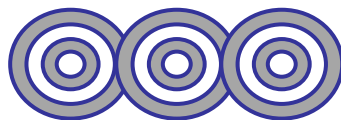




*Chemistry, The Central Science*, 11th edition  
Theodore L. Brown; H. Eugene LeMay, Jr.;  
Bruce E. Bursten; Catherine J. Murphy



# Chapter 1

## Introduction:

# Matter and Measurement

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# What is Chemistry?

**Chemistry** is the study of the **properties** and **behavior of matter**.

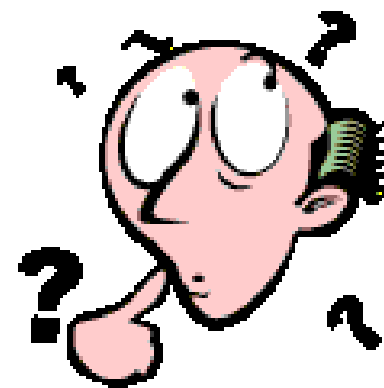


# Why Study Chemistry?

## Realms

Macroscopic  
Ordinary sized objects

Sub-microscopic  
Atoms and molecules

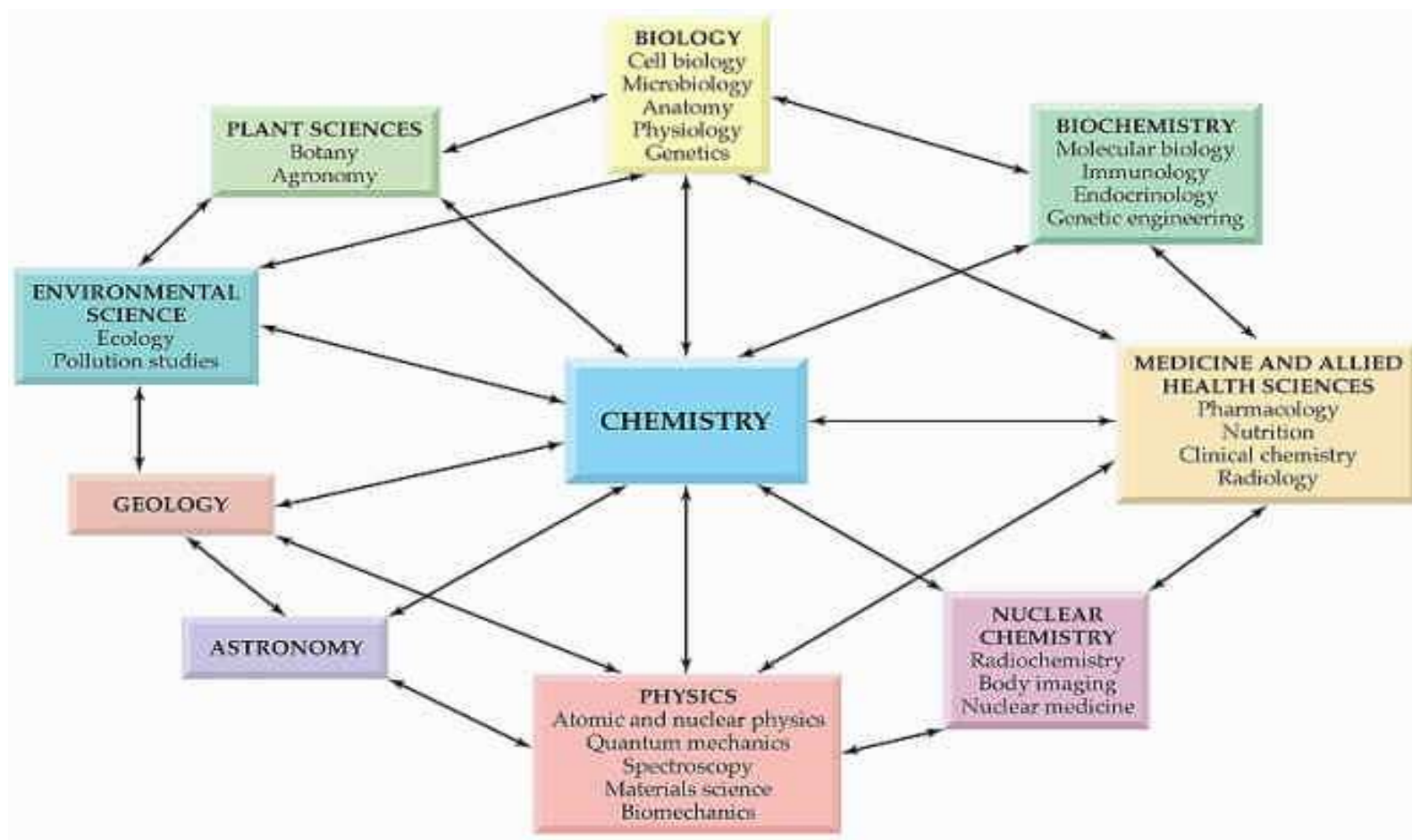


Chemistry is the science that seeks to understand the properties and behavior of matter by studying the properties of atoms and molecules.

Chemistry provides important understanding of our world and how it works.

# Chemistry = the central science

Chemistry is often called **the central science** because of its role in connecting the other sciences, with the life and applied sciences.



