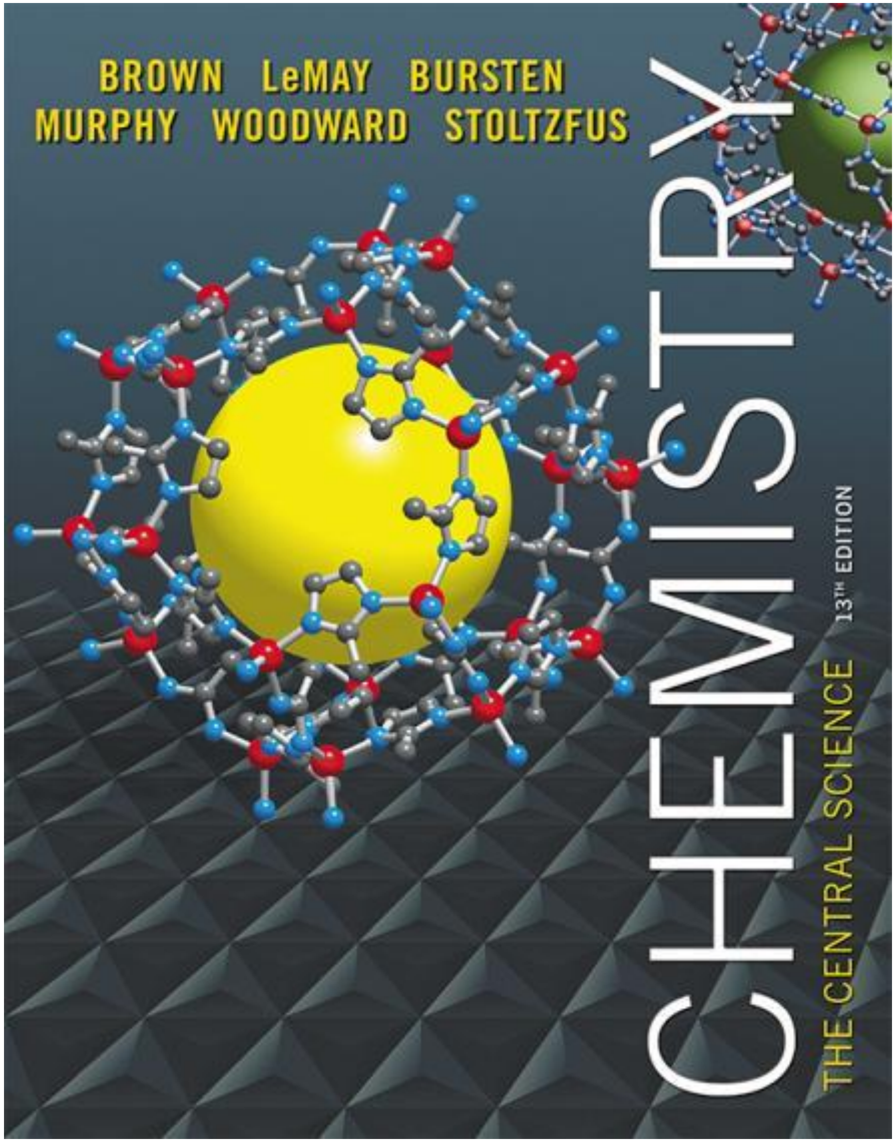

Lecture Presentation

Chapter 1

Introduction: Matter and Measurement

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