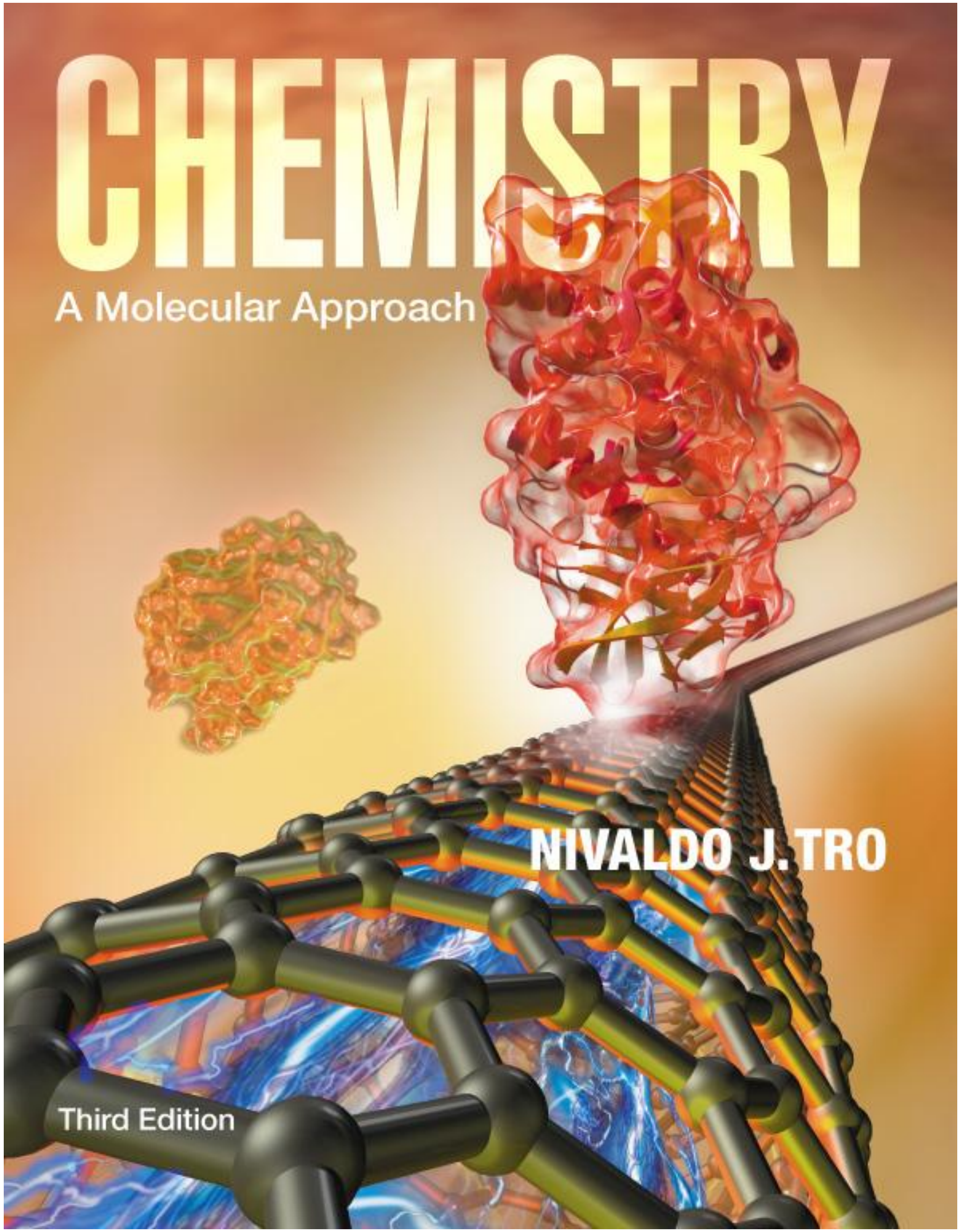

Lecture Presentation

Chapter 1

Matter, Measurement, and Problem Solving

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What Do You Think?

- **What do you think** is the most important idea in all of human knowledge?

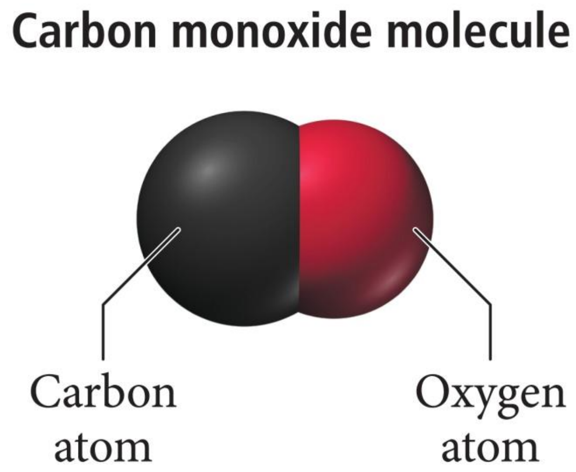
- If we limit ourselves only to scientific answers, it would be this:

  – **The properties of matter are determined by the properties of molecules and atoms**.

What Do You Think?

- Atoms and molecules determine how matter behaves; if they were different, matter would be different.
  - The properties of water molecules determine how water behaves; the properties of sugar molecules determine how sugar behaves.
- The understanding of matter at the molecular level gives us unprecedented control over that matter.

Atoms and Molecules



- The air contains carbon monoxide pollutant.
- Carbon monoxide gas is composed of carbon monoxide **molecules**.
- Each molecule contains a carbon **atom** and an oxygen **atom** held together by a chemical bond.

Atoms and Molecules

- **Atoms** are the submicroscopic particles that constitute the fundamental building blocks of ordinary matter.

- Free atoms are rare in nature; instead they bind together in specific geometrical arrangements to form **molecules**.

Atoms and Molecules

- If we want to understand the substances around us, we must understand the atoms and molecules that compose them—this is the central goal of chemistry.
  - **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

- The approach to scientific knowledge is empirical—it is based on ***observation*** and ***experiment***.
- The scientific method is a process for understanding nature by observing nature and its behavior, and by conducting experiments to test our ideas.
- Key characteristics of the scientific method include **observation**, formulation of **hypotheses**, **experimentation**, and formulation of **laws and theories**.

Observations

- Observations are also known as **data.**
- They are the descriptions about the characteristics or behavior of nature.
  - Antoine Lavoisier (1743–1794) noticed that there was no change in the total mass of material within the container during combustion.
- Observations often lead scientists to formulate a **hypothesis.**

Hypothesis

- A hypothesis is a tentative interpretation or explanation of the observations.
    - For example, Lavoisier explained his observations on combustion by hypothesizing that when a substance burns, it combines with a component of air.
- A good hypothesis is *falsifiable.*
    - The results of an experiment may support a hypothesis or—prove it wrong—in which case the scientist must modify or discard the hypothesis.

A Scientific Law

- A brief statement that summarizes past observations and predicts future ones
  - **Law of conservation of mass**— "In a chemical reaction matter is neither created nor destroyed."
- Allows you to predict future observations
  - So you can test the law with experiments
- Unlike state laws, you cannot choose to violate a scientific law.

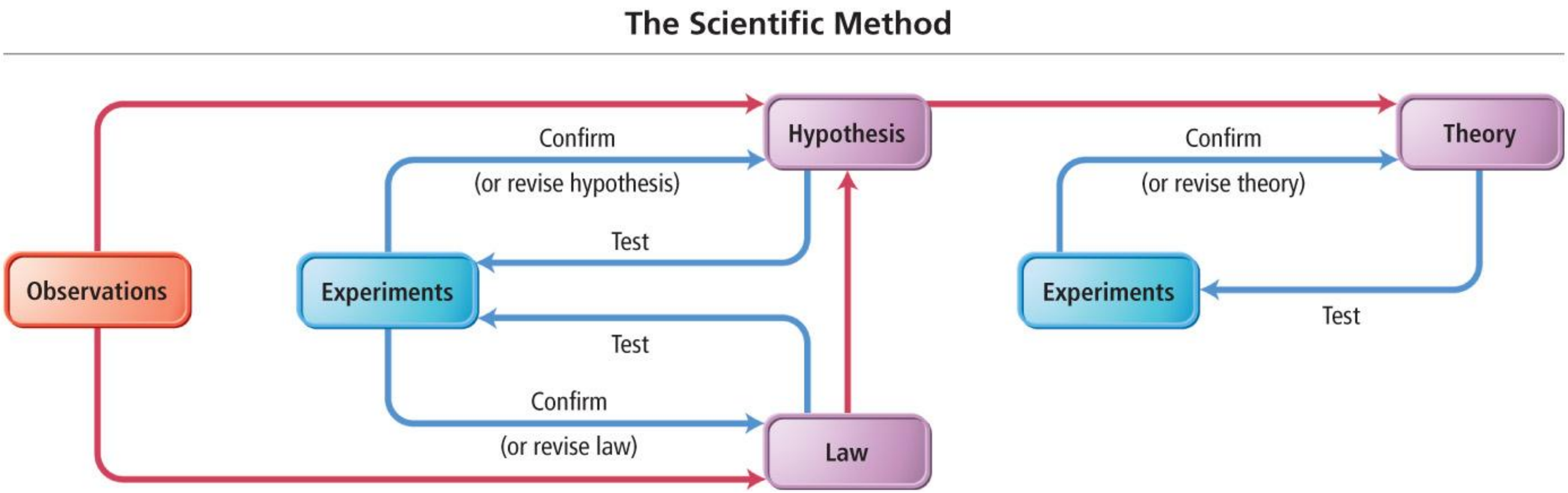
Theory

- One or more well-established hypotheses may form the basis for a scientific **theory**.
- A scientific theory is a model for the way nature is and tries to explain not merely what nature does, but why.
- Theories are validated by experiments.
- Theories can never be conclusively proven because some new observation or experiment always has the potential to reveal a flaw.

