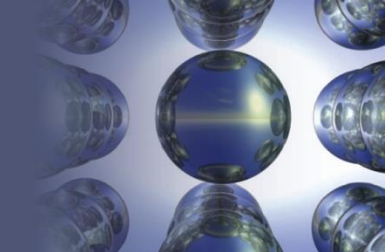


## Chapter 1

### *Chemical Foundations*

# Section 1.1

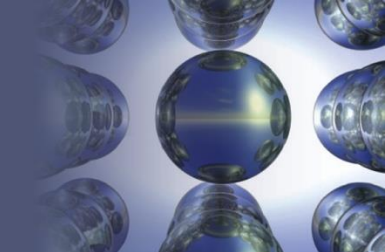
## *Chemistry: An Overview*



- 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.

# Section 1.1

## *Chemistry: An Overview*

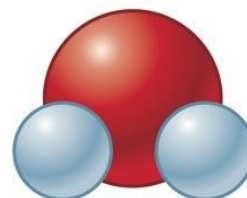


### Atoms vs. Molecules

- 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 atom



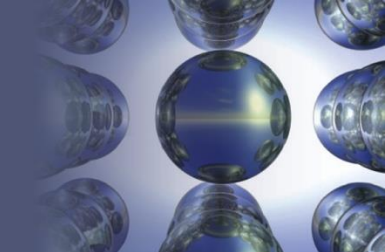
water molecule



hydrogen atom

# Section 1.1

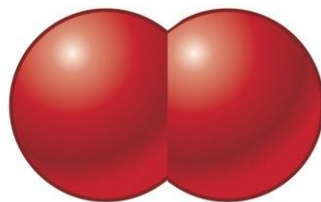
## *Chemistry: An Overview*



### Oxygen and Hydrogen Molecules

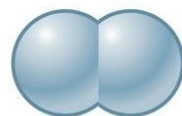
- Use subscripts when more than one atom is in the molecule.