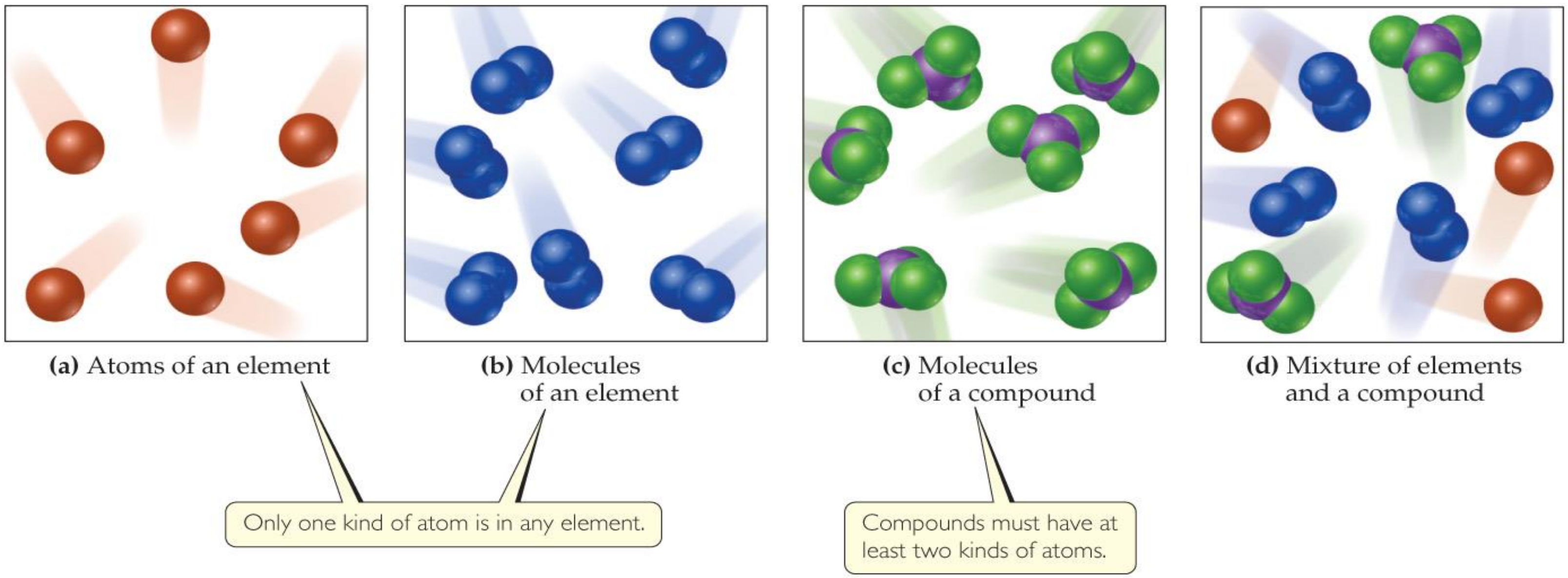
Chemistry

- Chemistry is the study of the **properties** and **behavior** of **matter**.
- **Matter** is anything that has mass and takes up space.



