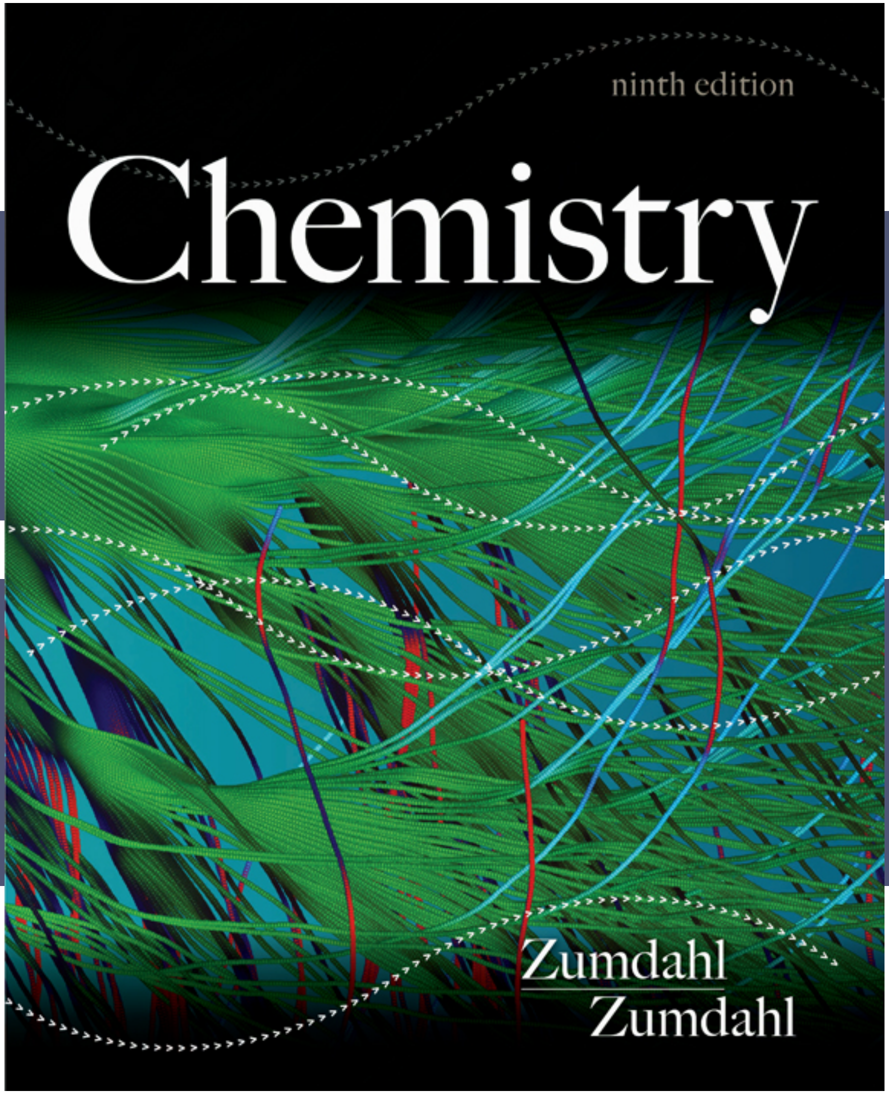


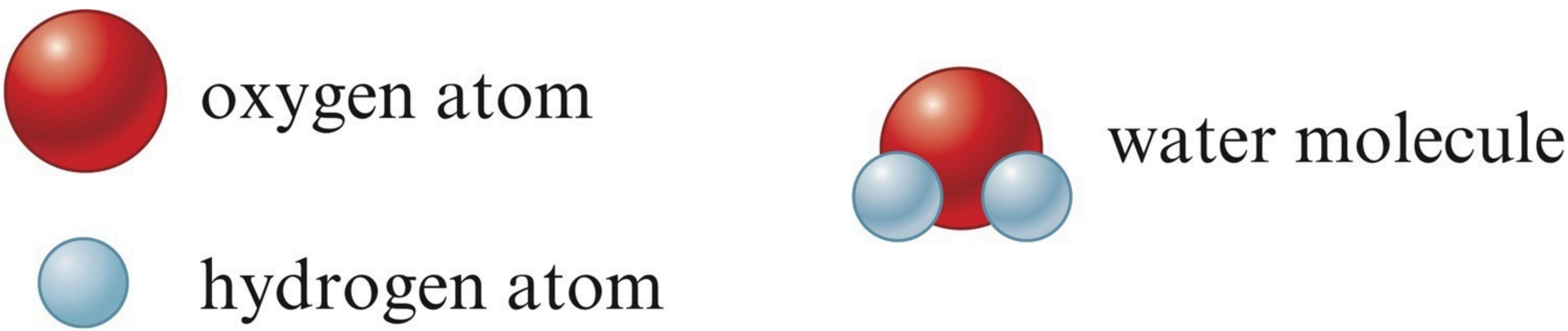
Chapter 1

Chemical Foundations

- A main challenge of chemistry is to understand the connection between the macroscopic world that we experience and the microscopic world of atoms and molecules.
- You must learn to think on the atomic level.

