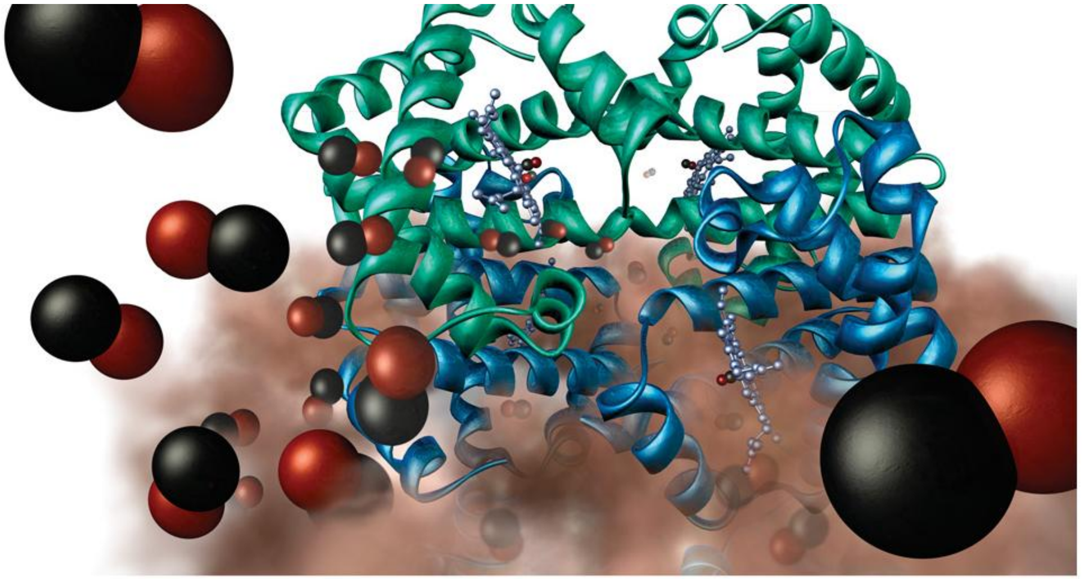
Chemistry: A Molecular Approach, 1st Ed.
Nivaldo Tro

Chapter 1
Matter, Measurement, and Problem Solving



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2008, Prentice Hall

Composition of Matter

Atoms and Molecules

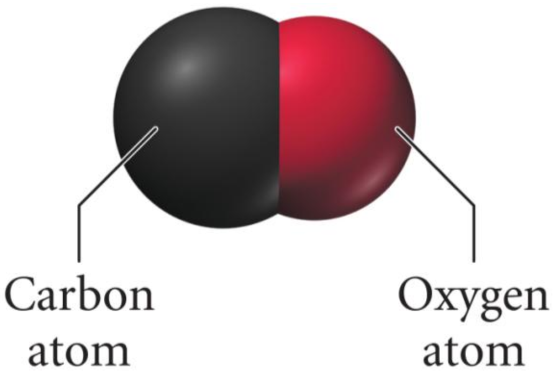
Scientific Method

Structure Determines Properties

- the properties of matter are determined by the atoms and molecules that compose it

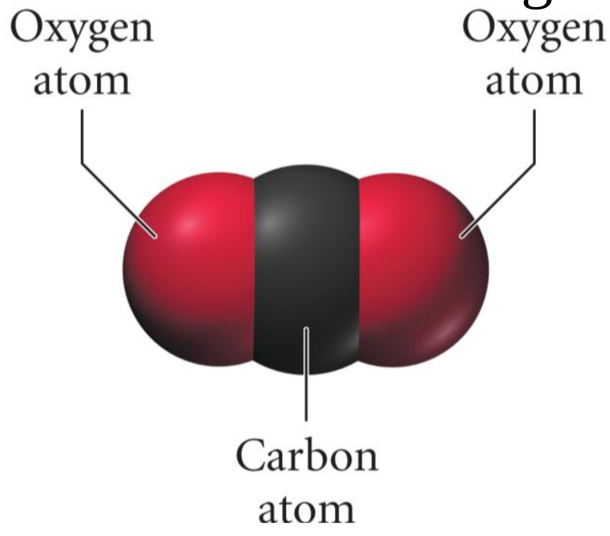
carbon monoxide

1. composed of one carbon atom and one oxygen atom
2. colorless, odorless gas
3. burns with a blue flame
4. binds to hemoglobin



carbon dioxide

1. composed of one carbon atom and two oxygen atoms
2. colorless, odorless gas
3. incombustible
4. does not bind to hemoglobin



Atoms and Molecules

- **atoms**
  - ✓ are submicroscopic particles
  - ✓ are the fundamental building blocks of all matter
- **molecules**
  - ✓ two or more atoms attached together
    - ➢ attachments are called bonds
    - ➢ attachments come in different strengths
  - ✓ molecules come in different shapes and patterns
- **Chemistry** is the science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules

The Scientific Approach to Knowledge

- philosophers try to understand the universe by reasoning and thinking about “ideal” behavior
- scientists try to understand the universe through empirical knowledge gained through observation and experiment

From Observation to Understanding

- **Hypothesis** – a tentative interpretation or explanation for an observation
  - ✓ falsifiable – confirmed or refuted by other observations
  - ✓ tested by experiments – validated or invalidated
- when similar observations are consistently made, it can lead to a **Scientific Law**
  - ✓ a statement of a behavior that is always observed
  - ✓ summarizes past observations and predicts future ones
  - ✓ Law of Conservation of Mass

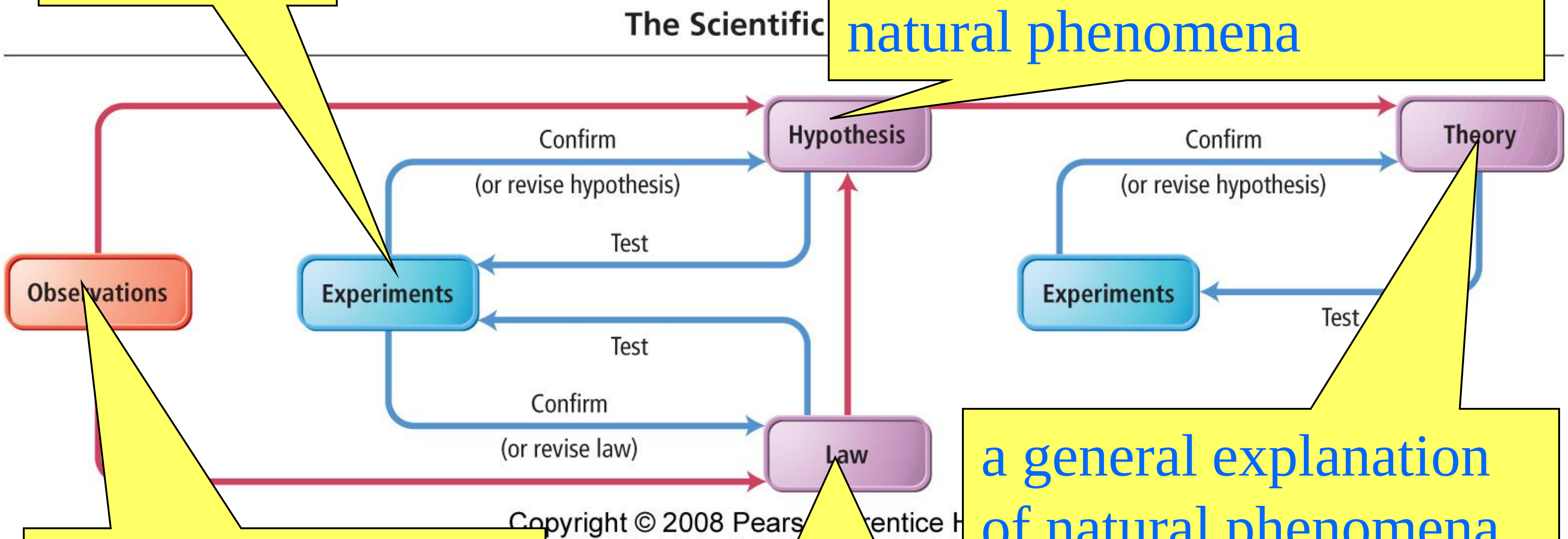
From Specific to General Understanding

- a hypothesis is a potential explanation for a single or small number of observations
- a **theory** is a general explanation for the manifestation and behavior of all nature
  - ✓ models
  - ✓ pinnacle of scientific knowledge
  - ✓ validated or invalidated by experiment and observation

Scientific Method

a test of a hypothesis or theory

a tentative explanation of a single or small number of natural phenomena



a general explanation of natural phenomena

the careful noting and recording of natural phenomena

a generally observed natural phenomenon

Classification of Matter

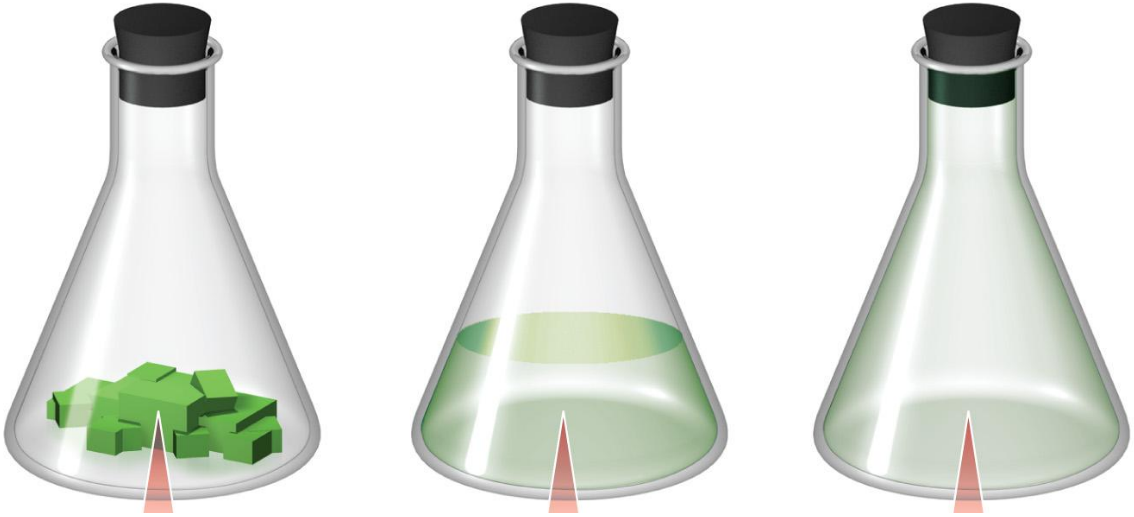
States of Matter

Physical and Chemical Properties

Physical and Chemical Changes

Classification of Matter

- **matter** is anything that has mass and occupies space
- we can classify matter based on whether it's **solid, liquid,** or **gas**



Classifying Matter by Physical State

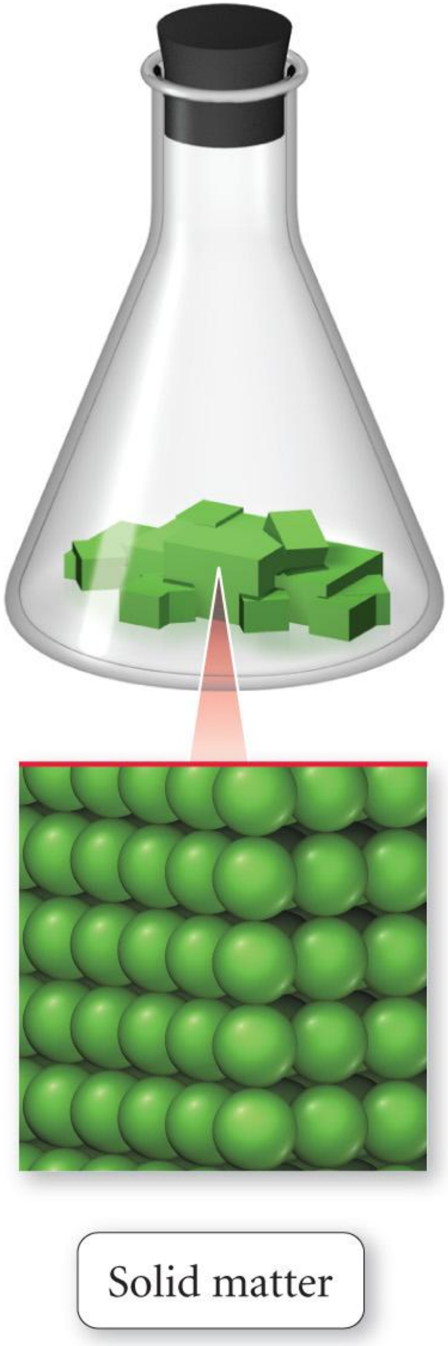
- matter can be classified as solid, liquid, or gas based on the characteristics it exhibits