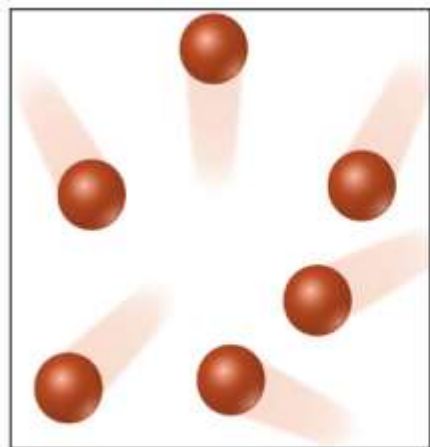
# Matter

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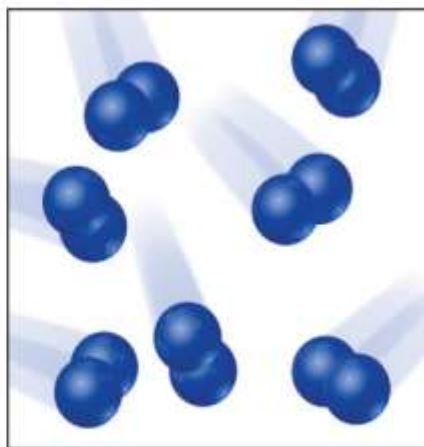
Is the physical material of the universe

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

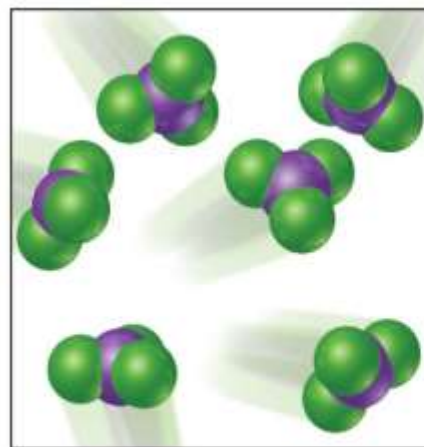
# Classification of matter according to its composition



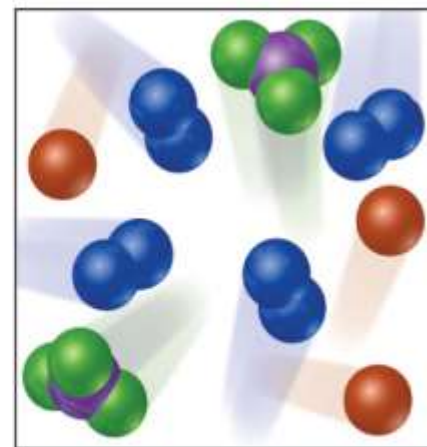
(a) Atoms of an element



(b) Molecules of an element



(c) Molecules of a compound



(d) Mixture of elements and a compound



(b) Water



(c) Carbon dioxide



(d) Ethanol



(e) Ethylene glycol



(f) Aspirin

**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.



# Classification of matter according to its physical state

## **Gas**

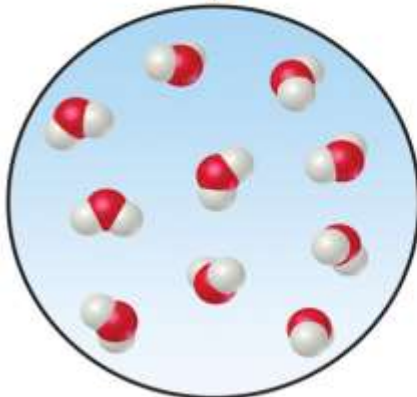
has no fixed volume or shape, it conforms to the volume and shape of its containers, It can be compressed or expand

## **Liquid**

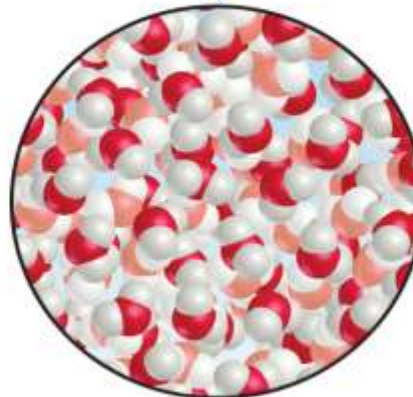
has distinct volume but has no specific shape

