Chapter 1

Introduction:
Matter and Measurement

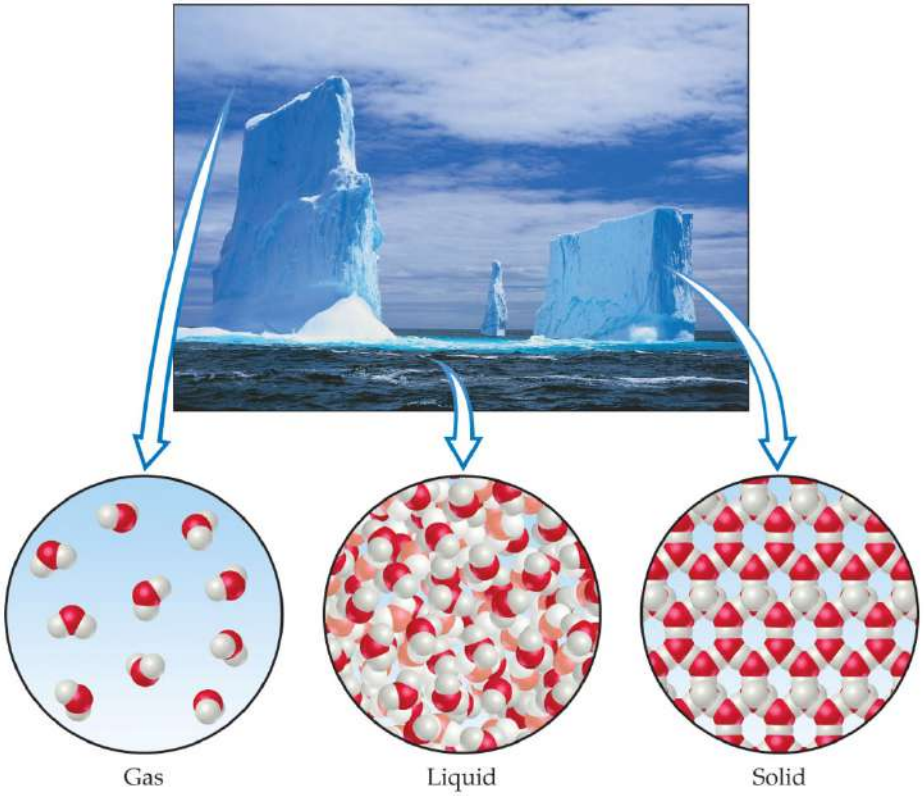
1.1 The Study of Chemistry

You should be able to:

- Distinguish between physical and chemical properties and also between physical and chemical changes.
- Differentiate between the 3 states of matter.
- Distinguish between elements, compounds, and mixtures.
- Know the symbols of the elements.
- List the basic SI and metric units and the commonly used prefixes.
- Work with numbers in exponential notation.

- Determine the number of significant figures in a measured quantity.
- Express the result of a calculation with the proper number of significant figures.
- Convert temperatures among the Fahrenheit, Celsius, and Kelvin scales.
- Perform calculations involving density.
- Convert between units using dimensional analysis.

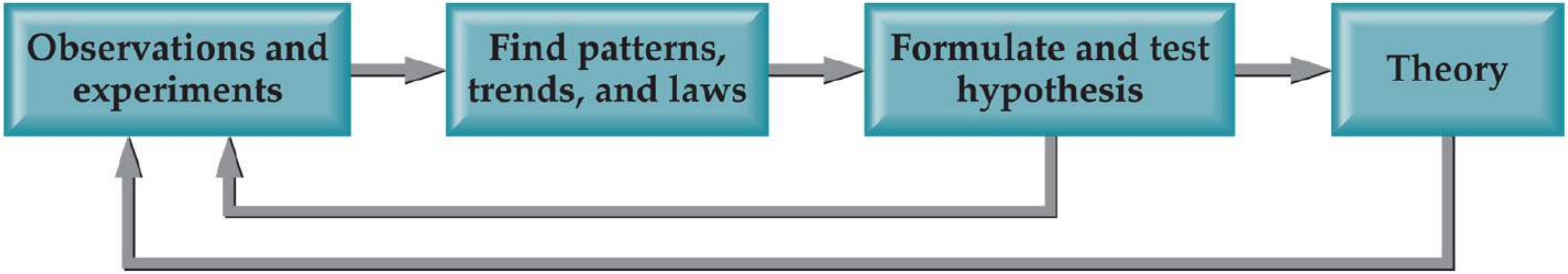
Chemistry



In this science we study matter and the changes it undergoes.

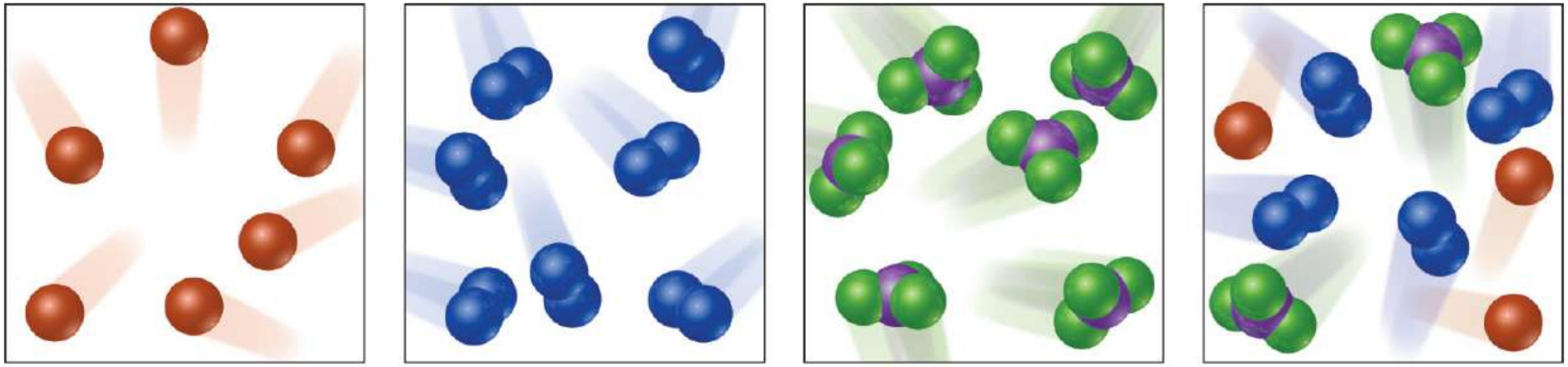
Scientific Method

The scientific method is simply a systematic approach to solving problems.