State	***Shape***	***Volume***	***Compress***	***Flow***
Solid	Fixed	Fixed	No	No
Liquid	Indef.	Fixed	No	Yes
Gas	Indef.	Indef.	Yes	Yes

- Fixed = keeps shape when placed in a container
- Indefinite = takes the shape of the container

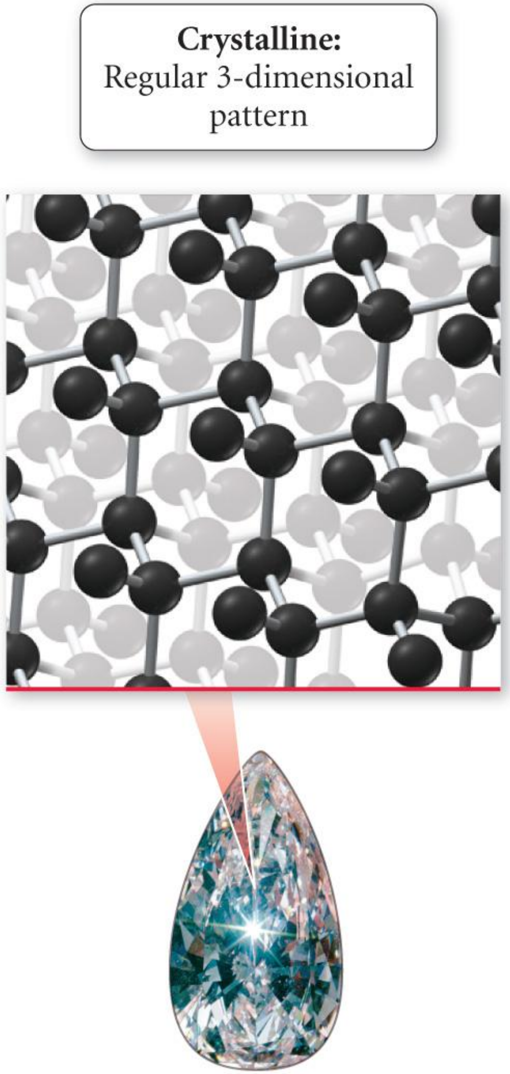
Solids

- the particles in a solid are packed close together and are fixed in position
  - ✓ though they may vibrate
- the close packing of the particles results in solids being incompressible
- the inability of the particles to move around results in solids retaining their shape and volume when placed in a new container, and prevents the particles from flowing



Crystalline Solids

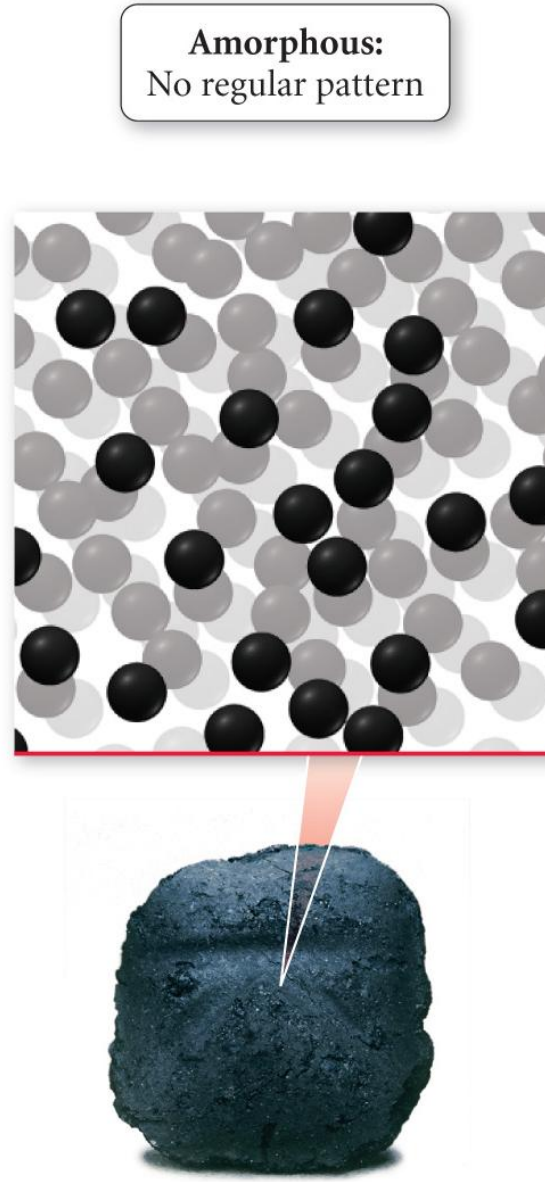
- some solids have their particles arranged in an orderly geometric pattern – we call these **crystalline solids**
  - ✓salt and diamonds



Diamond
C (*s*, diamond)

Amorphous Solids

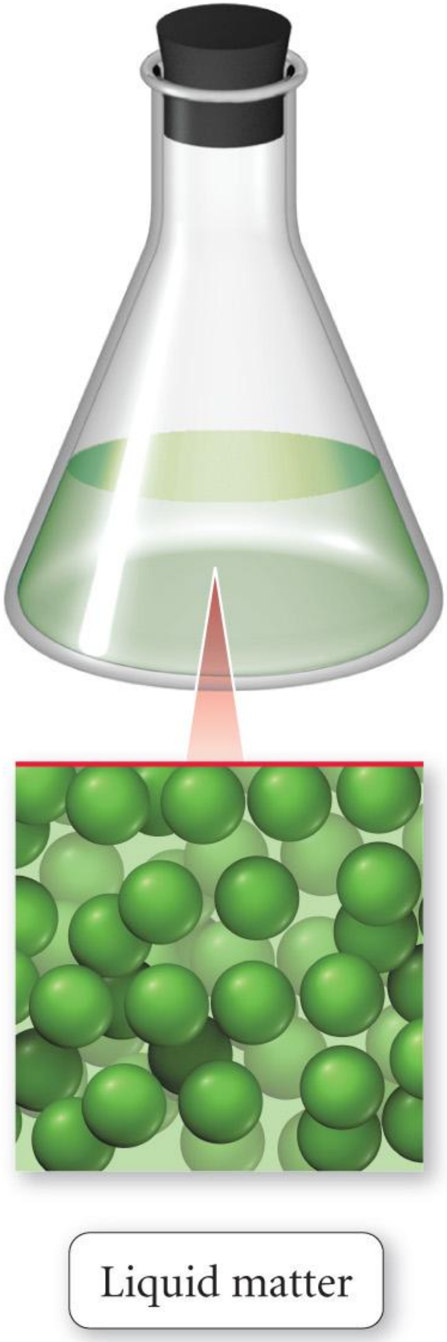
- some solids have their particles randomly distributed without any long-range pattern – we call these **amorphous solids**
  - ✓plastic
  - ✓glass
  - ✓charcoal



Charcoal
C (*s*, amorphous)

Liquids

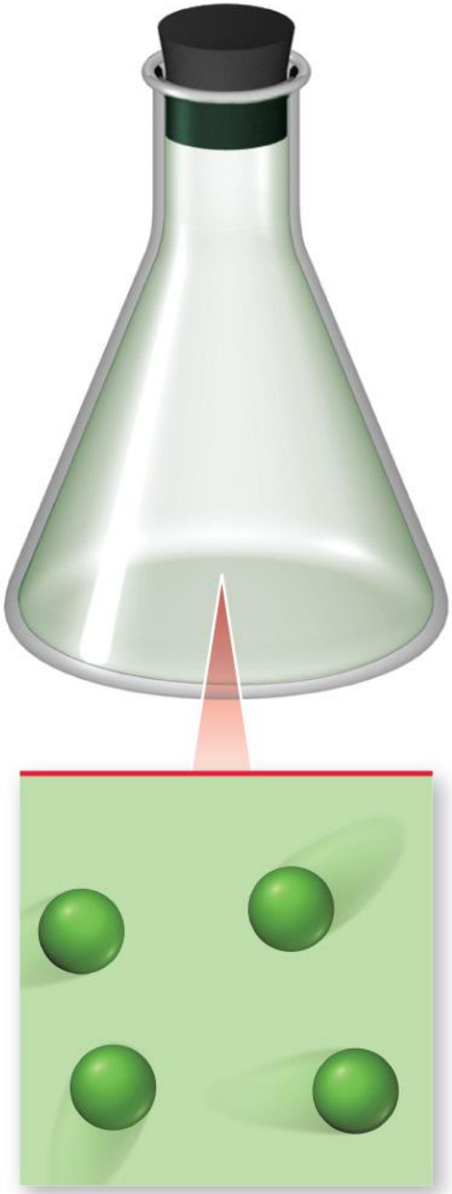
- the particles in a liquid are closely packed, but they have some ability to move around
- the close packing results in liquids being incompressible
- but the ability of the particles to move allows liquids to take the shape of their container and to flow – however, they don’t have enough freedom to escape and expand to fill the container



Tro, Chemistry: A Molecular Approach

Gases

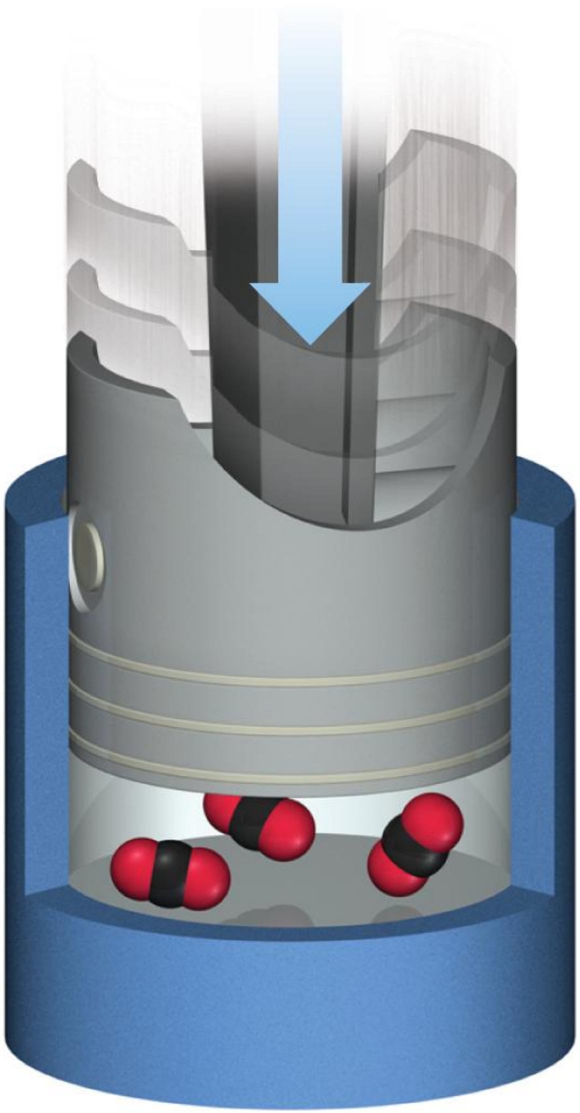
- in the gas state, the particles have complete freedom from each other
- the particles are constantly flying around, bumping into each other and the container
- in the gas state, there is a lot of empty space between the particles
  - ✓ on average



Gaseous matter

Gases

- because there is a lot of empty space, the particles can be squeezed closer together – therefore gases are compressible
- because the particles are not held in close contact and are moving freely, gases expand to fill and take the shape of their container, and will flow

