

Chemistry, To Accompany the OWL Electronic Learning System

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How to use this text:

This textbook, in trial format, is intended for use in conjunction with the OWL electronic homework system. You have a login on the system and will do your course homework, as well as much of your studying, using the system. To login, go to:

<http://owl.oit.umass.edu>

Choose: Chemistry General (UMass Amherst) under University of Massachusetts & Affiliates

login: SUCO + your last name

password = your student number

For me, it would be SUCOVining

You should be enrolled in a Course named "SUNY Oneonta Chem 111 Fall 2006"

The OWL system presents you with a set of assignments, allowing you to master the material within. This is connected closely with the text. Throughout the chapter you will encounter OWL assignment icons in a box:



OWL Example Problems

1.11 Temperature: Tutor

1.12 Temperature

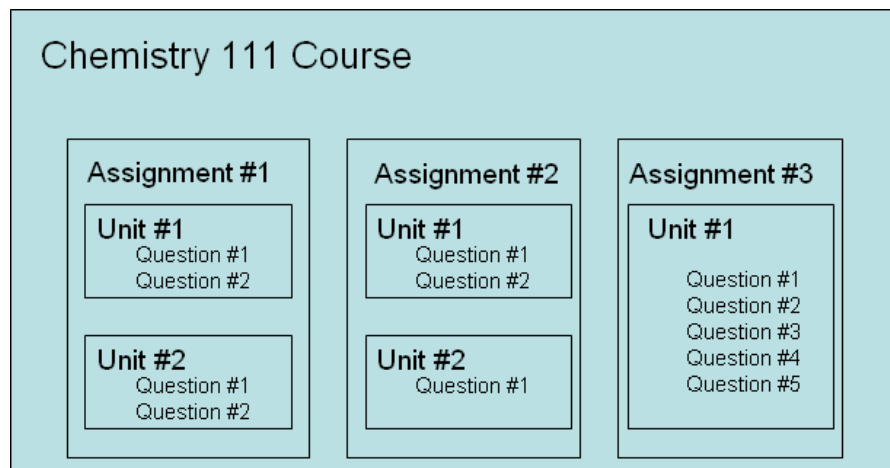
These point to an assignment of the same number and name in the OWL system.

The text can be used in two ways:

1. Text First: Read the text chapter and then go to the OWL system to do your graded work. As you run into problems, go back to the text and find the appropriate section for extra help.
2. Homework First: Go straight to your graded work on the OWL system. As you run into problems, find the appropriate section in the text and read through it, examining the worked examples.

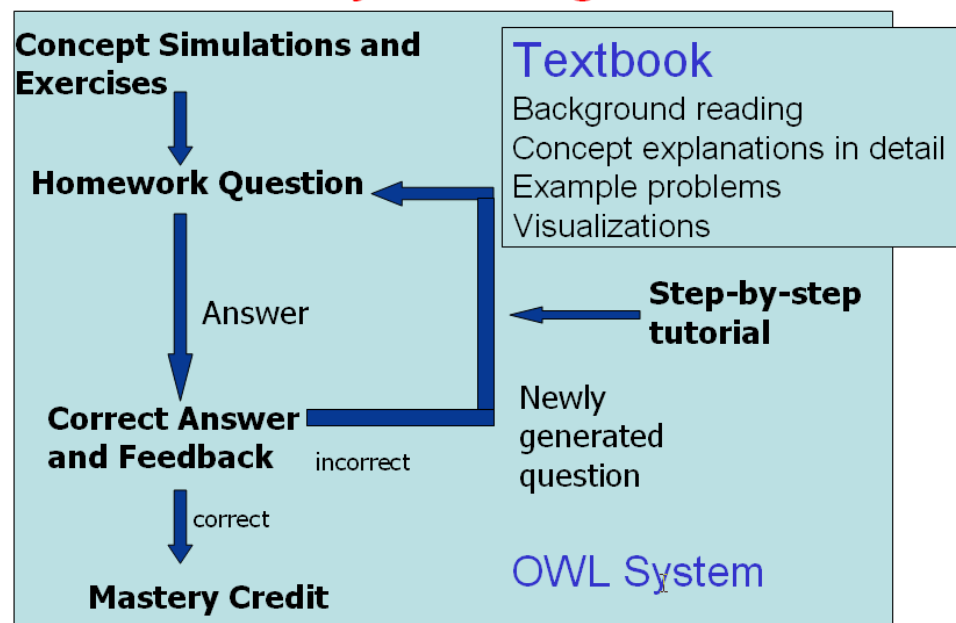
Either method is appropriate, and the package is intended to allow you to work in the way you are most comfortable.

Your work between OWL and the text follows a certain flow. The system is “mastery based.” This means that it is setup to allow you to work and practice at your own pace, and for as long as you need, to master each piece of material. Assignments in OWL contain units, which in turn contain questions.



To get credit for assignments, you must pass (that is, master) each unit within the assignment. The good news is that you get to try each section as many times as you need, and get feedback and help along the way. The Mastery Loop allows you to cycle through an area, exploring concepts, trying questions, and getting help as needed. Throughout, this text is used as the content repository, where you will find your most detailed explanations.

OWL Mastery Learning Schematic

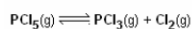


Problem Types in OWL

There are four types of assignments in OWL, each designed for a different purpose.

Homework: Questions that need to be answered. These are repeatable and give feedback on how to solve the problem. this type of assignment is most similar to traditional homework.

Consider the following reaction where $K_c = 1.20 \times 10^{-2}$ at 500 K:



A reaction mixture was found to contain 0.121 moles of $\text{PCl}_5(\text{g})$, 2.50×10^{-2} moles of $\text{PCl}_3(\text{g})$, and 3.85×10^{-2} moles of $\text{Cl}_2(\text{g})$, in a 1.00 Liter container.

Is the reaction at equilibrium?

If not, what direction must it run in order to reach equilibrium?

The reaction quotient, $Q_c =$.

The reaction:

A. Must run in the forward direction to reach equilibrium.

B. Must run in the reverse direction to reach equilibrium.

C. Is at equilibrium

CHECK ANSWER

Simulation: These are conceptual exercises designed to give you a “gut feel” for chemical systems.