oxygen molecule



written O<sub>2</sub>

hydrogen molecule



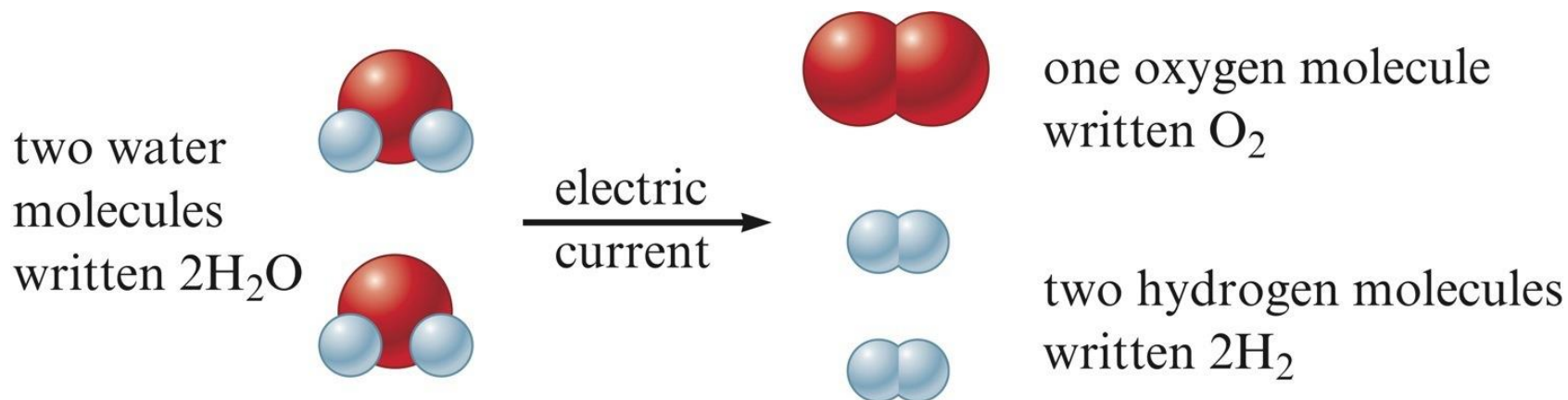
written H<sub>2</sub>

# Section 1.1

## *Chemistry: An Overview*

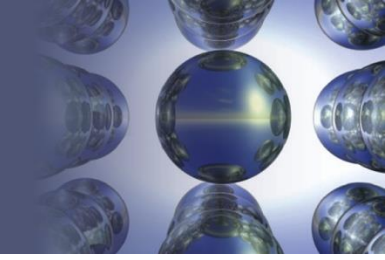
### A Chemical Reaction

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



# Section 1.2

## *The Scientific Method*



### Science

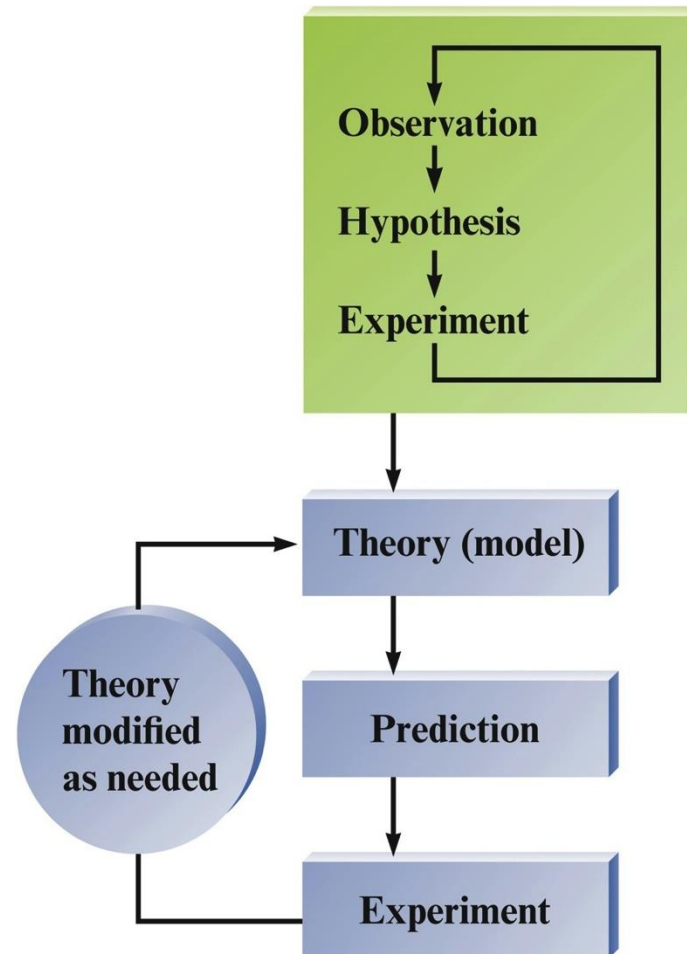
- Science is a framework for gaining and organizing knowledge.
- Science is a plan of action — a procedure for processing and understanding certain types of information.
- Scientists are always challenging our current beliefs about science, asking questions, and experimenting to gain new knowledge.
-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

# Section 1.2

## *The Scientific Method*

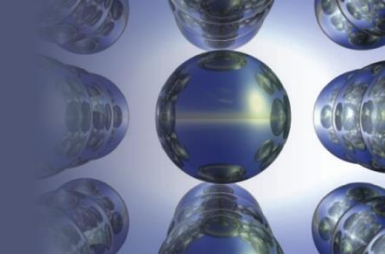
### Fundamental Steps of the Scientific Method

- Process that lies at the center of scientific inquiry.



# Section 1.2

## *The Scientific Method*



### Scientific Models

#### **Law**

- A summary of repeatable observed (measurable) behavior.

#### **Hypothesis**

- A possible explanation for an observation.

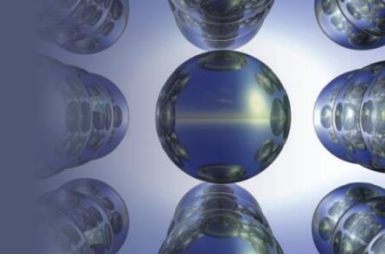
#### **Theory (Model)**

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



# Section 1.3

## *Units of Measurement*



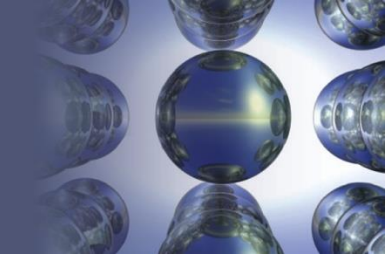
### Nature of Measurement

#### **Measurement**

- Quantitative observation consisting of two parts.
  - number
  - scale (unit)
- Examples
  - 20 grams
  - $6.63 \times 10^{-34}$  joule·second

# Section 1.3

## *Units of Measurement*



### The Fundamental SI Units

<u><i>Physical Quantity</i></u>	<u><i>Name of Unit</i></u>	<u><i>Abbreviation</i></u>
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

# Section 1.3

## Units of Measurement

### Prefixes Used in the SI System

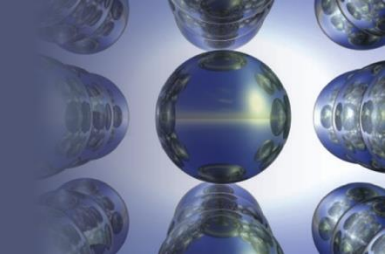
- Prefixes are used to change the size of the unit.

