that element.
  - The blocks are arranged in columns known as groups and rows known as periods.
  - The elements in Groups 1A through 8A are known as the main-group elements and those in the B groups are known as the transition elements. A staircase-shaped line on the right side of the periodic table separates the elements that are metals from those that are nonmetals.
  - Compounds are chemical combinations of elements represented by chemical formulas, which provide the identity and number of each element in the compound.

Chapter One Summary (continued)

- **1.3 Math Counts**
  - Some basic mathematical concepts apply to chemistry.
  - SI (includes metric) units is a system based on powers of 10, and the prefixes used reflect the powers of 10.
  - Conversion factors are used to convert SI units.
  - Significant figures allow us to designate the certainty of measurements in chemistry.
  - Rounding answers to no more certainty than measured ensures meaningful answers.
  - Scientific notation is used to write very large and very small numbers economically. This notation also allows a direct comparison of extremely large and very small numbers.
  - Direct comparison of different sample sizes can also be examined using percent.

Chapter One Summary (continued)

- **1.4 Matter: The "Stuff" of Chemistry**
  - Matter is anything that takes up space and has mass.
  - Chemists measure properties such as mass, volume, density, temperature, energy, heat, and specific heat.
  - Mass measures the amount of matter and can be measured on a balance.
  - Volume is a three-dimensional measure of the space that matter occupies.
  - Density is a property of matter and is a ratio of mass to volume.
  - Measuring the temperature of matter is useful because it indicates the amount of energy present.

Chapter One Summary (continued)

- **1.4 Matter: The "Stuff" of Chemistry**
  - Temperature units include Fahrenheit, Celsius, and kelvin. The kelvin and Celsius degree are the same size unit, but offset by 273. The Fahrenheit degree is smaller. There are nine Fahrenheit degrees for every five degrees Celsius.
  - Energy in matter can be either potential (stored) or kinetic (moving).
  - Energy is neither created nor destroyed but simply changes form. Heat is kinetic energy that flows from a warmer body to a colder one. The specific heat of a particular material measures how much heat energy it takes to raise its temperature.
  - Most matter exists in one of three different states: solid, liquid, or gas. The states of matter differ in the motion, kinetic energy, and positioning of their particles.

Chapter One Summary (continued)

- **1.5 Measuring Matter**
  - This section applies conversion factors and the units for measuring matter to solve problems in health.
  - The U.S. system of units is introduced and compared to the SI system.
  - Practical dosing calculations show how conversions, units, and percent can be applied in health care.

Chapter One Summary (continued)

- **1.6 How Matter Changes**
  - Matter can undergo two types of change: a physical change and a chemical change called a reaction.
  - In a physical change, the substance changes states, but its identity remains the same.
  - In a chemical reaction, the identity of the reacting substance (or substances) is changed.
  - A chemical reaction is represented using a chemical equation, which identifies the reacting substance(s) and the product(s), the physical state of all substances in the reaction, and any conditions necessary for the reaction to occur.
  - The law of conservation of mass dictates that a balanced chemical equation must have equal numbers of each atom in both the reactants and products.
  - An outline for balancing chemical equations is given as three steps: (1) Examine. (2) Balance. (3) Check.

Chapter One Study Guide

- **1.1 Classifying Matter**
  - Classify matter as a pure substance or a mixture.
  - Classify mixtures as homogeneous or heterogeneous.
  - Classify pure substances as elements or compounds.
- **1.2 Elements, Compounds, and the Periodic Table**
  - Distinguish between groups and periods.
  - Locate metals and nonmetals on the periodic table.
  - Identify the number of elements in a chemical formula.

Chapter One Study Guide (continued)

- **1.3 Math Counts**
  - Convert between metric units.
  - Apply the appropriate number of significant figures to a calculation.
  - Convert numbers to scientific notation.
  - Convert numbers and fractions to percent.
- **1.4 Matter: The “Stuff” of Chemistry**
  - Define mass and its measurement.
  - Define volume and its measurement.
  - Calculate and solve problems using density.
  - Convert temperatures between the three temperature scales.
  - Distinguish between kinetic and potential energy.
  - Convert between energy units.
  - Compare specific heat values of various materials.
  - Contrast the properties of solids, liquids, and gases.

Chapter One Study Guide (continued)

- **1.5 Measuring Matter**
  - Distinguish between accuracy and precision.
  - Convert between SI and U.S. units.
  - Apply conversion factors, units, and percent to measurements in health.
- **1.6 How Matter Changes**
  - Distinguish between physical changes and chemical reactions.
  - Balance a given chemical equation.