6.9 Heat Transfer

Block 1
Material: Wood
Mass: 5.0 g (selected), 10.0 g
Temperature: 89 °C

Block 2
Material: Copper
Mass: 5.0 g (selected), 10.0 g
Temperature: 8 °C

Buttons: Combine, Reset

Temperature scale: 0 °C to 89 °C, with $T_f = 74 °C$ marked.

Introduction
When two objects of different temperature are brought into contact, thermal energy will flow from the warmer object to the cooler object. This will occur until the temperatures of the two objects are equal. At this point, the objects are said to be in thermal equilibrium.

Five substances are available for study. To perform an experiment, select the material of each of the two blocks. The mass of each can be set to either 5.0 g or 10.0 g. Set the initial temperature of each material using the slider below the box. Press **Combine** to bring them together. The temperature that results when thermal equilibrium is reached is reported.

Continue

Tutorials: These are step-by-step walk-throughs of important problems.

4.10 Chemical Analysis

Question
Bromine trifluoride reacts quantitatively with metal oxides. For example, the reaction with titanium dioxide,
 $3 \text{TiO}_2(\text{s}) + 4 \text{BrF}_3(\text{l}) \rightarrow 3 \text{TiF}_4(\text{s}) + 2 \text{Br}_2(\text{l}) + 3 \text{O}_2(\text{g})$
can be used to analyze the content of TiO_2 in an ore.
Determine the mass percent of TiO_2 in an ore if a sample of ore with a mass of 2.22 g is treated with BrF_3 and 0.236 g of Br_2 .

Answer
Enter a response, then Submit.
 %
INCORRECT. Answer tutorial questions.

Question 1 of 4
The first step is to convert the mass of Br_2 to units of moles.
What is the amount of Br_2 produced, in moles?
Enter the result of the operation in scientific notation. For example, if the result is 1.234×10^{-4} , enter 1.234e-5.

Answer
Enter a response, then Submit.
 mol
INCORRECT. This conversion is
quantity $\text{Br}_2(\text{g}) \times \frac{1}{\text{molar mass}} \left(\frac{\text{mol}}{\text{g}} \right) = \text{amount Br}_2(\text{mol})$
Notice that the molar mass is used inverted, with units of grams in the denominator.

Chemistry Interactive: Dissolution of KMnO_4

Introduction
Description
The dissolution of ionic compounds involves the dissociation of the cations and anions of the ionic solid in concert with the hydration of these ions as they enter solution. This module examines this process on the macroscopic and nano scales, using dissolution of potassium permanganate as an example.

Macroscopic Scale
Image showing a beaker with a solution and a small container of solid KMnO_4 .

Nano Scale
Diagram showing the dissociation of $\text{KMnO}_4(\text{s})$ into K^+ and MnO_4^- ions in $\text{H}_2\text{O}(\text{l})$.

Intro Questions

Exercises: These are visually-oriented modules designed to help connect the macroscopic world to our understanding of the world on the atomic scale.

1.1 Introduction: What Is Chemistry?

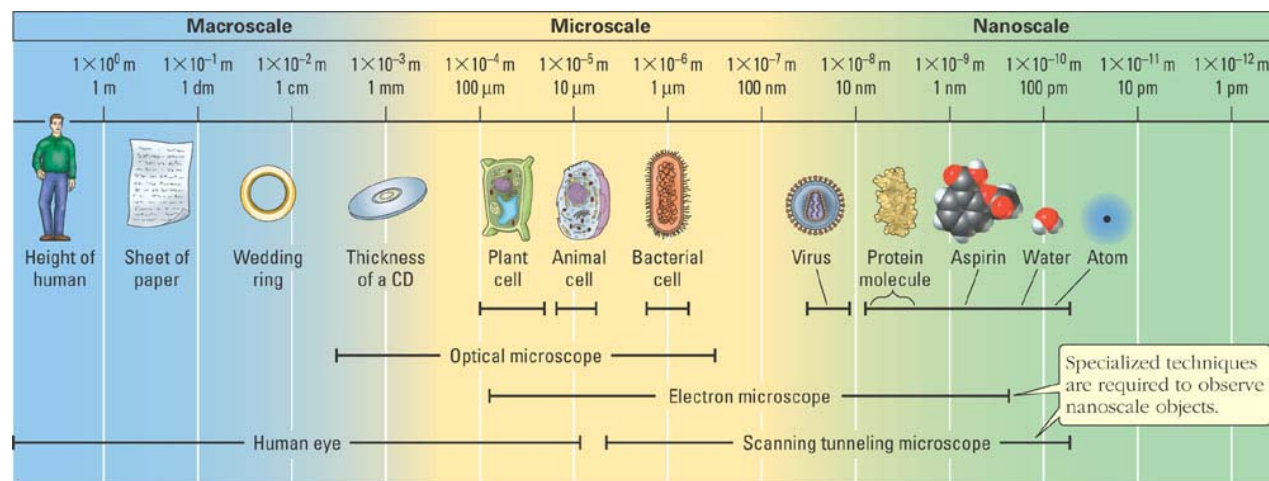


OWL Opening Exploration
1.1 Title

Chemistry is the study of matter, its transformations and how it behaves. **Matter**, any physical substance that occupies space and has mass, consists of atoms and molecules, and it is at the atomic and molecular level where chemical transformations take place.

The Scale of Chemistry

Different fields of science examine the world at different levels of detail (Figure 1.1).



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Figure 1.1 The macroscopic, microscopic, and atomic scales.

When describing matter that can be seen with the naked eye, scientists are working on the **macroscopic scale**. Chemists use the **atomic scale**, (sometimes called the nanoscale or the molecular scale) when describing individual atoms or molecules. In general, in chemistry we make observations at the macroscopic level and we describe and explain chemical processes on the atomic level. That is, we use our macroscopic scale observations to explain atomic scale properties.

Measuring Matter

Chemistry is an experimental science that involves designing thoughtful experiments and making careful observations of macroscopic amounts of matter. Everything that is known about how atoms and molecules interact has been learned through making careful observations on the macroscopic scale and inferring what those observations must mean about atomic scale objects.

For example, careful measurement of the mass of a chemical sample before and after it is heated provides information about the chemical composition of a substance. Observing how a chemical sample behaves in the presence of a strong magnetic field such as that found in an MRI scanner provides information about how molecules and atoms are arranged in human tissues.