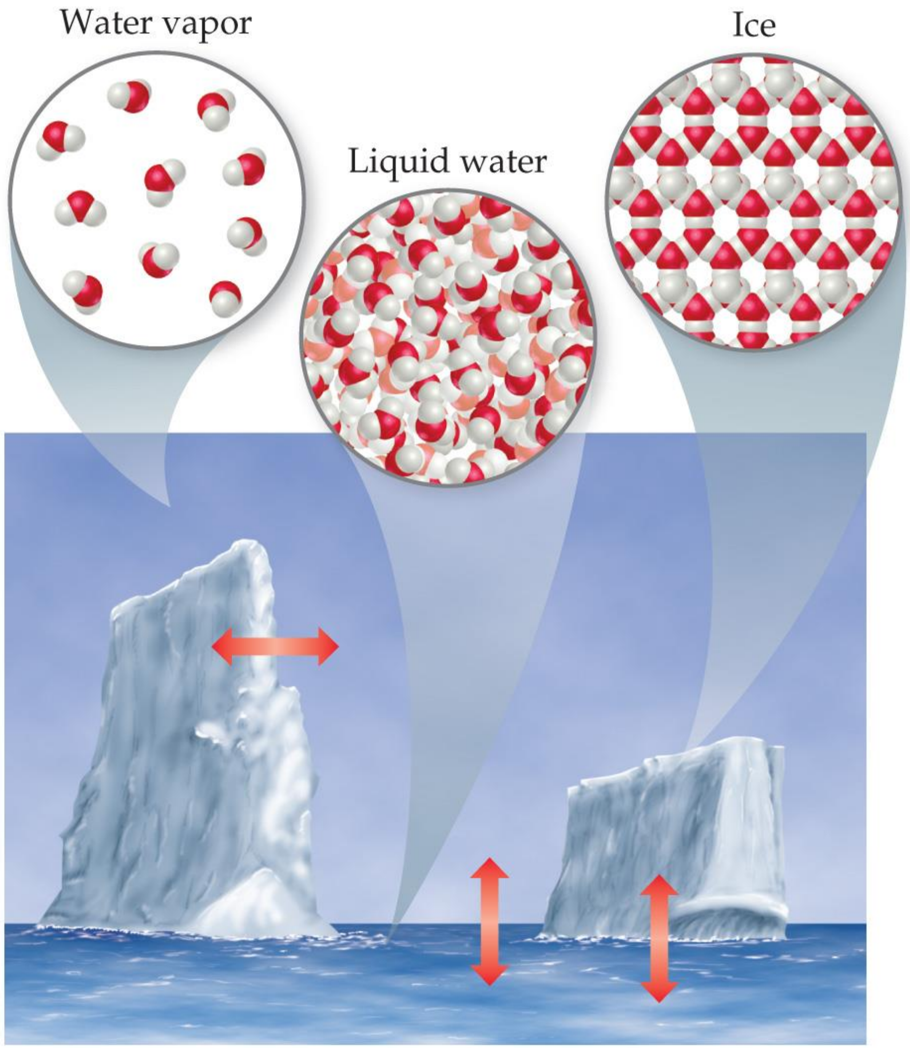
Note: Balls of different colors are used to represent atoms of different elements. Attached balls represent connections between atoms that are seen in nature. These groups of atoms are called **molecules**.

Matter

- **Atoms** are the building blocks of matter.
-
- Each **element** is made of a unique kind of atom. Can be monatomic, diatomic or polyatomic
- Ex: Ne, O_2, O_3
- An **element** is a substance which can *not* be decomposed to simpler substances.