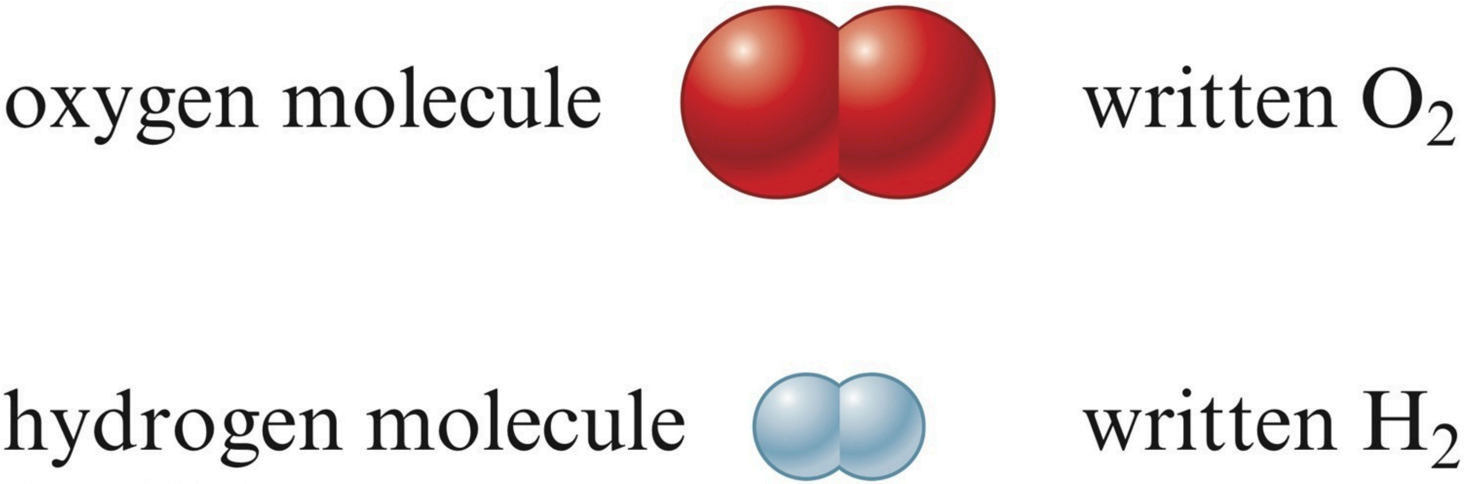
Atoms vs. Molecules

- Matter is composed of tiny particles called atoms.
- Atom: smallest part of an element that is still that element.
- Molecule: Two or more atoms joined and acting as a unit.



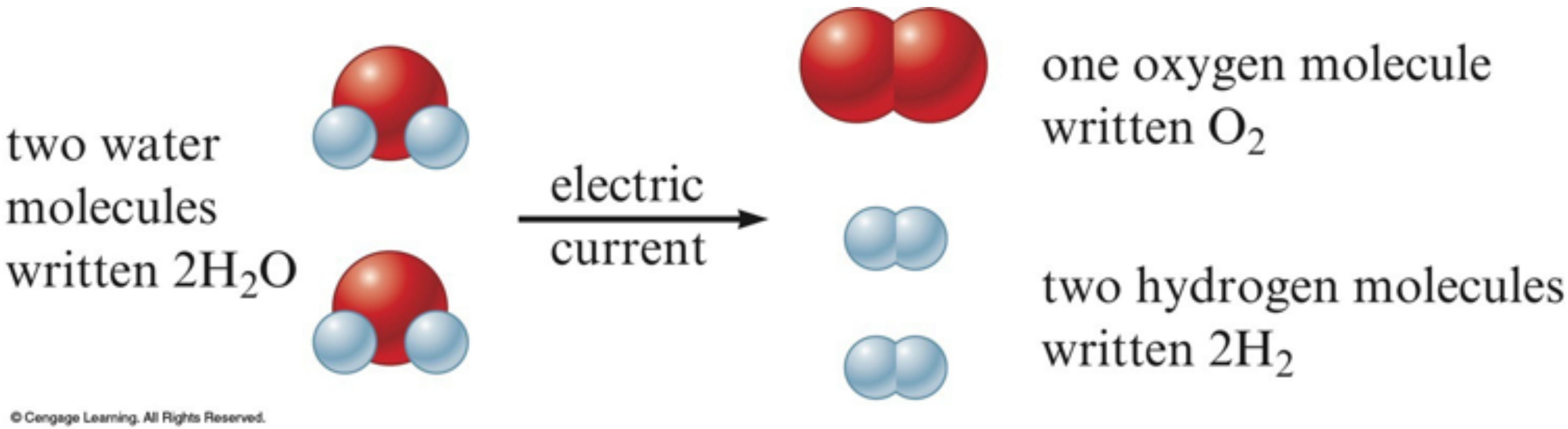
Oxygen and Hydrogen Molecules

- Use subscripts when more than one atom is in the molecule.



A Chemical Reaction

- One substance changes to another by reorganizing the way the atoms are attached to each other.