## **Solid**

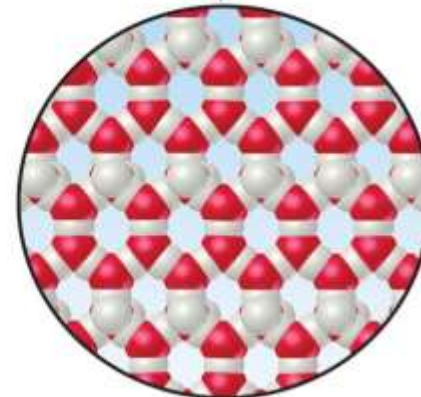
has both definite shape and volume



Gas



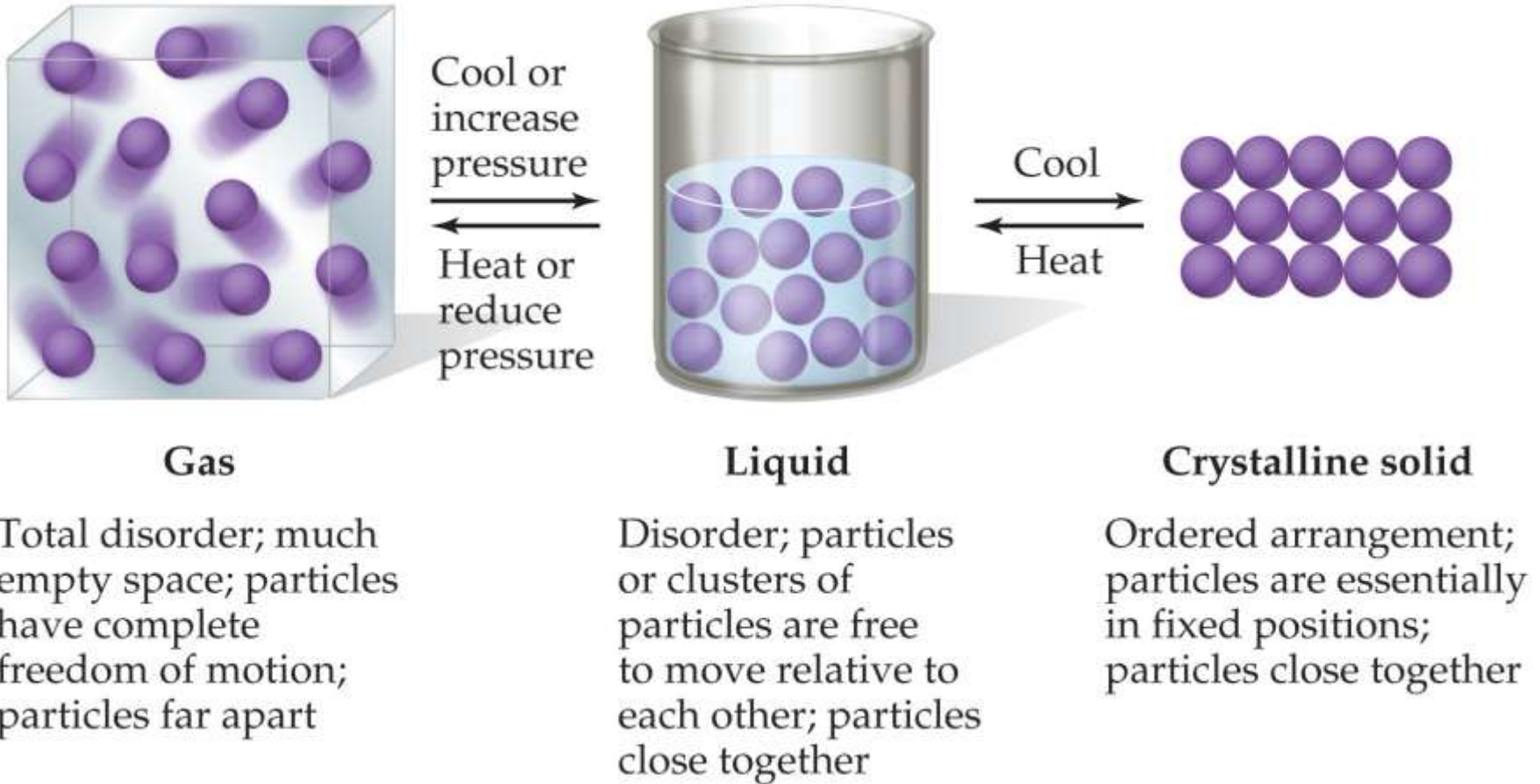
Liquid



Solid

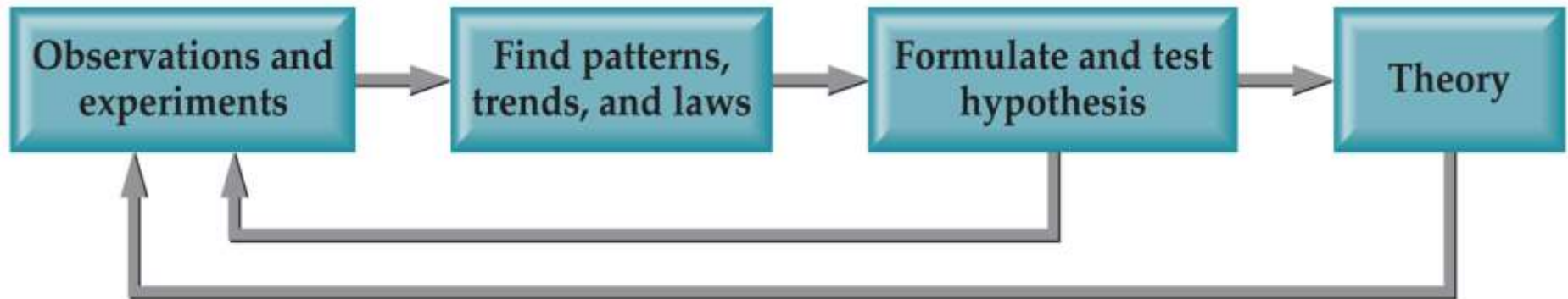


# States of Matter



# Scientific Method

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



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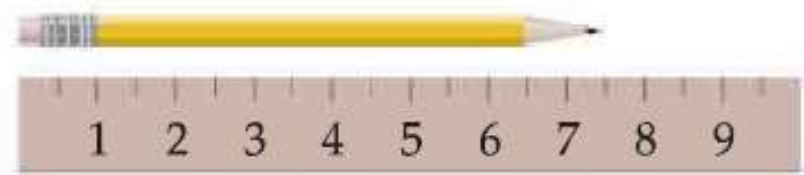
**1.4**

**Units of Measurement**

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# Units of Measurement

7.5 meaningless  
7.5 cm specifies length



The units used for scientific measurements are those of the **metric units**.