Theory

- General explanation for the characteristics and behavior of nature
- Models of nature
    - Dalton′s atomic theory
- Can be used to predict future observations
    - So they can be tested by experiments

The Scientific Approach to Knowledge



Conceptual Connection 1.1

Which statement best explains the difference between a law and a theory?

(a) A law is truth whereas a theory is a mere speculation.

(b) A law summarizes a series of related observations, while a theory gives the underlying reasons for them.

(c) A theory describes *what* nature does; a law describes *why* nature does it.

Conceptual Connection 1.1

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The Classification of Matter

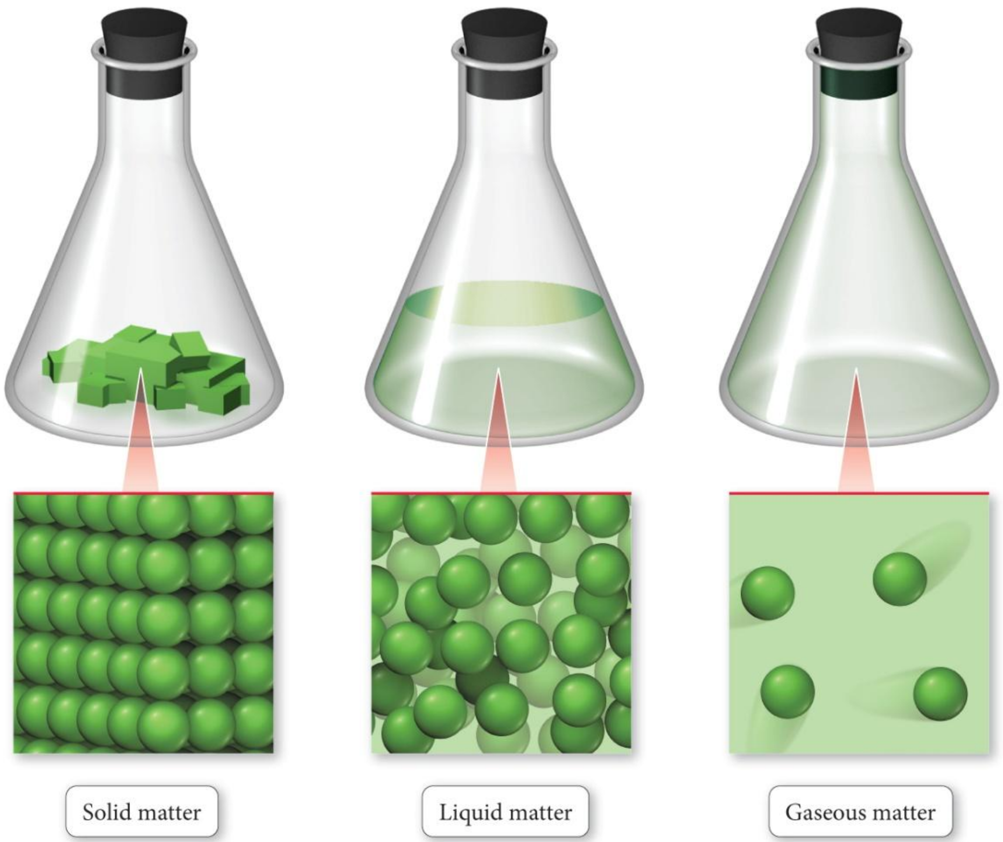
- **Matter** is anything that occupies space and has mass.
    - Your textbook, your desk, your chair, and even your body are all composed of matter.
- We can classify matter according to its **state** (its physical form) and its **composition** (the basic components that make it up).

The States of Matter

- **Matter** can be classified as solid, liquid, or gas based on what properties it exhibits.
- The state of matter changes from solid to liquid to gas with increasing temperature.

Structure Determines Properties

- The atoms or molecules have different structures in solids, liquids, and gases—leading to different properties.

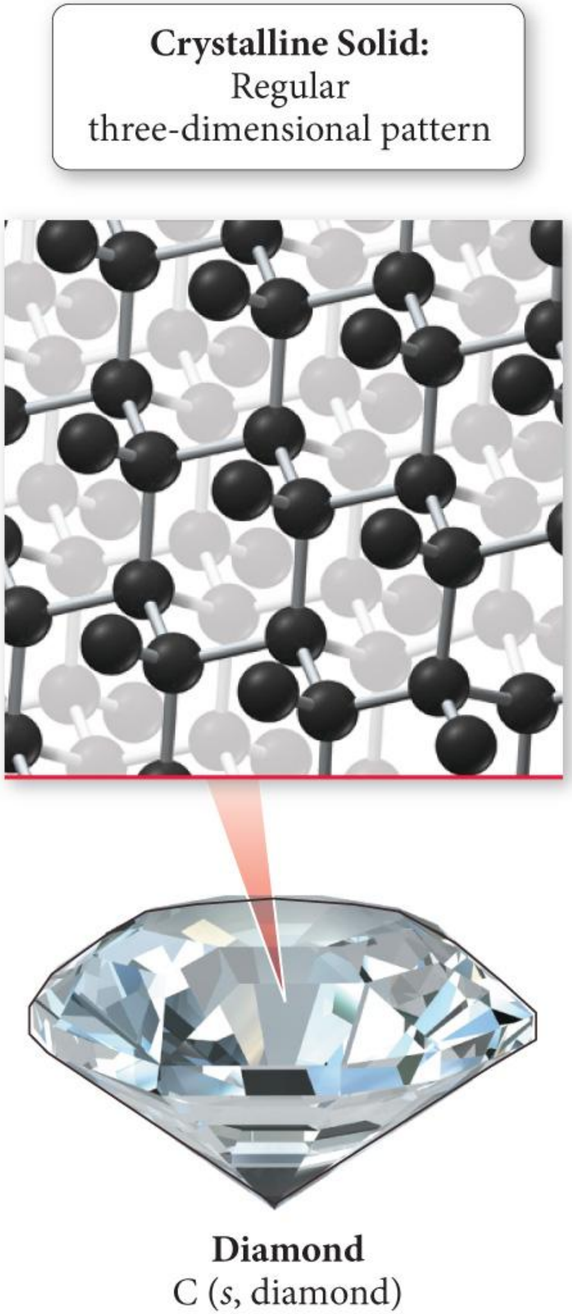


Solid Matter

- In *solid matter*, atoms or molecules pack close to each other in fixed locations.
- Although the atoms and molecules in a solid vibrate, they do not move around or past each other.
- Consequently, a solid has a fixed volume and rigid shape.
  - Ice, aluminum, and diamond are good examples of solids.

Solid Matter

- Solid matter may be **crystalline**—in which case its atoms or molecules are in patterns with long-range, repeating order.
    - Table salt and diamond are examples of solid matter.
- Others may be **amorphous**, in which case its atoms or molecules do not have any long-range order.
    - Examples of *amorphous* solids include glass and plastic.



Liquid Matter

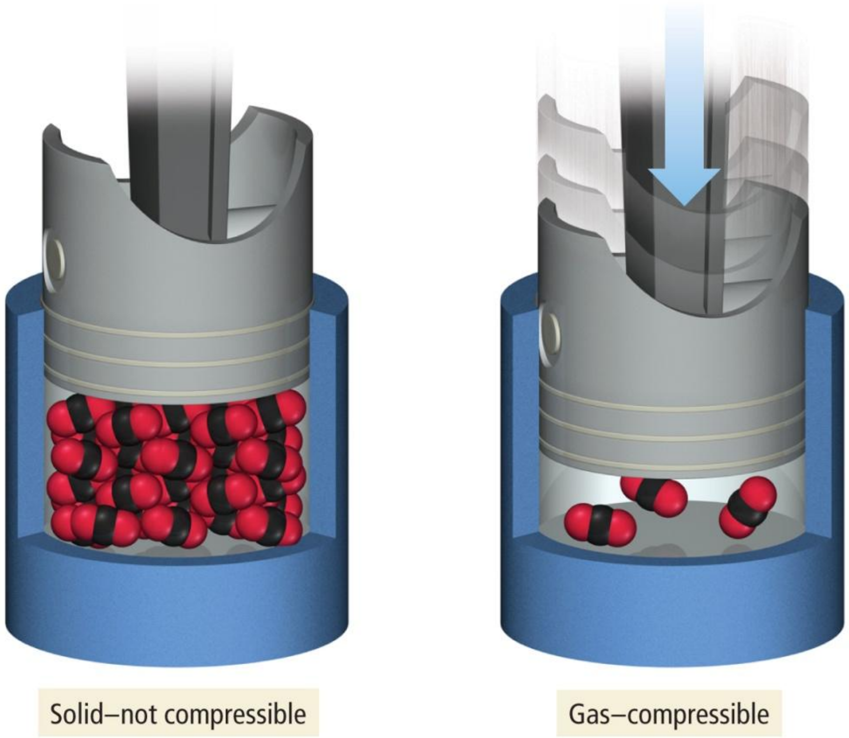
- In *liquid matter*, atoms or molecules pack about as closely as they do in solid matter, but they are free to move relative to each other.