**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	$10^0$

# Section 1.3

## *Units of Measurement*



### Prefixes Used in the SI System

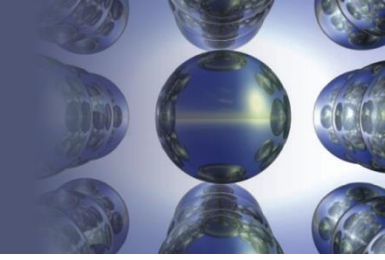
**Table 1.2** | Prefixes Used in the SI System (The most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
deci	d	0.1	$10^{-1}$
centi	c	0.01	$10^{-2}$
milli	m	0.001	$10^{-3}$
micro	$\mu$	0.000001	$10^{-6}$
nano	n	0.000000001	$10^{-9}$
pico	p	0.0000000000001	$10^{-12}$
femto	f	0.00000000000000001	$10^{-15}$
atto	a	0.0000000000000000001	$10^{-18}$

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

# Section 1.3

## *Units of Measurement*

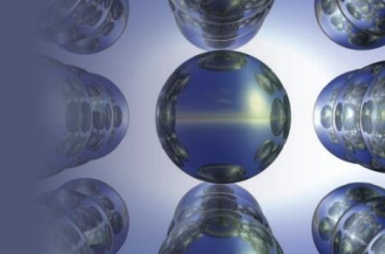


### Mass $\neq$ Weight

- 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*



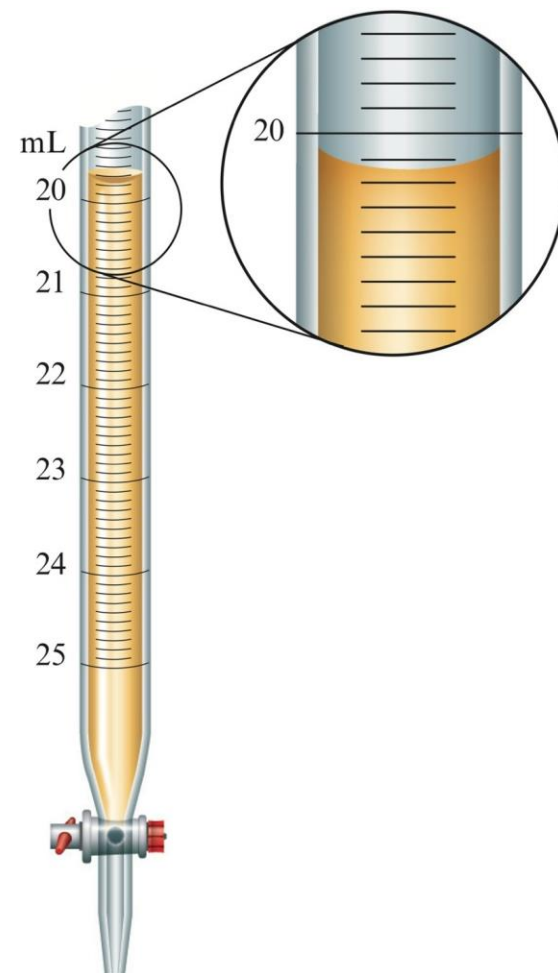
- A digit that must be estimated in a measurement is called uncertain.
- A measurement always has some degree of uncertainty. It is dependent on the precision of the measuring device.
- Record the certain digits and the first uncertain digit (the estimated number).

# Section 1.4

## *Uncertainty in Measurement*

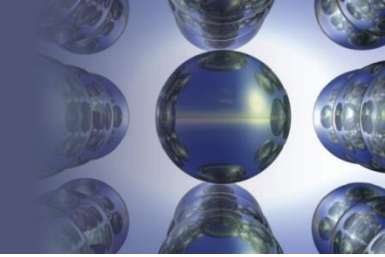
### Measurement of Volume Using a Buret

- 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



# Section 1.4

## *Uncertainty in Measurement*



### Precision and Accuracy

#### **Accuracy**

- Agreement of a particular value with the true value.

#### **Precision**

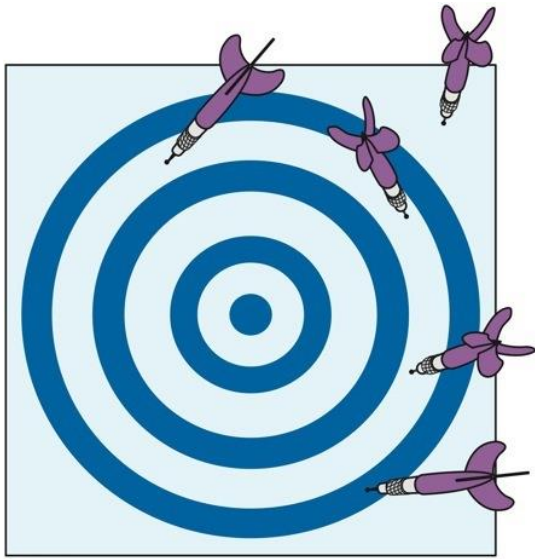
- Degree of agreement among several measurements of the same quantity.