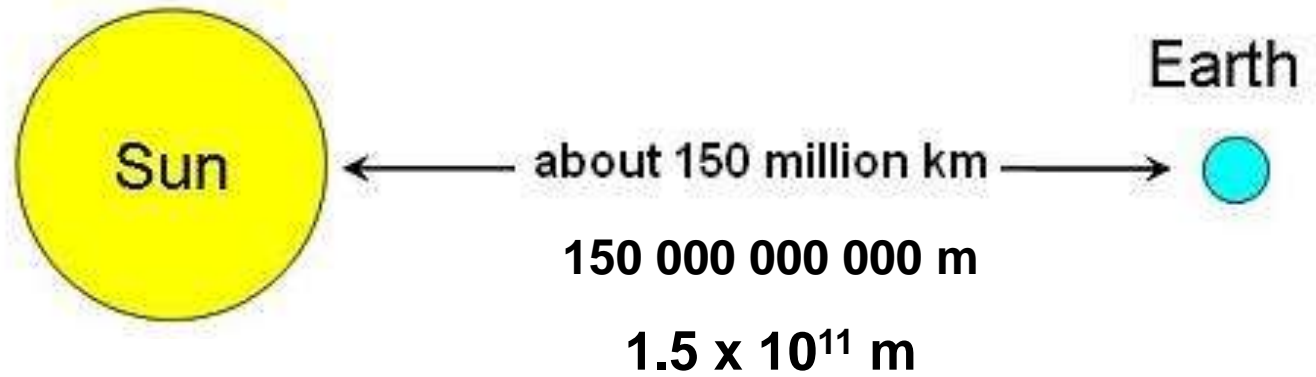
The metric system is an internationally agreed decimal system of measurement that was originally based on the mètre des Archives and the kilogramme des Archives introduced by France in 1799.

# SI Units

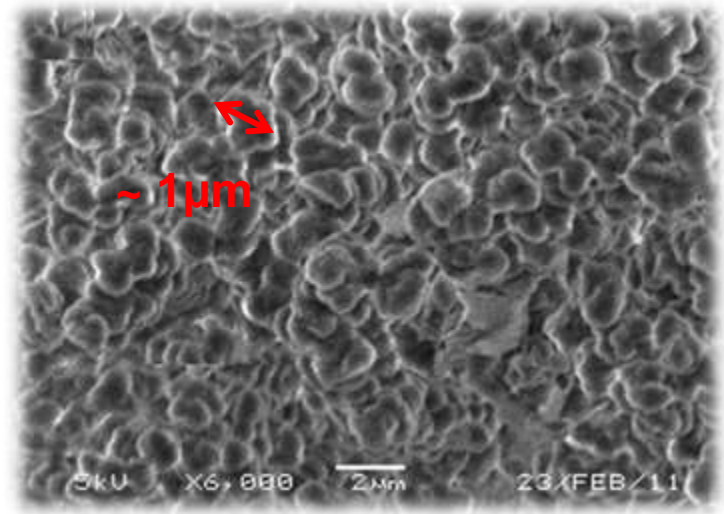
Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s <sup>a</sup>
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

<sup>a</sup>The abbreviation sec is frequently used.

- Système International d'Unités.
- A different base unit is used for each quantity.
- Seven base units from which all other units are derived.



$1 \mu\text{m}$   
 $0.000001 \text{ m}$   
 $1.0 \times 10^{-6} \text{ m}$



**Prefixes** are used to indicate decimal **fractions** or **multiples** of various units. 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	$\mu^a$	$10^{-6}$	1 micrometer ( $\mu$ m) = $1 \times 10^{-6}$ 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

<sup>a</sup>This is the Greek letter mu (pronounced “mew”).

**Exercise:** Which of the following quantities is the smallest?  
 1 mg, 1  $\mu$ g, or 1 pg

# Length, Mass and Temperature

**Length:** is the distance between two objects (SI unit m).

**Mass:** is a measure of the amount of material in an object (SI unit kg).

## Sample Exercise 1.2 Using Metric Prefixes

What is the name given to the unit that equals (a)  $10^{-9}$  gram, (b)  $10^{-6}$  second, (c)  $10^{-3}$  meter?

### Solution

(a) nanogram, ng, (b) microsecond,  $\mu\text{s}$ , (c) millimeter, mm.