- Liquids have fixed volume but not a fixed shape.

- Liquids' ability to flow makes them assume the shape of their container.
  - Water, alcohol, and gasoline are all substances that are liquids at room temperature**.**

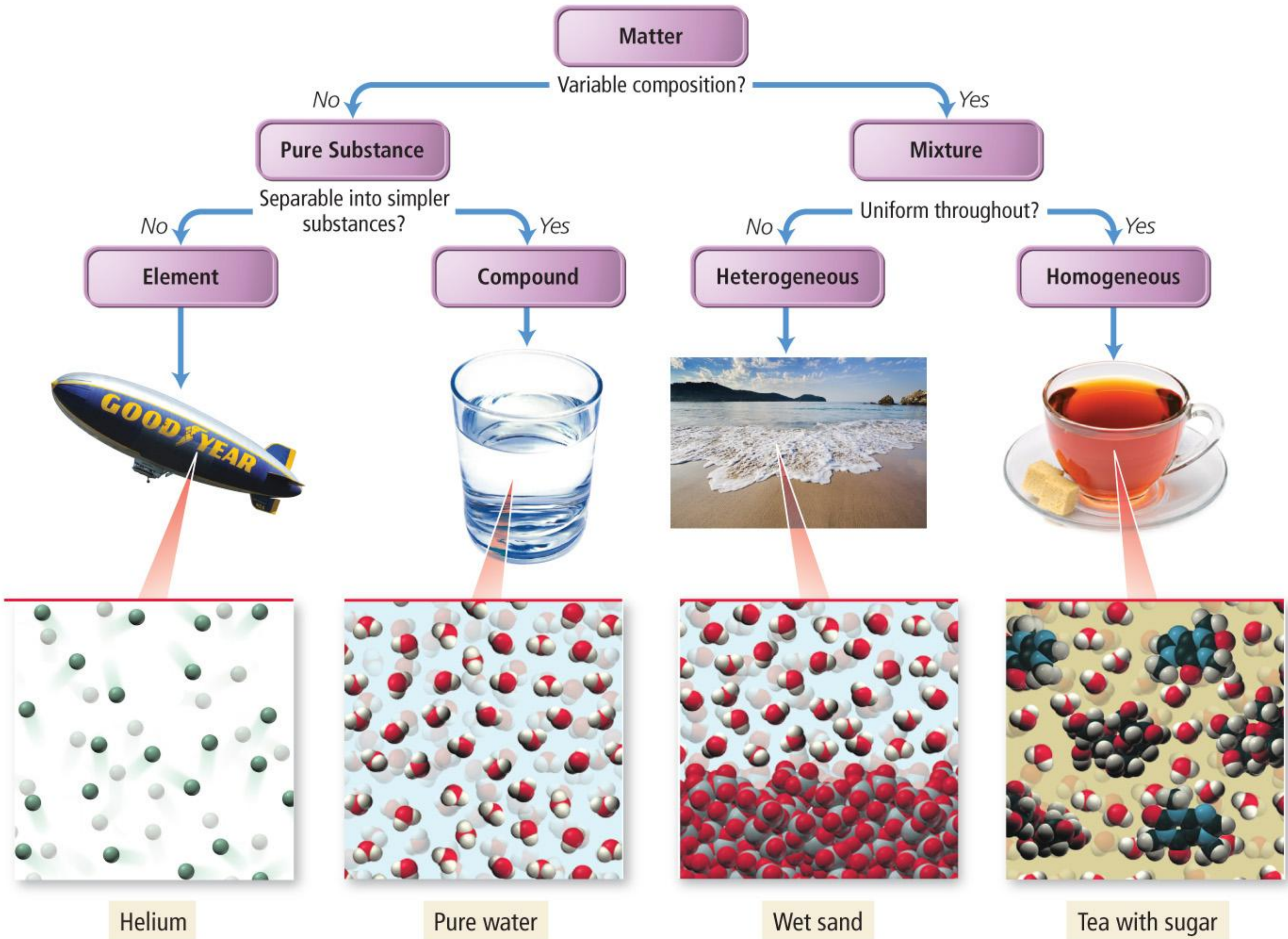
Gaseous Matter

- In *gaseous matter*, atoms or molecules have a lot of space between them.
- They are free to move relative to one another.
- These qualities make gases *compressible*.



The Classification of Matter by Components

- Matter can also be classified according to its composition: elements, compounds, and mixtures.



Classification of Matter by Components

- The first division in the classification of matter is between a *pure substance* and a *mixture*.
- A **pure substance** is made up of only one component and its composition is invariant.
- A **mixture**, by contrast, is a substance composed of two or more components in proportions that can vary from one sample to another.

Classification of Pure Substances

- Pure substances categorize into two types:
    - Elements
    - Compounds
- This categorization depends on whether or not they can be broken down (or decomposed) into simpler substances.

Classification of Pure Substances

- An **element** is a substance that cannot be chemically broken down into simpler substances.
    - Basic building blocks of matter
    - Composed of single type of atom, like helium
- A **compound** is a substance composed of two or more elements in fixed definite proportions.
- Most elements are chemically reactive and combine with other elements to form compounds like water, sugar, etc.

Classification of Mixtures

- **Mixtures** can be categorized into two types:
  - Heterogeneous mixtures
  - Homogeneous mixtures
- This categorization of mixture depends on how *uniformly* the substances within them mix.

Heterogeneous Mixture

- A **heterogeneous mixture** is one in which the composition varies from one region of the mixture to another.
    - Made of multiple substances, whose presence can be seen (Example: a salt and sand mixture)
        - Portions of a sample of heterogeneous mixture have different composition and properties.

Homogeneous Mixture

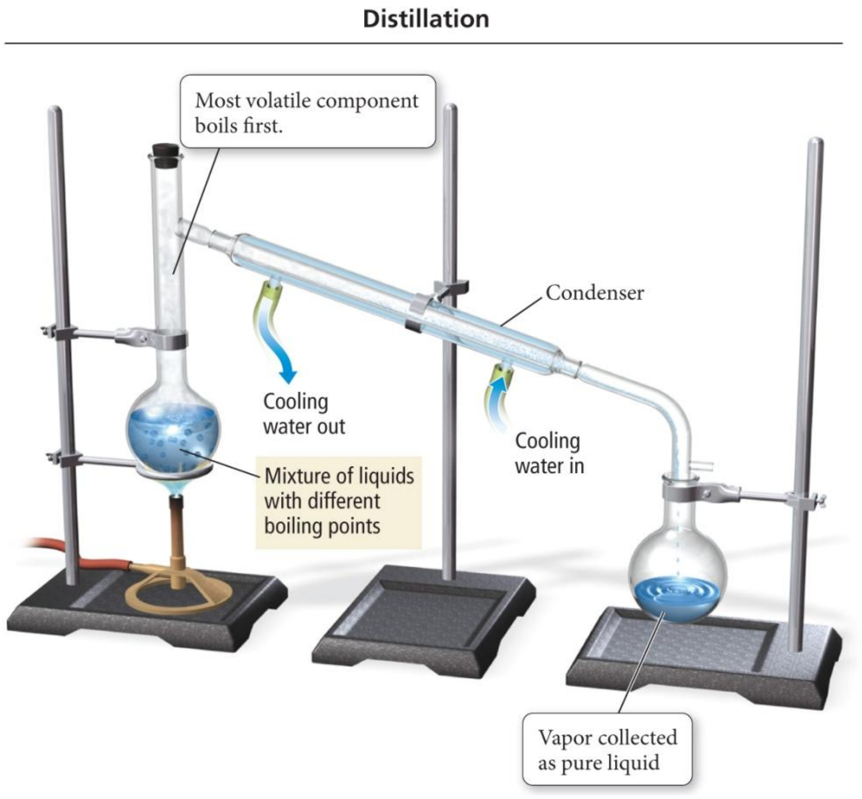
- A **homogeneous mixture** is one made of multiple substances, but appears to be one substance.
- All portions of a sample have the same composition and properties (like sweetened tea).
- Homogeneous mixtures have uniform compositions because the atoms or molecules that compose them mix uniformly.

Separating Mixtures

- Mixtures are separable because the different components have different physical or chemical properties.
- Various techniques that exploit these differences are used to achieve separation.
- A mixture of sand and water can be separated by **decanting**—carefully pouring off the water into another container.