# Section 1.4

## *Uncertainty in Measurement*

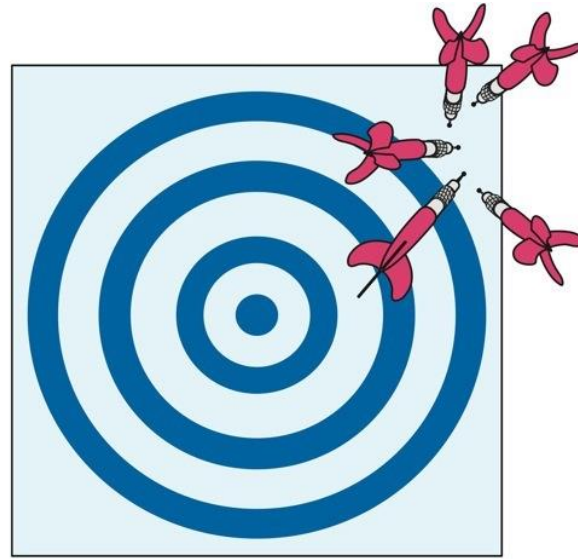
### Precision and Accuracy



**a**

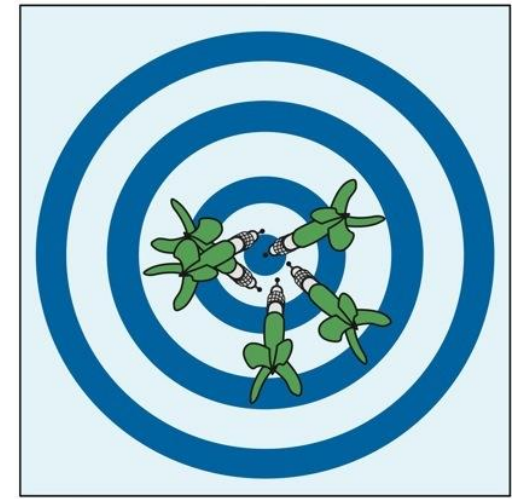
Neither accurate nor precise.

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**b**

Precise but not accurate.

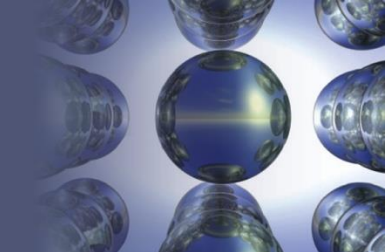


**c**

Accurate and precise.

## Section 1.5

# *Significant Figures and Calculations*

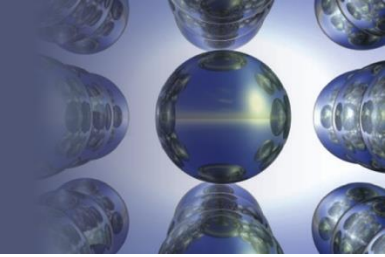


### Rules for Counting Significant Figures

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

## Section 1.5

# *Significant Figures and Calculations*



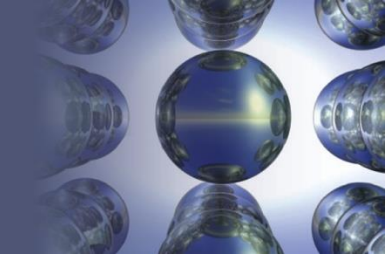
## Rules for Counting Significant Figures

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.

## Section 1.5

# *Significant Figures and Calculations*

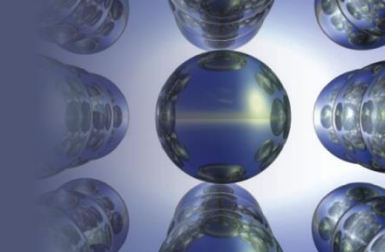


### Rules for Counting Significant Figures

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

## Section 1.5

# *Significant Figures and Calculations*

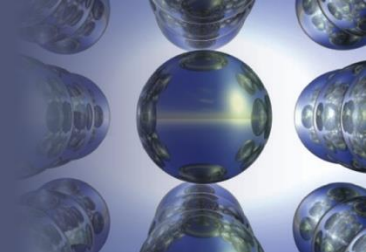


### Rules for Counting Significant Figures

- 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.

## Section 1.5

# *Significant Figures and Calculations*

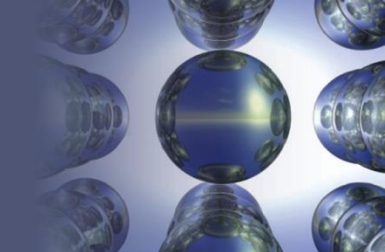


### Rules for Counting Significant Figures

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

# Section 1.5

## *Significant Figures and Calculations*

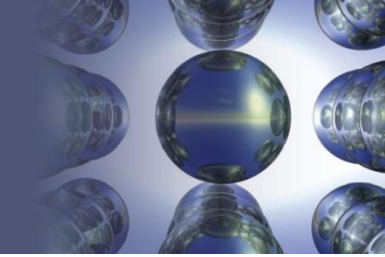


### Exponential Notation

- Example
  - 300. written as  $3.00 \times 10^2$
  - 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.