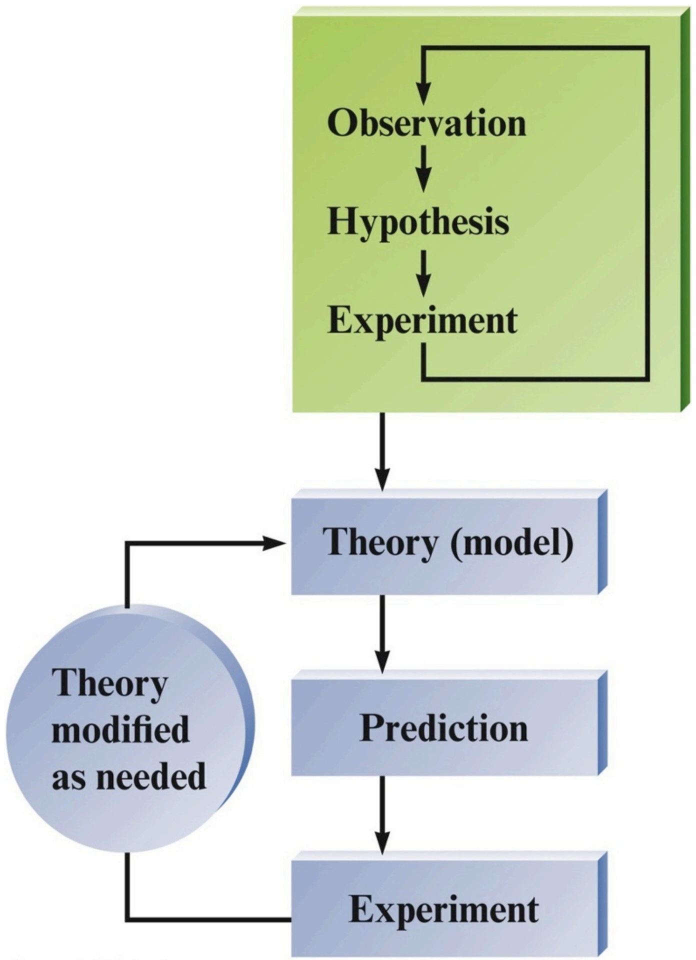


Science

- Science is a framework for gaining and organizing knowledge.
- Science is a plan of action — a procedure for processing and understanding certain types of information.
- Scientists are always challenging our current beliefs about science, asking questions, and experimenting to gain new knowledge.
- Scientific method is needed.

Fundamental Steps of the Scientific Method

- Process that lies at the center of scientific inquiry.



Scientific Models

Law

- A summary of repeatable observed (measurable) behavior.

Hypothesis

- A possible explanation for an observation.

Theory (Model)

- Set of tested hypotheses that gives an overall explanation of some natural phenomenon.

Nature of Measurement

Measurement

- Quantitative observation consisting of two parts.
  - number
  - scale (unit)

- Examples
  - 20 grams
  - 6.63×10^{-34} joule·second

The Fundamental SI Units

Physical Quantity	*Name of Unit*	*Abbreviation*
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

Prefixes Used in the SI System

- Prefixes are used to change the size of the unit.

Table 1.2 | Prefixes Used in the SI System (The most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}

Prefixes Used in the SI System

Table 1.2 | Prefixes Used in the SI System (The most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.

Mass ≠ Weight

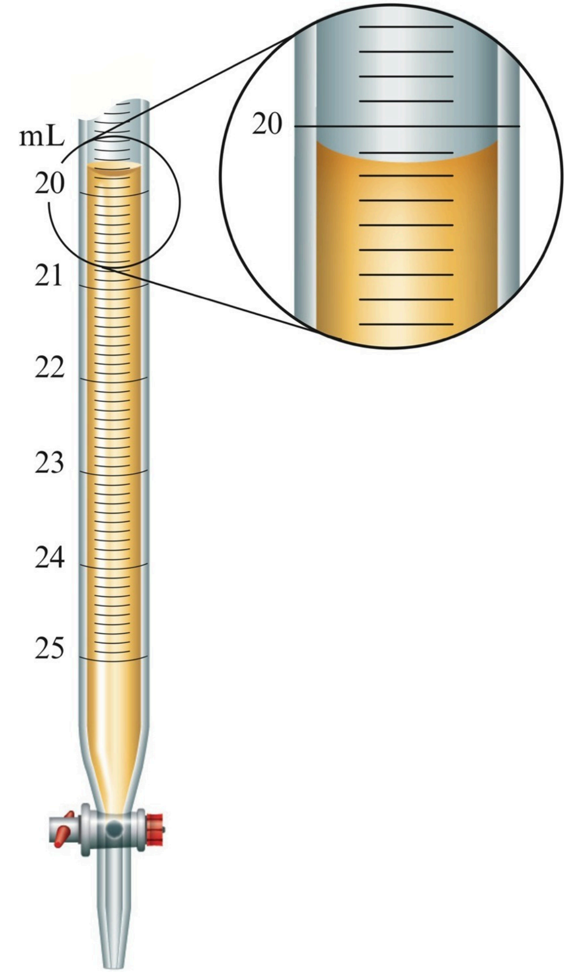
- Mass is a measure of the resistance of an object to a change in its state of motion. Mass does not vary.
- Weight is the force that gravity exerts on an object. Weight varies with the strength of the gravitational field.