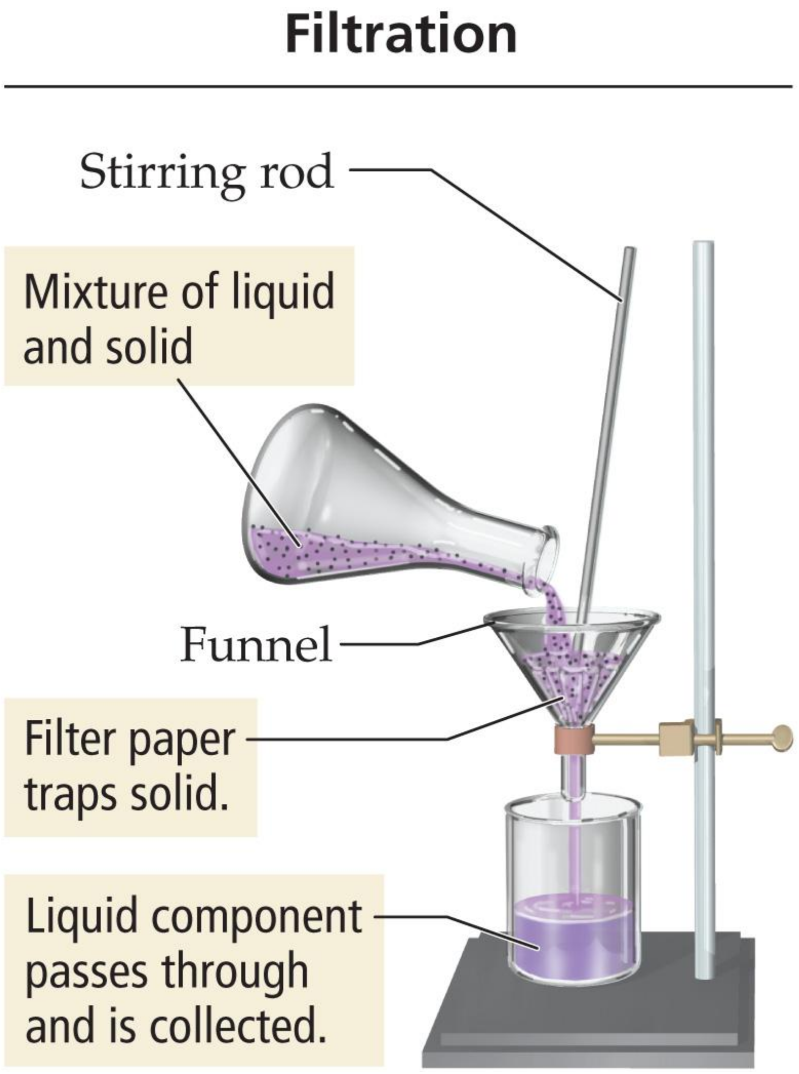
Separating Mixtures

- A homogeneous mixture of liquids can usually be separated by **distillation**, a process in which the mixture is heated to boil off the more **volatile** (easily vaporizable) liquid. The volatile liquid is then re-condensed in a condenser and collected in a separate flask.



Separating Mixtures

- A mixture of an insoluble solid and a liquid can be separated by **filtration**—process in which the mixture is poured through filter paper in a funnel.



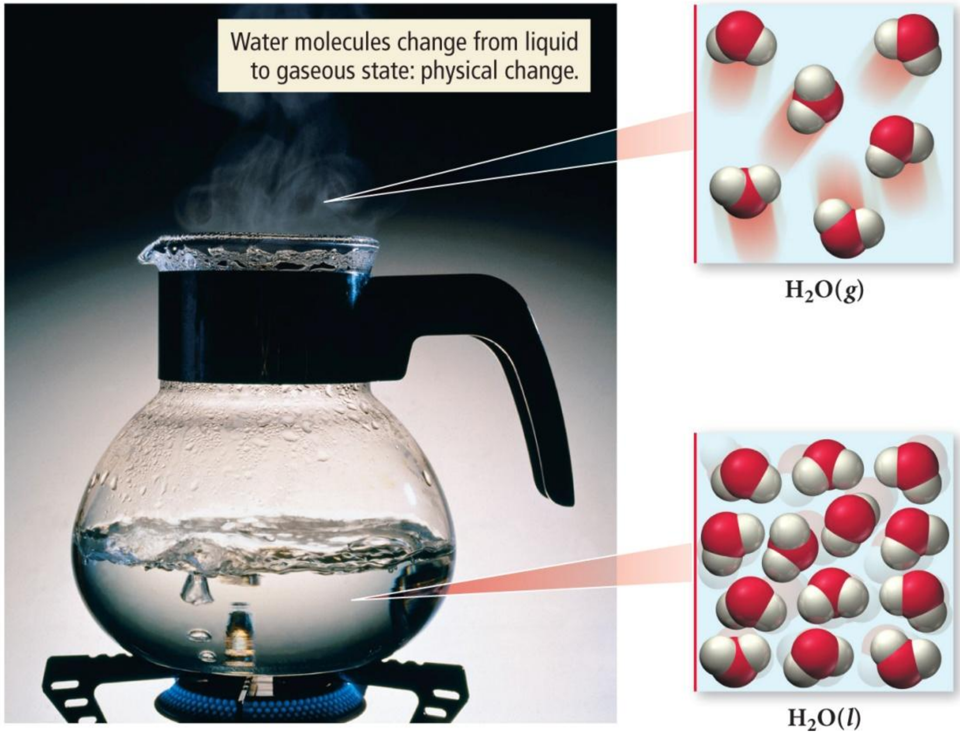
Physical and Chemical Changes

Physical Change:

- Changes that alter only the state or appearance, but not composition, are **physical changes**.
- The atoms or molecules that compose a substance *do not change* their identity during a physical change.

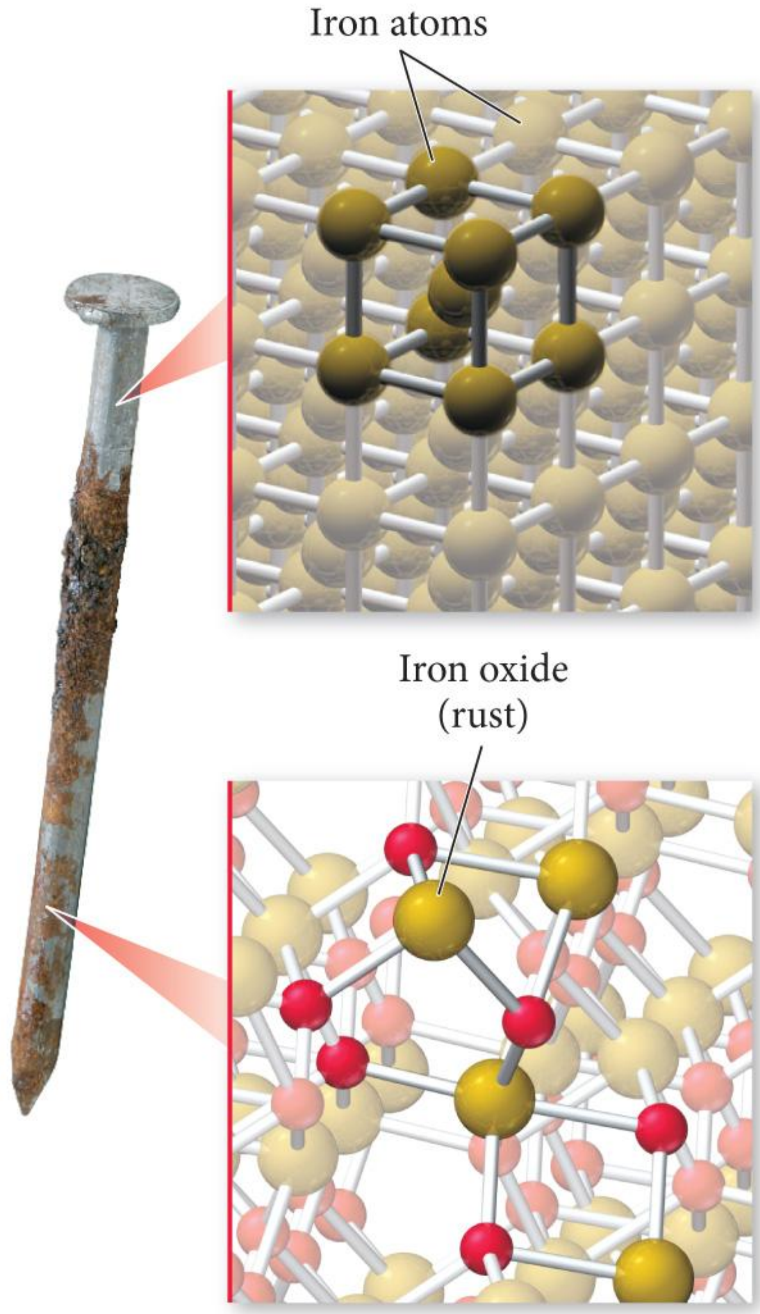
Physical Change

- When water boils, it changes its state from a liquid to a gas.
- The gas remains composed of water molecules, so this is a physical change.