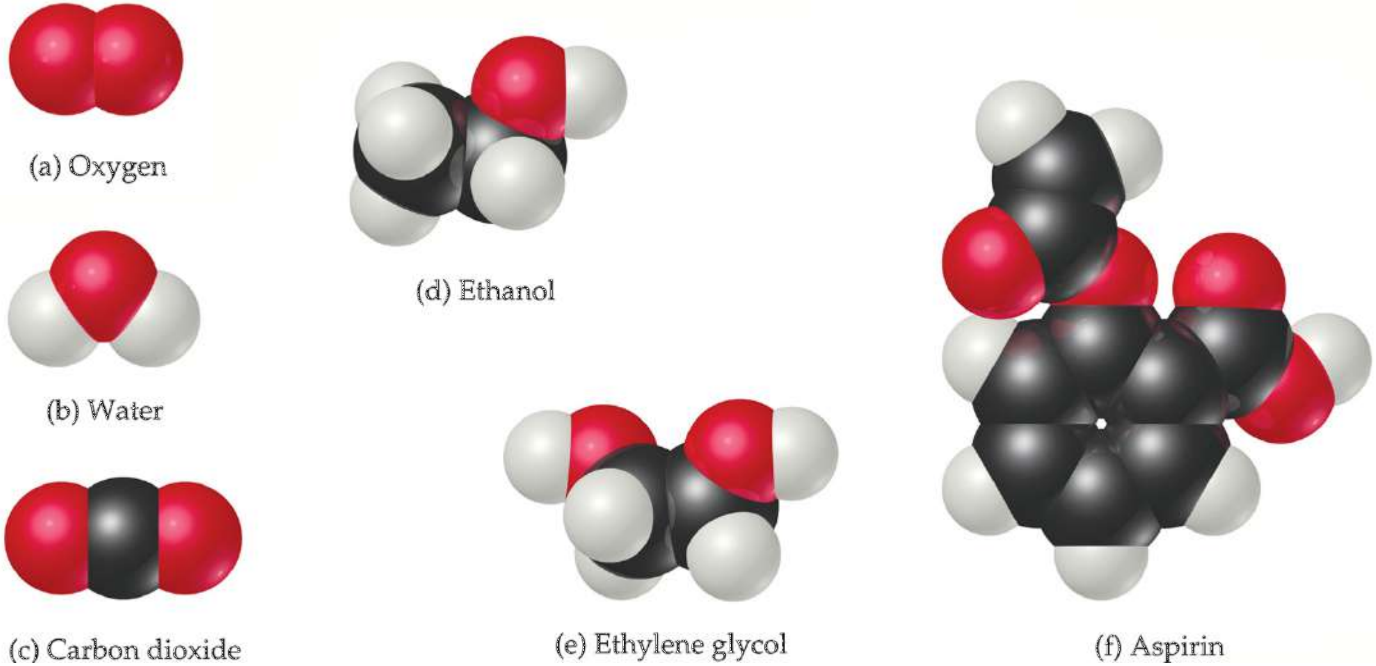
Matter

We define matter as anything that has mass and takes up space.



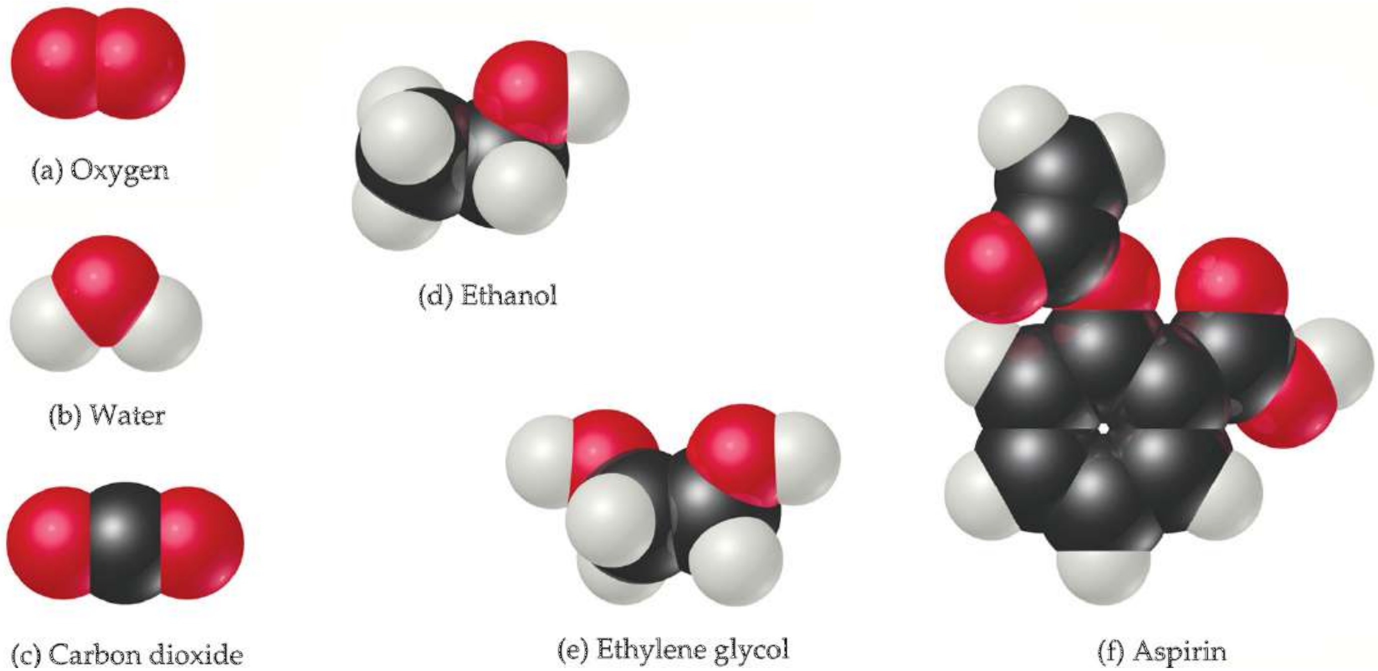
(a) Atoms of an element (b) Molecules of an element (c) Molecules of a compound (d) Mixture of elements and a compound

Matter



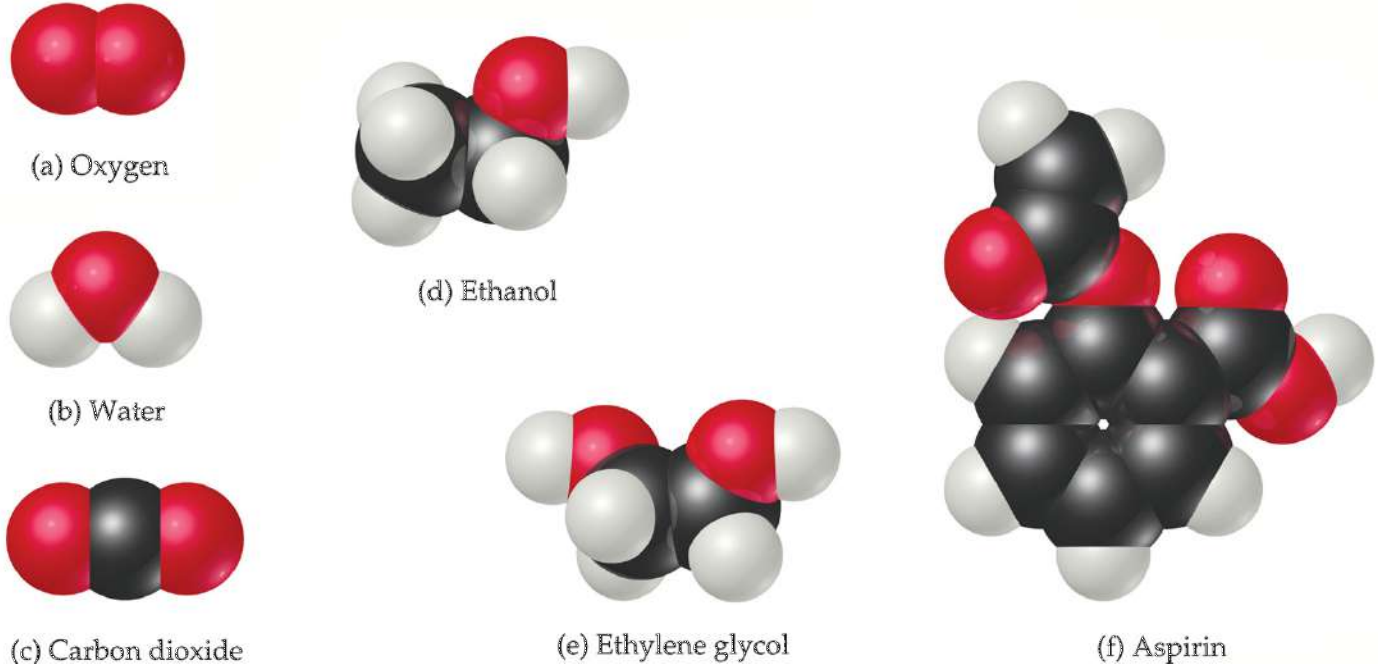
- Atoms are the building blocks of matter.