Section 1.4
Uncertainty in Measurement

- A digit that must be estimated in a measurement is called uncertain.
- A measurement always has some degree of uncertainty. It is dependent on the precision of the measuring device.
- Record the certain digits and the first uncertain digit (the estimated number).

Measurement of Volume Using a Buret

- The volume is read at the bottom of the liquid curve (meniscus).
- Meniscus of the liquid occurs at about 20.15 mL.
  - Certain digits: 20.15
  - Uncertain digit: 20.15



Precision and Accuracy

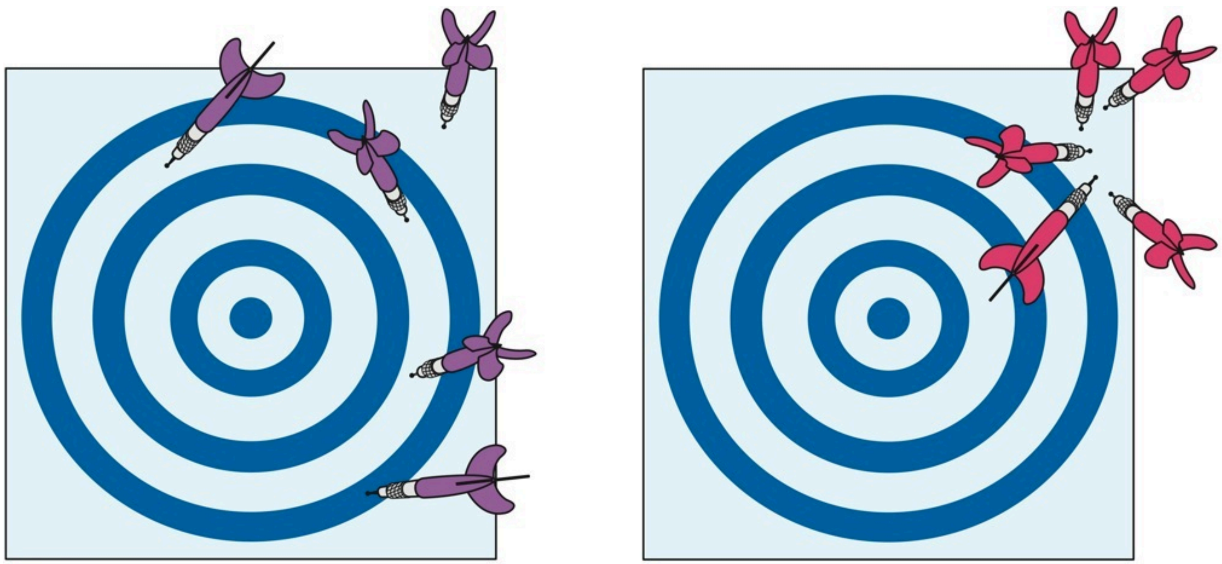
Accuracy

- Agreement of a particular value with the true value.

Precision

- Degree of agreement among several measurements of the same quantity.

Precision and Accuracy



a Neither accurate nor precise.

b Precise but not accurate.

c Accurate and precise.

Rules for Counting Significant Figures

1. Nonzero integers always count as significant figures.
   - 3456 has 4 sig figs (significant figures).

Rules for Counting Significant Figures

2. There are three classes of zeros.

a. Leading zeros are zeros that precede all the nonzero digits. These do not count as significant figures.
   - 0.048 has 2 sig figs.

Rules for Counting Significant Figures

b. Captive zeros are zeros between nonzero digits. These always count as significant figures.
   - 16.07 has 4 sig figs.

Rules for Counting Significant Figures

c. Trailing zeros are zeros at the right end of the number. They are significant only if the number contains a decimal point.
   - 9.300 has 4 sig figs.
   - 150 has 2 sig figs.

Rules for Counting Significant Figures

3. Exact numbers have an infinite number of significant figures.
   - 1 inch = 2.54 cm, exactly.
   - 9 pencils (obtained by counting).

Exponential Notation

- Example
  - 300. written as 3.00×10^2
  - Contains three significant figures.
- Two Advantages
  - Number of significant figures can be easily indicated.
  - Fewer zeros are needed to write a very large or very small number.

Significant Figures in Mathematical Operations

1. For multiplication or division, the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation.

$$1.342 \times \underline{5.5} = 7.381 \rightarrow \underline{7.4}$$

Significant Figures in Mathematical Operations

2. For addition or subtraction, the result has the same number of decimal places as the least precise measurement used in the calculation.

$$\begin{array}{r} 23.445 \\ 7.8\underline{3} \\ 31.275 \end{array} \quad \text{Corrected} \quad 31.2\underline{8}$$