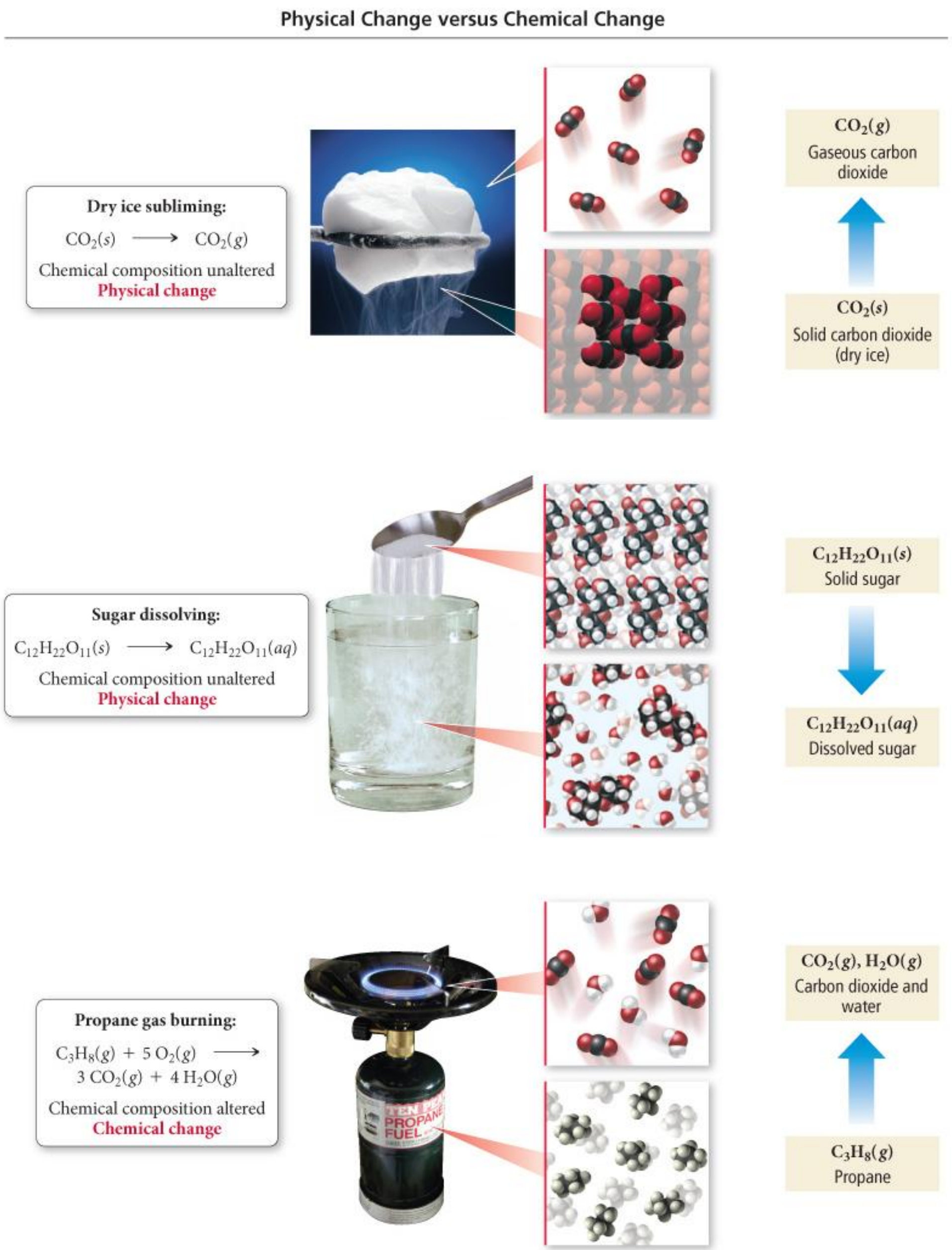


Chemical Change

- Changes that alter the composition of matter are **chemical changes**.
- During a chemical change, atoms rearrange, transforming the original substances into different substances.
- Rusting of iron is a chemical change.



Physical and Chemical Changes



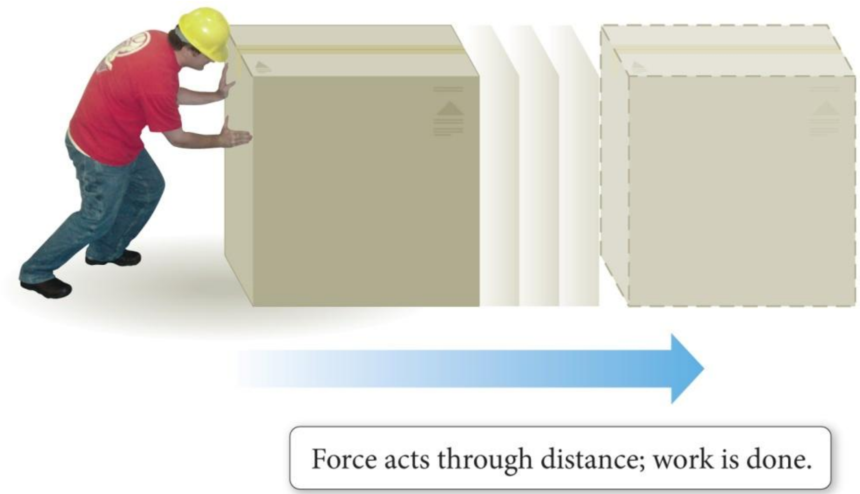
Physical and Chemical Properties

- A **physical property** is a property that a substance displays without changing its composition.
  - The smell of gasoline is a physical property.
  - Odor, taste, color, appearance, melting point, boiling point, and density are all physical properties.

- A **chemical property** is a property that a substance displays only by changing its composition via a chemical change (or chemical reaction).
  - The flammability of gasoline, in contrast, is a chemical property.
  - Chemical properties include corrosiveness, acidity, and toxicity.

Energy: A Fundamental Part of Physical and Chemical Change

- **Energy** is the *capacity to do work.*
- **Work** is defined as the action of a force through a distance.

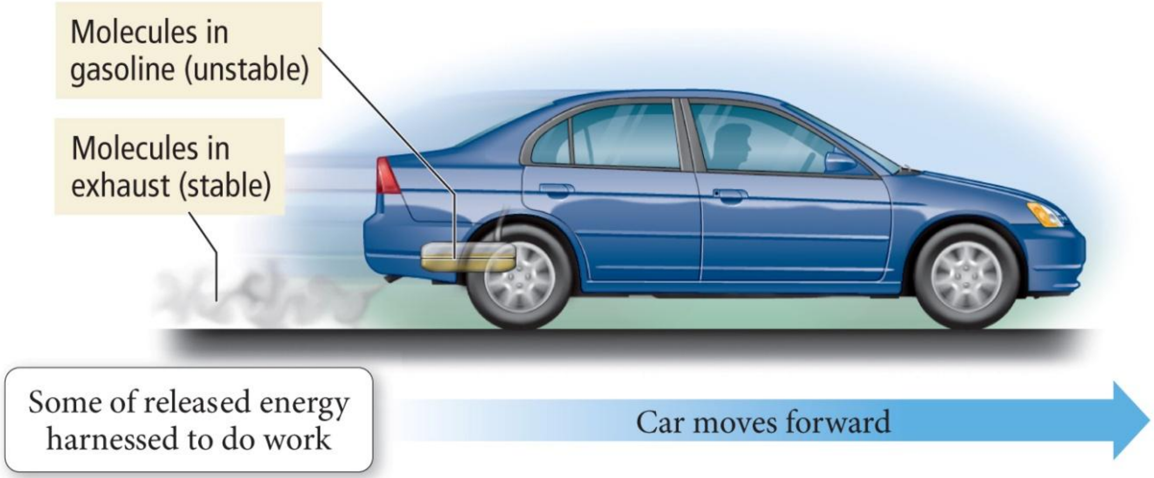


- When you push a box across the floor or pedal your bicycle across the street, you have done work.

Energy

- **Kinetic energy** is the energy associated with the motion of an object.
- **Potential energy** is the energy associated with the position or composition of an object.
- **Thermal energy** is the energy associated with the temperature of an object.
    - Thermal energy is actually a type of kinetic energy because it arises from the motion of the individual atoms or molecules that make up an object.

Summarizing Energy



- Energy is always conserved in a physical or chemical change; it is neither created nor destroyed (law of conservation of energy).
- Systems with high potential energy tend to change in a direction that lowers their potential energy, releasing energy into the surroundings.

The Units of Measurement

- In chemistry, **units**—standard quantities used to specify measurements—are critical.
- **The two most common unit systems are as follows:**
    - **Metric system**, used in most of the world
    - **English system**, used in the United States
- Scientists use the **International System of Units (SI)**, which is based on the metric system.
    - The abbreviation *SI* comes from the French, phrase *Système International d'Unités.*

The Standard Units

TABLE 1.1 SI Base Units

Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

The Meter: A Measure of Length

- The **meter (m)** is slightly longer than a yard (1 yard is 36 inches, while 1 meter is 39.37 inches).

- 1 meter = 1/10,000,000 of the distance from the equator to the North Pole (through Paris).
  - The International Bureau of Weights and Measures now defines it more precisely as the distance light travels through a vacuum in a certain period of time, 1/299,792,458 second.

The Kilogram: A Measure of Mass

- The **mass** of an object is a measure of the quantity of matter within it.
- The SI unit of mass = kilogram (kg).
  - 1 kg = 2 lb 3 oz
- A second common unit of mass is the gram (g).
  - One gram is 1/1000 kg.
- **Weight** of an object is a measure of the *gravitational pull* on its matter.