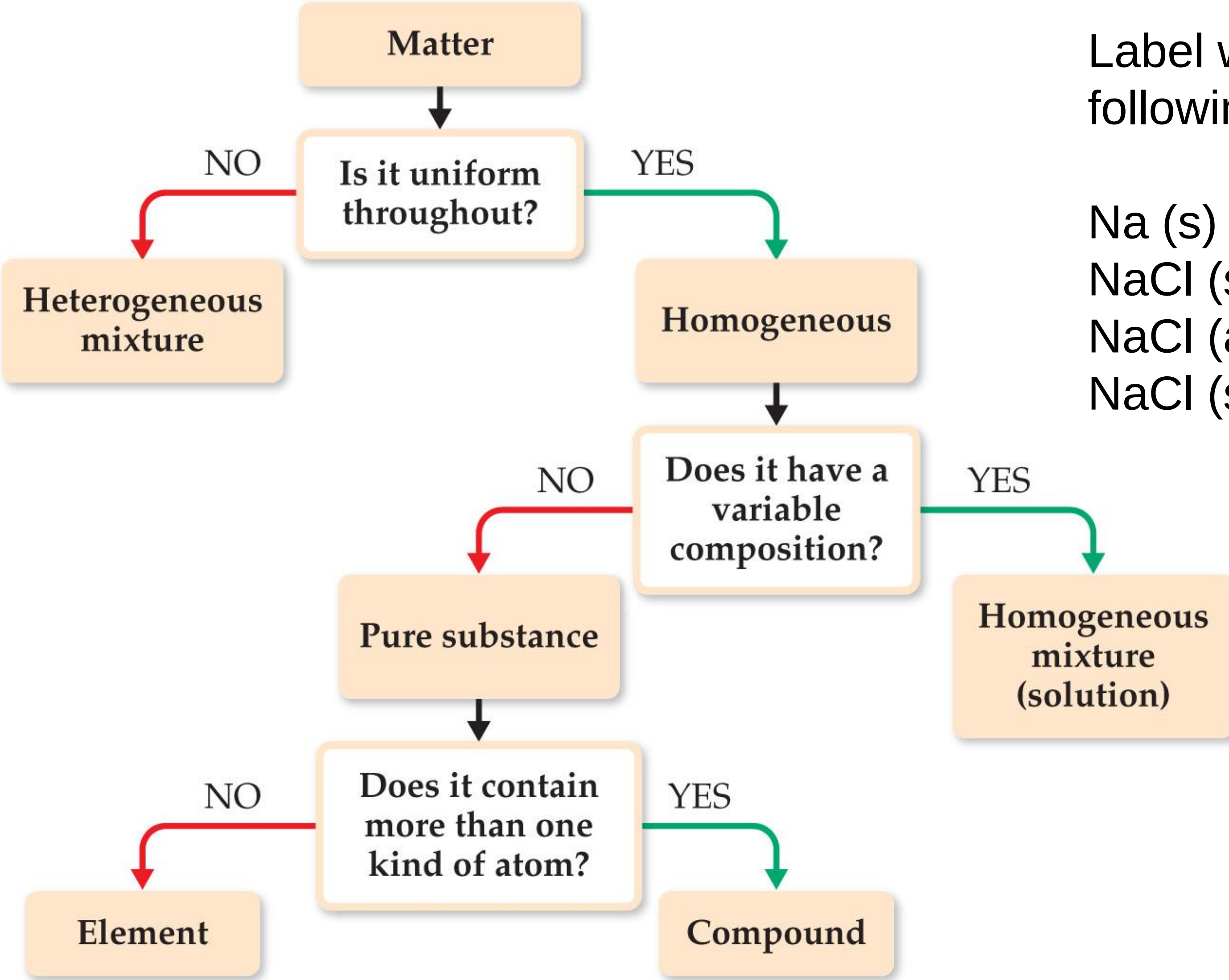
- A **compound** is made of two or more different kinds of elements. Can be ionic or molecular.
- Ex: NaCl, CO_2
- A **compound** is a substance which *can* be decomposed to simpler substances.

States of Matter



- The three states of matter are
  1) solid.
  2) liquid.
  3) gas.
- In this figure, those states are ice, liquid water, and water vapor. Substances that are liquid or solid at room temperature are called vapor when in gaseous form

Classification of Matter Based on Composition



Label where the following would go.

Na (s)
NaCl (s)
NaCl (aq)
NaCl (s) and SiO_2 (sand)

Matter And Measurement

Compounds and Composition

- Compounds have a definite composition. That means that the relative number of atoms of each element that makes up the compound is the same in any sample. They are pure substances (elements are pure as well).
- This is **The Law of Constant Composition** (or **The Law of Definite Proportions**).
- Ex: water is 2:1, carbon dioxide is 1:2

- Remember compounds have their own set of properties that are different from their component elements