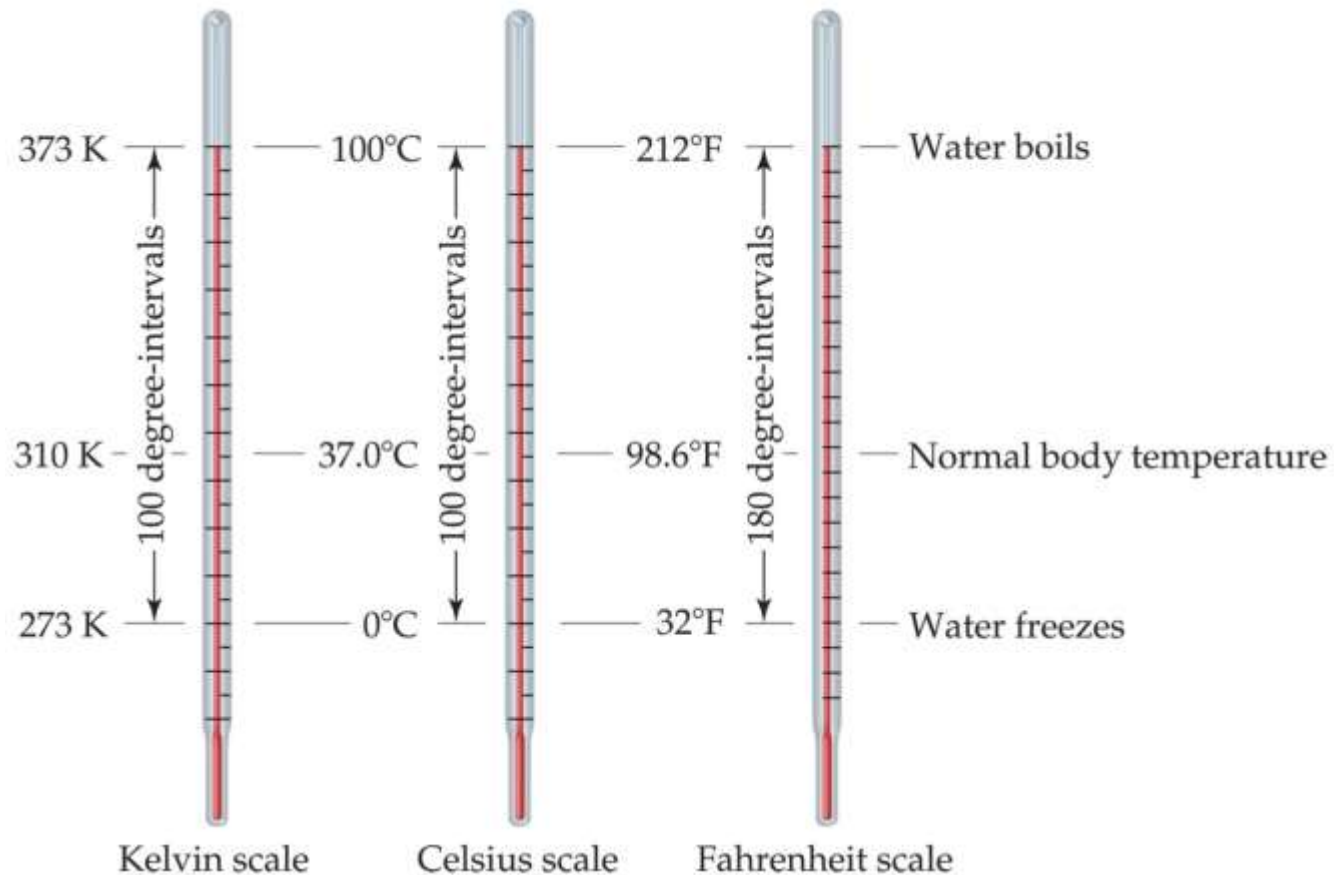
### Practice Exercise

(a) What decimal fraction of a second is a picosecond, ps? (b) Express the measurement  $6.0 \times 10^3\text{m}$  using a prefix to replace the power of ten. (c) Use exponential notation to express 3.76 mg in grams.