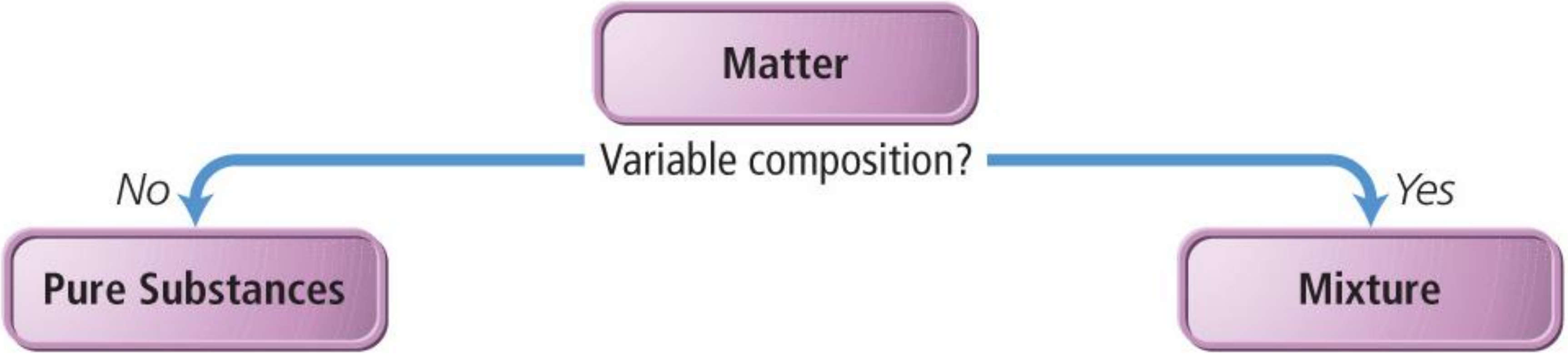


Gas–compressible

Classification of Matter by Composition

- matter whose composition does not change from one sample to another is called a **pure substance**
  - ✓ made of a single type of atom or molecule
  - ✓ because composition is always the same, all samples have the same characteristics
- matter whose composition may vary from one sample to another is called a **mixture**
  - ✓ two or more types of atoms or molecules combined in variable proportions
  - ✓ because composition varies, samples have the different characteristics

Classification of Matter by Composition



Pure Substances

1) made of one type of particle
2) all samples show the same intensive properties

Mixture

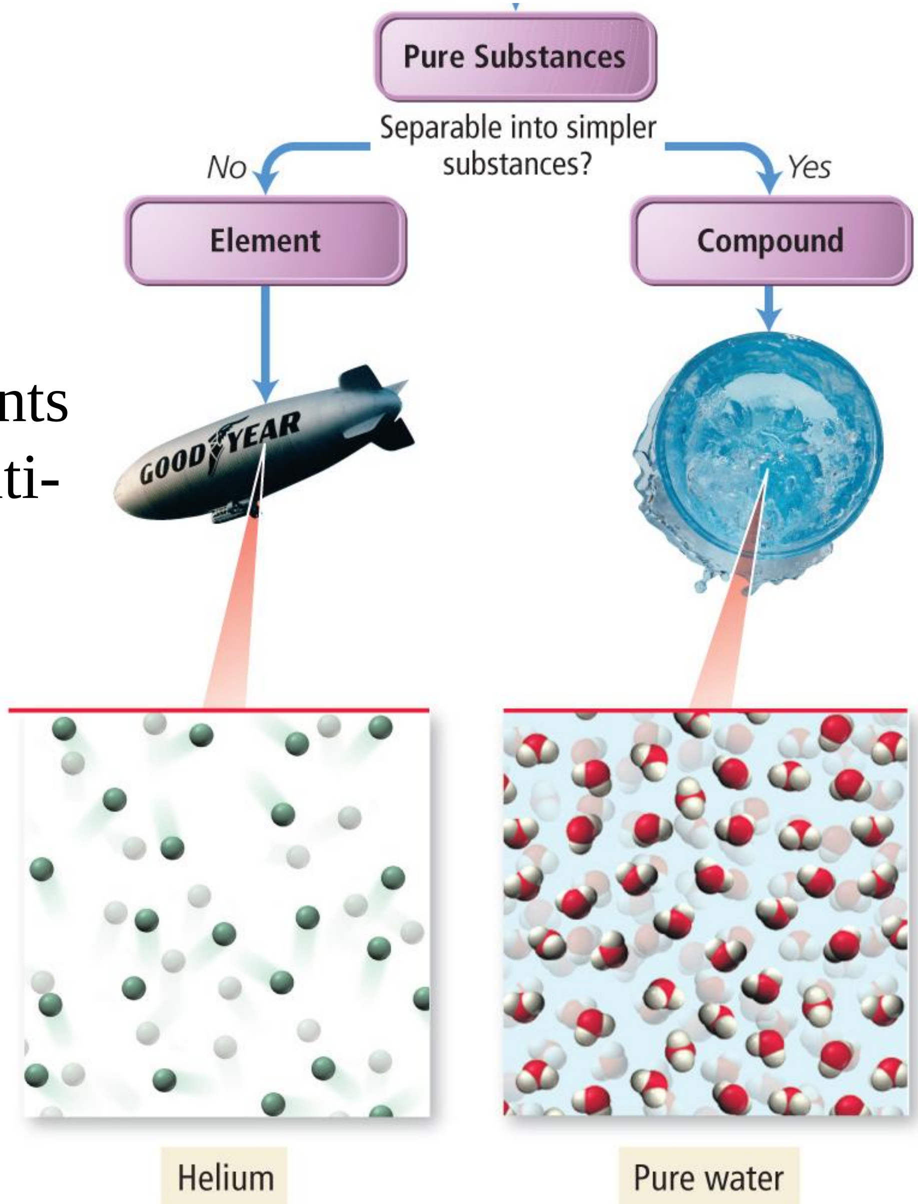
1) made of multiple types of particles
2) samples may show different intensive properties

Classification of Pure Substances

- substances that cannot be broken down into simpler substances by chemical reactions are called **elements**
  - ✓ basic building blocks of matter
  - ✓ composed of single type of atom
    - ➢ though those atoms may or may not be combined into molecules
- substances that can be decomposed are called **compounds**
  - ✓ chemical combinations of elements
  - ✓ composed of molecules that contain two or more different kinds of atoms
  - ✓ all molecules of a compound are identical, so all samples of a compound behave the same way
- most natural pure substances are compounds

Classification of Pure Substances

1) made of one type of atom (some elements found as multi-atom molecules in nature)
2) combine together to make compounds



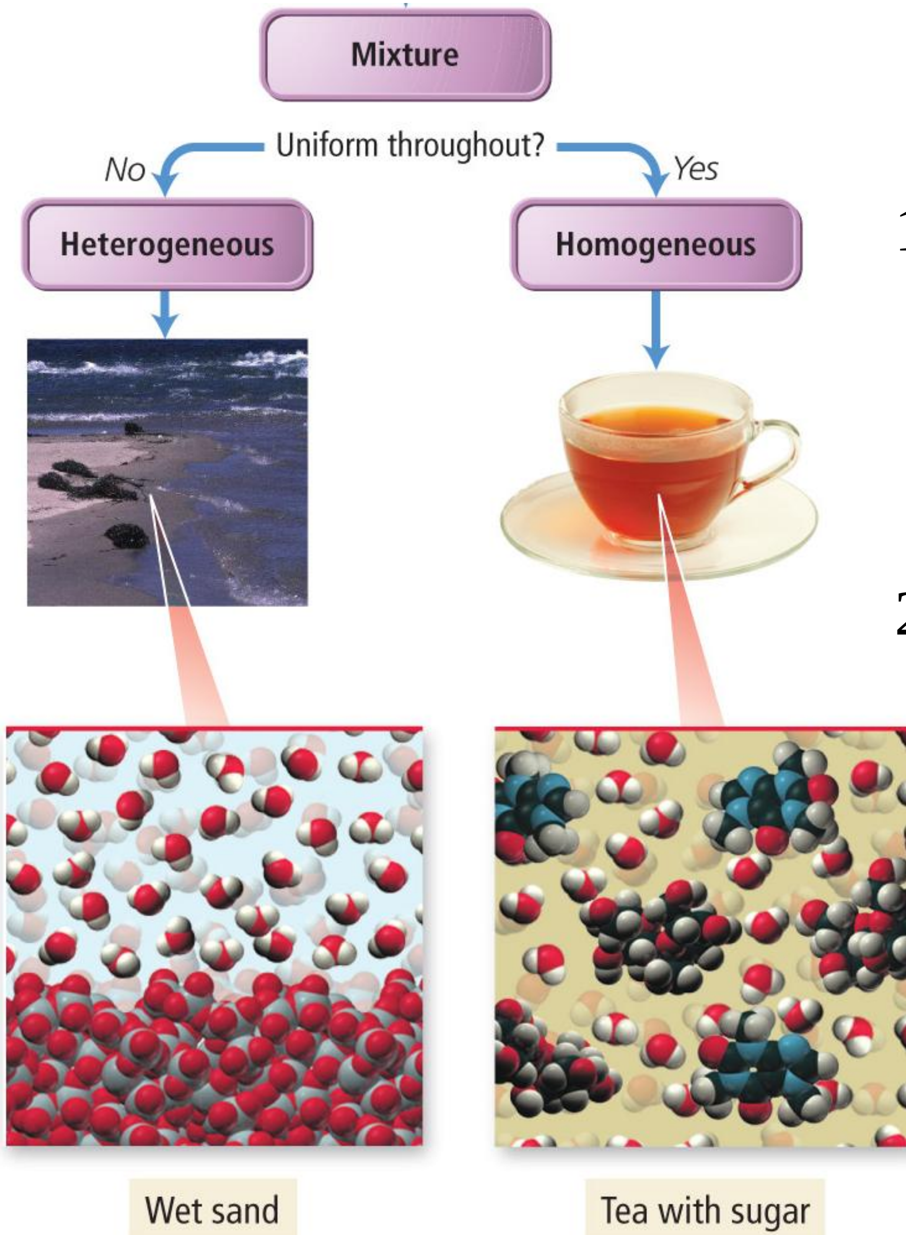
1) made of one type of molecule, or array of ions
2) molecules contain 2 or more different kinds of atoms

Classification of Mixtures

- **homogeneous** = mixture that has uniform composition throughout
  - ✓ every piece of a sample has identical characteristics, though another sample with the same components may have different characteristics
  - ✓ atoms or molecules mixed uniformly
- **heterogeneous** = mixture that does not have uniform composition throughout
  - ✓ contains regions within the sample with different characteristics
  - ✓ atoms or molecules not mixed uniformly

Classification of Mixtures

1) made of multiple substances, whose presence can be seen
2) portions of a sample have different composition and properties



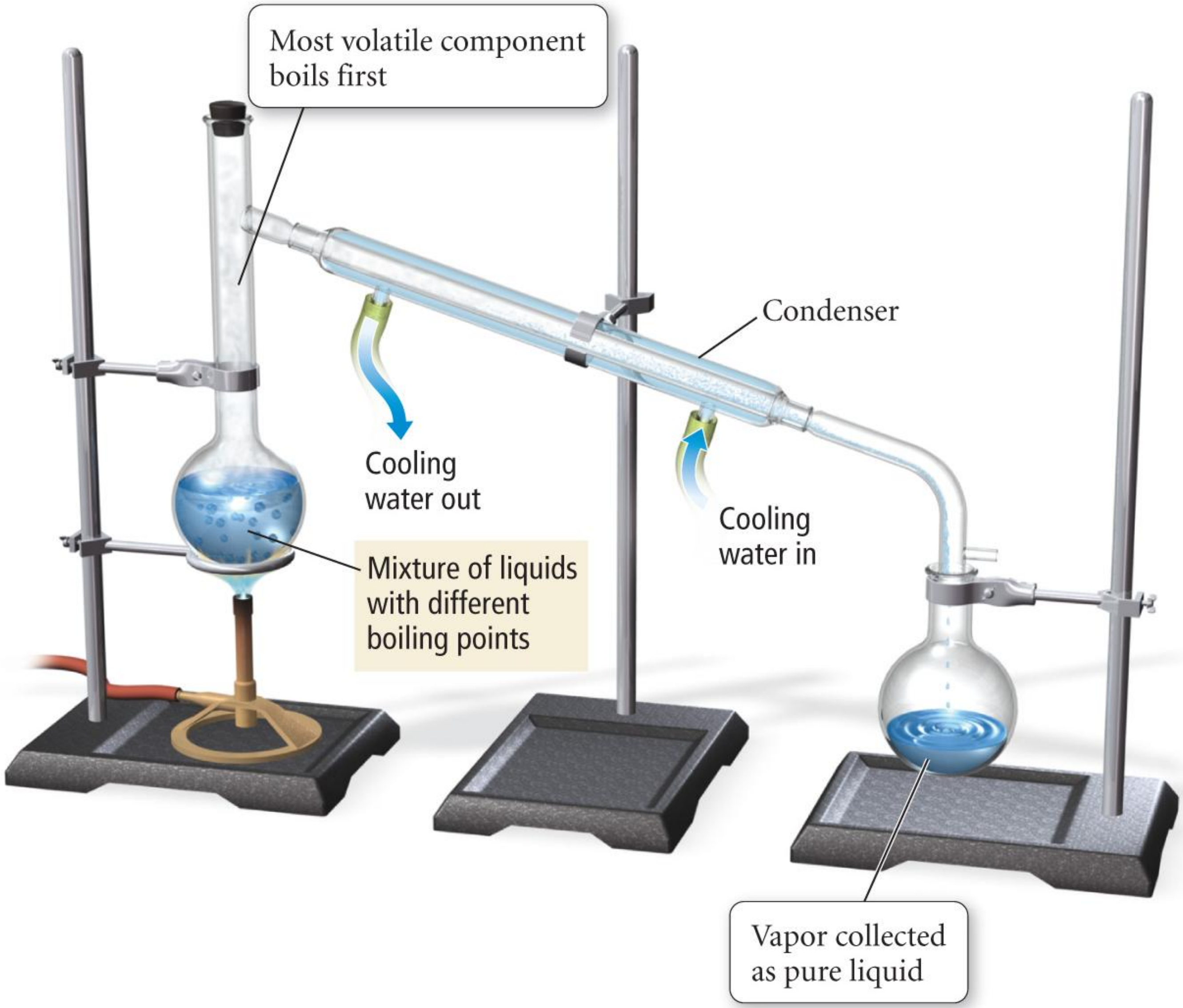
1) made of multiple substances, but appears to be one substance
2) all portions of a sample have the same composition and properties