## Section 1.5

# *Significant Figures and Calculations*



## Significant Figures in Mathematical Operations

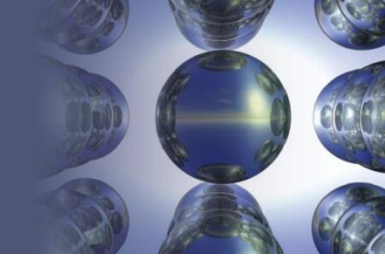
1. For multiplication or division, the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation.

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



## Section 1.5

# *Significant Figures and Calculations*



## 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.

$$\begin{array}{r} 23.445 \\ + \quad 7.83 \\ \hline 31.275 \end{array} \xrightarrow{\text{Corrected}} 31.28$$

# Section 1.5

## *Significant Figures and Calculations*

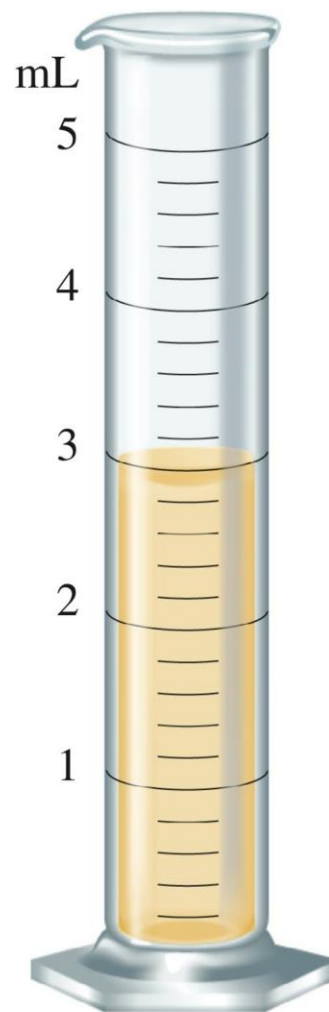
### **CONCEPT CHECK!**

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?

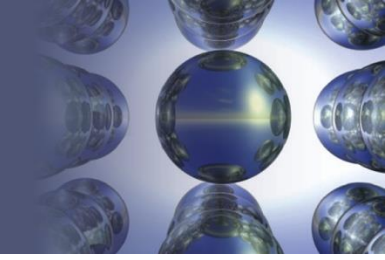
**3.1 mL**

What **limits** the precision of the total volume?



# Section 1.7

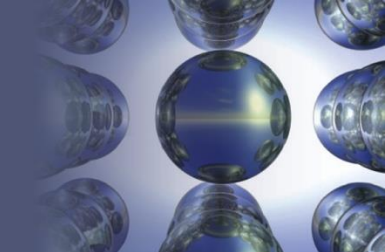
## *Dimensional Analysis*



- 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.

# Section 1.7

## *Dimensional Analysis*



### 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.

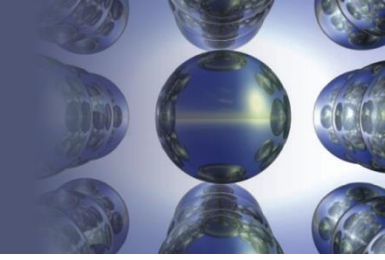
$$1 \text{ ft} = 12 \text{ in}$$

The two unit factors are:

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

# Section 1.7

## *Dimensional Analysis*



### Example #1

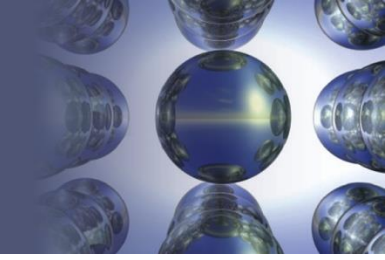
**A golfer putted a golf ball 6.8 ft across a green. How many inches does this represent?**

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

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

# Section 1.7

## *Dimensional Analysis*



### Example #1

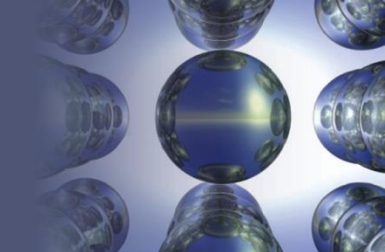
**A golfer putted a golf ball 6.8 ft across a green. How many inches does this represent?**

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

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

# Section 1.7

## *Dimensional Analysis*



### Example #2

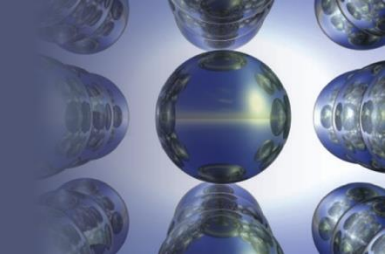
**An iron sample has a mass of 4.50 lb. What is the mass of this sample in grams?**

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

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

## Section 1.7

# *Dimensional Analysis*



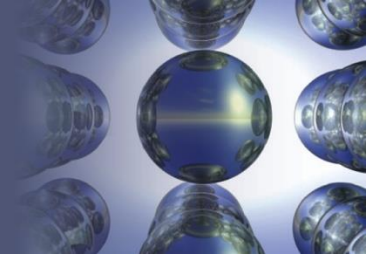
### **CONCEPT CHECK!**

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**.