The Second: A Measure of Time

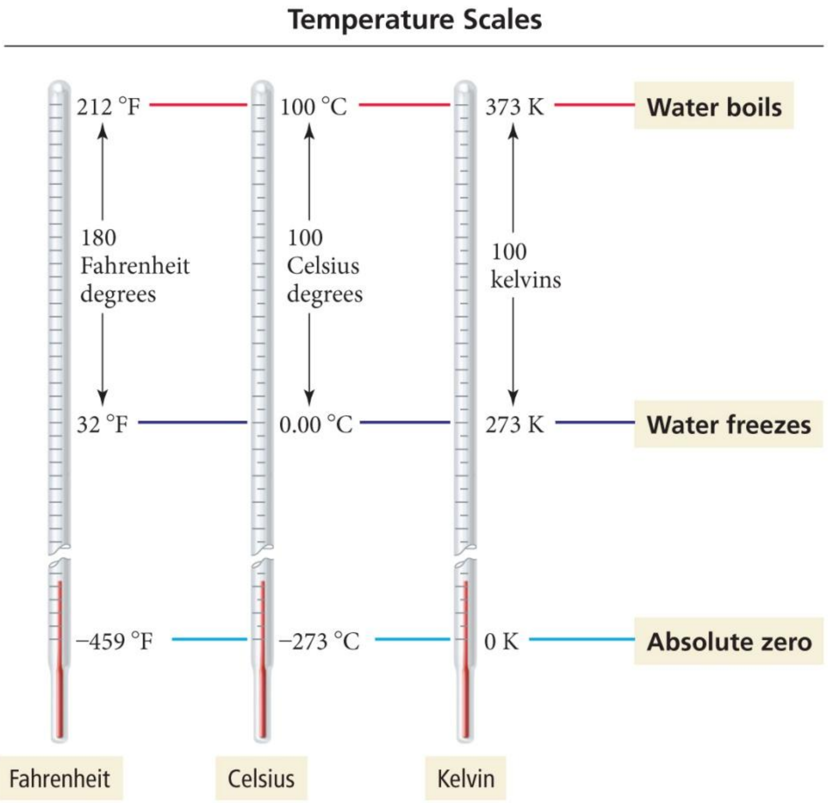
- Measure of the duration of an event
- SI units = second (s)
- 1 s is defined as the period of time it takes for a specific number of radiation events of a specific transition from cesium-133.

The Kelvin: A Measure of Temperature

- The **Kelvin** **(K)** is the SI unit of **temperature**.
- The temperature is a measure of the average amount of kinetic energy of the atoms or molecules that compose the matter.
- Temperature also determines the direction of thermal energy transfer, or what we commonly call heat.
- Thermal energy transfers from hot to cold objects.

The Kelvin: A Measure of Temperature

- **Kelvin scale** (*absolute scale*) assigns 0 K (absolute zero) to the coldest temperature possible.
- Absolute zero (–273 °C or –459 °F) is the temperature at which molecular motion virtually stops. Lower temperatures do not exist.



A Measure of Temperature

- The Fahrenheit degree is five-ninths the size of a Celsius degree.
- The Celsius degree and the Kelvin degree are the same size.
- Temperature scale conversion is done with these formulas:

$$°C = \frac{(°F - 32)}{1.8}$$

$$K = °C + 273.15$$

Prefix Multipliers

- The International System of Units uses the **prefix multipliers** shown in Table 1.2 with the standard units.
- These multipliers change the value of the unit by the powers of 10 (just like an exponent does in scientific notation).
- For example, the kilometer has the prefix *kilo* meaning 1000 or 10^3.

Counting Significant Figures

- Significant figures deal with writing numbers to reflect precision.
- The precision of a measurement depends on the instrument used to make the measurement.
- The preservation of this precision during calculations can be accomplished by using *significant figures.*

Counting Significant Figures

- *The greater the number of significant figures, the greater the certainty of the measurement.*

- **To determine the number of significant figures in a number, follow these rules (examples are on the right).**

Significant Figure Rules		Examples
1. All nonzero digits are significant	28.03	0.0540
2. Interior zeroes (zeroes between two nonzero digits) are significant.	408	7.0301

Counting Significant Figures

Significant Figure Rules	Examples	
3. Leading zeroes (zeroes to the left of the first nonzero digit) are not significant. They only serve to locate the decimal point.	0.0032	0.00006
	not significant	
4. Trailing zeroes (zeroes at the end of a number) are categorized as follows:	45.000	3.5600
▪ Trailing zeroes after a decimal point are always significant.		
▪ Trailing zeroes before a decimal point (and after a nonzero number) are always significant.	140.00	2500.55
▪ Trailing zeroes before an *implied* decimal point are ambiguous and should be avoided by using scientific notation.	1200	Ambiguous
	1.2×10^3	2 significant figures
	1.20×10^3	3 significant figures
	1.200×10^3	4 significant figures
▪ Decimal points are placed after one or more trailing zeroes if the zeroes are to be considered significant.	1200.	4 significant figures

Exact Numbers

- *Exact numbers have an unlimited number of significant figures.*

- Exact counting of discrete objects
- Integral numbers that are part of an equation
- Defined quantities

- *Some conversion factors are defined quantities, while others are not.*

Significant Figures in Calculations

- In calculations using measured quantities, the results of the calculation must reflect the precision of the measured quantities.

- We should not lose or gain precision during mathematical operations.

Significant Figure: Rules for Calculations

Multiplication and Division Rule:

- In multiplication or division, the result carries the same number of significant figures as the factor with the fewest significant figures.

$$\underset{\text{(4 sig. figures)}}{1.052} \times \underset{\text{(5 sig. figures)}}{12.504} \times \underset{\text{(2 sig. figures)}}{0.53} = 6.7208 = \underset{\text{(2 sig. figures)}}{6.7}$$

$$\underset{\text{(5 sig. figures)}}{2.0035} \div \underset{\text{(3 sig. figures)}}{3.20} = 0.626094 = \underset{\text{(3 sig. figures)}}{0.626}$$

Rules for Calculations

Addition and Subtraction Rule:

- In addition or subtraction the result carries the same number of decimal places as the quantity with the fewest decimal places.

$$
\begin{array}{r}
2.345 \\
0.07 \\
2.9975 \\
\hline
5.4125
\end{array}
= 5.41
\qquad\qquad
\begin{array}{r}
5.9 \\
-\,0.221 \\
\hline
5.679
\end{array}
= 5.7
$$

It is helpful to draw a line next to the number with the fewest decimal places. This line determines the number of decimal places in the answer.

Rules for Calculations

Rules for Rounding:

- When rounding to the correct number of significant figures,
  - round down if the last (or leftmost) digit dropped is four or less;
  - round up if the last (or leftmost) digit dropped is five or more.

Rules for Rounding

- Round to two significant figures:

  5.37 rounds to 5.4

  5.34 rounds to 5.3

  5.35 rounds to 5.4

  5.349 rounds to 5.3

- Notice in the last example that only the *last (or leftmost) digit being dropped* determines in which direction to round—ignore all digits to the right of it.

Rounding in Multistep Calculations

- To avoid rounding errors in multistep calculations round only the final answer.

- Do not round intermediate steps. If you write down intermediate answers, keep track of significant figures by underlining the least significant digit.

$$
\begin{aligned}
6.78 \times 5.903 \times (5.489 - 5.01) \\
= 6.78 \times 5.903 \times 0.4\underline{7}9 \\
= 19.1707 \\
= 19
\end{aligned}
$$

underline least
significant digit

Precision and Accuracy

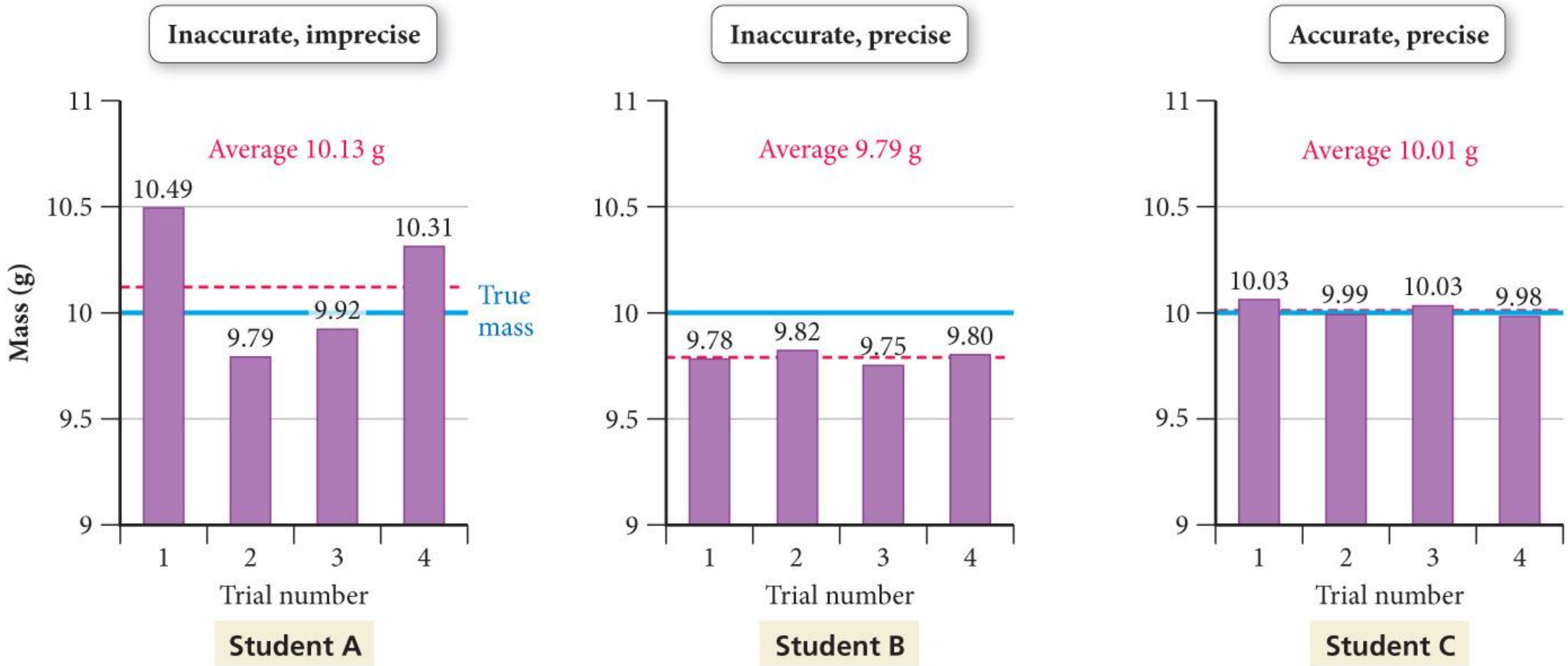
- **Accuracy** refers to how close the measured value is to the actual value.
- **Precision** refers to how close a series of measurements are to one another or how reproducible they are.