Matter



- Atoms are the building blocks of matter.
- Each element is made of the same kind of atom.

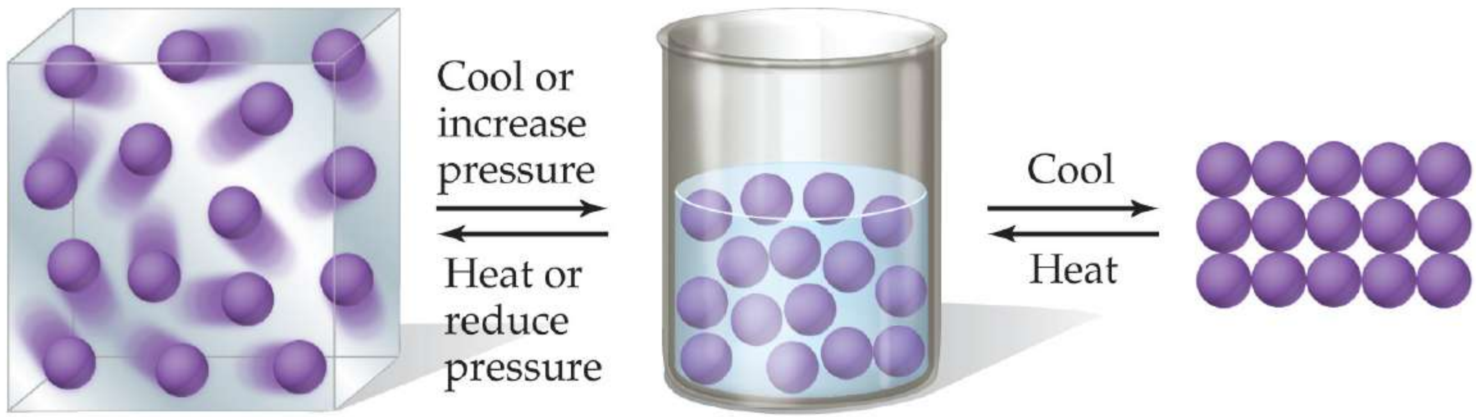
Matter



- Atoms are the building blocks of matter.
- Each element is made of the same kind of atom.
- A compound is made of two or more different kinds of elements.

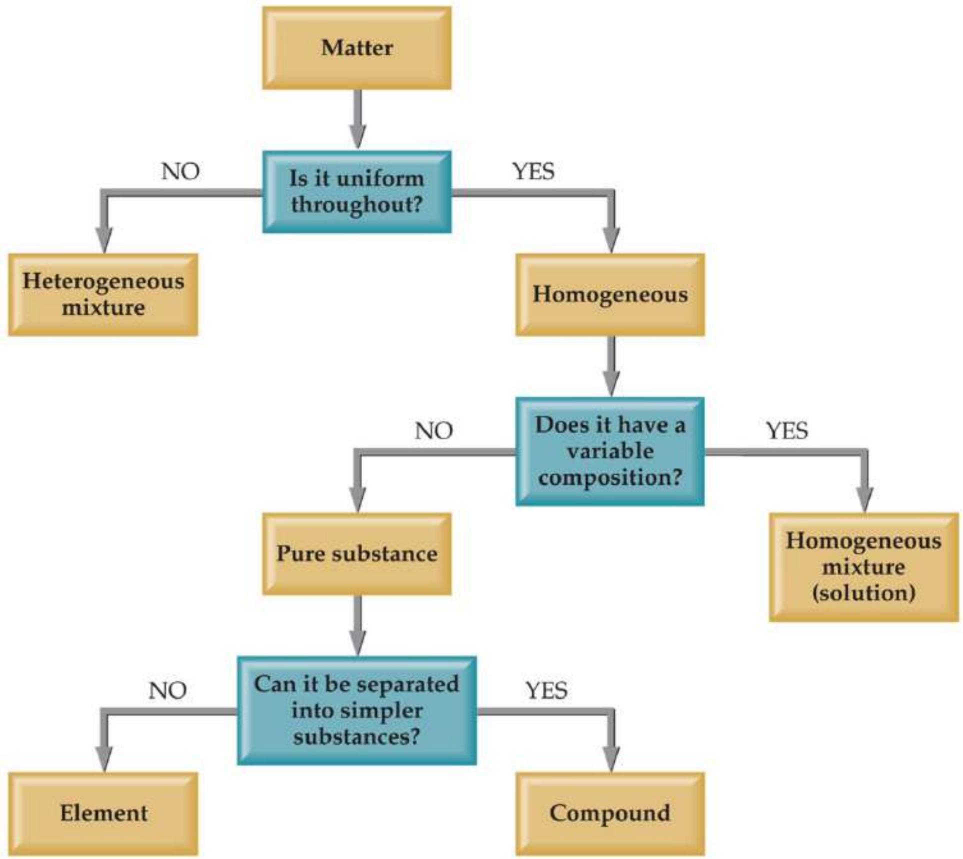
1.2 Classification of Matter

States of Matter



Gas	Liquid	Crystalline solid
Total disorder; much empty space; particles have complete freedom of motion; particles far apart	Disorder; particles or clusters of particles are free to move relative to each other; particles close together	Ordered arrangement; particles are essentially in fixed positions; particles close together

Classification of Matter



Law of constant composition

- AKA : law of definite proportions
- Joseph Proust 1800
- Elemental composition of a compound is always the same

Sample Exercise 1.1 Distinguishing Among Elements, Compounds, and Mixtures

"White gold," used in jewelry, contains gold and another "white" metal such as palladium. Two different samples of white gold differ in the relative amounts of gold and palladium that they contain. Both samples are uniform in composition throughout. Without knowing any more about the materials, classify white gold.

Answer

Because the material is uniform throughout, it is homogeneous. Because its composition differs for the two samples, it cannot be a compound. Instead, it must be a homogeneous mixture.

Practice Exercise

Aspirin is composed of 60.0% carbon, 4.5% hydrogen, and 35.5% oxygen by mass, regardless of its source. Classify aspirin.

1.3 Properties of Matter

Types of Properties

- Physical Properties…
  - Can be observed without changing a substance into another substance.
    - Boiling point, density, mass, volume, etc.
- Chemical Properties…
  - Can *only* be observed when a substance is changed into another substance.
    - Flammability, corrosiveness, reactivity with acid, etc.