An important part of chemistry and science in general is the idea that all ideas are open to challenge. When we perform measurements on chemical substances and interpret the results in terms of atomic scale properties, the results are always examined to see if there are alternative ways to interpret the data. This method of investigation leads to chemical information about the properties and behavior of matter that is supported by the results of many different experiments.

EXAMPLE PROBLEM: The Scale of Chemistry

Classify each of the following as matter that can be measured or observed on either the macroscopic or atomic scale.

- (a) An RNA molecule
- (b) A mercury atom
- (c) A sample of liquid mercury

SOLUTION:

- (a) Atomic scale. An RNA molecule is too small to be seen with the naked eye or with an optical microscope.
- (b) Atomic scale. Individual atoms cannot be seen with the naked eye or with an optical microscope.
- (c) Macroscopic scale. Liquid mercury can be seen with the naked eye.



OWL Example Problems
 1.3 The Scale of Chemistry
 Macroscopic, Microscopic & Particulate (1403)

1.2 Classification of Matter

OWL Opening Exploration
 1.X Title

Matter is characterized by a collection of characteristics that describes the material, called **properties**. One of the fundamental properties of matter is its composition, or the specific types of atoms or molecules that make it up.

Classifying Matter on the Atomic Scale

An **element** is the simplest building block of matter. You are already familiar with some of the most common elements such as gold, silver and copper, used in making coins and jewelry, and oxygen, nitrogen and argon, the three most abundant gases in our atmosphere. There are 116 elements that have been identified, 90 of which exist in nature (the rest have been synthesized in the laboratory). Elements are represented using a one- or two-letter element symbol and they are organized in the periodic table that is shown on the inside front cover of this textbook. A few common elements and their symbols are shown in Table 1.1.

An **atom** is the smallest indivisible unit of an element. For example, the element aluminum is made up entirely of aluminum atoms as shown in Figure 1.2. Although individual atoms are too small to see directly, many experiments enable them to be visualized. Both experimental observations and theoretical studies show that isolated atoms are spherical, and that atoms of different elements are different in size. The model used to represent isolated atoms therefore consists of spheres of different sizes.

Elements are made up of only one type of atom. For example, the element oxygen is found in two forms, as O_2 and O_3 . The most common form of oxygen is O_2 , dioxygen, a gas that makes up about 22% of the air we breathe. Ozone, O_3 , is a gas with a distinct odor that can be toxic to humans. Both dioxygen and ozone are elemental forms of oxygen because they consist of only one type of atom.



Dioxygen, O_2



Ozone, O_3

A **chemical compound** is a substance formed when two or more elements are combined in a defined ratio. Compounds differ from elements in that they can be broken down chemically into simpler substances. You have encountered chemical compounds in many common substances such as table salt, a compound consisting of the elements

Flashforward

2.2 Elements and the Periodic Table

Table 1.1 Some Common Elements and their Symbols

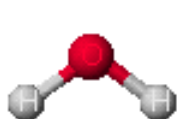
Name	Symbol
Hydrogen	H
Carbon	C
Oxygen	O
Sodium	Na
Silicon	Si
Iron	Fe

Notice that the symbol for an element consists of one or two letters. In two-letter element symbols, only the first letter is capitalized.

Pictures of some elements and compounds

sodium and chlorine, and phosphoric acid in soft drinks, a compound containing hydrogen, oxygen and phosphorus.

Molecules are small collections of atoms that are held together by chemical bonds. In models used to represent molecules, chemical bonds are often represented using cylinders or lines that connect atoms, represented as spheres. The composition and arrangement of elements in molecules affects the properties of a substance. For example, molecules of both water (H_2O) and hydrogen peroxide (H_2O_2) contain only the elements hydrogen and oxygen.



Water, H_2O



Hydrogen peroxide, H_2O_2

Water is a relatively inert substance that is safe to drink in its pure form. Hydrogen peroxide, however, is a reactive liquid that is used to disinfect wounds and can cause severe burns if swallowed.

EXAMPLE PROBLEM: Elements and Compounds

Classify each of the following substances as either an element or a compound.

- (a) Nitrogen
- (b) CO_2
- (c) P_4

SOLUTION:

- (a) Element. Nitrogen is an example of an element because it consists of only one type of atom.
- (b) Compound. This compound contains both carbon and oxygen.
- (c) Element. While this is an example of a molecular substance, it only consists of a single type of atom.

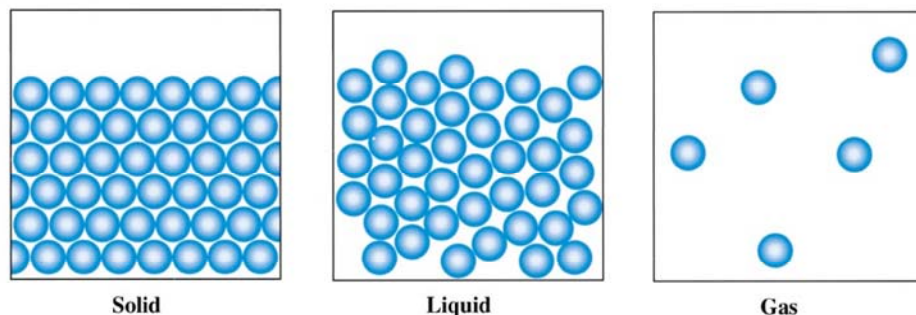


OWL Example Problems
1.4 Elements and Compounds

Classifying Matter on the Macroscopic Scale – Pure Substances

A **pure substance** contains only one type of element or compound and has fixed chemical composition. A pure substance also has characteristic properties, measurable qualities that are independent of the sample size. The **physical properties** of a chemical substance are those that do not change the chemical composition of the material when they are measured. Some examples of physical properties include physical state, color, viscosity (resistance to flow), opacity, density, conductivity, and melting and boiling points.

One of the most important physical properties is the physical state of a material. The three physical **states of matter** are solid, liquid and gas (Figure 1.6). The macroscopic properties of these states are directly related to the arrangement and properties of particles at the atomic level.