# Section 1.8

## *Temperature*

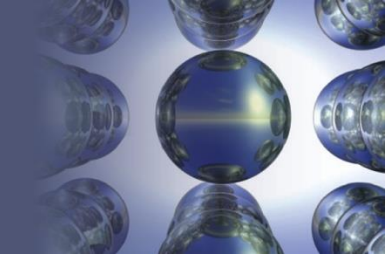


### Three Systems for Measuring Temperature

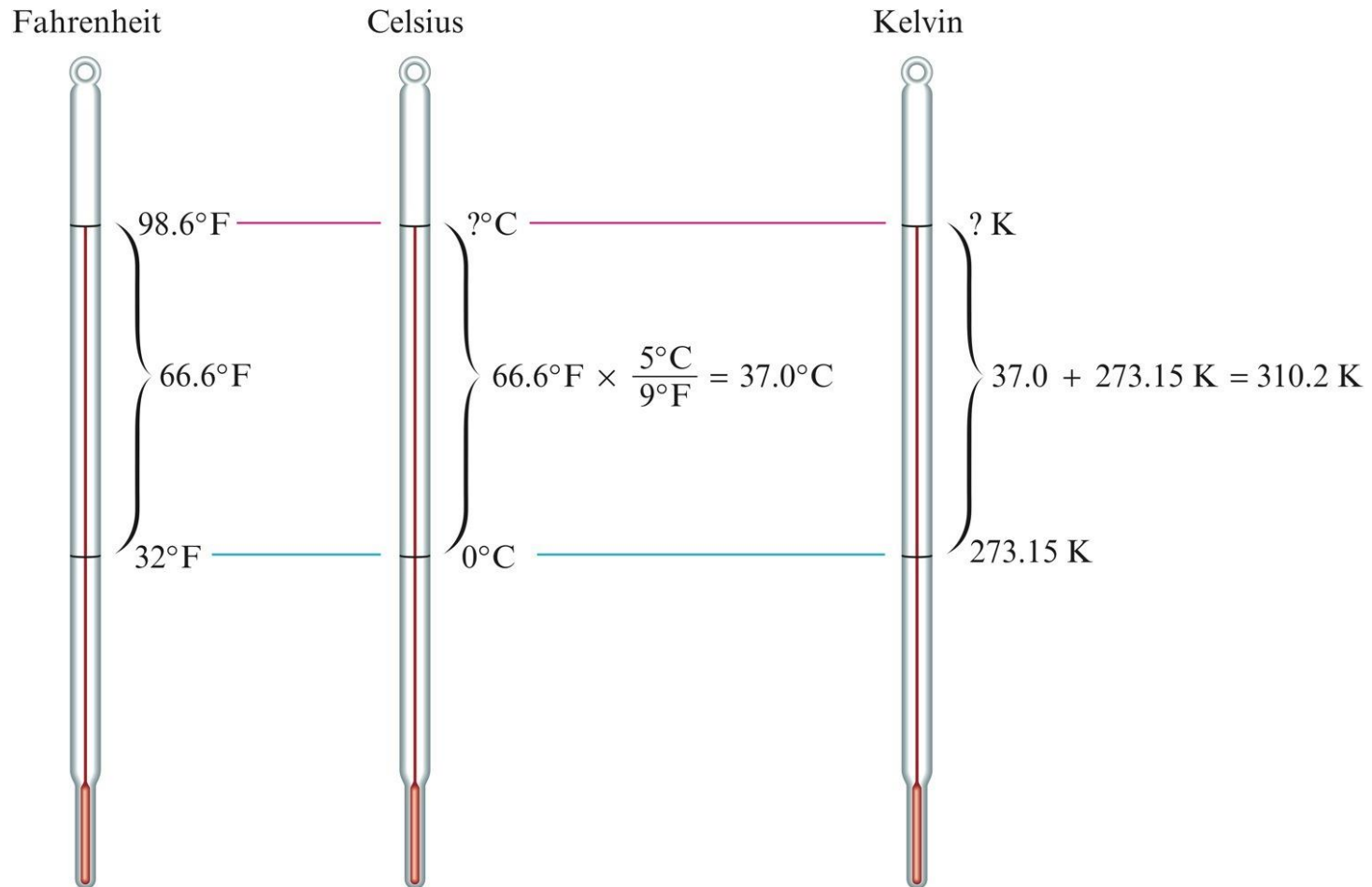
- Fahrenheit
- Celsius
- Kelvin

# Section 1.8

## Temperature

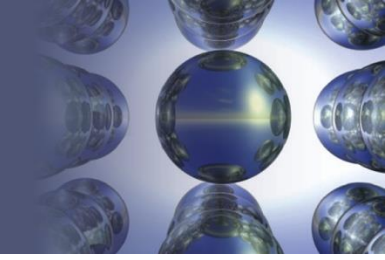


### The Three Major Temperature Scales



# Section 1.8

## *Temperature*



### Converting Between Scales

$$T_K = T_C + 273.15$$

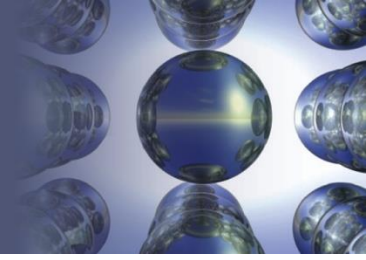
$$T_C = T_K - 273.15$$

$$T_C = (T_F - 32^\circ\text{F}) \frac{5^\circ\text{C}}{9^\circ\text{F}}$$

$$T_F = T_C \times \frac{9^\circ\text{F}}{5^\circ\text{C}} + 32^\circ\text{F}$$

# Section 1.8

## *Temperature*



### ***EXERCISE!***

At what temperature does  $^{\circ}\text{C} = ^{\circ}\text{F}$ ?

# Section 1.8

## *Temperature*

### **EXERCISE!**

- Since  $^{\circ}\text{C}$  equals  $^{\circ}\text{F}$ , they both should be the same value (designated as variable  $x$ ).
- Use one of the conversion equations such as:

$$T_{\text{C}} = (T_{\text{F}} - 32^{\circ}\text{F}) \frac{5^{\circ}\text{C}}{9^{\circ}\text{F}}$$