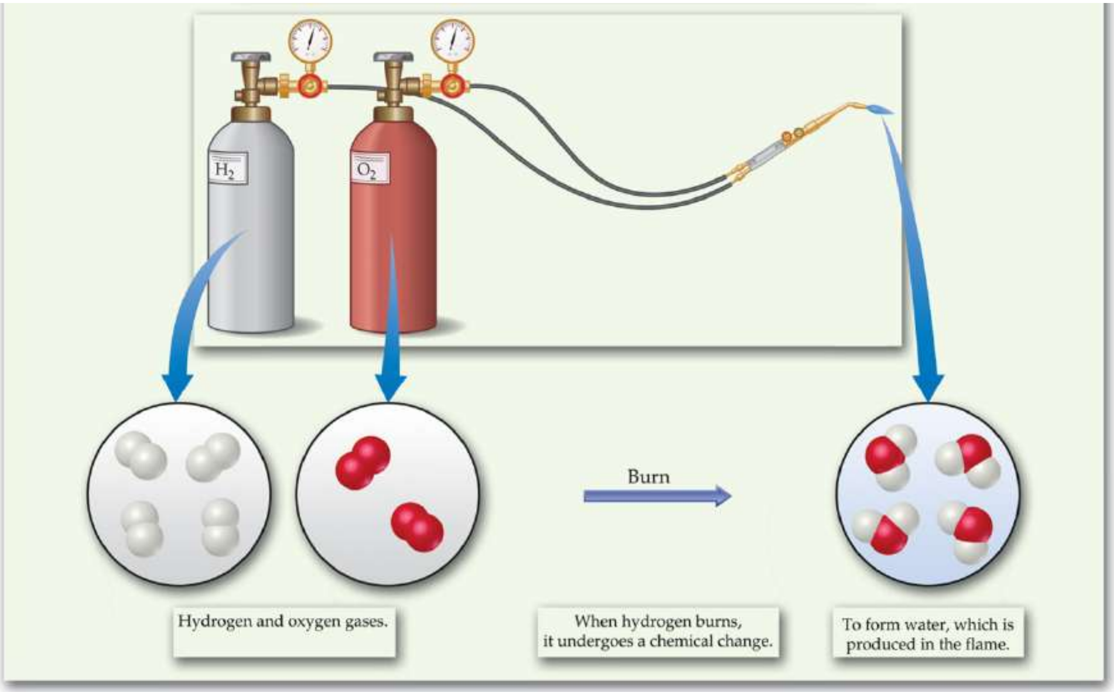
Types of Properties

- Intensive Properties…
  - Are independent of the amount of the substance that is present.
    - Density, boiling point, color, etc.
- Extensive Properties…
  - Depend upon the amount of the substance present.
    - Mass, volume, energy, etc.

Types of Changes

- Physical Changes
  - These are changes in matter that do not change the composition of a substance.
    - Changes of state, temperature, volume, etc.
- Chemical Changes
  - Chemical changes result in new substances.
    - Combustion, oxidation, decomposition, etc.

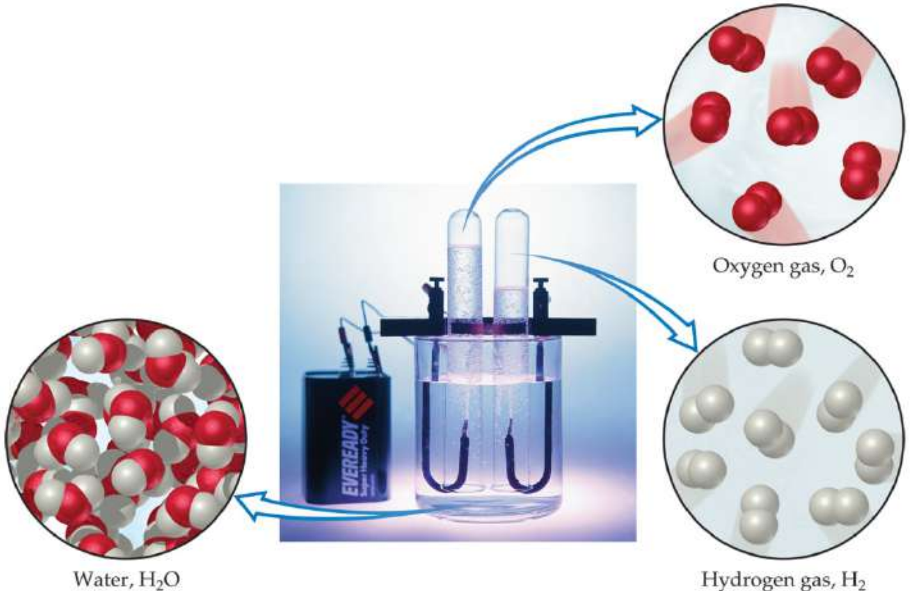
Chemical Reactions



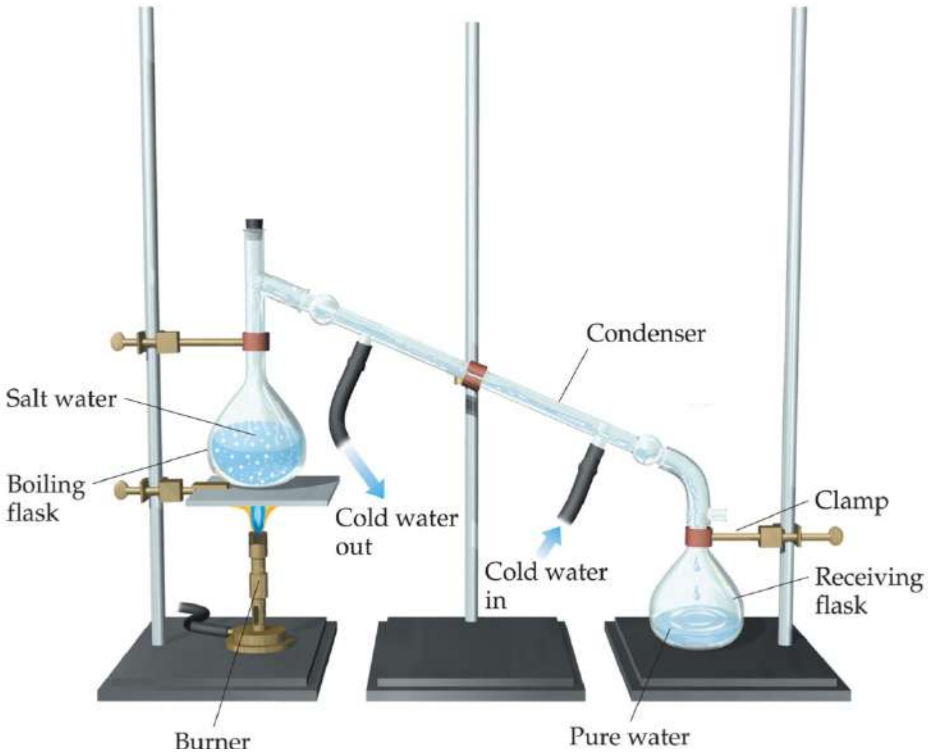
In the course of a chemical reaction, the reacting substances are converted to new substances.

Compounds

Compounds can be broken down into more elemental particles.

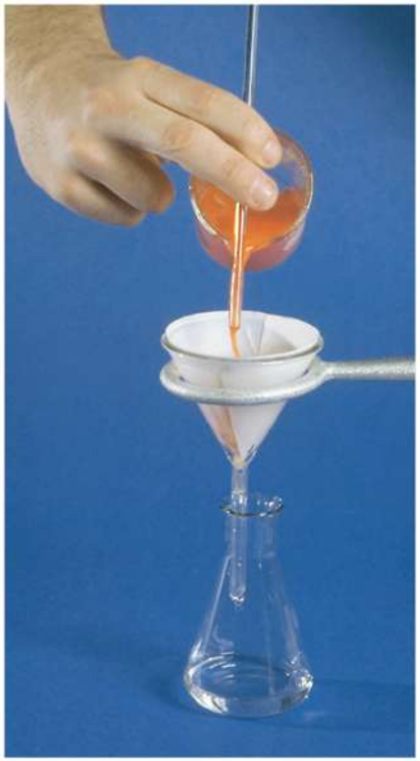


Separation of Mixtures: Distillation



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

Filtration



In filtration solid substances are separated from liquids and solutions.

Chromatography

- This technique separates substances on the basis of differences in solubility in a solvent.

1.4 The Nature of Energy

Energy

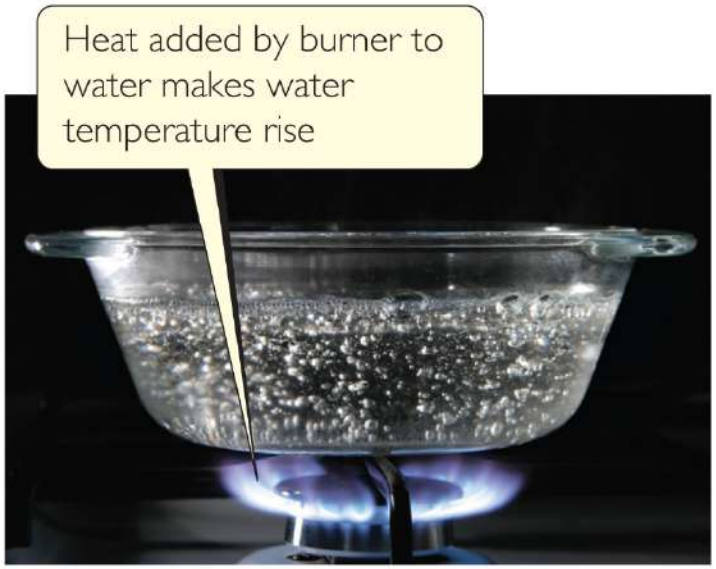
- **Energy** is the ability to do work or transfer heat.
- Energy used to cause an object that has mass to move is called **work**.
  - Energy used to cause the temperature of an object to rise is called **heat**.