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Figure 1.6 The three physical states of matter

At the macroscopic level, a **solid** is a dense material with a defined shape. At the atomic level, the atoms or molecules of a solid are packed together closely. The atoms or molecules are vibrating but they do not move past one another. At the macroscopic level, a **liquid** is also dense, but unlike a solid it flows and takes on the shape of its container. At the atomic level, the atoms or molecules of a liquid are close together but they move more than the particles in a solid and can flow past one another. Finally, at the macroscopic level a **gas** is diffuse and has no fixed shape or volume. At the atomic level, the atoms or molecules of a gas are spaced widely apart and are moving rapidly past one another. The particles of a gas do not strongly interact with one another and they move freely until they collide with one another or with the walls of the container.

The physical state of a substance can change when energy, often in the form of heat, is added or removed. When energy is added to a solid, the temperature at which the solid is converted to a liquid is the **melting point** of the substance. The conversion of liquid to solid occurs at the same temperature as energy is removed (the temperature falls) and is called the **freezing point**. A liquid is converted to a gas at the **boiling point** of a substance. As you will see in the following section, melting and boiling points are measured in Celsius (°C) or Kelvin (K) temperature units.

A change in the physical property of a substance is called a **physical change**. Physical changes may change the appearance or the physical state of a substance, but they do not change its chemical composition. For example, a change in the physical state of hydrogen—changing from a liquid to a gas—involves a change in how the particles are packed together at the atomic level, but does not change the chemical makeup of the material.

The **chemical properties** of a substance are those that involve a chemical change in the material and often involve a substance interacting with other chemicals. For example, a chemical property of hydrogen is that it is highly flammable because the element burns in air (it reacts with oxygen in the air) to form water. A **chemical change** involves a change in the chemical composition of the material. The flammability of hydrogen is a chemical property and demonstrating this chemical property involves a chemical change.

Not all materials can exist in all three physical states. Polyethylene, for example, does not exist as a gas. Heating a solid polyethylene milk bottle at high temperatures causes it to decompose into other substances. Helium, a liquid at room temperature, can be liquified at very low temperatures, but it is not possible to solidify helium.

EXAMPLE PROBLEM: Physical and Chemical Properties and Changes

- (a) When aluminum foil is placed into liquid bromine a white solid forms. Is this a chemical or physical property of aluminum?
- (b) Iodine is a purple solid. Is this a chemical or physical property of iodine?
- (c) Classify each of the following changes as chemical or physical.
- (i) Boiling water
 - (ii) Baking bread

SOLUTION:

- (a) Chemical property. Chemical properties are those that involve a chemical change in the material and often involve a substance interacting with other chemicals. In this example, one substance (the aluminum) is converted into something new (a white solid).
- (b) Physical property. A physical property such as color is observed without changing the chemical identity of the substance.
- (c) (i) Physical change. A physical change alters the physical form of a substance without changing its chemical identity. Boiling does not change the chemical composition of water.
- (ii) Chemical change. When a chemical change takes place, the original substances (the bread ingredients) are destroyed and a new substance (bread) is formed.

*OWL Example Problems**1.9 Physical and Chemical Properties and Changes***Classifying Matter on the Macroscopic Scale – Mixtures**

As you observe when you look around you, the world is made of complex materials. Much of what surrounds us is made up of mixtures of different substances. A **mixture** is defined as a substance made up of two or more elements or compounds (Figure 1.4) that have not reacted chemically.

Unlike compounds, where the ratio of elements is fixed, the relative amounts of different components in a mixture can vary. Mixtures that have a constant composition throughout the material are called **homogeneous mixtures**. For example, dissolving table salt in water creates a mixture of the two chemical compounds water (H_2O) and table salt (NaCl). Because the mixture is uniform, meaning that the same ratio of water to table salt is found no matter where it is sampled, it is a homogeneous mixture.

A mixture in which the composition is not uniform is called a **heterogeneous mixture**. For example, a cold glass of freshly squeezed lemonade is a heterogeneous mixture because you can see the individual components (ice cubes, lemonade, and pulp) and the relative amounts of each component will depend on where the lemonade is sampled (from the top of the glass or from the bottom).

Homogeneous and heterogeneous mixtures can usually be physically separated into individual components. For example, a homogeneous mixture of salt and water is separated by heating the mixture to evaporate the water, leaving behind the salt. A heterogeneous mixture of sand and water is separated by pouring the mixture through filter paper. The sand is trapped in the filter while the water passes through.

Like pure substances, mixtures have physical and chemical properties. These properties, however, depend on the composition of the mixture. For example, a mixture of 10. grams of table sugar and 100. grams of water has a boiling point of 100.15°C while a mixture of 20. grams of table sugar and 100. grams of water has a boiling point of 100.30°C .

Figure 1.4 Homogeneous and Heterogeneous Mixtures

Figure to come

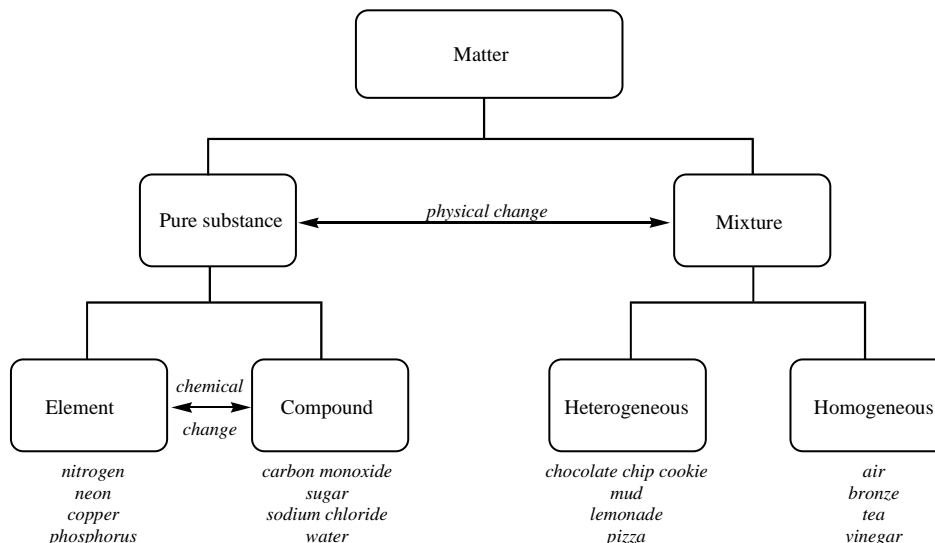


Figure 1.5 Matter flow chart

EXAMPLE PROBLEM: Pure Substances and Mixtures

Classify each of the following as a pure substance, a homogeneous mixture or a heterogeneous mixture:

- (a) Copper wire
- (b) Oil and vinegar salad dressing
- (c) Gasoline

SOLUTION:

- (a) Pure substance. Copper is an element.
- (b) Heterogeneous mixture. The salad dressing is a mixture that does not have a uniform composition.
- (c) Homogeneous mixture. Gasoline, a transparent liquid, is a uniform mixture of different compounds.