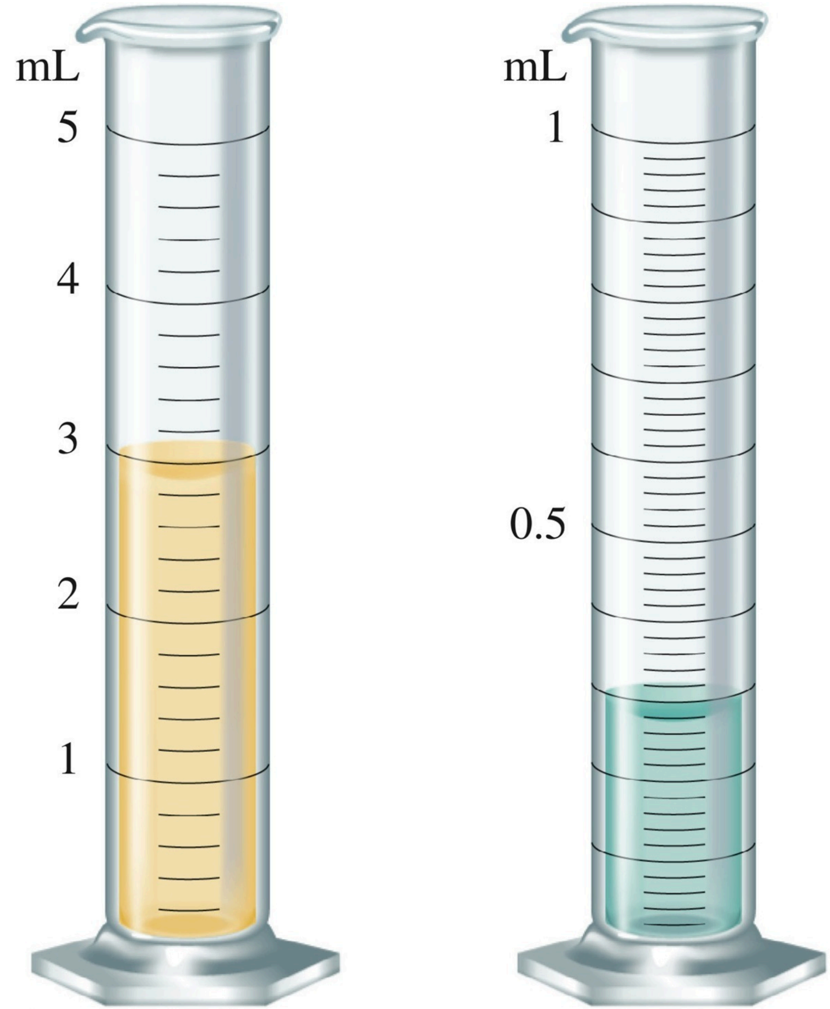
CONCEPT CHECK!

You have water in each graduated cylinder shown. You then add both samples to a beaker (assume that all of the liquid is transferred).

How would you write the number describing the total volume?

3.1 mL

What limits the precision of the total volume?



Questions to ask when approaching a problem

- What is my goal?
- What do I know?
- How do I get there?

- Use when converting a given result from one system of units to another.
  - To convert from one unit to another, use the equivalence statement that relates the two units.
  - Derive the appropriate unit factor by looking at the direction of the required change (to cancel the unwanted units).
  - Multiply the quantity to be converted by the unit factor to give the quantity with the desired units.

Example #1

A golfer putted a golf ball 6.8 ft across a green. How many inches does this represent?

- To convert from one unit to another, use the equivalence statement that relates the two units.

1 ft = 12 in

The two unit factors are:

$$\frac{1 \text{ ft}}{12 \text{ in}} \text{ and } \frac{12 \text{ in}}{1 \text{ ft}}$$

Example #1

A golfer putted a golf ball 6.8 ft across a green. How many inches does this represent?

- Derive the appropriate unit factor by looking at the direction of the required change (to cancel the unwanted units).

$$6.8\ \cancel{\text{ft}} \quad \frac{12\ \text{in}}{1\ \cancel{\text{ft}}} \qquad\qquad \text{in}$$

Example #1

A golfer putted a golf ball 6.8 ft across a green. How many inches does this represent?

- Multiply the quantity to be converted by the unit factor to give the quantity with the desired units.

$$6.8\ \cancel{\text{ft}} \quad \frac{12\ \text{in}}{1\ \cancel{\text{ft}}} \quad 82\ \text{in}$$

Section 1.7

Dimensional Analysis

Example #2

An iron sample has a mass of 4.50 lb. What is the mass of this sample in grams?

(1 kg = 2.2046 lbs; 1 kg = 1000 g)

$$4.50\ \cancel{\text{lbs}} \quad \frac{1\ \cancel{\text{kg}}}{2.2046\ \cancel{\text{lbs}}} \quad \frac{1000\ \text{g}}{1\ \cancel{\text{kg}}} = 2.04 \quad 10^{3}\ \text{g}$$