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

## Section 1.8

### *Temperature*

#### **EXERCISE!**

$$T_C = (T_F - 32^\circ\text{F}) \frac{5^\circ\text{C}}{9^\circ\text{F}}$$

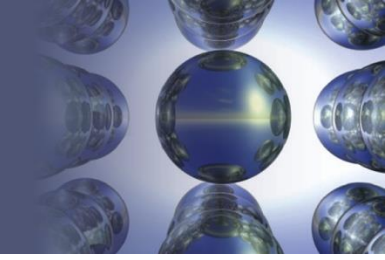
$$x = (x - 32^\circ\text{F}) \frac{5^\circ\text{C}}{9^\circ\text{F}}$$

$$x = -40$$

$$\text{So } -40^\circ\text{C} = -40^\circ\text{F}$$

# Section 1.9

## *Density*

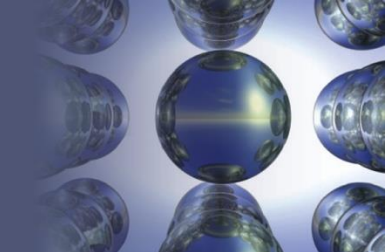


- Mass of substance per unit volume of the substance.
- Common units are  $\text{g/cm}^3$  or  $\text{g/mL}$ .

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

## Section 1.9

### *Density*



### Example #1

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

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

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

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



## Section 1.9

### *Density*

#### Example #2

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

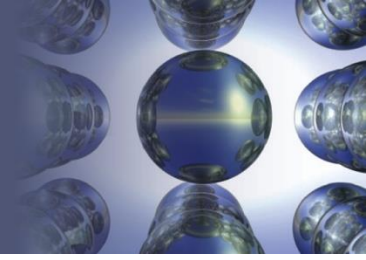
$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