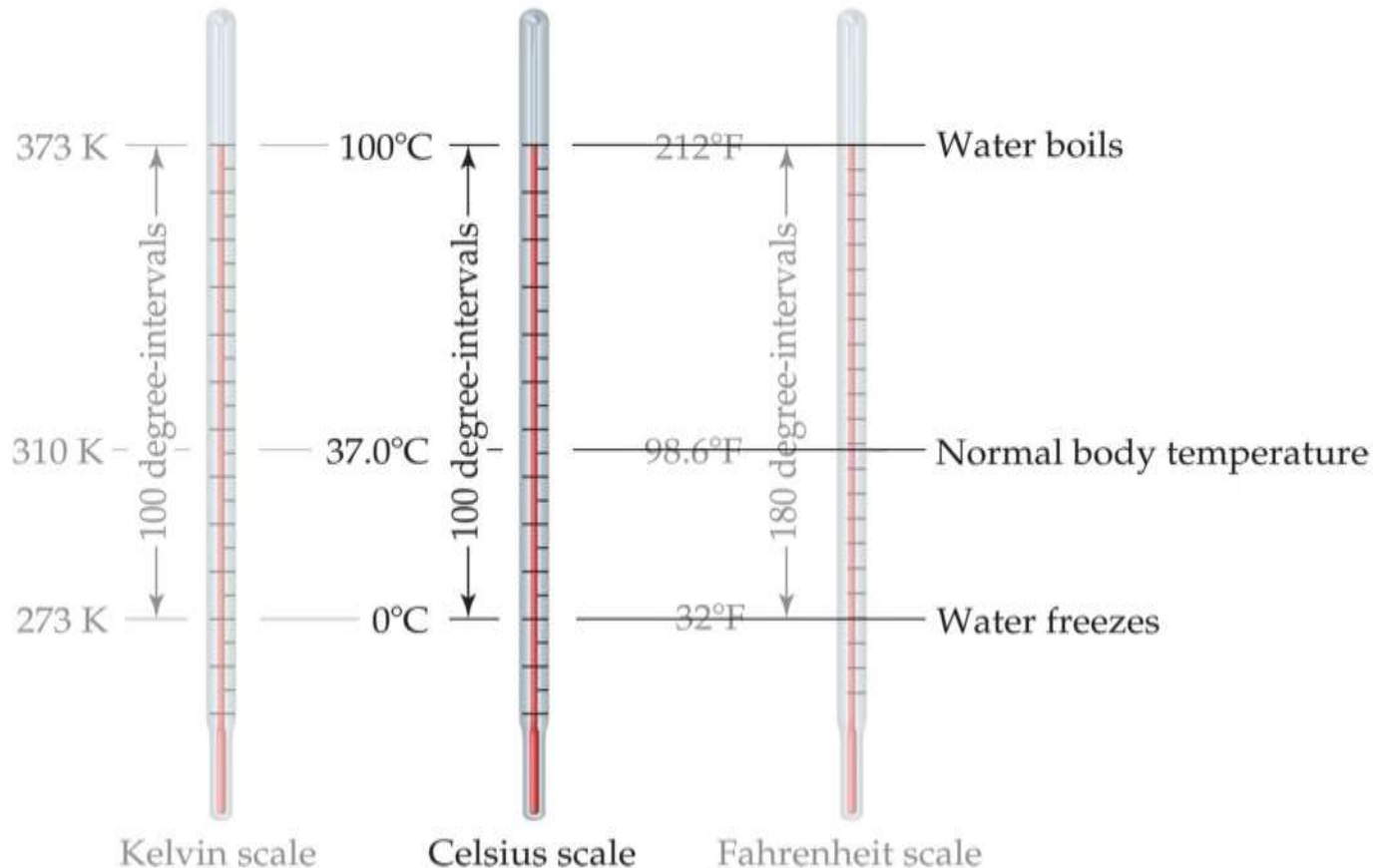
**Answer:** (a)  $10^{-12}$  second, (b) 6.0 km, (c)  $3.76 \times 10^{-3}$  g

# Temperature

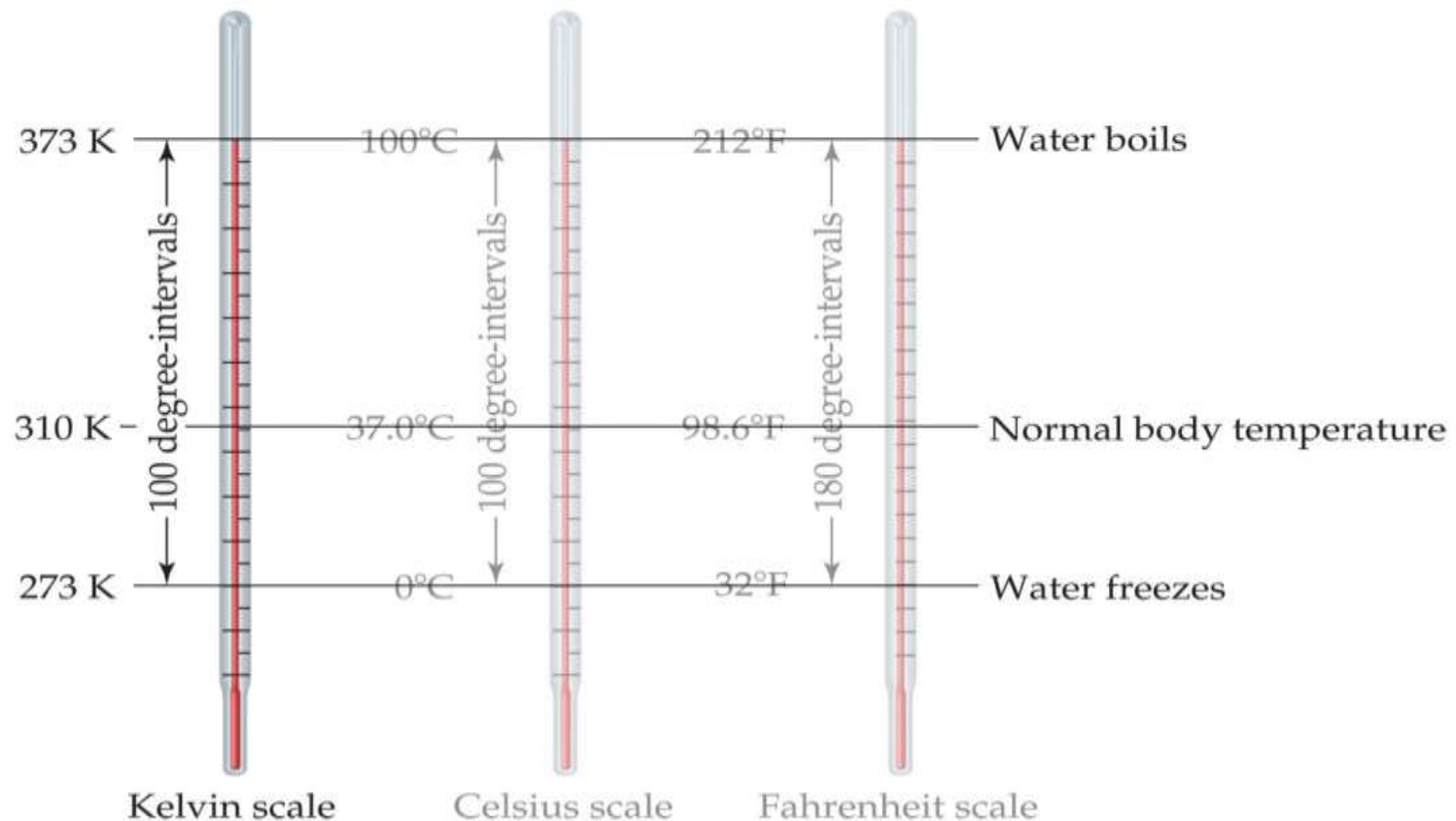
**Temperature:** is a measure of the hotness or coldness of an object (SI unit K).