Precision and Accuracy

- Consider the results of three students who repeatedly weighed a lead block known to have a true mass of 10.00 g (indicated by the solid horizontal blue line on the graphs).

	Student A	Student B	Student C
Trial 1	10.49 g	9.78 g	10.03 g
Trial 2	9.79 g	9.82 g	9.99 g
Trial 3	9.92 g	9.75 g	10.03 g
Trial 4	10.31 g	9.80 g	9.98 g
Average	10.13 g	9.79 g	10.01 g

Precision and Accuracy



- Measurements are said to be
  - precise if they are consistent with one another.
  - accurate only if they are close to the actual value.

Precision and Accuracy

- The results of student A are both inaccurate (not close to the true value) and imprecise (not consistent with one another).
    - ***Random error*** *is an error that has the equal probability of being too high or too low.*

- The results of student B are precise (close to one another in value), but inaccurate.
    - ***Systematic error*** *is an error that tends toward being either too high or too low*.

- The results of student C display little systematic error or random error—they are both accurate and precise.

Solving Chemical Problems

- Most chemistry problems you will solve in this course are **unit conversion problems.**
- Using units as a guide to solving problems is called **dimensional analysis.**
- Units should always be included in calculations; they are multiplied, divided, and canceled like any other algebraic quantity.

Dimensional Analysis

- A unit equation is a statement of two equivalent quantities, such as

2.54 cm = 1 in.

- A **conversion factor** is a fractional quantity of a unit equation with the units we are ***converting from*** on the bottom and the units we are ***converting to*** on the top.

Dimensional Analysis

- Most unit conversion problems take the following form:

$$\text{Information given} \times \text{conversion factor(s)} = \text{information sought}$$

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

Dimensional Analysis

Units Raised to a Power:

- When building conversion factors for units raised to a power, remember to raise both the number and the unit to the power. For example, to convert from in^2 to cm^2, we construct the conversion factor as follows:

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

$$(2.54\ \text{cm})^2 = (1\ \text{in})^2$$

$$(2.54)^2\ \text{cm}^2 = 1^2\ \text{in}^2$$

$$6.45\ \text{cm}^2 = 1\ \text{in}^2$$

$$\frac{6.45\ \text{cm}^2}{1\ \text{in}^2} = 1$$

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 call 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

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

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
• Because 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		$1.8(°C) = (°F - 32)$ $1.8(°C) + 32 = °F$
• Substitute and compute		$1.8(40.00) + 32 = °F$ $104.00\ °F = °F$

Practice – Convert 0.0°F into Kelvin

Practice – Convert 0.0°F into Kelvin

• Sort information	**Given:** **Find:**	0.0 °F Kelvin
• Strategize	**Concept Plan:** **Equations:**	$°F \rightarrow °C \rightarrow K$ $°C = \frac{(°F - 32)}{1.8}$ $K = °C + 273.15$
• Follow the concept plan to **solve** the problem	**Solution:** $°C = \frac{(0.0°F - 32.0)}{1.8}$ $°C = -1\underline{7}.8°C$	$K = °C + 273.15$ $K = -1\underline{7}.8 + 273.15$ $K = 25\underline{5}.37$
• Sig. figs. and round	**Round:**	$25\underline{5}.37$ K = 255 K
• Check	**Check:**	Because kelvin temperatures are always positive and generally between 250 and 300, the answer makes sense

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

Tro: Chemistry: A

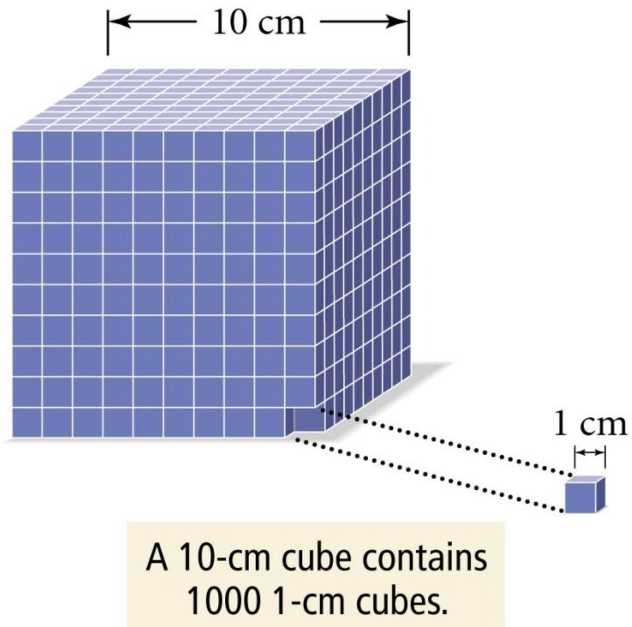
Common Prefix Multipliers in the SI System

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}

Tro: Chemistry: A

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)
  - $1\ m^3 = 10^6\ cm^3$
  - $1\ cm^3 = 10^{-6}\ m^3 = 0.000\ 001\ m^3$
- Commonly measure liquid or gas volume in milliliters (mL)
  - 1 L is slightly larger than 1 quart
  - $1\ L = 1\ dm^3 = 1000\ mL = 10^3\ mL$
  - $1\ mL = 0.001\ L = 10^{-3}\ L$
  - **$1\ mL = 1\ cm^3$**



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

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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)

Practice — which of the following units would be best used for measuring the diameter of a quarter?

a) kilometer

b) meter

c) centimeter

d) micrometer

e) megameters

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Density

Intensive and Extensive Properties

- **Extensive properties** are properties whose 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?
- **Intensive properties** are properties whose value is independent of the quantity of matter
  - intensive properties are often used to identify the type of matter

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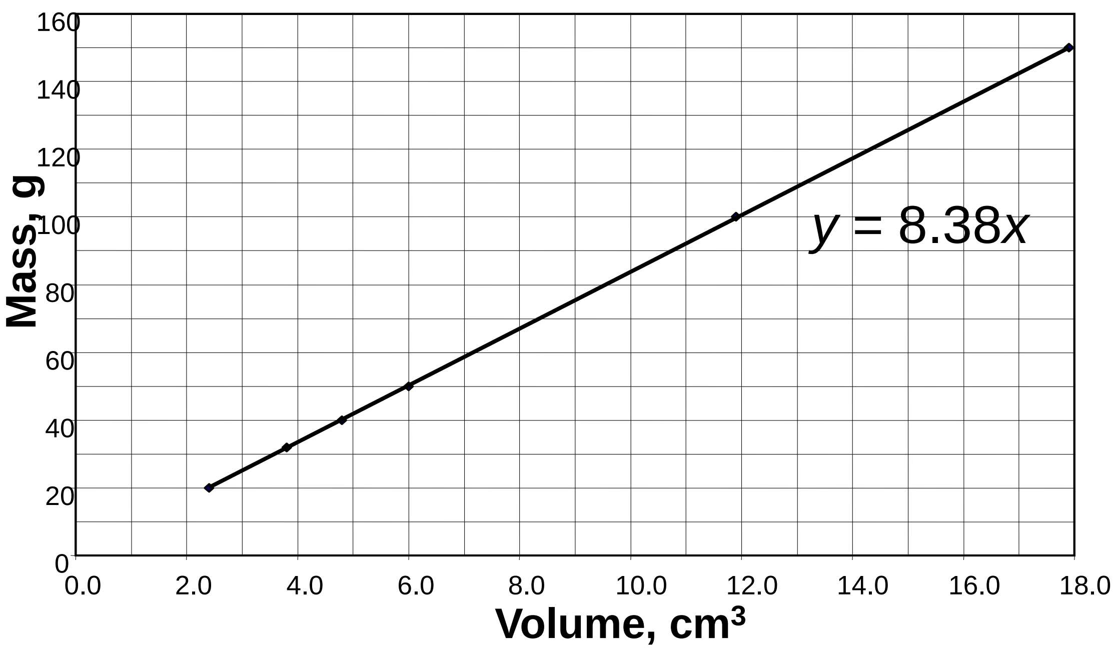
Mass & Volume

- Two main physical properties of matter
- Mass and volume are **extensive properties**
- Even though mass and volume are individual properties, for a given type of matter they are related to each other!

Measurements of the Mass and Volume of Several Samples 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



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Density

- Density is the ratio of mass to volume
  - is an **intensive property**
- 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!

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

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

- 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

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

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

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Calculating Density

- What is the density of a brass sample if 100.0 g added to a cylinder of water causes the water level to rise from 25.0 mL to 36.9 mL?

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Practice — What is the density of the brass sample?