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

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

# Section 1.10

## *Classification of Matter*



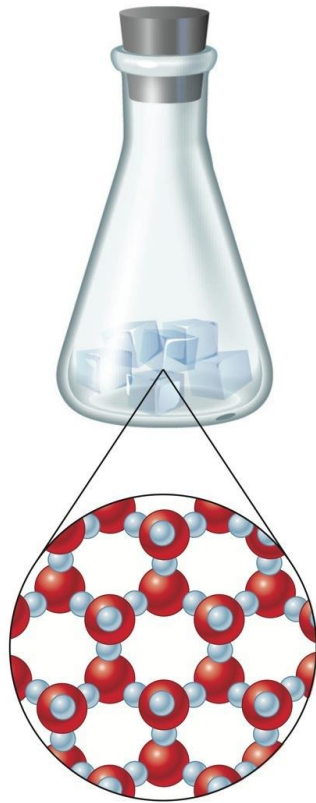
### Matter

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

# Section 1.10

## *Classification of Matter*

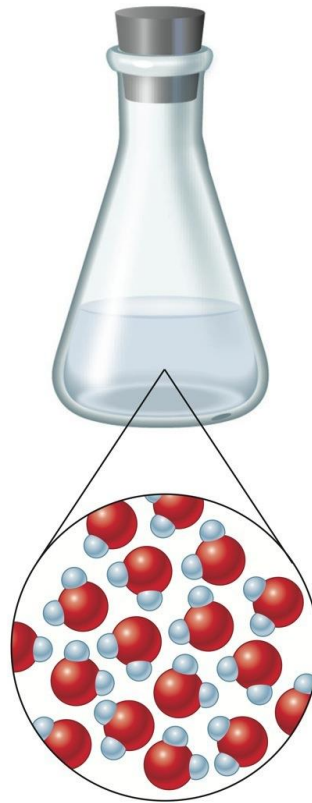
### The Three States of Water



**Ice**

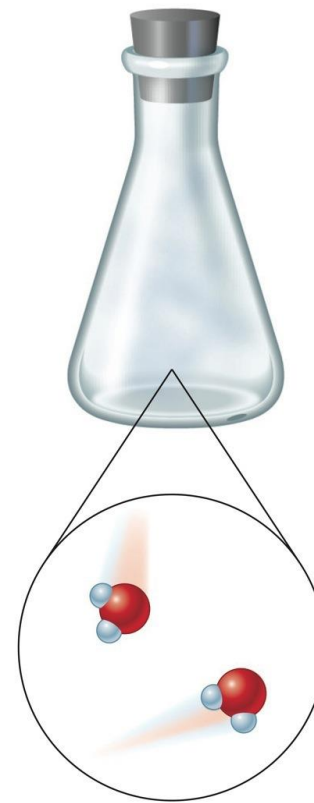
**Solid:** The water molecules are locked into rigid positions and are close together.

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**Water**

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

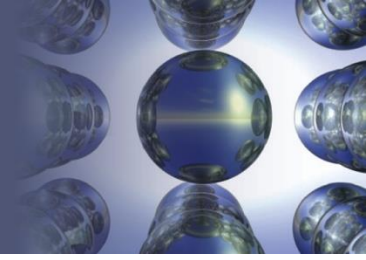


**Steam**

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

# Section 1.10

## *Classification of Matter*

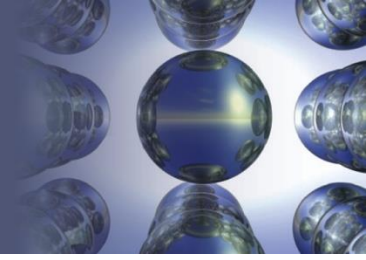


### Solid

- Rigid
- Has fixed volume and shape.

# Section 1.10

## *Classification of Matter*

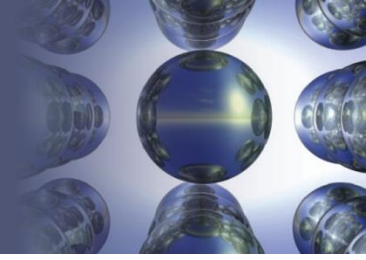


### Liquid

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

# Section 1.10

## *Classification of Matter*

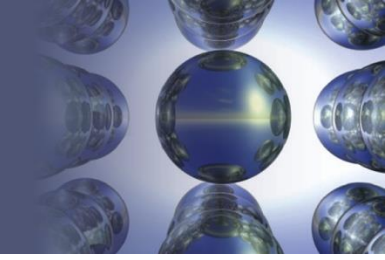


### Gas

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

# Section 1.10

## *Classification of Matter*



### Mixtures

- Have variable composition.

#### **Homogeneous Mixture**

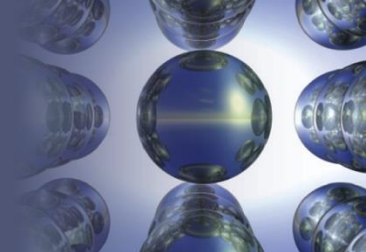
- Having visibly indistinguishable parts; solution.

#### **Heterogeneous Mixture**

- Having visibly distinguishable parts.

# Section 1.10

## *Classification of Matter*



### **CONCEPT CHECK!**

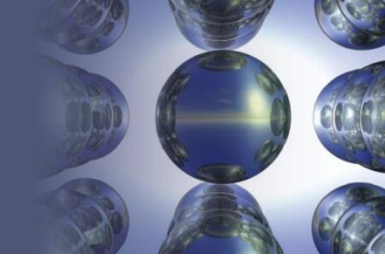
Which of the following is a **homogeneous mixture**?

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



# Section 1.10

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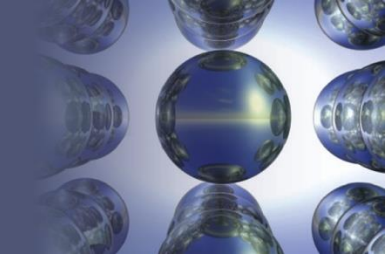


### Physical Change

- 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

# Section 1.10

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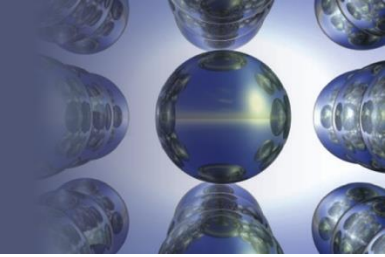


### Chemical Change

- 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)**

## Section 1.10

### *Classification of Matter*



#### **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

# Section 1.10

## *Classification of Matter*

### The Organization of Matter