OWL Example Problems
 1.5 Pure Substances and Mixtures: Tutor
 1.6 Pure Substances and Mixtures

1.4 Units and Measurement

OWL Opening Exploration
 1.4 Title

Chemistry involves observing matter, and our observations are substantiated by careful measurements of physical quantities. Chemists in particular need to make careful measurements because we use those measurements to infer the properties of matter on the atomic scale.

Scientific Units

Some of the most common measurements in chemistry are mass, volume, time, temperature, and density. Measuring these quantities allows us to describe the chemical and physical properties of matter and to study the chemical and physical changes that matter undergoes. When reporting a measurement, we use scientific units to indicate what was measured. **SI units**, abbreviated from the French *Système International d'Unités*, are used in scientific measurements in almost all countries. This unit system consists of seven base units. Other units are called derived units and are combinations of the base units (Table 1.2).

Table 1.2 SI Units**Base Units:**

Physical quantity	SI Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

Some Derived Units:

Physical quantity	Symbol
Volume	m^3 (cubic meter)
Density	kg/m^3
Energy	J (joule)

Metric prefixes are combined with SI units when reporting physical quantities in order to reflect the relative size of the measured quantity. Table 1.3 shows the metric prefixes most commonly used in scientific measurements.

Table 1.3*Common Prefixes Used in the SI and Metric Systems*

Prefix	Abbreviation	Meaning	Example
mega-	M	10^6	1 megameter (Mm) = 1×10^6 m
kilo-	k	10^3	1 kilometer (km) = 1×10^3 m
deci-	d	10^{-1}	1 decimeter (dm) = 1×10^{-1} m
centi-	c	10^{-2}	1 centimeter (cm) = 1×10^{-2} m
milli-	m	10^{-3}	1 milligram (mg) = 1×10^{-3} g
micro-	μ	10^{-6}	1 microgram (μg) = 1×10^{-6} g
nano-	n	10^{-9}	1 nanogram (ng) = 1×10^{-9} g
pico-	p	10^{-12}	1 picogram (pg) = 1×10^{-12} g

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When making and reporting measurements, it is important to use both the value and an appropriate unit. For example, in the United States, speed limits are reported in units of miles per hour (mph). A US citizen traveling to Canada might see a speed limit sign reading 100 and assume the units are mph. This might be an expensive mistake, however, because speed limits in Canada are reported in units of kilometers per hour (km/h), and so a 100 km/h speed limit is the equivalent of 62 mph.

Scientific Notation

Numbers that are very large or very small can be represented using **scientific notation**. A number written in scientific notation has the general form $N \times 10^x$, where N is a number between 1 and 10, and x is a positive or negative integer. For example, the number 13433 is written as 1.3433×10^4 and the number 0.0058 is written as 5.8×10^{-3} in scientific notation. Notice that x is positive for numbers greater than 1 and negative for numbers less than 1.

To convert a number from standard notation to scientific notation, count the number of places the decimal must be moved to the right (for numbers less than 1) or to the left (for numbers greater than 1) in order to result in a number between one and ten. For the number 13433,

1 3 4 3 3

the decimal is moved 4 times to the left and the number is written 1.3433×10^4 . When a number is less than 1, the decimal is moved to the right and the exponent (x) is negative. For the number 0.0058,

0.0058

the decimal is moved 3 times to the right and the number is written 5.8×10^{-3} . Notice that in both cases, moving the decimal is the equivalent of multiplying or dividing by 10.

To convert a number from exponential notation to standard notation, write the value of N and then move the decimal point x places to the right if x is positive or move the decimal point x places to the left if x is negative.

EXAMPLE PROBLEM: Scientific Notation

(a) Write the following numbers in scientific notation:

- (i) 0.0000422
- (ii) 9700000000

(b) Write the following numbers in standard notation:

- (i) 7.22×10^6
- (ii) 2.5×10^{-3}

SOLUTION

- (a) (i) Moving the decimal 5 times to the right results in a number between 1 and 10. The exponent is negative because this number is less than 1.
 4.22×10^{-5}
- (ii) Moving the decimal 9 times to the left results in a number between 1 and 10. The exponent is positive because this number is greater than 1.
 9.7×10^9
- (b) (i) Move the decimal 6 times to the right.
 7220000
- (ii) Move the decimal 3 times to the left.
 0.0025



OWL Example Problems

1.13a Scientific Notation 1: Tutor

1.13a Scientific Notation 2: Tutor

1.13a Scientific Notation 3: Tutor

1.13b Exponential (Scientific) Notation

SI Base Units: Length, Mass and Temperature

Length The SI unit of length is the meter, m. A pencil has a length of about 0.16 m, which is equivalent to 16 centimeters (cm). Atomic radii are measured using nanometer (nm) or picometer (pm) units. The definition of the meter is based on the speed of light in a vacuum, exactly 299,792,458 meters per second. One meter is therefore the length of the path traveled by light in vacuum during $1/299,792,458$ of a second.

Mass The SI unit of mass is the kilogram, kg. This is the only SI base unit that contains a metric prefix. One kg is equal to approximately 2.2 pounds (lb). In the chemistry lab, the mass of a sample is typically measured using units of grams (g) or milligrams (mg). The kilogram standard is the mass of a piece of platinum-iridium alloy that is kept at the International Bureau of Weights and Measures.

Temperature Temperature is a relative measure of how hot or cold a substance is, and is commonly reported using one of three temperature scales. In the United States, temperatures are commonly reported using the Fahrenheit temperature scale that has units of degrees Fahrenheit ($^{\circ}\text{F}$). In scientific measurements, the Celsius and Kelvin temperature scales are used, with units of degrees Celsius ($^{\circ}\text{C}$) and kelvins (K), respectively. Notice that for the Kelvin temperature scale, the name of the temperature unit (kelvin) is not capitalized but the abbreviation (K) is capitalized.