Separation of Mixtures

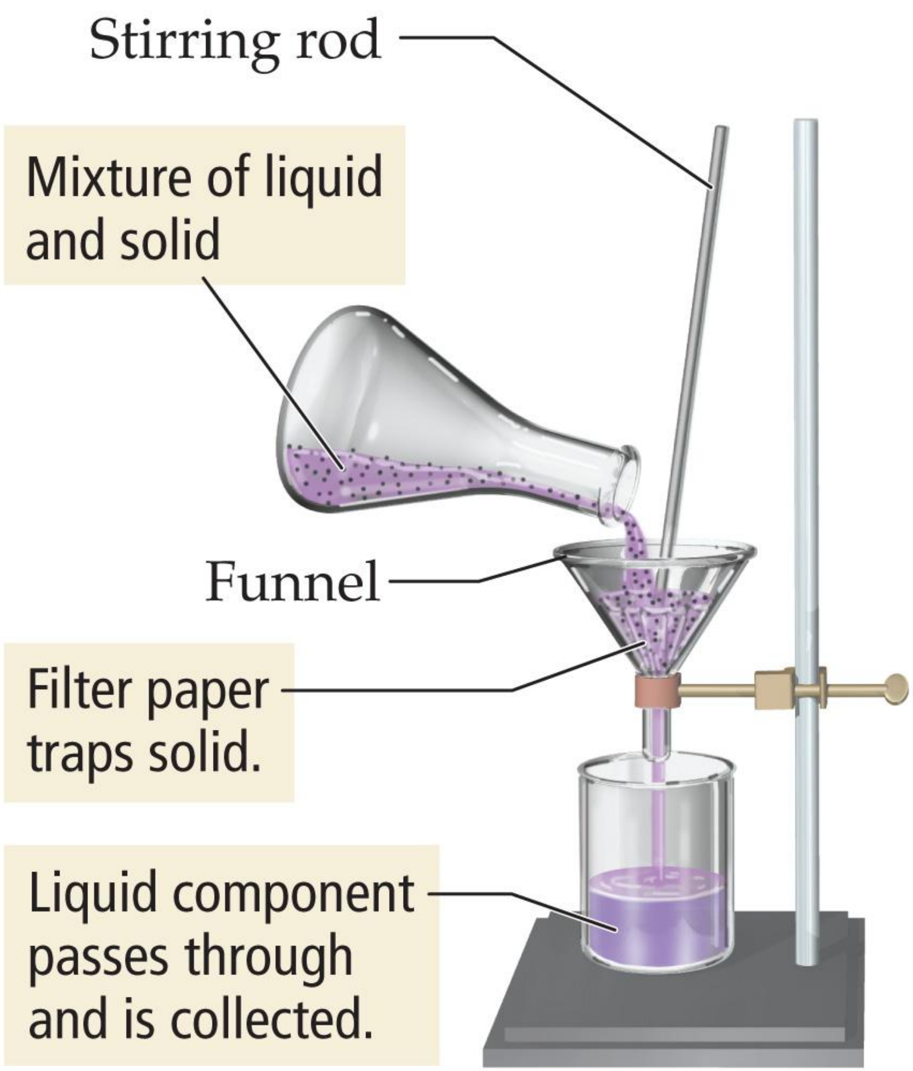
- separate mixtures based on different physical properties of the components
  - ✓Physical change

Different Physical Property	Technique
Boiling Point	Distillation
State of Matter (solid/liquid/gas)	Filtration
Adherence to a Surface	Chromatography
Volatility	Evaporation
Density	Centrifugation & Decanting

Distillation



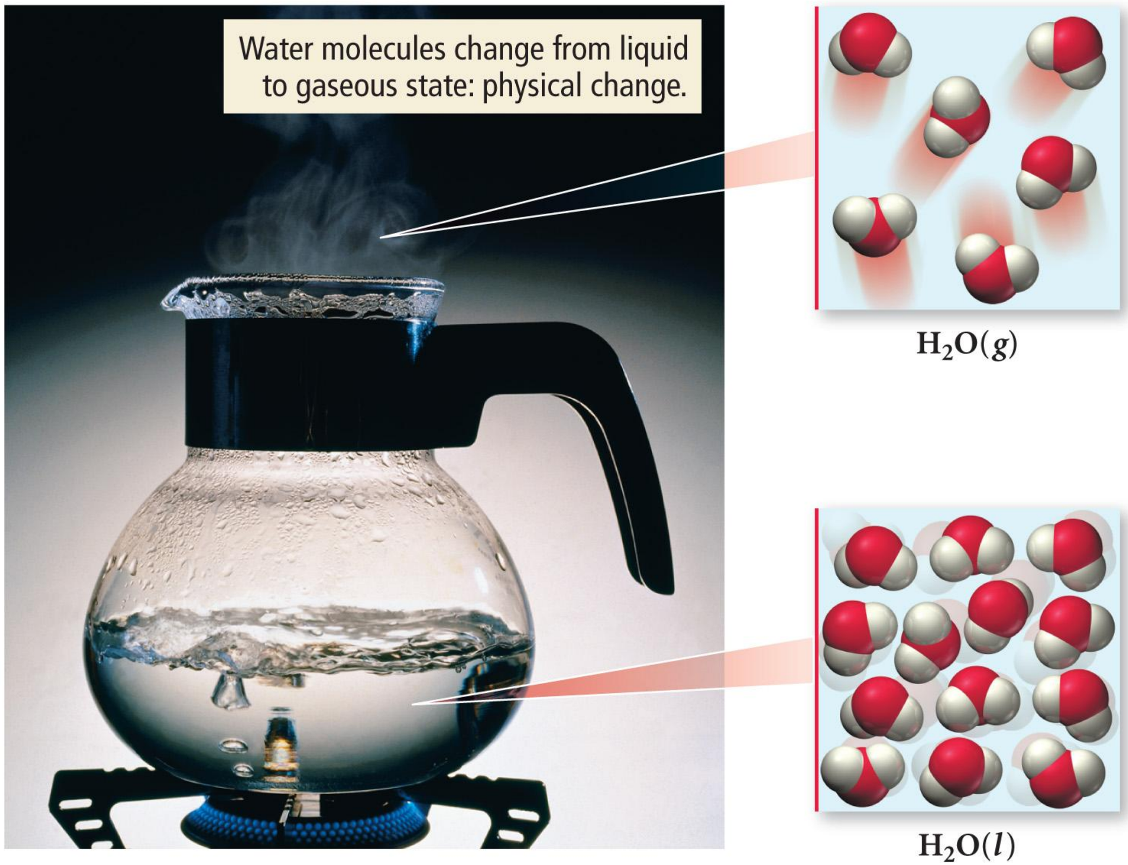
Filtration


Stirring rod
Mixture of liquid and solid
Funnel
Filter paper traps solid.
Liquid component passes through and is collected.

Changes in Matter

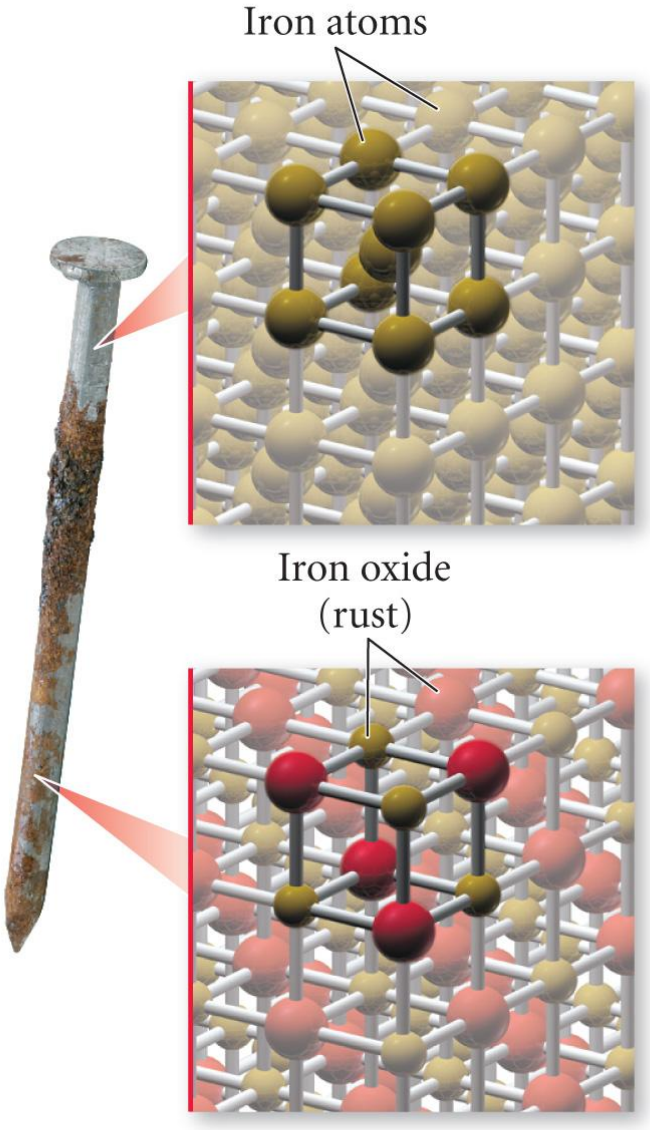
- changes that alter the state or appearance of the matter without altering the composition are called **physical changes**
- changes that alter the composition of the matter are called **chemical changes**
  - ✓ during the chemical change, the atoms that are present rearrange into new molecules, but all of the original atoms are still present

Physical Changes in Matter



The boiling of water is a physical change. The water molecules are separated from each other, but their structure and composition do not change.

Chemical Changes in Matter



The rusting of iron is a chemical change. The iron atoms in the nail combine with oxygen atoms from O_2 in the air to make a new substance, rust, with a different composition.

Properties of Matter

- **physical properties** are the characteristics of matter that can be changed without changing its composition
  - ✓characteristics that are directly observable
- **chemical properties** are the characteristics that determine how the composition of matter changes as a result of contact with other matter or the influence of energy
  - ✓characteristics that describe the behavior of matter

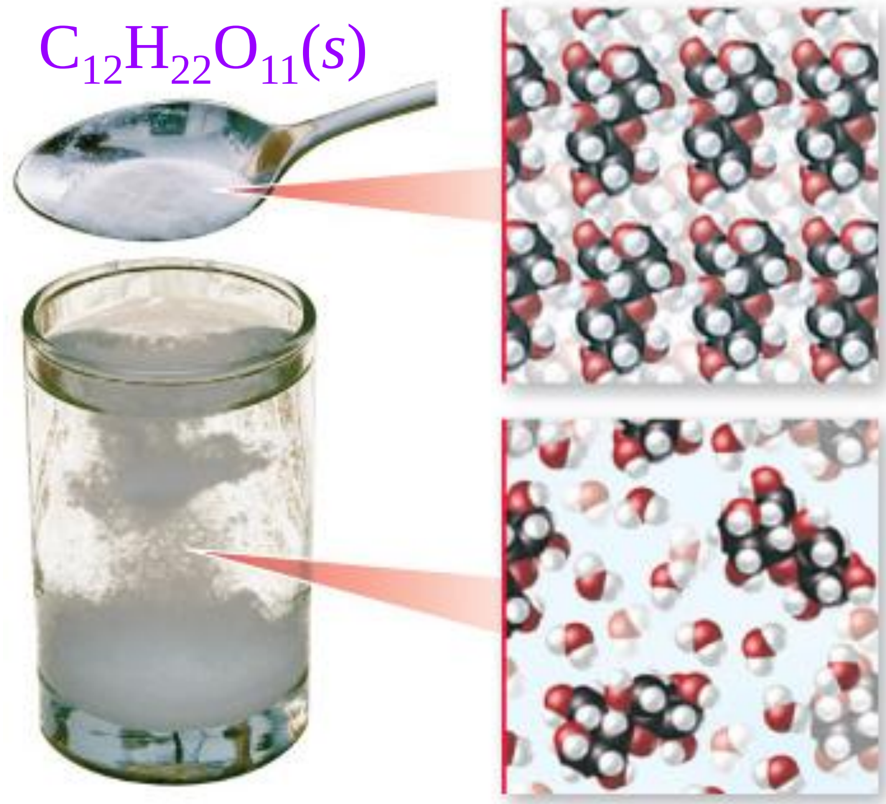
Common Physical Changes

- processes that cause changes in the matter that do not change its composition
- state changes
  - ✓boiling / condensing
  - ✓melting / freezing
  - ✓subliming
- dissolving