Definitions: Work

- Energy used to move an object over some distance is **work**:
- $w = F \times d$

where w is work, F is the force, and d is the distance over which the force is exerted.

Heat

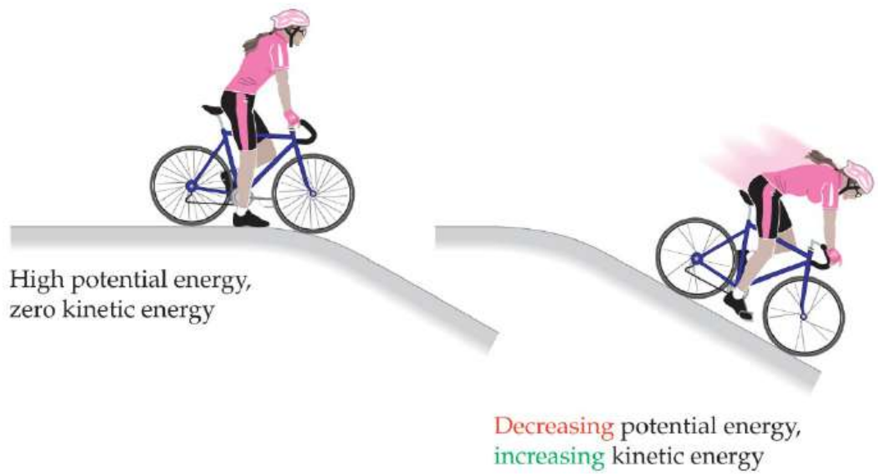


- Energy can also be transferred as heat.
- Heat flows from warmer objects to cooler objects.

Kinetic Energy

Kinetic energy is energy an object possesses by virtue of its motion:

$$E_k = \frac{1}{2}mv^2$$



Potential Energy

- **Potential energy** is energy an object possesses by virtue of its position or chemical composition.

Describing and Calculating Energy Changes: Sample exercise

- A bowler lifts a 5.4 kg (12-lb) bowling ball from ground level to a height of 1.6 m and then drops the ball back to the ground.
  - A) What happens to the potential energy of the ball as it is raised?
  - B) How much work is used to raise the ball?
  - C) If all of the work in (B) is converted to kinetic energy, what is the speed of the ball just before it hits the ground?

Practice Exercise 1

Which of the following objects has the greatest kinetic energy?

(a) a 500-kg motorcycle moving at 100 km / h

(b) a 1,000-kg car moving at 50 km / h

(c) a 1,500-kg car moving at 30 km / h

(d) a 5,000-kg truck moving at 10 km / h

(e) a 10,000-kg truck moving at 5 km / h

Practice Exercise 2

- What is the kinetic energy of:
  - A) an Ar atom moving with a speed of 650 m/s
  - B) a mole of Ar atoms moving with a speed of 650 m/s (1 amu = 1.66×10^{-27} kg)

Units of Energy

- The SI unit of energy is the **joule (*J*)**:

$$1\ \mathrm{J} = 1\ \frac{\mathrm{kg\ m^2}}{\mathrm{s^2}}$$

- An older, non-SI unit is still in widespread use, the **calorie (*cal*)**:

1 cal = 4.184 J

(Note: this is *not* the same as the calorie of foods; the food calorie is 1 kcal!)

Sample Exercise 1.5

- A standard propane (C_3H_8) tank used in an outdoor grill holds approximately 9.0 kg of propane. When the grill is operating, propane reacts with oxygen to form carbon dioxide and water. For every gram of propane that reacts with oxygen, 46 kJ of energy is released as heat.

a) How much energy (in J) is released if all the propane reacts?
b) As propane reacts, does the potential energy stored in chemical bonds increase or decrease?
c) If you were to store an equivalent amount of potential energy by pumping water to an elevation 75 m above the ground, what mass of water would be needed? ($F = m \times g$ where $g = 9.8\ m/s^2$)

1.5 Units of Measurement

SI Units

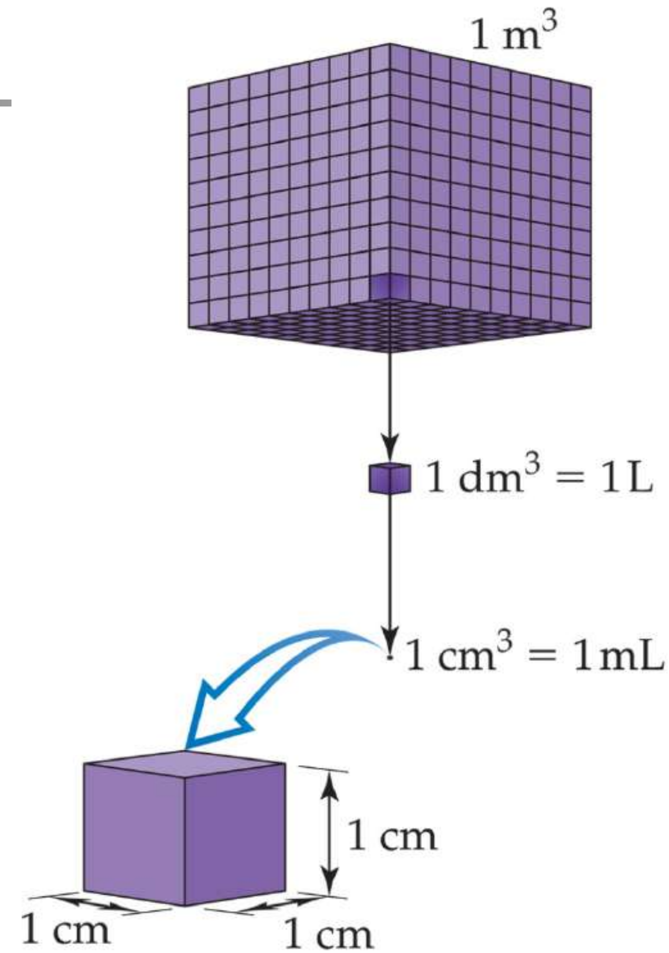
Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

[a]The abbreviation sec is frequently used.

- *Système International d'Unités*
- A different base unit is used for each quantity.

Volume

- The most commonly used metric units for volume are the liter (L) and the milliliter (mL).
  - A liter is a cube 1 dm long on each side.
  - A milliliter is a cube 1 cm long on each side.



Density

Density is an intensive physical property of a substance.

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

Sample Exercise 1.4

- 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 methanol 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?

Density Practice

- 1. Calculate the density of a 374.5 g sample of copper if it has a volume of 41.8 cm^3.
- 2. A student needs 15.0 g of ethanol for an experiment. If the density of ethanol is 0.789 g/mL, how many mililiters of ethanol are needed?
- 3. What is the mass of 25.0 mL of mercury if the density is 13.6 g/mL?