- In scientific measurements, the Celsius and Kelvin scales are most often used.
- The Celsius scale is based on the properties of water.
  - $0^{\circ}\text{C}$  is the freezing point of water.
  - $100^{\circ}\text{C}$  is the boiling point of water.

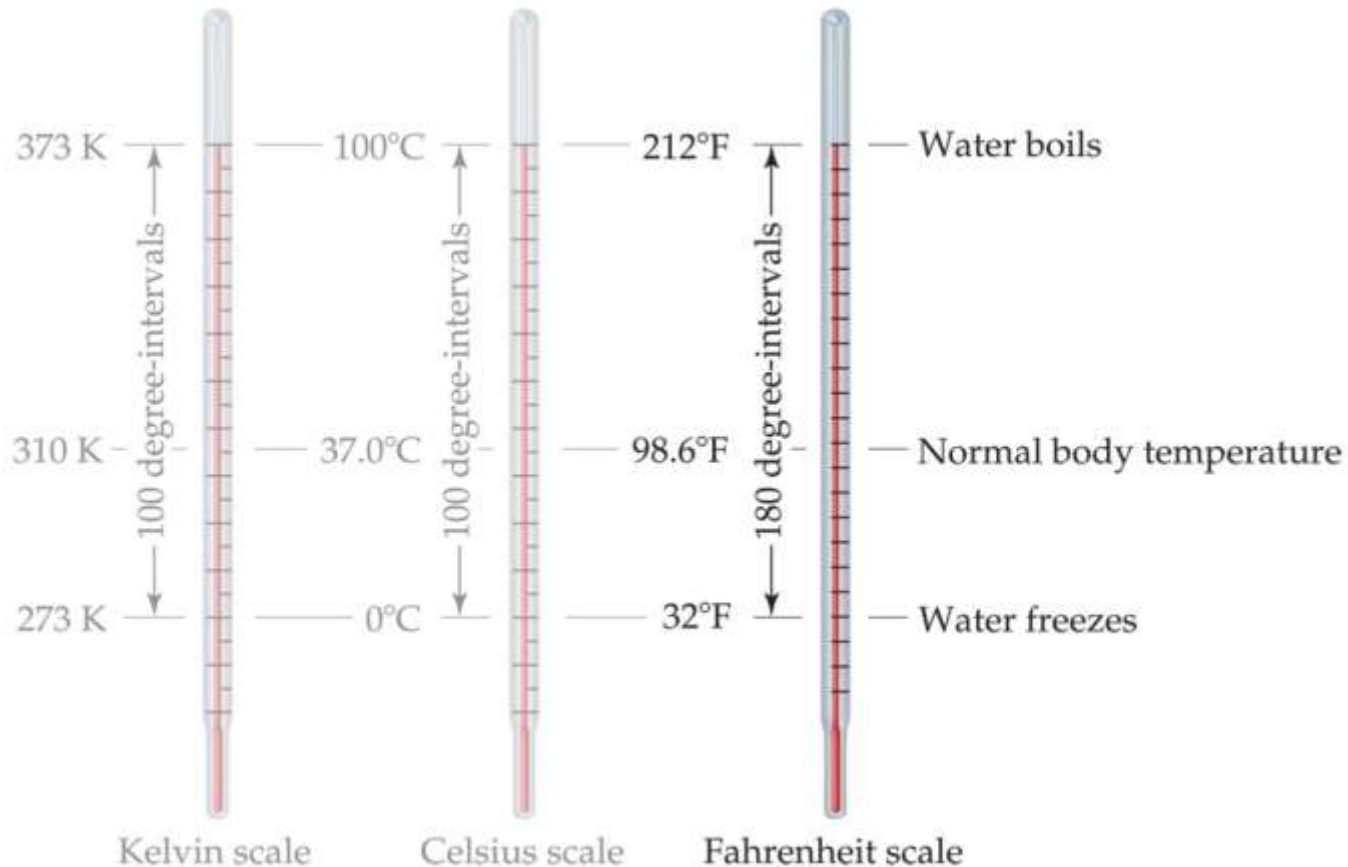


- The Kelvin is the SI unit of temperature.
- It is based on the properties of gases.
- There are no negative Kelvin temperatures.
- **$K = ^\circ C + 273.15$**
- At  $-273.15\text{ }^\circ C$  (0 K) called absolute zero



- The Fahrenheit scale is not used in scientific measurements.
- On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F.