Dissolving of Sugar

Subliming of Dry Ice

$C_{12}H_{22}O_{11}(s)$



$C_{12}H_{22}O_{11}(aq)$

Common Chemical Changes

- processes that cause changes in the matter that change its composition
- rusting
- processes that release lots of energy
- burning

$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$$

Energy

Energy Changes in Matter

- changes in matter, both physical and chemical, result in the matter either gaining or releasing energy
- **energy** is the capacity to do work
- work is the action of a force applied across a distance
  - ✓ a force is a push or a pull on an object
  - ✓ electrostatic force is the push or pull on objects that have an electrical charge

Energy of Matter

- all matter possesses energy
- energy is classified as either kinetic or potential
- energy can be converted from one form to another
- when matter undergoes a chemical or physical change, the amount of energy in the matter changes as well

Energy of Matter - Kinetic

- **kinetic energy** is energy of motion
  - ✓ motion of the atoms, molecules, and subatomic particles
  - ✓ thermal (heat) energy is a form of kinetic energy because it is caused by molecular motion

Energy of Matter - Potential

- **potential energy** is energy that is stored in the matter
  - ✓ due to the composition of the matter and its position in the universe
  - ✓ chemical potential energy arises from electrostatic forces between atoms, molecules, and subatomic particles

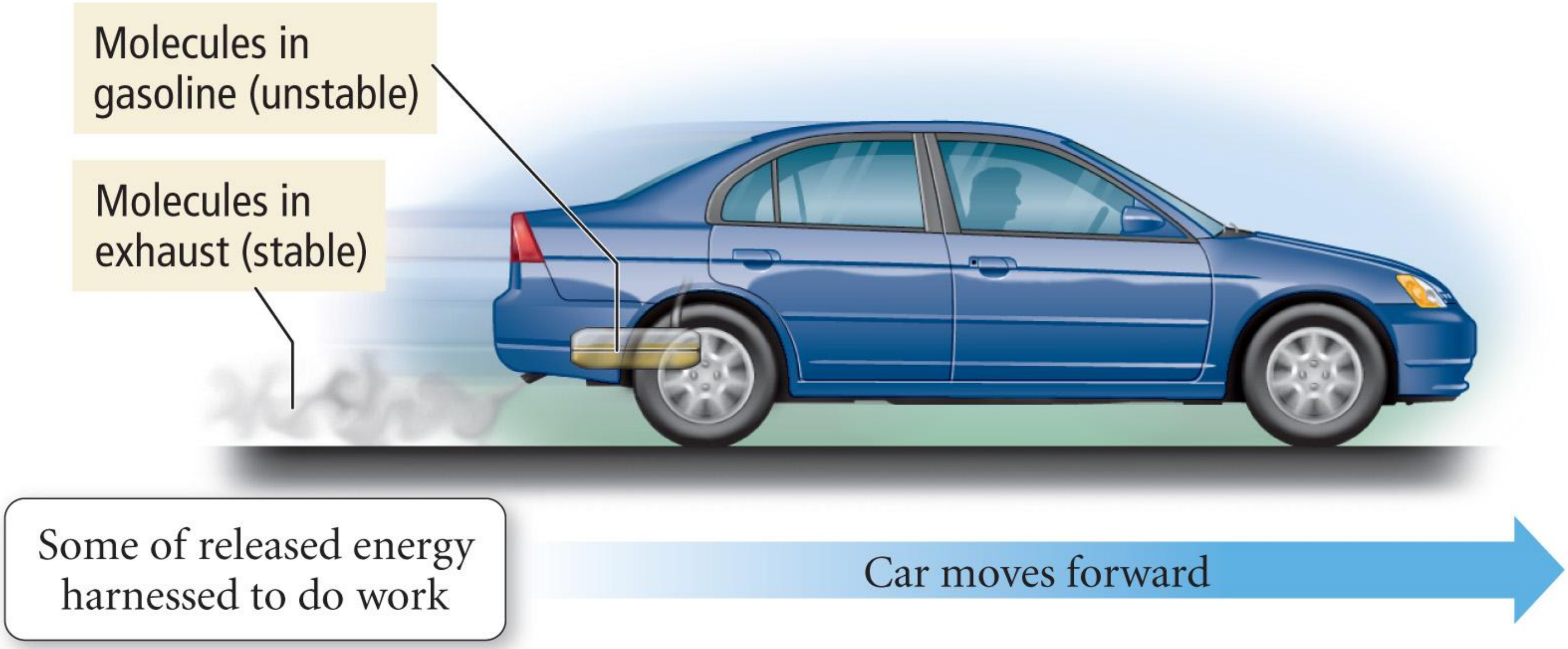
Conversion of Energy

- you can interconvert kinetic energy and potential energy
- whatever process you do that converts energy from one type or form to another, the total amount of energy remains the same
  - ✓**Law of Conservation of Energy**

Spontaneous Processes

- materials that possess high potential energy are less stable
- processes in nature tend to occur on their own when the result is material(s) with lower total potential energy
  - ✓ processes that result in materials with higher total potential energy can occur, but generally will not happen without input of energy from an outside source
- when a process results in materials with less potential energy at the end than there was at the beginning, the difference in energy is released into the environment

Potential to Kinetic Energy



Standard Units of Measure

The Standard Units

- Scientists have agreed on a set of international standard units for comparing all our measurements called the SI units
  - ✓ *Système International* = International System

Quantity	Unit	Symbol
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K

Length

- Measure of the two-dimensional distance an object covers
  - ✓ often need to measure lengths that are very long (distances between stars) or very short (distances between atoms)
- SI unit = meter
  - ✓ About 3.37 inches longer than a yard
    - ➢ 1 meter = one ten-millionth the distance from the North Pole to the Equator = distance between marks on standard metal rod = distance traveled by light in a specific period of time
- Commonly use centimeters (cm)
  - ✓ 1 m = 100 cm
  - ✓ 1 cm = 0.01 m = 10 mm
  - ✓ 1 inch = 2.54 cm (exactly)

Mass

- Measure of the amount of matter present in an object
  - ✓ weight measures the gravitational pull on an object, which depends on its mass
- SI unit = kilogram (kg)
  - ✓ about 2 lbs. 3 oz.
- Commonly measure mass in grams (g) or milligrams (mg)
  - ✓ 1 kg = 2.2046 pounds, 1 lbs. = 453.59 g
  - ✓ 1 kg = 1000 g = 10^3 g
  - ✓ 1 g = 1000 mg = 10^3 mg
  - ✓ 1 g = 0.001 kg = 10^{-3} kg
  - ✓ 1 mg = 0.001 g = 10^{-3} g

Time

- measure of the duration of an event
- SI units = second (s)
- 1 s is defined as the period of time it takes for a specific number of radiation events of a specific transition from cesium-133

Temperature

- measure of the average amount of kinetic energy
  - ✓ higher temperature = larger average kinetic energy
- heat flows from the matter that has high thermal energy into matter that has low thermal energy
  - ✓ until they reach the same temperature
  - ✓ heat is exchanged through molecular collisions between the two materials

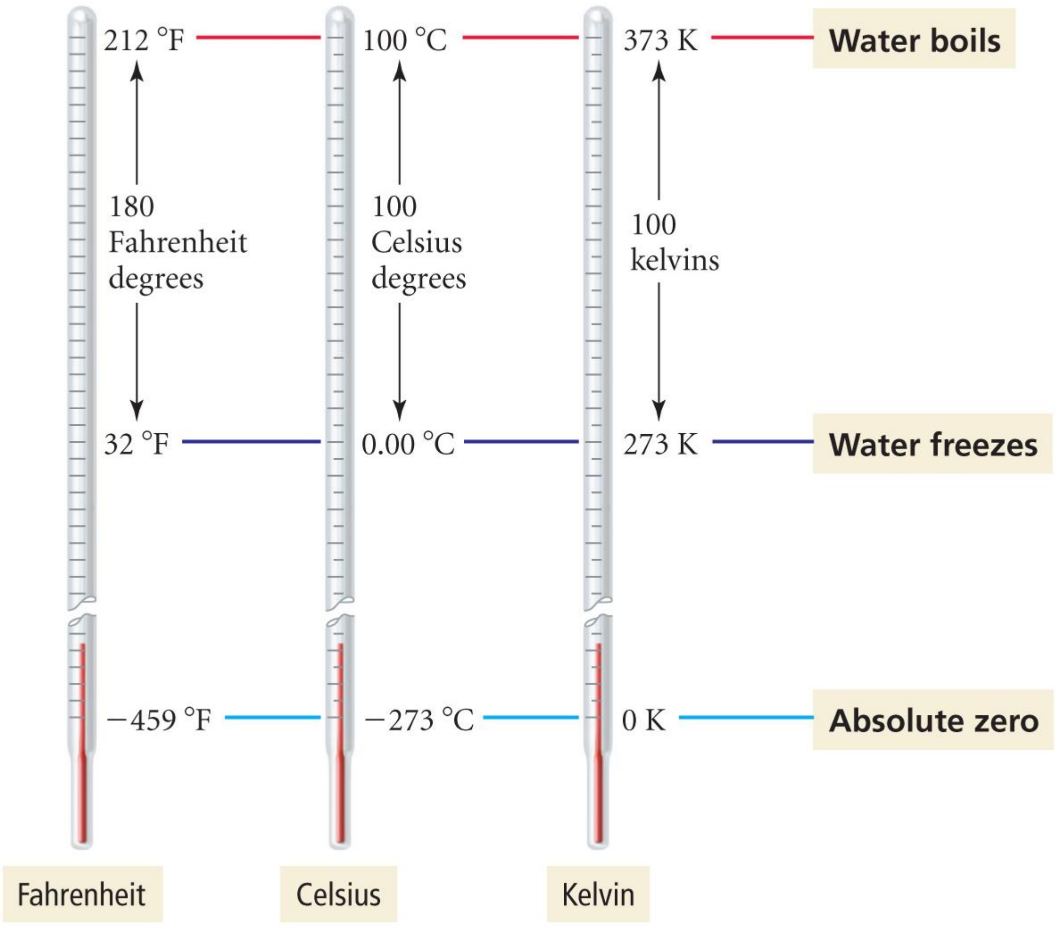
0 °C – Water freezes

22 °C – Room temperature

40 °C – Summer day in Death Valley

Temperature Scales

- Fahrenheit Scale, °F
  - ✓ used in the U.S.
- Celsius Scale, °C
  - ✓ used in all other countries
- Kelvin Scale, K
  - ✓ absolute scale
    - ➢ no negative numbers
  - ✓ directly proportional to average amount of kinetic energy
  - ✓ 0 K = absolute zero



Fahrenheit vs. Celsius

- a Celsius degree is 1.8 times larger than a Fahrenheit degree
- the standard used for 0° on the Fahrenheit scale is a lower temperature than the standard used for 0° on the Celsius scale

$$°C = \frac{(°F - 32)}{1.8}$$

Kelvin vs. Celsius

- the size of a “degree” on the Kelvin scale is the same as on the Celsius scale
  - ✓ though technically, we don’t call the divisions on the Kelvin scale degrees; we called them kelvins!
  - ✓ so 1 kelvin is 1.8 times larger than 1°F
- the 0 standard on the Kelvin scale is a much lower temperature than on the Celsius scale

$$K = °C + 273.15$$

Example 1.2 Convert 40.00 °C into K and °F

• Find the equation that relates the given quantity to the quantity you want to find	**Given:** **Find:** **Equation:**	40.00 °C K K = °C + 273.15
• Since the equation is solved for the quantity you want to find, substitute and compute		K = °C + 273.15 K = 40.00 + 273.15 K = 313.15 K
• Find the equation that relates the given quantity to the quantity you want to find	**Given:** **Find:** **Equation:**	40.00 °C °F $°C = \frac{(°F - 32)}{1.8}$
• Solve the equation for the quantity you want to find		$1.8 \times °C = (°F - 32)$ $1.8 \times °C + 32 = °F$
• Substitute and compute		$1.8 \times 40.00 + 32 = °F$ $104.00\ °F = °F$