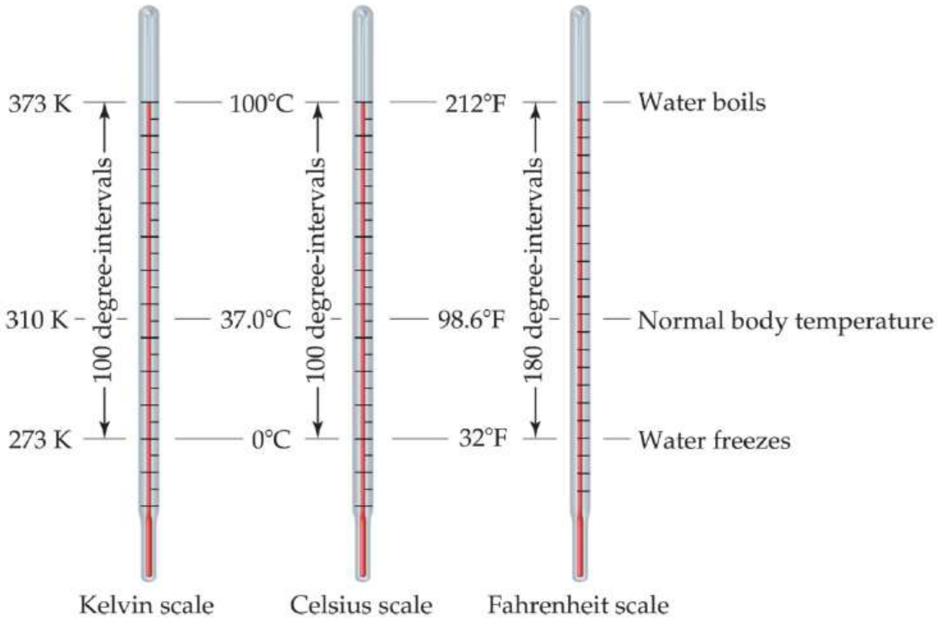
Metric System

Prefixes convert the base units into units that are appropriate for the item being measured.

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) $= 1 \times 10^9$ m
Mega	M	10^6	1 megameter (Mm) $= 1 \times 10^6$ m
Kilo	k	10^3	1 kilometer (km) $= 1 \times 10^3$ m
Deci	d	10^{-1}	1 decimeter (dm) = 0.1 m
Centi	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli	m	10^{-3}	1 millimeter (mm) = 0.001 m
Micro	μ[a]	10^{-6}	1 micrometer (μm) $= 1 \times 10^{-6}$ m
Nano	n	10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
Pico	p	10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m
Femto	f	10^{-15}	1 femtometer (fm) $= 1 \times 10^{-15}$ m

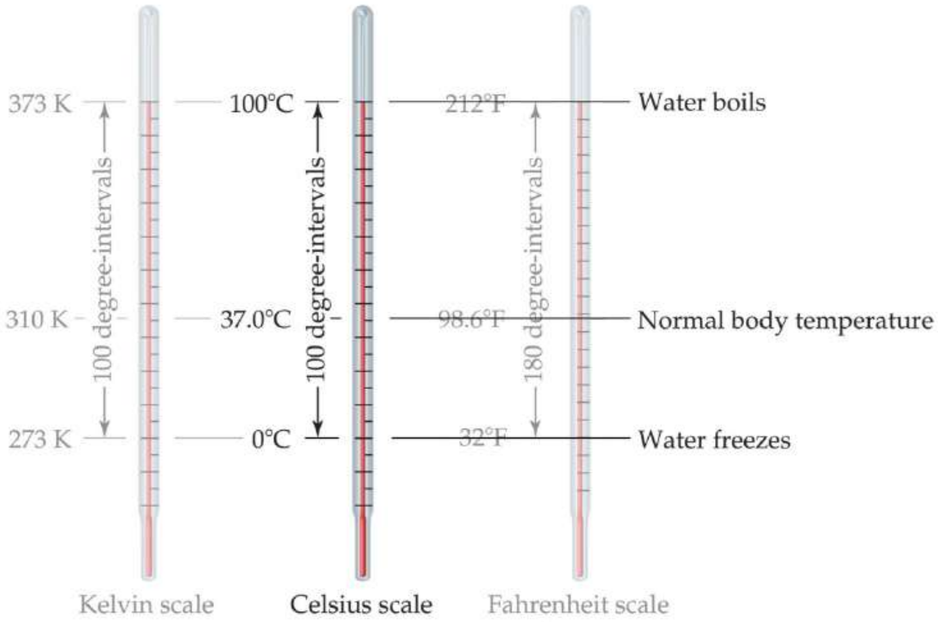
[a]This is the Greek letter mu (pronounced "mew").

Temperature



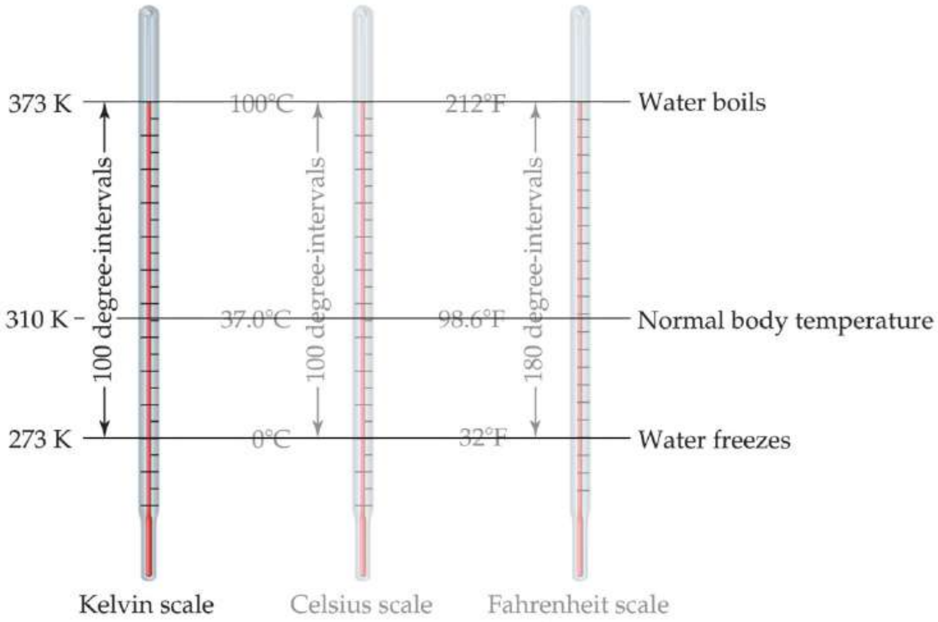
By definition temperature is a measure of the average kinetic energy of the particles in a sample.

Temperature



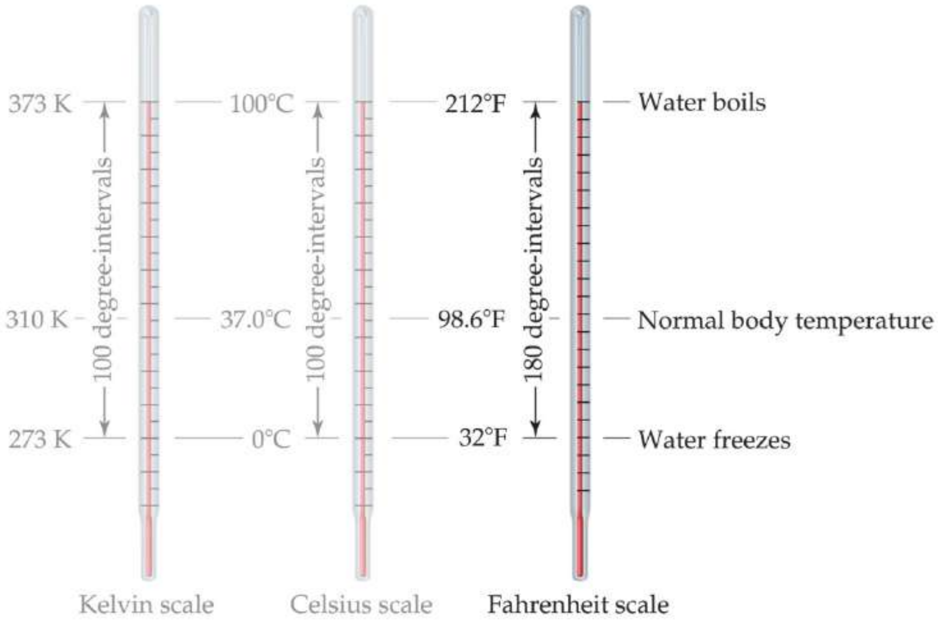
- In scientific measurements, the Celsius and Kelvin scales are most often used.
- The Celsius scale is based on the properties of water.
  - 0°C is the freezing point of water.
  - 100°C is the boiling point of water.