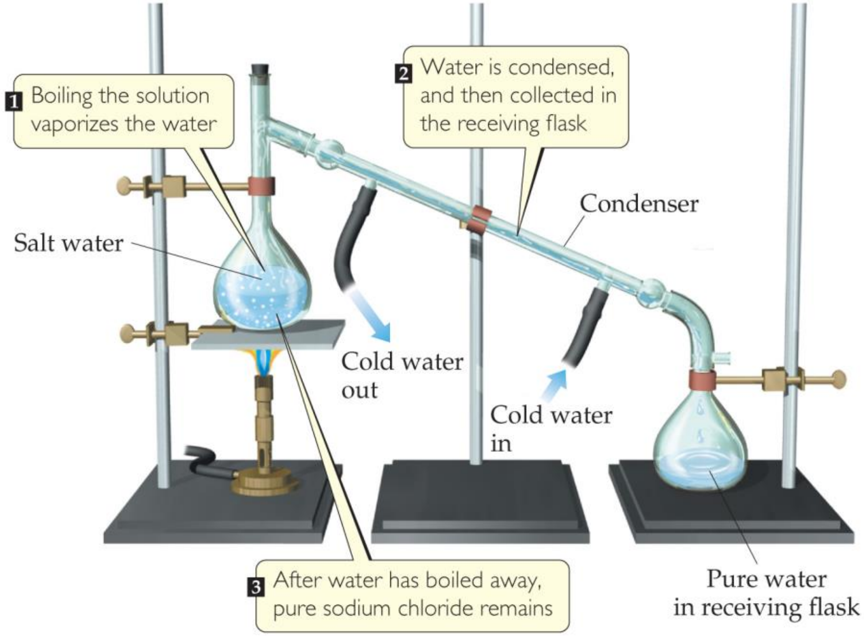
Classification of Matter—Mixtures

- **Mixtures** exhibit the properties of the substances that make them up. They keep the properties of the substances that make them up.
- They are NOT pure substances
- Mixtures can vary in composition throughout a sample (**heterogeneous**) or can have the same composition throughout the sample (**homogeneous**).
- Another name for a homogeneous mixture is **solution**.
  - When the solvent is water, it is an aqueous (aq) solution
- They can be separated by PHYSICAL means based on physical properties of the components of the mixture. Some methods used are

  Filtration, distillation, chromatography

Filtration

Distillation

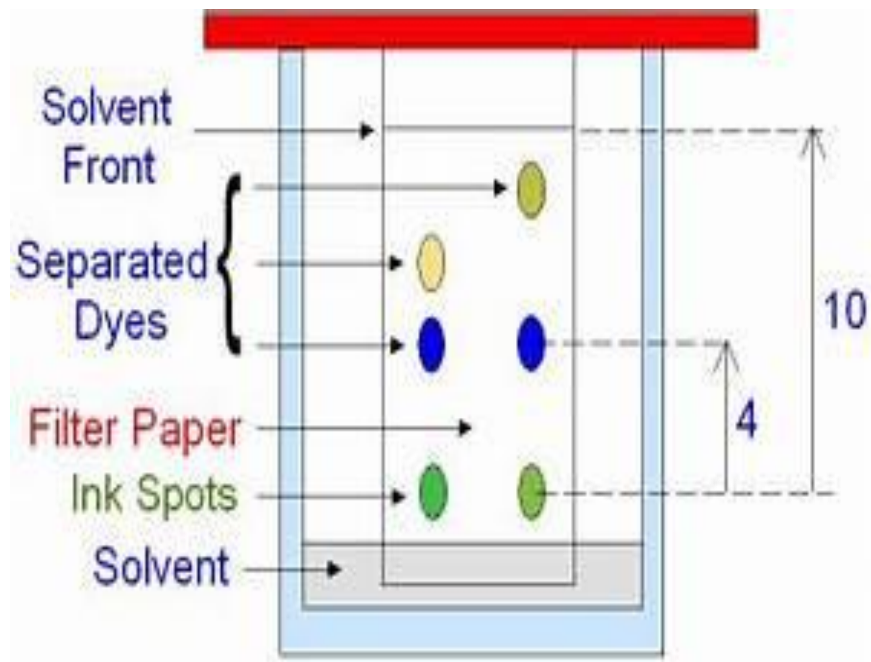


In filtration, solid substances are separated from liquids and solutions based on particle size. Does not work for homogeneous mixtures

Distillation uses differences in the boiling points of substances to separate a homogeneous mixture into its components.

Chromatography

- This technique separates substances on the basis of differences in the ability of substances to adhere to the solid surface, in this case, dyes to paper. Dyes that are more soluble will travel farther up the paper.   Remember substances that soluble in each other have the same polarity. Polar dissolves polar and nonpolar dissolves nonpolar



Types of Properties

Physical Properties can be observed <u>without</u> changing a substance into another substance.