Related Units in the SI System

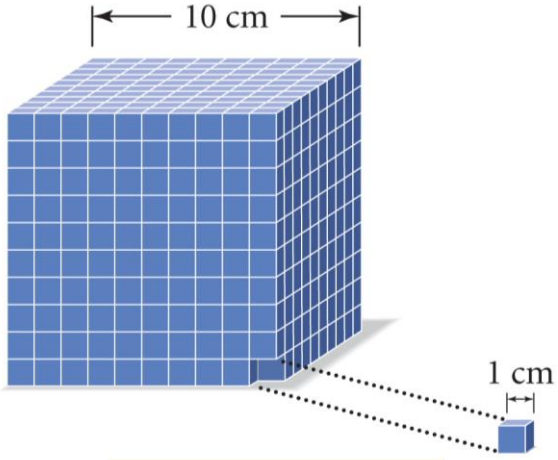
- All units in the SI system are related to the standard unit by a power of 10
- The power of 10 is indicated by a prefix multiplier
- The prefix multipliers are always the same, regardless of the standard unit
- Report measurements with a unit that is close to the size of the quantity being measured

Common Prefix Multipliers in the SI System

Prefix	Symbol	Decimal Equivalent	Power of 10
mega-	M	1,000,000	Base x 10^{6}
kilo-	k	1,000	Base x 10^{3}
deci-	d	0.1	Base x 10^{-1}
centi-	c	0.01	Base x 10^{-2}
milli-	m	0.001	Base x 10^{-3}
micro-	μ or mc	0.000 001	Base x 10^{-6}
nano-	n	0.000 000 001	Base x 10^{-9}
pico	p	0.000 000 000 001	Base x 10^{-12}

Volume

- Derived unit
  - ✓ any length unit cubed
- Measure of the amount of space occupied
- SI unit = cubic meter (m^3)
- Commonly measure solid volume in cubic centimeters (cm^3)
  - ✓ $1\ m^3 = 10^6\ cm^3$
  - ✓ $1\ cm^3 = 10^{-6}\ m^3 = 0.000001\ m^3$
- Commonly measure liquid or gas volume in milliliters (mL)
  - ✓ 1 L is slightly larger than 1 quart
  - ✓ $1\ L = 1\ dm^3 = 1000\ mL = 10^3\ mL$
  - ✓ $1\ mL = 0.001\ L = 10^{-3}\ L$
  - ✓ $1\ mL = 1\ cm^3$



A 10-cm cube contains 1000 1-cm cubes.

Common Units and Their Equivalents

Length

1 kilometer (km)	=	0.6214 mile (mi)
1 meter (m)	=	39.37 inches (in.)
1 meter (m)	=	1.094 yards (yd)
1 foot (ft)	=	30.48 centimeters (cm)
1 inch (in.)	=	2.54 centimeters (cm) exactly

Common Units and Their Equivalents

Mass

1 kilogram (km)	=	2.205 pounds (lb)
1 pound (lb)	=	453.59 grams (g)
1 ounce (oz)	=	28.35 grams (g)

Volume

1 liter (L)	=	1000 milliliters (mL)
1 liter (L)	=	1000 cubic centimeters (cm^3)
1 liter (L)	=	1.057 quarts (qt)
1 U.S. gallon (gal)	=	3.785 liters (L)

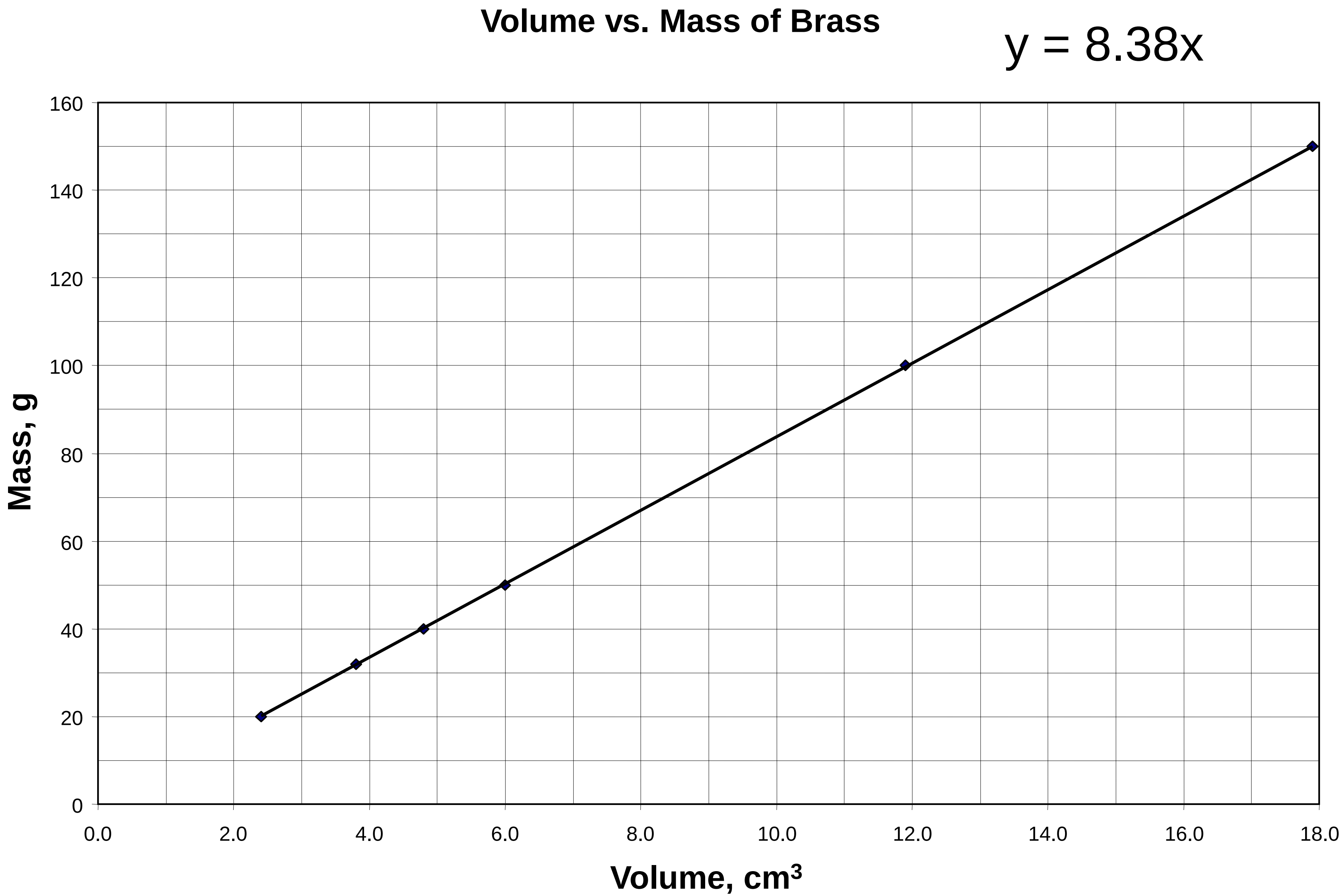
Density

Mass & Volume

- two main physical properties of matter
- mass and volume are **extensive properties**
  - ✓the value depends on the quantity of matter
  - ✓extensive properties cannot be used to identify what *type* of matter something is
    - ➢if you are given a large glass containing 100 g of a clear, colorless liquid and a small glass containing 25 g of a clear, colorless liquid - are both liquids the same stuff?
- even though mass and volume are individual properties, for a given type of matter they are related to each other!

Mass vs. Volume of Brass

Mass grams	**Volume cm^3**
20	2.4
32	3.8
40	4.8
50	6.0
100	11.9
150	17.9


Volume vs. Mass of Brass
y = 8.38x
Mass, g
160
140
120
100
80
60
40
20
0
0.0
2.0
4.0
6.0
8.0
10.0
12.0
14.0
16.0
18.0
Volume, cm3

Density

- Ratio of mass:volume is an **intensive property**
  - ✓value independent of the quantity of matter
- Solids = g/cm^3
  - ✓$1\ cm^3 = 1\ mL$
- Liquids = g/mL
- Gases = g/L
- Volume of a solid can be determined by water displacement – Archimedes Principle
- Density : solids > liquids >>> gases
  - ✓except ice is less dense than liquid water!

$$Density = \frac{Mass}{Volume}$$

$$Density = \frac{Mass}{Volume}$$

Density

- For equal volumes, denser object has larger mass
- For equal masses, denser object has smaller volume
- Heating an object generally causes it to expand, therefore the density changes with temperature

TABLE 1.4 The Density of Some Common Substances at 20 °C

Substance	Density (g/cm^3)
Charcoal (from oak)	0.57
Ethanol	0.789
Ice	0.917 (at 0 °C)
Water	1.00 (at 4 °C)
Sugar (sucrose)	1.58
Table salt (sodium chloride)	2.16
Glass	2.6
Aluminum	2.70
Titanium	4.51
Iron	7.86
Copper	8.96
Lead	11.4
Mercury	13.55
Gold	19.3
Platinum	21.4

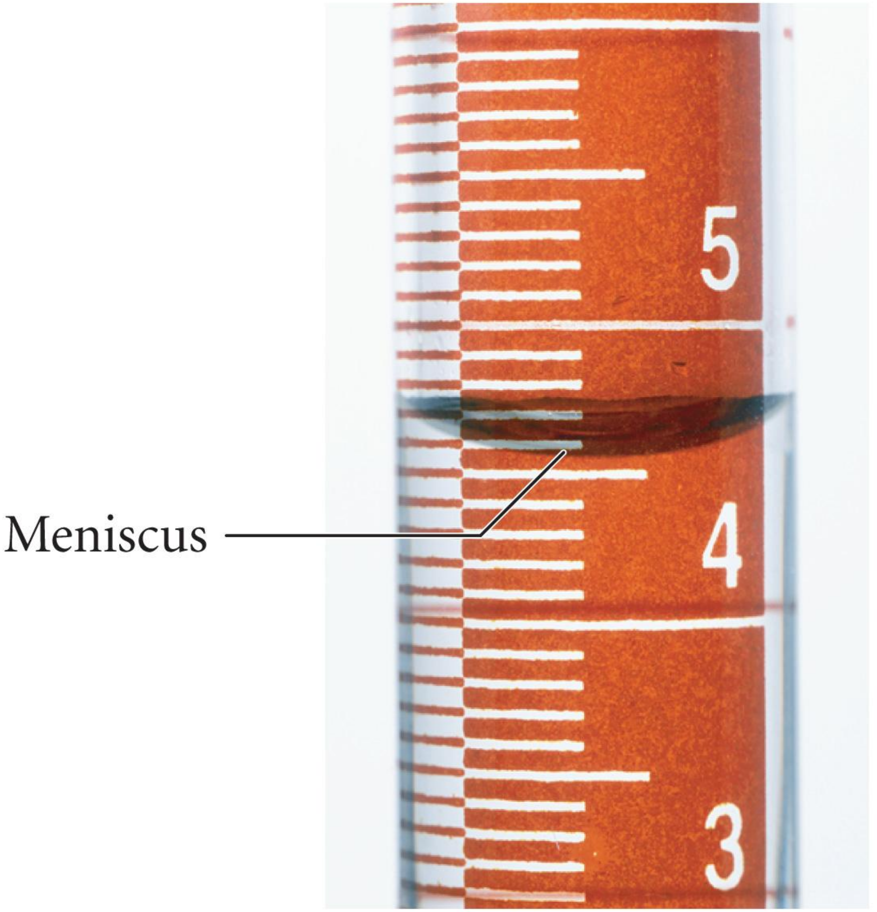
Example 1.3 Decide if a ring with a mass of 3.15 g that displaces 0.233 cm^3 of water is platinum

• Find the equation that relates the given quantity to the quantity you want to find	**Given:** **Find:** **Equation:**	mass = 3.15 g volume = 0.233 cm^3 density, g/cm^3 $Density = \frac{Mass}{Volume}$
• Since the equation is solved for the quantity you want to find, and the units are correct, substitute and compute		$d = \frac{m}{V} = \frac{3.15\ \text{g}}{0.233\ \text{cm}^3}$ $d = 13.5\ \text{g/cm}^3$
• Compare to accepted value of the intensive property		Density of platinum = 21.4 g/cm^3 therefore not platinum

Measurement
and Significant Figures

What Is a Measurement?