Temperature



- The Kelvin is the SI unit of temperature.
- It is based on the properties of gases.
- There are no negative Kelvin temperatures.
- K = °C + 273.15

Temperature



- The Fahrenheit scale is not used in scientific measurements.
- $°F = 9/5(°C) + 32$
- $°C = 5/9(°F - 32)$

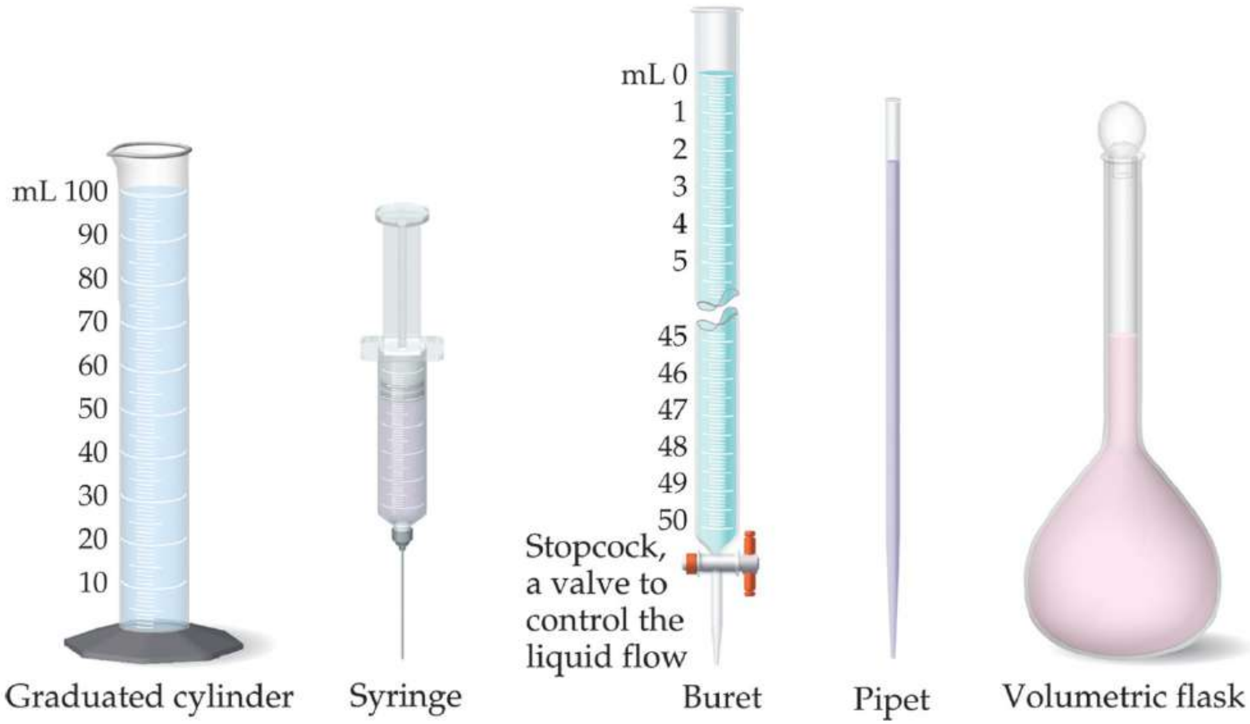
Temperature Practice

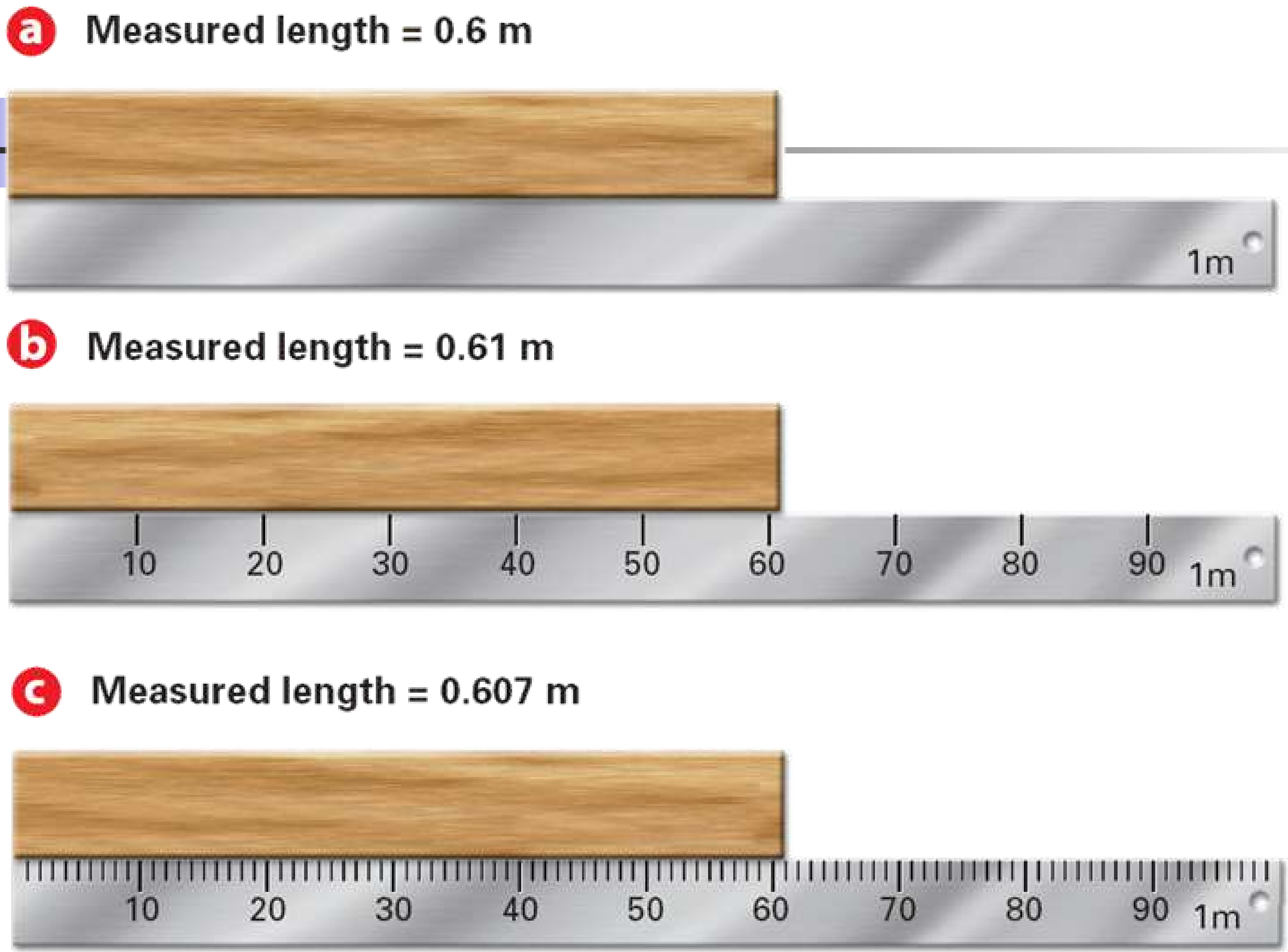
- Ethylene glycol, a major ingredient in antifreeze, freezes at -11.5ºC. What is the freezing point in Kelvin and ºF?