Ex. boiling point, density, mass, or volume, color, shape

Chemical Properties can *only* be observed when a substance is changed into another substance.

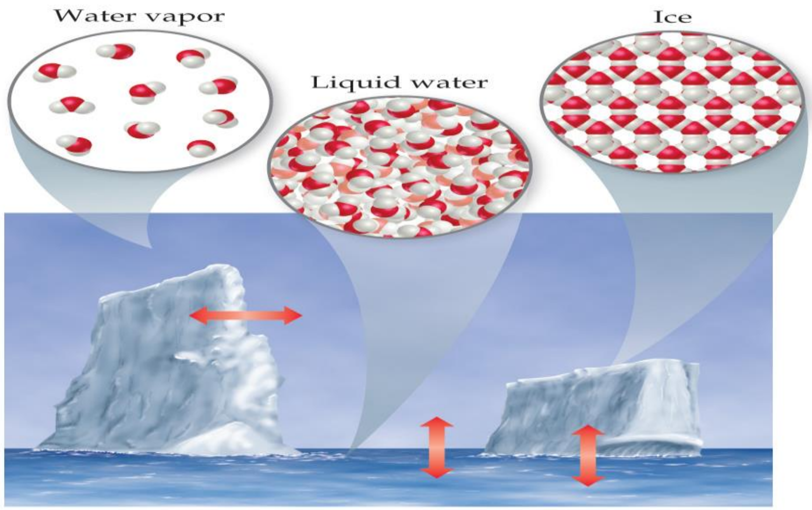
Ex. flammability, corrosiveness, or reactivity with acid.

- **Intensive Properties** are independent of the amount of the substance that is present. Can be used to identify a substance. Ex. density, boiling point, or color.
- **Extensive Properties** depend upon the amount of the substance present. Ex. mass, volume, or energy.

Types of Changes

Physical Changes

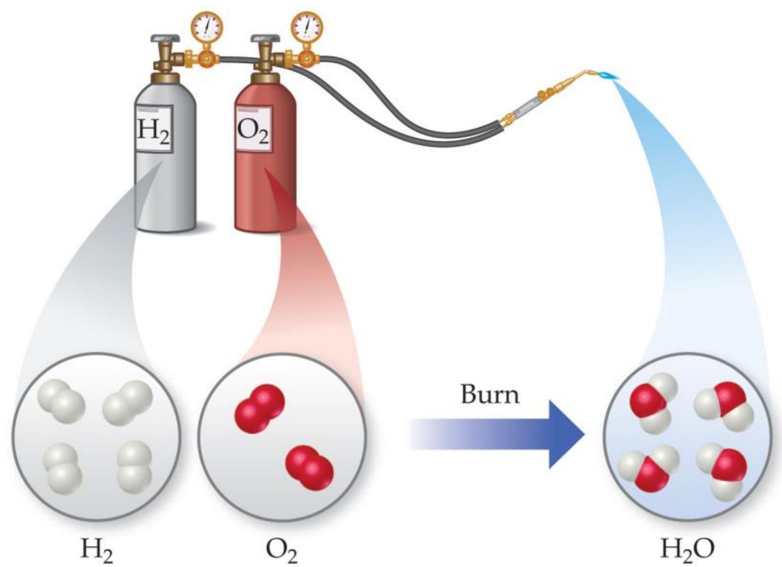
are changes in matter that do *not* change the composition of a substance. Converting between the three states of matter is a **physical change**. When ice melts or water evaporates, there are still 2 H atoms and 1 O atom in each molecule.



Chemical Changes

result in new substances with new chemical properties. Another name is a <u>chemical reaction</u>

Examples include combustion, oxidation, and decomposition.