As shown in Figure 1.X, the three temperature scales have different defined values for the melting and freezing points of water. In the Fahrenheit scale, the freezing point of water is set at 32°F and the boiling point is 180 degrees higher, 212°F . In the Celsius

temperature scale, the freezing point of water is assigned a temperature of 0 °C and the boiling point of water is assigned a temperature of 100 °C. The lowest temperature on the Kelvin temperature scale, 0 K, is 273.15 degrees lower than 0 °C.

$$0 \text{ K} = -273.15 \text{ }^{\circ}\text{C}$$

This temperature, known as **absolute zero**, is the lowest temperature possible. The Celsius and Kelvin temperature scales are similar in that a one-degree increment is the same on both scales. That is, an increase of 1 K is equal to an increase of 1 °C. Celsius and Kelvin temperatures are related as shown in the equations below:

$$T (\text{K}) = T (^{\circ}\text{C}) + 273.15 \quad (1.x)$$

$$T (^{\circ}\text{C}) = T (\text{K}) - 273.15 \quad (1.x)$$

Notice that the Fahrenheit and Celsius temperature scales differ in the size of a degree.

$$180 \text{ Fahrenheit degrees} = 100 \text{ Celsius degrees}$$

$$\frac{9}{5} \text{ Fahrenheit degrees} = 1 \text{ Celsius degree}$$

Fahrenheit and Celsius temperatures are related as shown in the equations below:

$$T (^{\circ}\text{F}) = \frac{9}{5} [T (^{\circ}\text{C})] + 32 \quad (1.x)$$

$$T (^{\circ}\text{C}) = \frac{5}{9} [T (^{\circ}\text{F}) - 32] \quad (1.x)$$

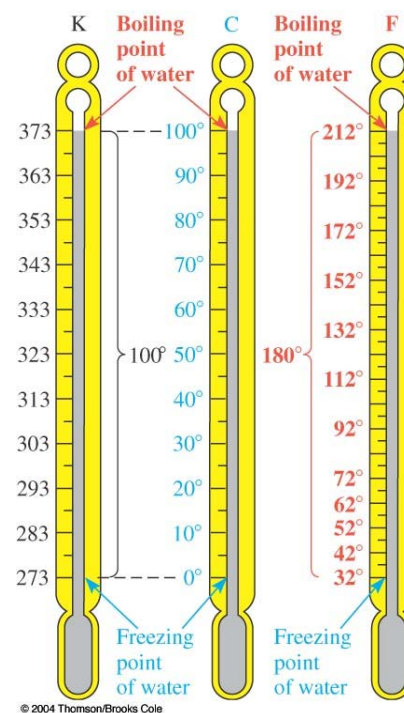


Figure 1.X Portions of the Kelvin, Celsius, and Fahrenheit Temperature Scales

EXAMPLE PROBLEM: Temperature

The boiling point of a liquid is 355.78 K. What is this temperature on the Celsius and Fahrenheit scales?

SOLUTION:

Convert the temperature to Celsius temperature units.

$$T (\text{K}) = T (^{\circ}\text{C}) + 273.15$$

$$355.78 \text{ K} = T (^{\circ}\text{C}) + 273.15$$

$$T (^{\circ}\text{C}) = 355.78 \text{ K} - 273.15 = 82.63 \text{ }^{\circ}\text{C}$$

Use the temperature in °C units to calculate the temperature on the Fahrenheit scale.

$$T (^{\circ}\text{F}) = \frac{9}{5} [T (^{\circ}\text{C})] + 32$$

$$T (^{\circ}\text{F}) = \frac{9}{5} (82.63 \text{ }^{\circ}\text{C}) + 32$$

$$T (^{\circ}\text{F}) = 180.73 \text{ }^{\circ}\text{F}$$



OWL Example Problems

1.13 Temperature: Tutor

1.14 Temperature

SI Derived Units: Volume, Density and Energy

Volume While the SI unit of volume is the cubic meter (m^3), a more common volume unit is the liter (L). Notice that the abbreviation for liter is a capital “L.” A useful relationship to remember is that one milliliter is equal to one cubic centimeter ($1 \text{ mL} = 1 \text{ cm}^3$).

Energy The SI unit of energy is the joule, J. Another common energy unit is the calorie (cal), and one calorie is equal to 4.184 joules ($1 \text{ cal} = 4.184 \text{ J}$). One dietary calorie (Cal) is equal to 1000 calories ($1 \text{ Cal} = 1000 \text{ cal} = 1 \text{ kcal}$).

Density The **density** of a substance is a physical property that relates the mass of a substance to its volume (Equation 1.x).

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad (1.x)$$

The densities of solids and liquids are reported in units of grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL) while the density of a gas is typically reported in units of grams per liter (g/L). Because the volume of most substances changes with a change in temperature, density also changes with temperature. Most density values are reported at a standard temperature, 25°C , close to room temperature. The densities of some common substances are listed in Table 1.x.

Density can be calculated from mass and volume data as shown in the following example.

EXAMPLE PROBLEM: Density

A 5.78-mL sample of a colorless liquid has a mass of 4.54 g. Calculate the density of the liquid and identify it as either ethanol (density = 0.785 g/mL) or benzene (density = 0.874 g/mL).

SOLUTION:

Use equation 1.x to calculate the density of the liquid.

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{4.54 \text{ g}}{5.78 \text{ mL}} = 0.785 \text{ g/mL}$$

The liquid is probably ethanol.

Table 1.X Densities of Some Common Substances at 25°C

to come



OWL Example Problems
1.17 Density 1: Tutor
1.19a Density

Significant Figures

The certainty in any measurement is limited by the instrument that is used to make the measurement. For example, an orange weighed on a grocery scale weighs about 157 g. A standard laboratory balance, however, might report the mass of the same orange as 156.719 g. In both cases, some uncertainty is present in the measurement. The grocery scale measurement is certain to the nearest 1 g, and the value is reported as $157 \pm 1 \text{ g}$. The laboratory scale has less uncertainty, and is reported as $156.719 \pm 0.001 \text{ g}$. In general, we will drop the \pm symbol and assume an uncertainty of one unit in the rightmost digit when reading a measurement. When using a measuring device, we always estimate the rightmost digit when reporting the measured value, as shown in Figure 1.x.