Sort information	**Given:** **Find:**	mass = 100 g vol displ: 25.0 → 36.9 mL *d*, g/cm^3
Strategize	**Concept Plan:** **Equation:**	$m, V \rightarrow d$ $d = \dfrac{m}{V}$
Solve the equation for the unknown variable	**Solution:** V = 36.9−25.0 = 11.9 mL = 11.9 cm^3	$d = \dfrac{100\ \text{g}}{11.9\ \text{cm}^3} = 8.4\underline{0}33\ \text{g/cm}^3$
Sig. figs. and round	**Round:**	$8.4\underline{0}33\ \text{g/cm}^3 = 8.40\ \text{g/cm}^3$
Check	**Check:**	units and number make sense

Example 1.7: Convert 1.76 yd to centimeters

• Sort the information	**Given:** **Find:**	1.76 yd length, cm
• Strategize	**Conceptual Plan:** **Relationships:**	yd → m → cm 1 m = 1.094 yd 1 m = 100 cm
• Follow the conceptual 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 are correct; number makes sense: cm << yd

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Practice – Convert 30.0 mL to quarts

(1 mL = 0.001 L; 1 L = 1.057 qt)

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Practice – Convert 30.0 mL to quarts

• Sort information	**Given:** **Find:**	30.0 mL volume, qts
• Strategize	**Conceptual Plan:** **Relationships:**	mL → L → qt 1 L = 1.057 qt 0.001 L = 1 mL
• Follow the conceptual plan to solve the problem	**Solution:**	$30.0\,\cancel{\text{mL}} \times \frac{0.001\,\cancel{\text{L}}}{1\,\cancel{\text{mL}}} \times \frac{1.057\,\text{qt}}{1\,\cancel{\text{L}}} = 0.03171\,\text{qt}$
• Sig. figs. and round	**Round:**	0.03171 qt = 0.0317 qt
• Check	**Check:**	units are correct; and number makes sense: mL << qt

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Conceptual 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

$$\text{in.}^3 \longrightarrow \text{cm}^3$$

  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	**Conceptual Plan:** **Relationships:**	$\text{L} \rightarrow \text{mL} \rightarrow \text{cm}^3 \rightarrow \text{in}^3$ $1\ \text{mL} = 1\ \text{cm}^3$, $1\ \text{mL} = 10^{-3}\ \text{L}$ $1\ \text{cm} = 2.54\ \text{in.}$
• Follow the conceptual plan to solve the problem	**Solution:**	$5.70\ \cancel{\text{L}} \times \frac{1\ \cancel{\text{mL}}}{10^{-3}\ \cancel{\text{L}}} \times \frac{1\ \cancel{\text{cm}^3}}{1\ \cancel{\text{mL}}} \times \frac{(1\ \text{in})^3}{(2.54\ \cancel{\text{cm}})^3} = 34\underline{7}.835\ \text{in}^3$
• Sig. figs. and round	**Round:**	$34\underline{7}.835\ \text{in.}^3 = 348\ \text{in.}^3$
• Check	**Check:**	units are correct; number makes sense: $\text{in.}^3 << \text{L}$

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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	**Conceptual Plan:** **Relationships:**	$yd^3 \rightarrow in^3 \rightarrow cm^3$ 1 yd = 36 in. 1 in. = 2.54 cm
• Follow the conceptual plan to solve the problem	**Solution:**	$2.11\,\cancel{yd^3} \times \frac{(36\,\cancel{in})^3}{(1\,\cancel{yd})^3} \times \frac{(2.54\,cm)^3}{(1\,\cancel{in})^3}$ $= 16\underline{1}3210.75\,cm^3$
• Sig. figs. and round	**Round:**	$16\underline{1}3210.75\,cm^3$ $= 1.61 \times 10^6\,cm^3$
• Check	**Check:**	units and number make sense

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Density as a Conversion Factor

- **Can use density as a conversion factor between mass and volume!!**
  - density of H_2O = 1.0 g/mL $\therefore$ 1.0 g H_2O = 1 mL H_2O
  - density of Pb = 11.3 g/cm^3 $\therefore$ 11.3 g Pb = 1 cm^3 Pb

How much does 4.0 cm^3 of lead weigh?

$$4.0\,\cancel{\text{cm}^3\,\text{Pb}} \times \frac{11.3\,\text{g Pb}}{1\,\cancel{\text{cm}^3\,\text{Pb}}} = 4\underline{5}.2\,\text{g Pb} = 45\,\text{g Pb}$$

Example 1.10: What is the mass in kg of 173,231 L of jet fuel whose density is 0.768 g/mL?

• Sort information	**Given:** **Find:**	173,231 L density = 0.768 g/mL mass, kg
• Strategize	**Conceptual Plan:** **Relationships:**	L → mL → g → kg 1 mL = 0.768 g, 1 mL = 10^{-3} L 1 kg = 1000 g
• Follow the conceptual plan to solve the problem	**Solution:** $173{,}231\,\cancel{\mathrm{L}} \times \frac{1\,\cancel{\mathrm{mL}}}{10^{-3}\,\cancel{\mathrm{L}}} \times \frac{0.768\,\cancel{\mathrm{g}}}{1\,\cancel{\mathrm{mL}}} \times \frac{1\,\mathrm{kg}}{1000\,\cancel{\mathrm{g}}}$ $= 1.3\underline{3}041 \times 10^{5}\ \mathrm{kg}$	
• Sig. figs. and round	**Round:**	$1.3\underline{3}041 \times 10^{5} = 1.33 \times 10^{5}$ kg
• Check	**Check:**	units and number make sense

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Practice − Calculate the Following

- How much does 3.0×10^2 mL of ether weigh? ($d = 0.71$ g/mL)

- What volume does 100.0 g of marble occupy? ($d = 4.0$ g/cm^3)

Practice - How much does 3.0 x 10^2 mL of ether weigh?

Sort information	**Given:** **Find:**	3.0 x 10^2 mL density = 0.71 g/mL mass, g
Strategize	**Conceptual Plan:** **Relationships:**	$\text{mL} \rightarrow \text{g}$ 1 mL = 0.71 g
Follow the conceptual plan to solve the problem	**Solution:**	$$3.0 \times 10^{-2}\,\cancel{\text{mL}} \times \frac{0.71\text{ g}}{1\,\cancel{\text{mL}}} = 2\underline{1}3\text{ g}$$
Sig. figs. and round	**Round:**	2.1 x 10^2 g
Check	**Check:**	units are correct; number makes sense: if density < 1, mass < volume

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Practice – What volume does 100.0 g of marble occupy?

Sort information	**Given:** **Find:**	m = 100.0 g density = 4.0 g/cm^3 volume, cm^3
Strategize	**Conceptual Plan:** **Relationships:**	g $\rightarrow$ cm^3 1 cm^3 = 4.0 g
Follow the conceptual plan to solve the problem	**Solution:** $1.000\times10^{2}\,\cancel{\mathrm{g}}\times\dfrac{1\,\mathrm{cm}^3}{4.0\,\cancel{\mathrm{g}}}=2\underline{5}\,\mathrm{cm}^3$	
Sig. figs. and round	**Round:**	25 cm^3
Check	**Check:**	units are correct; number makes sense: if density > 1, mass > volume

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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 x 10^5 atoms per second for a year. How many would you count?

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

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

1 hour = 6 x $10^1 \approx 10^2$ 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

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	**Conceptual Plan:** **Relationships:**	$l, r \rightarrow V$ $m, V \rightarrow d$ $V = \pi\, r^2\, l$ $d = m/V$
Follow the conceptual plan to solve the problem	**Solution:**	$V = \pi\,(0.55\text{ cm})^2\,(1.94\text{ cm})$ $V = 1.\underline{8}436\text{ cm}^3$
Sig. figs. and round		$d = \dfrac{8.3\text{ g}}{1.\underline{8}436\text{ cm}^3} = 4.\underline{5}0206\text{ g/cm}^3 = 4.5\text{ g/cm}^3$
Check	**Check:**	units and number make sense

Practice – What is the mass in kilograms of a cube of lead that measures 0.12 m on each side? (d_{Pb} = 11.3 g/cm^3)

Practice – What is the mass in kilograms of a cube of lead that measures 0.12 m on each side?

Sort information	**Given:** **Find:**	$l = 0.12$ m, $d = 11.3$ g/cm^3 mass, kg
Strategize	**Conceptual Plan:**	$l \rightarrow V$ $m^3 \rightarrow cm^3 \rightarrow g \rightarrow kg$
	Relationships:	$V = l^3$, 11.3 g = 1 cm^3, 1 cm = 10^{-2} m, 1 kg = 10^3 g
Follow the conceptual plan to solve the problem	**Solution:**	$V = (0.12\text{ m})^3$ $V = 1.\underline{7}28 \times 10^{-3}\text{ m}^3$
Sig. figs. and round		$1.\underline{7}28\times10^{-3}\,\cancel{\text{m}}^3\times\left(\frac{1\,\cancel{\text{cm}}}{10^{-2}\,\cancel{\text{m}}}\right)^3\times\frac{11.3\,\cancel{\text{g}}}{1\,\cancel{\text{cm}}^3}\times\frac{1\text{kg}}{10^3\,\cancel{\text{g}}}$ $=1\underline{9}.526\text{ kg}=2.0\times10^1\text{ kg}$
Check	**Check:**	units and number make sense