- quantitative observation
- comparison to an agreed-upon standard
- every measurement has a number and a unit



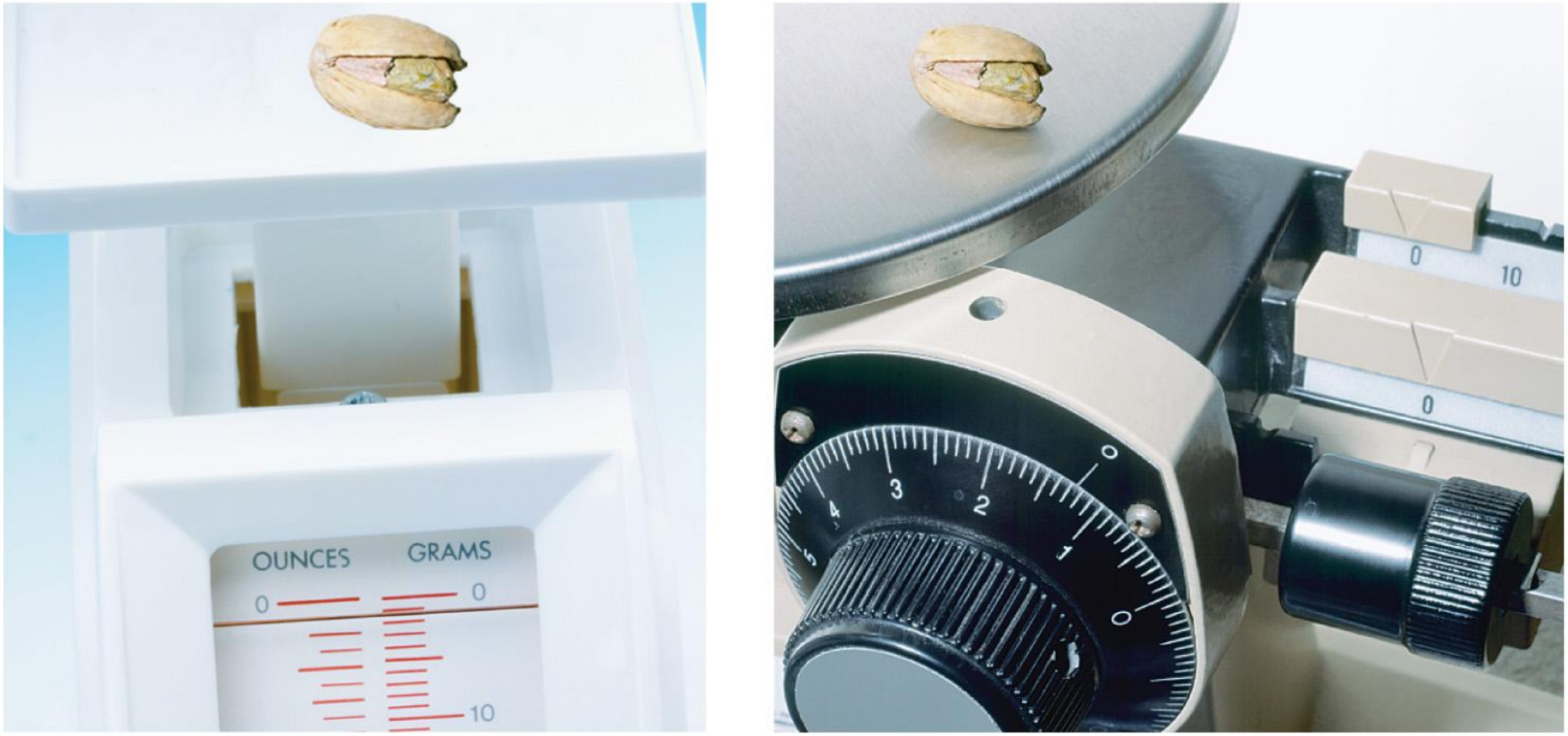
A Measurement

- the unit tells you what standard you are comparing your object to
- the number tells you
  1. what multiple of the standard the object measures
  2. the uncertainty in the measurement
- scientific measurements are reported so that every digit written is certain, except the last one which is estimated

Estimating the Last Digit

- for instruments marked with a scale, you get the last digit by estimating between the marks
  - ✓ if possible
- mentally divide the space into 10 equal spaces, then estimate how many spaces over the indicator mark is

Estimation in Weighing



(a)

(b)

Markings every 1 g
Estimated reading 1.2 g

Markings every 0.1 g
Estimated reading 1.27 g

Significant Figures

- the non-place-holding digits in a reported measurement are called **significant figures**
  - ✓ some zeros in a written number are only there to help you locate the decimal point
- significant figures tell us the range of values to expect for repeated measurements
  - ✓ the more significant figures there are in a measurement, the smaller the range of values is

12.3 cm
has 3 sig. figs.
and its range is
12.2 to 12.4 cm

12.30 cm
has 4 sig. figs.
and its range is
12.29 to 12.31 cm

Counting Significant Figures

1) All non-zero digits are significant
   - ✓ 1.5 has 2 sig. figs.
2) Interior zeros are significant
   - ✓ 1.05 has 3 sig. figs.
3) Leading zeros are **NOT** significant
   - ✓ 0.001050 has 4 sig. figs.
     - ➢ 1.050×10^{-3}

Counting Significant Figures

4) Trailing zeros may or may not be significant
    1) Trailing zeros after a decimal point are significant
        - 1.050 has 4 sig. figs.
    2) Zeros at the end of a number without a written decimal point are ambiguous and should be avoided by using scientific notation
        - if 150 has 2 sig. figs. then 1.5×10^2
        - but if 150 has 3 sig. figs. then 1.50×10^2

Significant Figures and Exact Numbers

- Exact numbers have an unlimited number of significant figures
- A number whose value is known with complete certainty is **exact**
  - ✓from counting individual objects
  - ✓from definitions
    - ➢1 cm is exactly equal to 0.01 m
  - ✓from integer values in equations
    - ➢ in the equation for the radius of a circle, the 2 is exact

$$\text{radius of a circle} = \frac{\text{diameter of a circle}}{2}$$

Example 1.5 Determining the Number of Significant Figures in a Number

How many significant figures are in each of the following?

0.04450 m	4 sig. figs.; the digits 4 and 5, and the trailing 0
5.0003 km	5 sig. figs.; the digits 5 and 3, and the interior 0’s
10 dm = 1 m	infinite number of sig. figs., exact numbers
1.000×10^5 s	4 sig. figs.; the digit 1, and the trailing 0’s
0.00002 mm	1 sig. figs.; the digit 2, not the leading 0’s
10,000 m	Ambiguous, generally assume 1 sig. fig.

Multiplication and Division with Significant Figures

- when multiplying or dividing measurements with significant figures, the result has the same number of significant figures as the measurement with the fewest number of significant figures

5.02	×	89,665	×	0.10	= 45.0118 = 45
3 sig. figs.		5 sig. figs.		2 sig. figs.	2 sig. figs.

5.892	÷	6.10	= 0.96590	= 0.966
4 sig. figs.		3 sig. figs.		3 sig. figs.

Addition and Subtraction with Significant Figures

- when adding or subtracting measurements with significant figures, the result has the same number of decimal places as the measurement with the fewest number of decimal places

5.74	+	0.823	+	2.651 = 9.214 = 9.21
2 dec. pl.		3 dec. pl.		3 dec. pl. 2 dec. pl.

4.8	-	3.965	=	0.835	=	0.8
1 dec. pl		3 dec. pl.				1 dec. pl.

Rounding

- when rounding to the correct number of significant figures, if the number after the place of the last significant figure is

1. 0 to 4, round down
   - ✓ drop all digits after the last sig. fig. and leave the last sig. fig. alone
   - ✓ add insignificant zeros to keep the value if necessary
2. 5 to 9, round up
   - ✓ drop all digits after the last sig. fig. and increase the last sig. fig. by one
   - ✓ add insignificant zeros to keep the value if necessary

- to avoid accumulating extra error from rounding, round only at the end, keeping track of the last sig. fig. for intermediate calculations

Rounding

- rounding to 2 significant figures
- 2.34 rounds to 2.3
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 4 or less
- 2.37 rounds to 2.4
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 5 or greater
- 2.349865 rounds to 2.3
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 4 or less

Rounding

- rounding to 2 significant figures
- 0.0234 rounds to 0.023 or 2.3×10^{-2}
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 4 or less
- 0.0237 rounds to 0.024 or 2.4×10^{-2}
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 5 or greater
- 0.02349865 rounds to 0.023 or 2.3×10^{-2}
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 4 or less

Rounding

- rounding to 2 significant figures
- 234 rounds to 230 or 2.3×10^2
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 4 or less
- 237 rounds to 240 or 2.4×10^2
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 5 or greater
- 234.9865 rounds to 230 or 2.3×10^2
  - ✓because the 3 is where the last sig. fig. will be and the number after it is 4 or less

Both Multiplication/Division and Addition/Subtraction with Significant Figures

- when doing different kinds of operations with measurements with significant figures, do whatever is in parentheses first, evaluate the significant figures in the intermediate answer, then do the remaining steps

$$3.489 \times (5.67 - 2.3) =$$

2 dp   1 dp

$$3.489 \times 3.\underline{3}7 = 12$$

4 sf   1 dp & 2 sf   2 sf

Example 1.6 Perform the following calculations to the correct number of significant figures

a) $1.10 \times 0.5120 \times 4.0015 \div 3.4555$

b)
$$\begin{array}{r} 0.355 \\ +105.1 \\ \underline{-100.5820} \end{array}$$

c) $4.562 \times 3.99870 \div (452.6755 - 452.33)$

d) $(14.84 \times 0.55) - 8.02$

Example 1.6 Perform the following calculations to the correct number of significant figures

a) $1.10 \times 0.5120 \times 4.0015 \div 3.4555 = 0.65219 = 0.652$

b)
$$\begin{array}{l} \quad 0.355 \\ +105.1 \\ \underline{-100.5820} \\ 4.8730 = 4.9 \end{array}$$

c) $4.562 \times 3.99870 \div (452.6755 - 452.33) = 52.79904 = 53$

d) $(14.84 \times 0.55) - 8.02 = 0.142 = 0.1$

Precision and Accuracy

Uncertainty in Measured Numbers

- uncertainty comes from limitations of the instruments used for comparison, the experimental design, the experimenter, and nature's random behavior
- to understand how reliable a measurement is we need to understand the limitations of the measurement
- **accuracy** is an indication of how close a measurement comes to the **actual** value of the quantity
- **precision** is an indication of how reproducible a measurement is

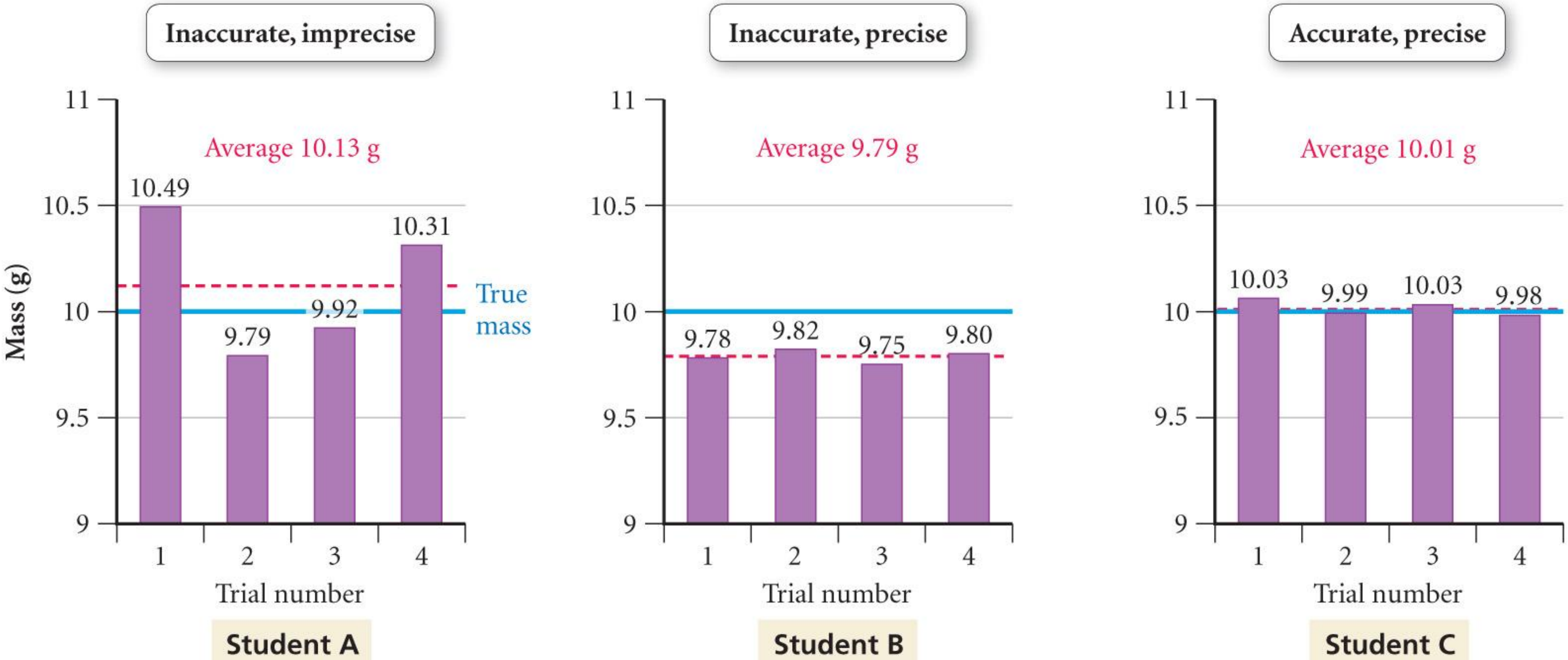
Precision

- imprecision in measurements is caused by **random errors**
  - ✓errors that result from random fluctuations
  - ✓no specific cause, therefore cannot be corrected
- we determine the precision of a set of measurements by evaluating how far they are from the actual value and each other
- even though every measurement has some random error, with enough measurements these errors should average out