Some measured quantities are infinitely certain, or exact. For example, the number of oranges you have purchased at the grocery store is an exact number. Some units are defined with exact numbers, such as the metric prefixes ($1 \text{ mm} = 0.001 \text{ m}$) and the relationship between inches and centimeters ($1 \text{ in} = 2.54 \text{ cm}$, exactly).

The digits in a measurement, both the certain and uncertain digits, are called **significant figures** or significant digits. For example, the mass of an orange has three significant figures when measured using a grocery scale (157 g) and six significant figures when measured on a laboratory balance (156.719 g). Some simple rules are used to determine the number of significant figures in a measurement (Table 1.x).

Figure 1.X Measuring Devices

Figure showing two scales or rulers or thermometers than can be read to different certainty

TABLE 1.X RULES FOR DETERMINING SIGNIFICANT FIGURES IN A NUMBER

1. All non-zero digits and zeros between non-zero digits are significant.
Example: In 0.03080, the digits 3 and 8 and the zero between 3 and 8 are significant. In 728060, the digits 7, 2, 8 and 6 and the zero between 8 and 6 are significant.
2. In numbers containing a decimal point,
 - (a) all zeros at the end of the number are significant.
Example: In 0.03080, the zero to the right of 8 is significant.
 - (b) all zeros at the beginning of the number are not significant.
Example: In 0.03080, the two zeros to the left of 3 are not significant.
3. In numbers with no decimal point, all zeros at the end of the number are not significant.
Example: In 728060, the zero to the right of 6 is not significant.
The number 0.03080 has 4 significant figures and the number 728060 has 5 significant figures

Notice that for numbers written in scientific notation, the number of significant figures is equal to the number of digits in the number written before the exponent. For example, the number 3.25×10^{-4} has three significant figures and 1.200×10^3 has four significant figures.

EXAMPLE PROBLEM: Significant Figures

Identify the number of significant figures in the following numbers.

- (a) 19.5400
- (b) 0.0095
- (c) 1030

SOLUTION:

- (a) All non-zero digits are significant (there are 4) and because this number has a decimal point, the zeros at the end of the number are also significant (there are 2). This number has 6 significant figures.
- (b) All non-zero digits are significant (there are 2) and because this number has a decimal point, the 3 zeros at the beginning of the number are not significant. This number has 2 significant figures.
- (c) All non-zero digits are significant (there are 2) and the zero between the non-zero digits is also significant (there is 1). Because this number has no decimal, the zero at the end of the number is not significant. This number has 3 significant figures.



OWL Example Problems
1.13c Significant Figures I

Significant Figures and Calculations

When performing calculations involving measurements, the final calculated result is no more certain than the least certain number in the calculation. If necessary, the answer is rounded off to the correct number of significant figures. For example, consider a density calculation involving a sample with a mass of 3.2 g and a volume of 25.67 cm³. A standard calculator reports a density of 0.124659135 g/cm³, but is this a reasonable number? In this case, the least certain number is the sample mass with two significant figures, and the calculated density therefore has two significant figures, or 0.12 g/cm³.

The final step in a calculation usually involves rounding the answer so that it has the correct certainty. When rounding numbers, the last digit retained is increased by 1 only if the digit that follows is 5 or greater. When performing calculations involving multiple steps, it is best to round off only at the final step in the calculation. That is, carry at least one extra significant figure during each step of the calculation so that you do not introduce rounding errors in the final calculated result.

When performing multiplication or division, the certainty in the answer is related to the significant figures in the numbers in the calculation. The answer has the same number of significant figures as the measurement with the fewest significant figures. For example,

$$\begin{array}{lcl} 39.485 \times 6.70 = 264.5495 & \text{Round off to } 265 \\ \text{Number of significant figures: } \begin{array}{cc} 5 & 3 \end{array} & & 3 \end{array}$$
$$\begin{array}{lcl} 0.029 / 1.285 = 0.022568093 & \text{Round off to } 0.023 \\ \text{Number of significant figures: } \begin{array}{cc} 2 & 4 \end{array} & & 2 \end{array}$$

When doing addition or subtraction, the certainty of the answer is related to the decimal places present in the numbers in the calculation. The answer has the same number of decimal places as the number with the fewest decimal places. For example,

$$\begin{array}{lcl} 12.50 + 6.080 = 18.580 & \text{Round off to } 18.58 \\ \text{Number of decimal places: } \begin{array}{cc} 2 & 3 \end{array} & & 2 \end{array}$$
$$\begin{array}{lcl} 125.2 - 0.07 = 125.13 & \text{Round off to } 125.1 \\ \text{Number of decimal places: } \begin{array}{cc} 1 & 2 \end{array} & & 1 \end{array}$$

Exact numbers do not limit the number of significant figures or the number of decimal places in a calculated result. For example, suppose you want to convert the amount of time it takes for a chemical reaction to take place, 7.2 minutes, to units of seconds. There are exactly 60 seconds in 1 minute, so it is the number of significant figures in the measured number that determines the number of significant figures in the answer.

$$\begin{array}{lcl} 7.2 \text{ minutes} \times 60 \text{ seconds/minute} = 430 \text{ seconds} \\ \text{Number of significant figures: } \begin{array}{ccc} 2 & \text{exact number} & 2 \end{array} \end{array}$$



OWL Example Problems
1.13d Significant Figures 2

Precision and Accuracy

Along with the number of significant figures in a value, the certainty of a measurement can also be described using the terms precision and accuracy. **Precision** is how close the values in a set of measurements are to each other. **Accuracy** is how close a measurement or a set of measurements is to a real value. Figure 1.x demonstrates the relationship between precision and accuracy in a set of laboratory measurements.

1.5 Unit Conversions



OWL Opening Exploration
1.5 Title

Quantities of matter, energy or time can be expressed using different units. It is very useful to be able to convert the representation of a quantity from one unit (or set of units) to another. To make these conversions we use conversion factors derived from equalities. For example, the equality that relates hours and minutes is 1 hour = 60 minutes. Both 1 hour and 60 minutes represent the same amount of time but with different units. When an equality is written as a ratio, it is called a **conversion factor**. We use conversion factors to change the units of a quantity without changing the amount using the method called **dimensional analysis**. In dimensional analysis, the original value is multiplied by the conversion factor. Because both the numerator and the denominator of a conversion factor are equal, the ratio is equal to one and multiplying by this ratio does not change the original value, it only changes the units in which it is expressed.