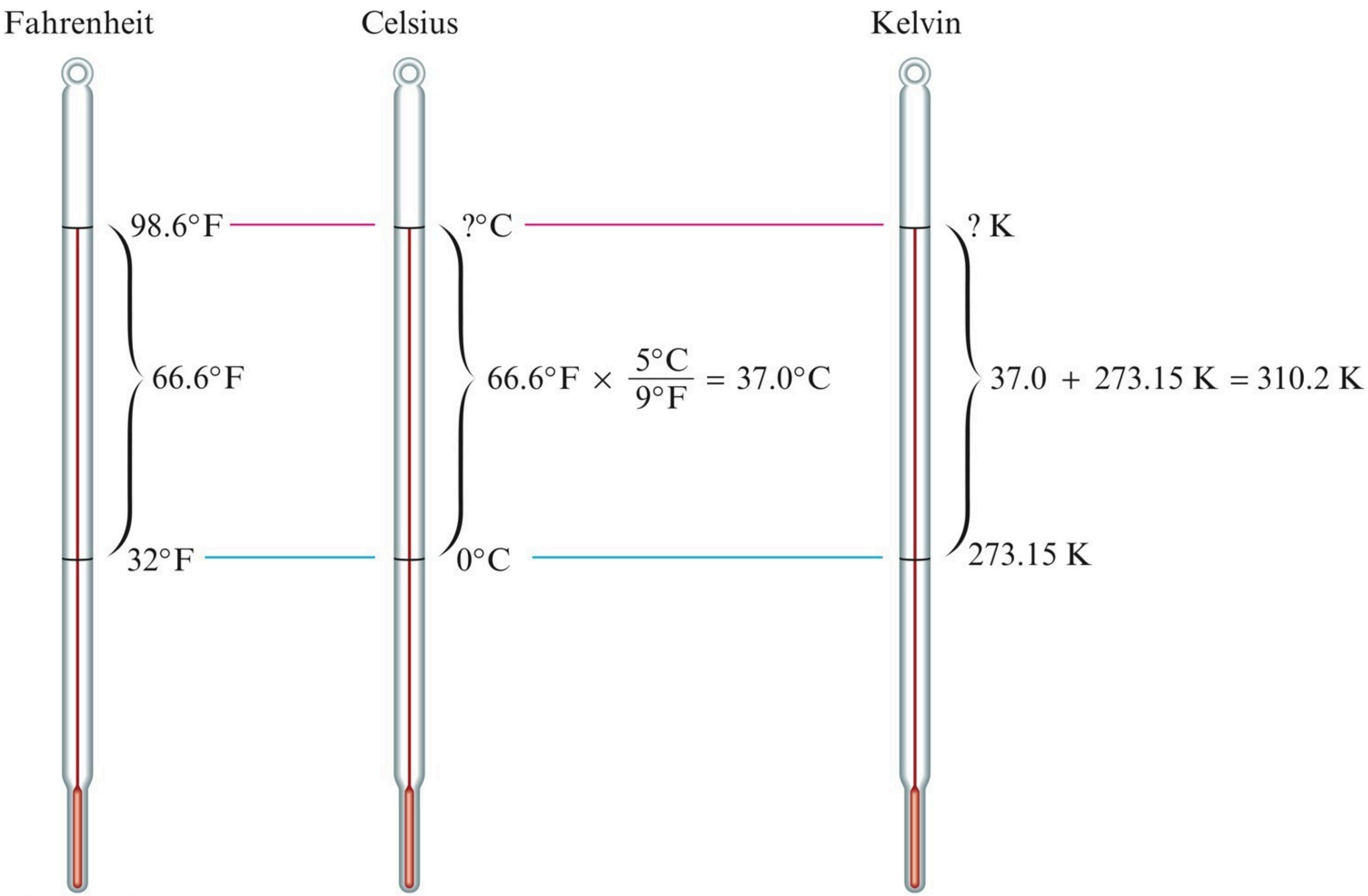
CONCEPT CHECK!

What data would you need to estimate the money you would spend on gasoline to drive your car from New York to Los Angeles? Provide estimates of values and a sample calculation.

Three Systems for Measuring Temperature

- Fahrenheit
- Celsius
- Kelvin

The Three Major Temperature Scales



Converting Between Scales

$$T_K \qquad T_C + 273.15 \qquad\qquad T_C \qquad T_K \qquad 273.15$$

$$T_C \qquad T_F \qquad 32^{\circ}\mathrm{F} \; \frac{5^{\circ}\mathrm{C}}{9^{\circ}\mathrm{F}} \qquad\qquad T_F \qquad T_C \; \frac{9^{\circ}\mathrm{F}}{5^{\circ}\mathrm{C}} + 32^{\circ}\mathrm{F}$$

EXERCISE!

At what temperature does °C = °F?

Section 1.8
Temperature

EXERCISE!

- Since ° C equals ° F, they both should be the same value (designated as variable *x*).
- Use one of the conversion equations such as:

$$T_{\mathrm{C}} \qquad T_{\mathrm{F}} \qquad 32^{\mathrm{o}}\mathrm{F} \; \frac{5^{\mathrm{o}}\mathrm{C}}{9^{\mathrm{o}}\mathrm{F}}$$

- Substitute in the value of *x* for both T_{C} and T_{F}. Solve for *x*.

Section 1.8
Temperature

EXERCISE!

$$T_C \quad T_F \quad 32^{\mathrm{o}}\mathrm{F} \quad \frac{5^{\mathrm{o}}\mathrm{C}}{9^{\mathrm{o}}\mathrm{F}}$$

$$x \quad x \quad 32^{\mathrm{o}}\mathrm{F} \quad \frac{5^{\mathrm{o}}\mathrm{C}}{9^{\mathrm{o}}\mathrm{F}}$$

$$x \quad 40$$

So 40 C = 40 F

Section 1.9
Density

- Mass of substance per unit volume of the substance.
- Common units are g/cm^3 or g/mL.