Accuracy

- inaccuracy in measurement caused by **systematic errors**
  - ✓errors caused by limitations in the instruments or techniques or experimental design
  - ✓can be reduced by using more accurate instruments, or better technique or experimental design
- we determine the accuracy of a measurement by evaluating how far it is from the actual value
- systematic errors do not average out with repeated measurements because they consistently cause the measurement to be either too high or too low

Accuracy vs. Precision



Solving Chemical Problems

Equations &

Dimensional Analysis

Units

- Always write every number with its associated unit
- Always include units in your calculations
  - ✓ you can do the same kind of operations on units as you can with numbers
    - $\text{cm} \times \text{cm} = \text{cm}^2$
    - $\text{cm} + \text{cm} = \text{cm}$
    - $\text{cm} \div \text{cm} = 1$
  - ✓ using units as a guide to problem solving is called ***dimensional analysis***

Problem Solving and Dimensional Analysis

- Many problems in chemistry involve using relationships to convert one unit of measurement to another
- Conversion factors are relationships between two units
  - ✓ May be exact or measured
- Conversion factors generated from equivalence statements
  - ✓ e.g., 1 inch = 2.54 cm can give $\frac{2.54\text{cm}}{1\text{in}}$ or $\frac{1\text{in}}{2.54\text{cm}}$

Problem Solving and Dimensional Analysis

- Arrange conversion factors so given unit cancels
  - ✓Arrange conversion factor so given unit is on the bottom of the conversion factor
- May string conversion factors
  - ✓So we do not need to know every relationship, as long as we can find something else the given and desired units are related to

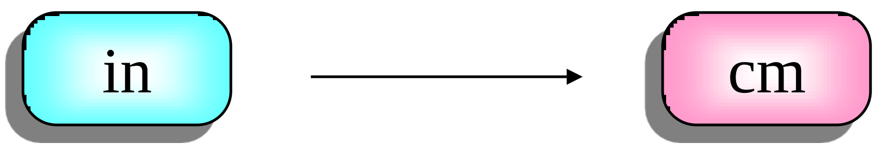
$$\cancel{\text{given}}\ \text{unit} \times \frac{\text{desired unit}}{\cancel{\text{given}}\ \text{unit}} = \text{desired unit}$$

Conceptual Plan

- a conceptual plan is a visual outline that shows the strategic route required to solve a problem
- for unit conversion, the conceptual plan focuses on units and how to convert one to another
- for problems that require equations, the conceptual plan focuses on solving the equation to find an unknown value

Concept Plans and Conversion Factors

- Convert inches into centimeters
  1) Find relationship equivalence: 1 in = 2.54 cm
  2) Write concept plan



  3) Change equivalence into conversion factors with starting units on the bottom

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

Systematic Approach

- Sort the information from the problem
  - ✓ identify the given quantity and unit, the quantity and unit you want to find, any relationships implied in the problem
- Design a strategy to solve the problem
  - ✓ Concept plan
    - ➢ sometimes may want to work backwards
    - ➢ each step involves a conversion factor or equation
- Apply the steps in the concept plan
  - ✓ check that units cancel properly
  - ✓ multiply terms across the top and divide by each bottom term
- Check the answer
  - ✓ double check the set-up to ensure the unit at the end is the one you wished to find
  - ✓ check to see that the size of the number is reasonable
    - ➢ since centimeters are smaller than inches, converting inches to centimeters should result in a larger number

Example 1.7 Convert 1.76 yd. to centimeters

• Sort information	**Given:** **Find:**	1.76 yd length, cm
• Strategize	**Concept Plan:** **Relationships:**	yd → m → cm 1 yd = 1.094 m 1 m = 100 cm
• Follow the concept plan to **solve** the problem	**Solution:**	$1.79\ \cancel{\text{yd}} \times \dfrac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \dfrac{100\ \text{cm}}{1\ \cancel{\text{m}}} = 160.8775\ \text{cm}$
• Sig. figs. and round	**Round:**	160.8775 cm = 161 cm
• Check	**Check:**	Units & magnitude are correct

Practice – Convert 30.0 mL to quarts

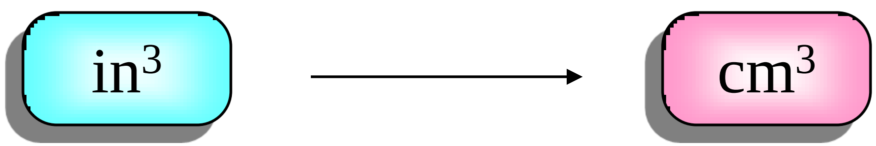
(1 L = 1.057 qt)

Convert 30.0 mL to quarts

• Sort information	**Given:** **Find:**	30.0 mL volume, qts
• Strategize	**Concept Plan:** **Relationships:**	mL → L → qt 1 L = 1.057 qt 1 L = 1000 mL
• Follow the concept plan to **solve** the problem	**Solution:**	$30.0\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{1.057\ \text{qt}}{1\ \cancel{\text{L}}} = 0.03171\ \text{qt}$
• Sig. figs. and round	**Round:**	0.03171 qt = 0.0317 qt
• Check	**Check:**	Units & magnitude are correct

Concept Plans for Units Raised to Powers

- Convert cubic inches into cubic centimeters
  1) Find relationship equivalence: 1 in = 2.54 cm
  2) Write concept plan



  3) Change equivalence into conversion factors with given unit on the bottom

$$\left(\frac{2.54\ \text{cm}}{1\ \text{in}}\right)^3 = \frac{2.54^3\ \text{cm}^3}{1^3\ \text{in}^3} = \frac{16.4\ \text{cm}^3}{1\ \text{in}^3}$$

Example 1.9 Convert 5.70 L to cubic inches

• Sort information	**Given:** **Find:**	5.70 L volume, in^3
• Strategize	**Concept Plan:** **Relationships:**	L → mL → cm^3 → in^3 1 mL = 1 cm^3, 1 mL = 10^{-3} L 1 cm = 2.54 in
• Follow the concept plan to **solve** the problem	**Solution:**	$$5.70\ \text{L} \times \frac{1\ \text{mL}}{10^{-3}\ \text{L}} \times \frac{1\ \text{cm}^3}{1\ \text{mL}} \times \frac{(1\ \text{in})^3}{(2.54\ \text{cm})^3}$$ $$= 347.835\ \text{in}^3$$
• Sig. figs. and round	**Round:**	347.835 in^3 = 348 in^3
• Check	**Check:**	Units & magnitude are correct

Practice 1.9 How many cubic centimeters are there in 2.11 yd^3?

Practice 1.9 Convert 2.11 yd^3 to cubic centimeters

• Sort information	**Given:** **Find:**	2.11 yd^3 volume, cm^3
• Strategize	**Concept Plan:** **Relationships:**	$yd^3 \rightarrow in^3 \rightarrow cm^3$ 1 yd = 36 in 1 in = 2.54 cm
• Follow the concept plan to **solve** the problem	**Solution:**	$2.11\,\cancel{yd^3} \times \frac{(36\,\cancel{in})^3}{(1\,\cancel{yd})^3} \times \frac{(2.54\,cm)^3}{(1\,\cancel{in})^3}$ $= 1613210.75\,cm^3$
• Sig. figs. and round	**Round:**	1613210.75 cm^3 $= 1.61 \times 10^6\ cm^3$
• Check	**Check:**	Units & magnitude are correct

Density as a Conversion Factor

- **can use density as a conversion factor between mass and volume!!**
  - ✓ density of H_2O = 1.0 g/mL ∴ 1.0 g H_2O = 1 mL H_2O
  - ✓ density of Pb = 11.3 g/cm^3 ∴ 11.3 g Pb = 1 cm^3 Pb

How much does 4.0 cm^3 of lead weigh?

$$4.0\ \text{cm}^3\ \text{Pb} \times \frac{\mathbf{11.3\ g\ Pb}}{\mathbf{1\ cm^3\ Pb}} = 45\ \text{g Pb}$$

Example 1.10 What is the mass in kg of 173,231 L of jet fuel whose density is 0.738 g/mL?

• Sort information	**Given:** **Find:**	173,231 L density = 0.738 g/mL mass, kg
• Strategize	**Concept Plan:** **Relationships:**	L → mL → g → kg 1 mL = 0.738 g, 1 mL = 10^{-3} L 1 kg = 1000 g
• Follow the concept plan to **solve** the problem	**Solution:** $173{,}231\ \text{L} \times \frac{1\ \text{mL}}{10^{-3}\ \text{L}} \times \frac{0.738\ \text{g}}{1\ \text{mL}} \times \frac{1\ \text{kg}}{1000\ \text{g}} = 1.33 \times 10^{5}\ \text{kg}$	
• Sig. figs. and round	**Round:**	1.33×10^{5} kg
• Check	**Check:**	Units & magnitude are correct

Order of Magnitude Estimations

- using scientific notation
- focus on the exponent on 10
- if the decimal part of the number is less than 5, just drop it
- if the decimal part of the number is greater than 5, increase the exponent on 10 by 1
- multiply by adding exponents, divide by subtracting exponents

Estimate the Answer

- Suppose you count 1.2×10^5 atoms per second for a year. How many would you count?

1 s = $1.2 \times 10^5 \approx 10^5$ atoms

1 minute = $6 \times 10^1 \approx 10^2$ s

1 hour = $6 \times 10^1 \approx 10^2$ min

1 day = $24 \approx 10^1$ hr

1 yr = $365 \approx 10^2$ days

$$1\,\text{yr} \times \frac{10^2\ \text{days}}{1\,\text{yr}} \times \frac{10^1\ \text{hr}}{1\,\text{day}} \times \frac{10^2\ \text{min}}{1\,\text{hr}} \times \frac{10^2\ \text{s}}{1\,\text{min}} \times \frac{10^5\ \text{atoms}}{1\,\text{s}}$$

$$\approx 10^{12}\ \text{atoms}$$

Problem Solving with Equations

- When solving a problem involves using an equation, the concept plan involves being given all the variables except the one you want to find
- Solve the equation for the variable you wish to find, then substitute and compute

Using Density in Calculations

Concept Plans:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

m, V → D

$$\text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

m, D → V

$$\text{Mass} = \text{Density} \times \text{Volume}$$

V, D → m

Example 1.12 Find the density of a metal cylinder with mass 8.3 g, length 1.94 cm, and radius 0.55 cm

• Sort information	**Given:** **Find:**	m = 8.3 g l = 1.94 cm, r = 0.55 cm density, g/cm^3
• Strategize	**Concept Plan:** **Relationships:**	$l, r \rightarrow V$ $\quad$ $m, V \rightarrow d$ $V = \pi\, r^2\, l$ $d = m/V$
• Follow the concept plan to **solve** the problem • Sig. figs. and round	**Solution:**	$V = \pi\,(0.55\text{ cm})^2\,(1.94\text{ cm})$ $V = 1.\underline{8}436\text{ cm}^3$ $d = \dfrac{8.3\text{ g}}{1.\underline{8}436\text{ cm}^3} = 4.50206\text{ g/cm}^3$ $d = 4.5\text{ g/cm}^3$
• Check	**Check:**	Units & magnitude OK