Figure 1.X Precision and Accuracy

Figure showing sets of data

For example, you can use the conversion factor 1 hour = 60 minutes to express 48 hours in units of minutes, as shown in the following example.

EXAMPLE PROBLEM: Dimensional Analysis

Using dimensional analysis, express 48 hours in units of minutes (1 hour = 60 minutes).

SOLUTION:

Step 1. Write the quantity in its current units and identify the desired new units:

$$48 \text{ hours} \times \text{conversion factor} = \text{_____ minutes}$$

Step 2. Identify an equality that relates the current units to the desired units and write it as a ratio.

The equality 1 hour = 60 minutes can be written as a ratio in two ways:

$$\frac{1 \text{ hour}}{60 \text{ min}} \quad \text{or} \quad \frac{60 \text{ min}}{1 \text{ hour}}$$

Because our goal is to convert units of hours to units of minutes we will use the second form of the conversion factor. Remember that we will multiply the original value by the conversion factor. The second form of the conversion factor will allow the hours units to be cancelled in the multiplication.

Step 3. Multiply the original value by the conversion factor and make sure that the original units are cancelled by the same units in the denominator. The desired units should appear in the numerator.

$$48 \text{ hours} \times \frac{60 \text{ min}}{1 \text{ hour}} = \text{_____ minutes}$$

Notice that the amount of time will not change when we perform this multiplication because the conversion factor is derived from an equality and is therefore equal to 1 (60 minutes is equal to 1 hour).

Step 4. Perform the multiplication and cancel any units that appear in both the numerator and denominator.

$$48 \cancel{\text{ hours}} \times \frac{60 \text{ min}}{1 \cancel{\text{ hour}}} = 2900 \text{ minutes}$$

Notice that the answer is reported to 2 significant figures. The equality 1 hour = 60 minutes is exact and therefore has an infinite number of significant figures. The initial value, 48 hours, has two significant figures.

It is common to combine conversion factors in sequence to perform a unit conversion if you are not provided with a single conversion factor that converts from the given unit to the desired unit, as shown in the following example.

EXAMPLE PROBLEM: Dimensional Analysis Using More Than One Conversion Factor

Using dimensional analysis, express the volume of soda in a soft drink can (12 oz.) in units of mL.

(1 gallon = 128 oz; 1 gallon = 3.7854 L; 1 L = 1000 mL)

SOLUTION:

Step 1. Write the quantity in its current units and identify the desired new units

$$12 \text{ oz.} \times \text{conversion factor} = \text{_____ mL}$$

Step 2. Identify equalities that will convert the given units to the desired units and write them as ratios.

First convert units of ounces (oz.) to units of gallons (gal). The next conversion changes units of gallons (gal.) to units of liters (L). Finally, convert units of liters (L) to units of milliliters (mL).

$$\text{Conversion factor 1: } \frac{1 \text{ gal}}{128 \text{ oz.}} \quad \text{Conversion factor 2: } \frac{3.7854 \text{ L}}{1 \text{ gal}} \quad \text{Conversion factor 3: } \frac{1000 \text{ mL}}{1 \text{ L}}$$

Step 3. Multiply the original value by the conversion factors and make sure that the original units are cancelled by the same units in the denominator. The desired units should appear in the numerator.

$$12 \text{ oz.} \times \frac{1 \text{ gal}}{128 \text{ oz.}} \times \frac{3.7854 \text{ L}}{1 \text{ gal}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = \text{_____ mL}$$

Step 4. Perform the multiplication and division and cancel any units that appear in both the numerator and denominator.

$$12 \cancel{\text{ oz.}} \times \frac{1 \cancel{\text{ gal}}}{128 \cancel{\text{ oz.}}} \times \frac{3.7854 \cancel{\text{ L}}}{1 \cancel{\text{ gal}}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{ L}}} = 350 \text{ mL}$$

Notice that the initial value limits the answer to 2 significant figures.

*OWL Example Problems**1.15 Dimensional Analysis: Tutor**1.15a Metric System Prefixes: Tutor**1.16 Dimensional Analysis***Unit Conversions using Density**

Density, the relationship between mass and volume, is also an equality that relates the mass and volume of a substance. As shown in the following example, density can be used to convert units of mass to units of volume.

EXAMPLE PROBLEM: Dimensional Analysis Using Density

Use the density of diamond ($d = 3.52 \text{ g/cm}^3$) to determine the volume of a 1.2-carat diamond.

(1 carat = 0.200 g)

SOLUTION:

Step 1. Write the quantity in its current units and identify the desired new units

$$1.2 \text{ carat} \times \text{conversion factor} = \text{_____ cm}^3$$

Step 2. Identify equalities that will convert the given units to the desired units and write them as ratios.

First convert from carats to units of grams. Next, use density to convert mass units (g) to volume units (cm^3).

$$\text{Conversion factor 1: } \frac{0.200 \text{ g}}{1 \text{ ct.}} \quad \text{Conversion factor 2: } \frac{1 \text{ cm}^3}{3.52 \text{ g}}$$

Notice that the density conversion factor is inverted. Density can be expressed as an equality ($3.52 \text{ g} = 1 \text{ cm}^3$), so it can be written with either mass or volume in the numerator. In this example, the desired unit, cubic centimeters (cm^3), appears in the numerator.

Step 3. Multiply the original value by the correct conversion factors and make sure that the original unit is cancelled by the same unit in the denominator. The desired unit should appear in the numerator.

$$1.2 \text{ carat} \times \frac{0.200 \text{ g}}{1 \text{ ct.}} \times \frac{1 \text{ cm}^3}{3.52 \text{ g}} = \text{_____ cm}^3$$

Step 4. Perform the multiplication and cancel any units that appear in both the numerator and denominator.

$$1.2 \cancel{\text{carat}} \times \frac{0.200 \cancel{\text{g}}}{1 \cancel{\text{ct.}}} \times \frac{1 \text{ cm}^3}{3.52 \cancel{\text{g}}} = 0.68 \text{ cm}^3$$

*OWL Example Problems**1.18 Density 2: Tutor**1.19b Density*