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Example #1

A certain mineral has a mass of 17.8 g and a volume of 2.35 cm^3. What is the density of this mineral?

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

$$\text{Density} = \frac{17.8\ \text{g}}{2.35\ \text{cm}^3}$$

$$\text{Density} = 7.57\ \text{g/cm}^3$$

Example #2

What is the mass of a 49.6-mL sample of a liquid, which has a density of 0.85 g/mL?

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

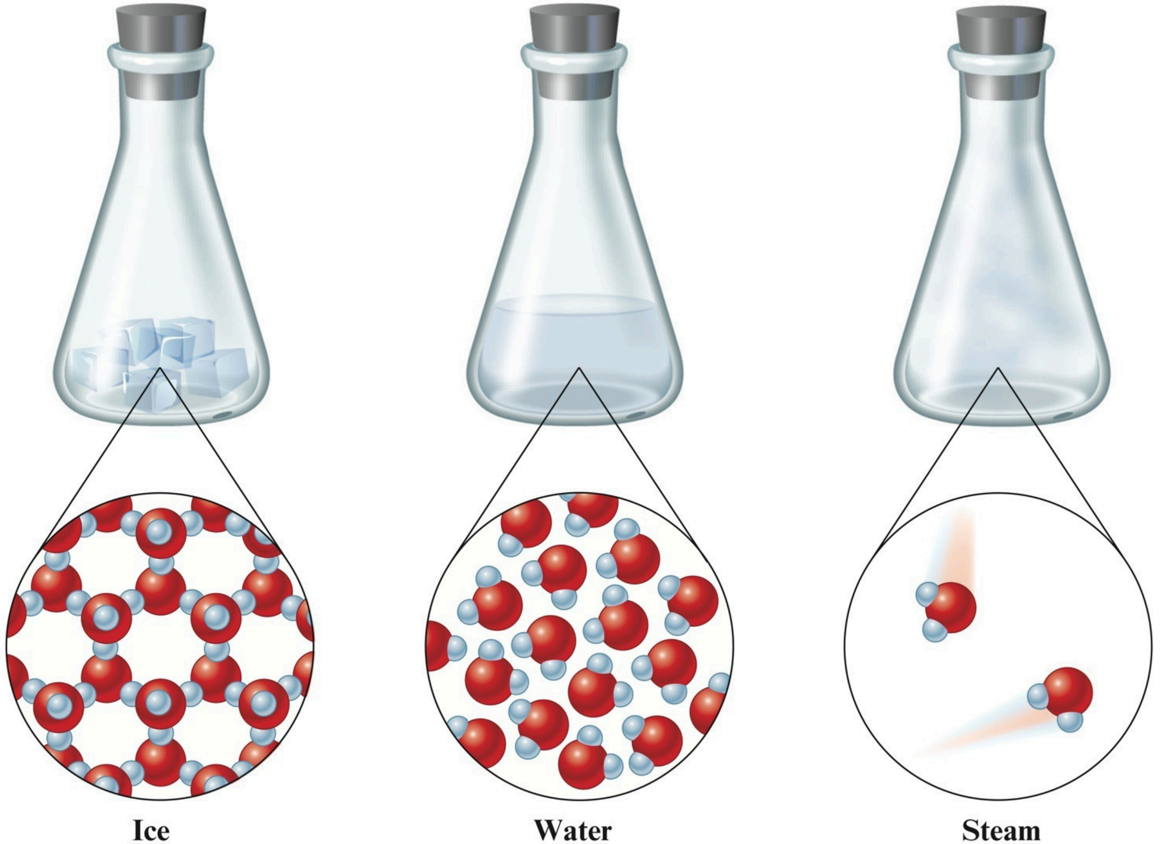
$$0.85 \text{ g/mL} = \frac{x}{49.6 \text{ mL}}$$

$$\text{mass} = x = 42 \text{ g}$$

Matter

- Anything occupying space and having mass.
- Matter exists in three states.
  - Solid
  - Liquid
  - Gas

The Three States of Water



Solid: The water molecules are locked into rigid positions and are close together.

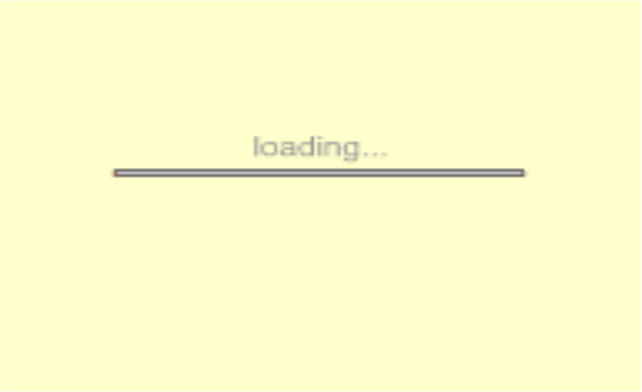
Liquid: The water molecules are still close together but can move around to some extent.

Gas: The water molecules are far apart and move randomly.

Solid

- Rigid
- Has fixed volume and shape.