Units of Measurement—Metric System

- Mass: gram (g)
- Length: meter (m)
- Time: second (s or sec)
- Temperature: degrees Celsius (^oC) *or* Kelvins (K)
  $^oC + 273 = K$
- Amount of a substance: mole (mol)

  $6.02 \times 10^{23} = 1$ mole
- Volume: cubic centimeter (cm^3) *or* liter (l)

  $1\ mL = 1\ cm^3$ $1\ L = 1\ dm^3$

Metric System Prefixes

Table 1.5 Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

[a]The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1\ J = 1\ kg \cdot m^2/s^2$ and $1\ W = 1\ J/s$.
[b]Greek letter mu, pronounced "mew."

Temperature

- In scientific measurements, the Celsius and Kelvin scales are most often used.
- The Celsius scale is based on the properties of water.
  - 0 °C is the freezing point of water.
  - 100 °C is the boiling point of water.
- The kelvin is the SI unit of temperature.
  - It is based on the properties of gases.
  - There are no negative Kelvin temperatures.
  - The lowest possible temperature is called absolute zero (0 K).
- $K = {}^{\circ}C + 273.15$

Density

- Density is a physical property of a substance.
- It has units that are derived from the units for mass and volume.
- The most common units are g/mL or g/cm^3.
- $D = m/V$

Sample Exercise 1.4 Determining Density and Using Density to Determine Volume or Mass

(a) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3.

(b) Calculate the volume of 65.0 g of liquid methanol (wood alcohol) if its density is 0.791 g/mL.

(c) What is the mass in grams of a cube of gold (density = 19.32 g/cm^3) if the length of the cube is 2.00 cm?

Solution

(a) We are given mass and volume, so Equation 1.3 yields

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.00 \times 10^2 \text{ g}}{7.36 \text{ cm}^3} = 13.6 \text{ g/cm}^3$$

(b) Solving Equation 1.3 for volume and then using the given mass and density gives

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{65.0 \text{ g}}{0.791 \text{ g/mL}} = 82.2 \text{ mL}$$

(c) We can calculate the mass from the volume of the cube and its density. The volume of a cube is given
by its length cubed:

$$\text{Volume} = (2.00 \text{ cm})^3 = (2.00)^3 \text{ cm}^3 = 8.00 \text{ cm}^3$$

Solving Equation 1.3 for mass and substituting the volume and density of the cube, we have

$$\text{Mass} = \text{volume} \times \text{density} = (8.00 \text{ cm}^3)(19.32 \text{ g/cm}^3) = 155 \text{ g}$$

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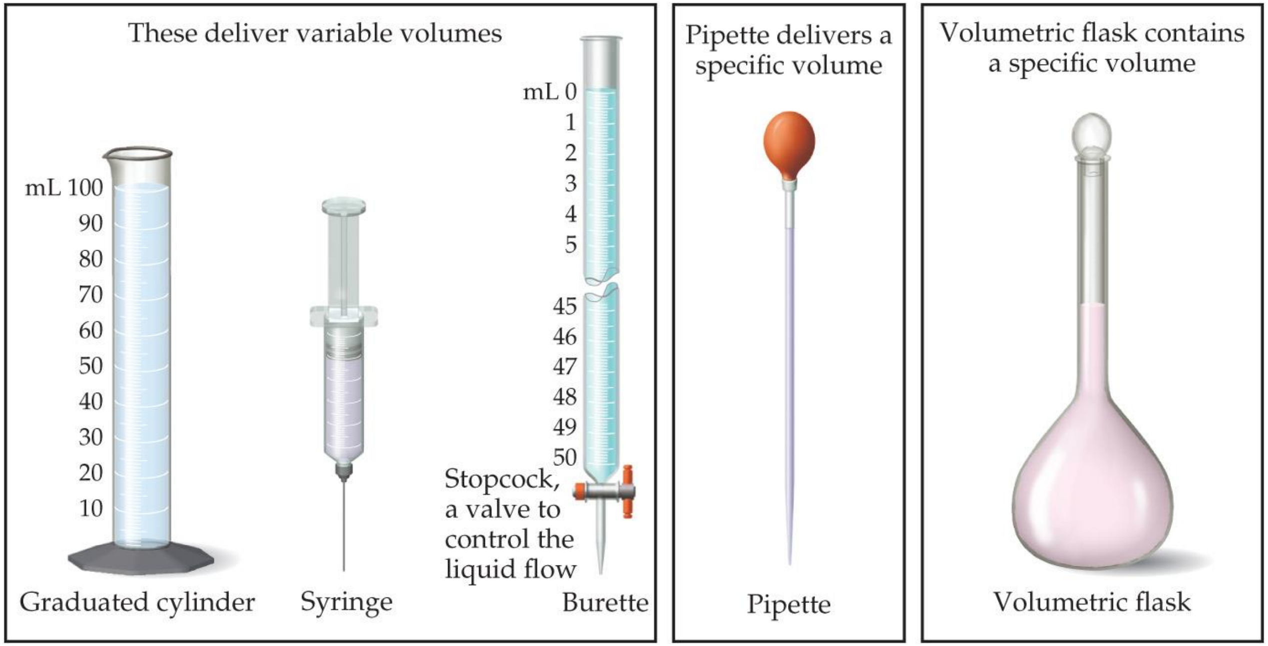
Numbers Encountered in Science