1.6 Uncertainty in Measurement

Uncertainty in Measurements

Different measuring devices have different uses and different degrees of accuracy.




a
Measured length = 0.6 m
1m
b
Measured length = 0.61 m
10
20
30
40
50
60
70
80
90
1m
c
Measured length = 0.607 m
10
20
30
40
50
60
70
80
90
1m

Significant Figures

- The term significant figures refers to digits that were measured.
- When rounding calculated numbers, we pay attention to significant figures so we do not overstate the accuracy of our answers.

Significant Figures

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.
5. Exact numbers are assumed to have an infinite number of sig figs

Sample Exercise 1.5

- What difference exists between the measured values 4.0 g and 4.00 g?
- A balance has a precision of ±0.001 g. A sample that has a mass of about 25 g is placed on this balance. How many significant figures should be reported for this measurement?

Significant Figure Practice

- How many significant figures are in the following:
- A) 4.003
- B) 5000
- C) 0.00135
- D) 0.1270
- E) 6.023 x 10^{23}
- F) 3.549
- G) 2.3 x 10^{4}

Significant Figures

- 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.

Operations with Significant Figures Practice

1) Calculate the volume of a box with the following dimensions: 15.5 cm, 27.3 cm and 5.4 cm. Report your answer to the correct number of significant figures.
2) Calculate: 212.2 + 26.7 + 402.09
3) It takes 10.5 s for a sprinter to run 100.00 m. Calculate the average speed in m/s.

Sample Exercise 1.8

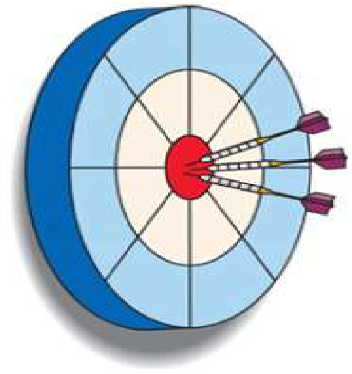
- A gas at 25ºC fills a container whose volume is 1.05×10^3 cm^3. The container plus the gas have a mass of 837.6 g. The container when emptied of all gas has a mass of 836.2 g. What is the density of the gas at 25ºC?
- To how many significant figures should the mass of the container be measured (with and without the gas) for density to be calculated to three significant figures?

Accuracy, Precision, and Error

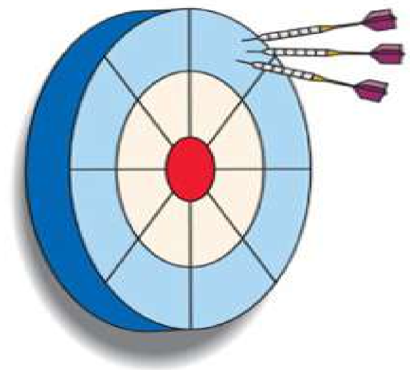
- Just because a measuring device works, you cannot assume it is accurate. The scale below has not been properly zeroed, so the reading obtained for the person's weight is inaccurate.

Accuracy versus Precision

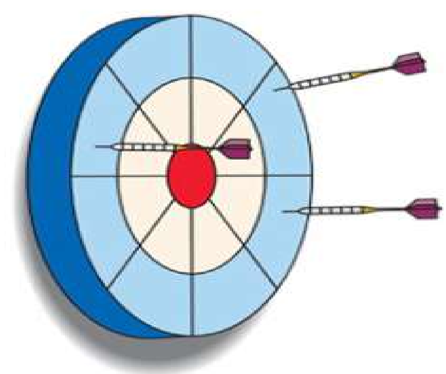
- Accuracy refers to the proximity of a measurement to the true value of a quantity.
- Precision refers to the proximity of several measurements to each other.



Good accuracy
Good precision



Poor accuracy
Good precision



Poor accuracy
Poor precision

$$\text{Error} = \text{experimental value} - \text{accepted value}$$

- Determining Error
  - The **accepted value** is the correct value.
  - The **experimental value** is the value measured in the lab.
  - The difference between the experimental value and the accepted value is called the **error**.

Accuracy, Precision, and Error

- The **percent error** is the absolute value of the error divided by the accepted value, multiplied by 100%.

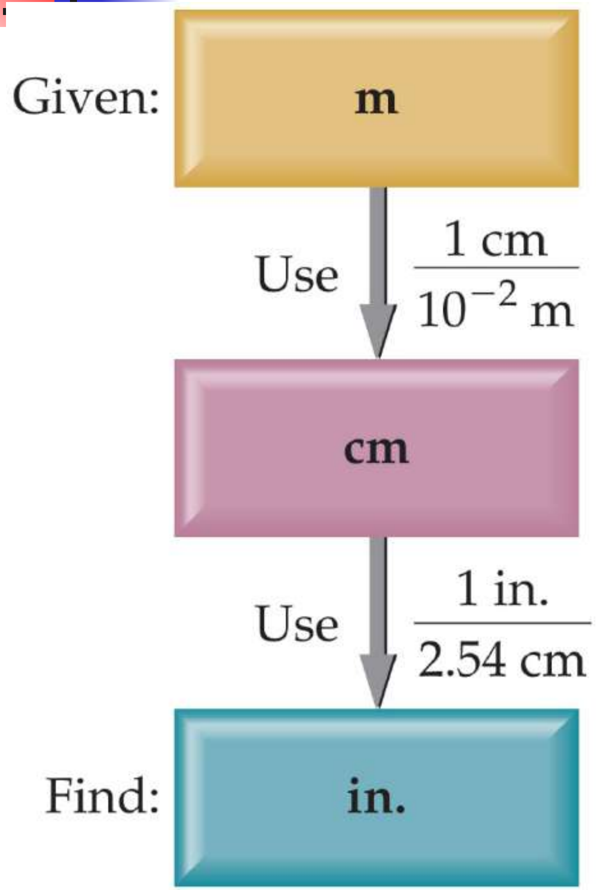
$$\text{Percent error} = \frac{|\text{error}|}{\text{accepted value}} \times 100\%$$

Practice Problem

- A technician experimentally determined the boiling point of octane to be 124.1ºC. The actual boiling point of octane is 125.7ºC. Calculate the error and percent error.

1.7 Dimensional Analysis

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)

$$\frac{1\ \text{in.}}{2.54\ \text{cm}} \quad \text{or} \quad \frac{2.54\ \text{cm}}{1\ \text{in.}}$$

Dimensional Analysis

- Use the form of the conversion factor that puts the sought-for unit in the numerator.

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

Conversion factor

Sample Exercise 1.9

- If a woman has a mass of 115 lb, what is her mass in grams?
- Determine the length in kilometers of a 500.0 mile automobile race.

Dimensional Analysis

- For example, to convert 8.00 m to inches,
  - convert m to cm
  - convert cm to in.

$$8.00\ \text{m} \times \frac{100\ \text{cm}}{1\ \text{m}} \times \frac{1\ \text{in.}}{2.54\ \text{cm}} = 315\ \text{in.}$$

Sample Exercise 1.10

- The average speed of a nitrogen molecule in air at 25ºC is 515 m/s. Convert this speed into miles per hour.

Practice

1) Earth's oceans contain approximately 1.36×10^9 km^3 of water. Calculate the volume in liters.
2) If the volume of an object is reported as 5.0 ft^3, what is the volume in cubic meters?
3) A car travels 28 mi per gallon of gasoline. How many kilometers per liter will it go?

Sample Exercise 1.12

- What is the mass in grams of 1.00 gal of water? The density of water is 1 g/mL.
- The density of benzene is 0.879 g/mL. Calculate the mass in grams of 1.00 qt of benzene.