Structure of a Solid



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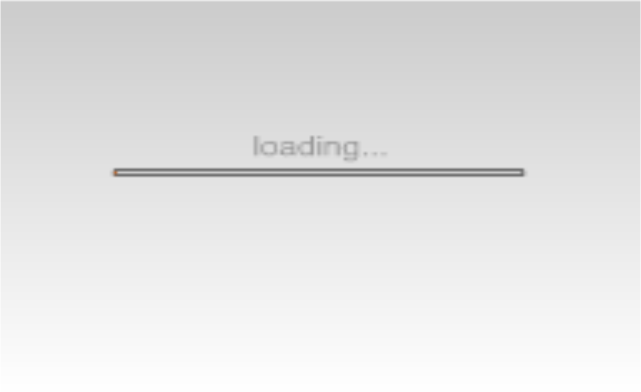
Liquid

- Has definite volume but no specific shape.
- Assumes shape of container.

Structure of a liquid



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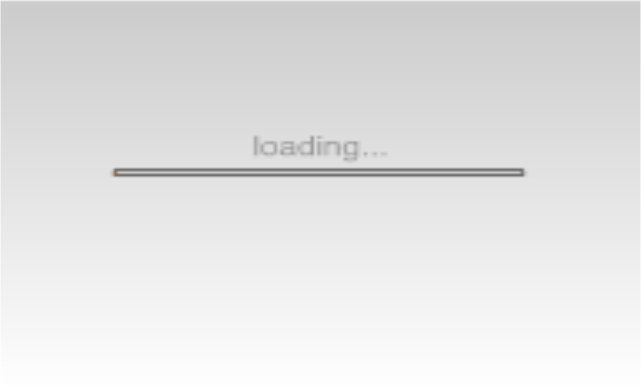
Gas

- Has no fixed volume or shape.
- Takes on the shape and volume of its container.

Structure of a gas



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Mixtures

- Have variable composition.

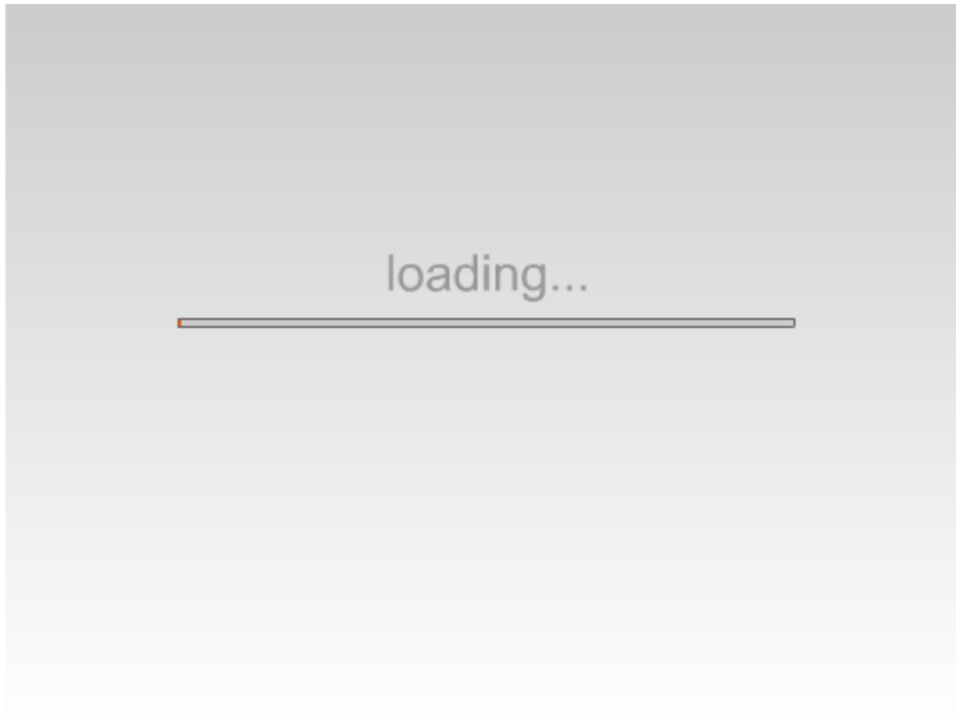
Homogeneous Mixture

- Having visibly indistinguishable parts; solution.

Heterogeneous Mixture

- Having visibly distinguishable parts.

Homogeneous Mixtures



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CONCEPT CHECK!

Which of the following is a homogeneous mixture?

- Pure water
- Gasoline
- Jar of jelly beans
- Soil
- Copper metal

Physical Change

- Change in the form of a substance, not in its chemical composition.
  - Example: boiling or freezing water
- Can be used to separate a mixture into pure compounds, but it will not break compounds into elements.
  - Distillation
  - Filtration
  - Chromatography

Chemical Change

- A given substance becomes a new substance or substances with different properties and different composition.
  - Example: Bunsen burner (methane reacts with oxygen to form carbon dioxide and water)

CONCEPT CHECK!

Which of the following are examples of a chemical change?

- Pulverizing (crushing) rock salt
- Burning of wood
- Dissolving of sugar in water
- Melting a popsicle on a warm summer day

The Organization of Matter