- **Exact** numbers are counted or given by definition. For example, there are 12 eggs in 1 dozen.
- **Inexact** (or **measured**) numbers depend on how they were determined. Scientific instruments have limitations. Some balances measure to ±0.01 g; others measure to ±0.0001g.

Uncertainty in Measurements

- Different measuring devices have different uses and different degrees of accuracy.
- All measured numbers have some degree of inaccuracy.



Accuracy versus Precision

- **Accuracy** refers to the proximity of a measurement to the true value of a quantity.
- **Precision** refers to the proximity of several measurements to each other.

Good accuracy
Good precision

Poor accuracy
Good precision

Poor accuracy
Poor precision

Significant Figures

When rounding calculated numbers, we pay attention to significant figures so we do not overstate the accuracy of our answers.

1. All nonzero digits are significant.
2. Zeroes between two significant figures are themselves significant.
3. Zeroes at the beginning of a number are never significant.
4. Zeroes at the end of a number are significant if a decimal point is written in the number.

- When addition or subtraction is performed, answers are rounded to the least significant **decimal place**.
- When multiplication or division is performed, answers are rounded to the number of digits that corresponds to the ***least* number of significant figures** in any of the numbers used in the calculation.

Dimensional Analysis

- We use **dimensional analysis** to convert one quantity to another.
- Most commonly, dimensional analysis utilizes **conversion factors** (e.g., 1 in. = 2.54 cm).
- We can set up a ratio of comparison for the equality either 1 in/2.54 cm *or* 2.54 cm/1 in.
- We use the ratio which allows us to change units (puts the units we have in the denominator to cancel).

The average speed of a nitrogen molecule in air at 25 ° C is 515 m/s. Convert this speed to miles per hour

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Solution

To go from the given units, m/s, to the desired units, mi/hr, we must convert meters to miles and seconds to hours. From our knowledge of SI prefixes we know that 1 km = 10^3 m. From the relationships given on the back inside cover of the book, we find that 1 mi = 1.6093 km.

Thus, we can convert m to km and then convert km to mi. From our knowledge of time we know that 60 s = 1 min and 60 min = 1 hr. Thus, we can convert s to min and then convert min to hr. The overall process is

Applying first the conversions for distance and then those for time, we can set up one long equation in which unwanted units are canceled:

$$\text{Speed in mi/hr} = \left(515\frac{\cancel{\text{m}}}{\cancel{\text{s}}}\right)\left(\frac{1\ \cancel{\text{km}}}{10^3\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{mi}}{1.6093\ \cancel{\text{km}}}\right)\left(\frac{60\ \cancel{\text{s}}}{1\ \cancel{\text{min}}}\right)\left(\frac{60\ \cancel{\text{min}}}{1\ \text{hr}}\right)$$

$$= 1.15 \times 10^3\ \text{mi/hr}$$