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





## Sample Exercise 1.3 Converting Units of Temperature

If a weather forecaster predicts that the temperature for the day will reach  $31\text{ }^{\circ}\text{C}$ , what is the predicted temperature **(a)** in K, **(b)** in  $^{\circ}\text{F}$ ?

### Solution

$$\text{(a) } K = 31 + 273 = 304\text{ K}$$

$$\text{(b) } ^{\circ}\text{F} = \frac{9}{5}(31) + 32 = 56 + 32 = 88\text{ }^{\circ}\text{F}$$

### Practice Exercise

Ethylene glycol, the major ingredient in antifreeze, freezes at  $-11.5^{\circ}\text{F}$ . What is the freezing point in **(a)** K, **(b)**  $^{\circ}\text{F}$ ?

**Answer:** **(a)** 261.7 K, **(b)**  $11.3\text{ }^{\circ}\text{F}$

# Derived SI Units

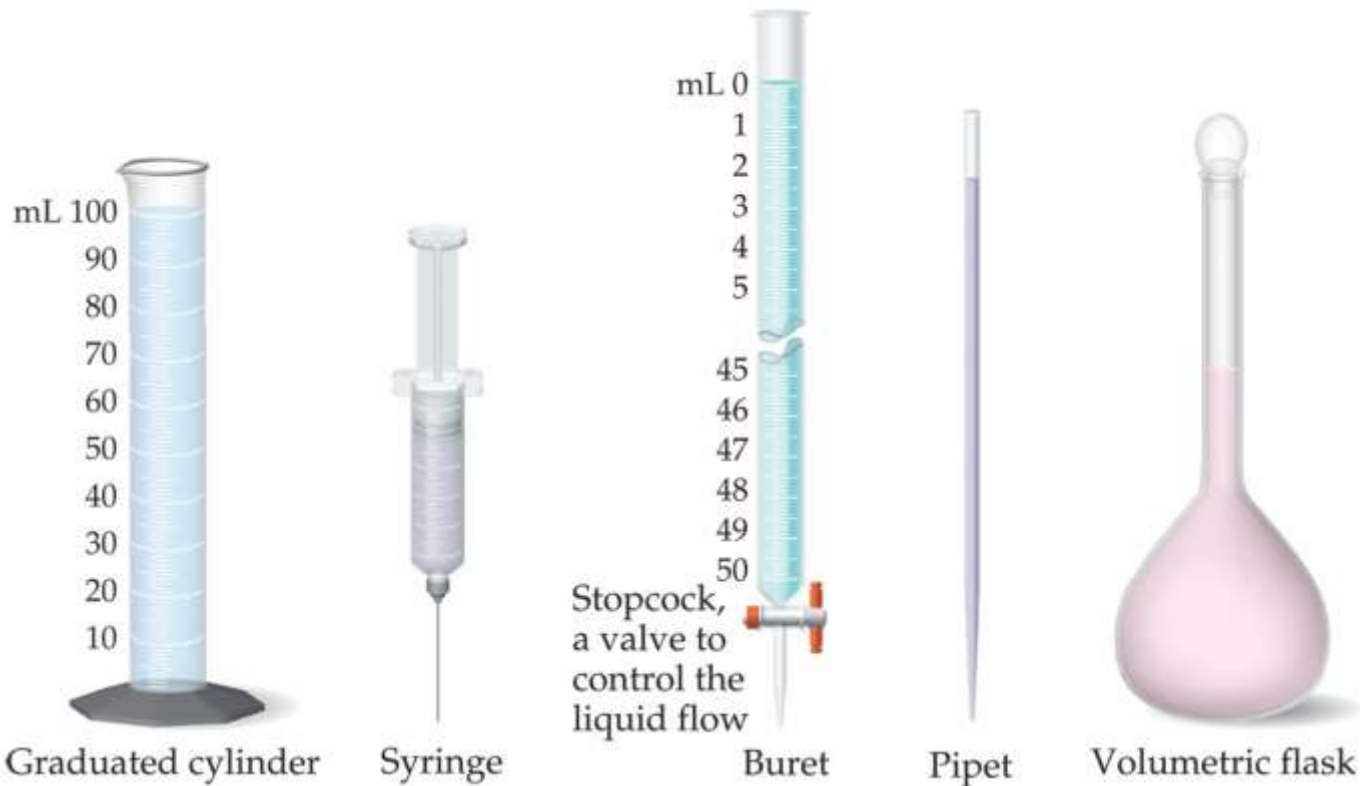
The SI are used to derive the units of other quantities.

**For example:**

Speed is defined as the ratio of distance traveled to elapsed time. Thus, the SI unit for speed is meters per second (m/s).

# Volume

The devices used most frequently in chemistry to measure volume are Syringes, burets, pipets, graduated cylinders and volumetric flasks. Different measuring devices have different uses and different degrees of accuracy



# The SI unit of volume is $\text{m}^3$

The most commonly used metric units for volume are the liter (L) and the milliliter (mL) (not SI units).

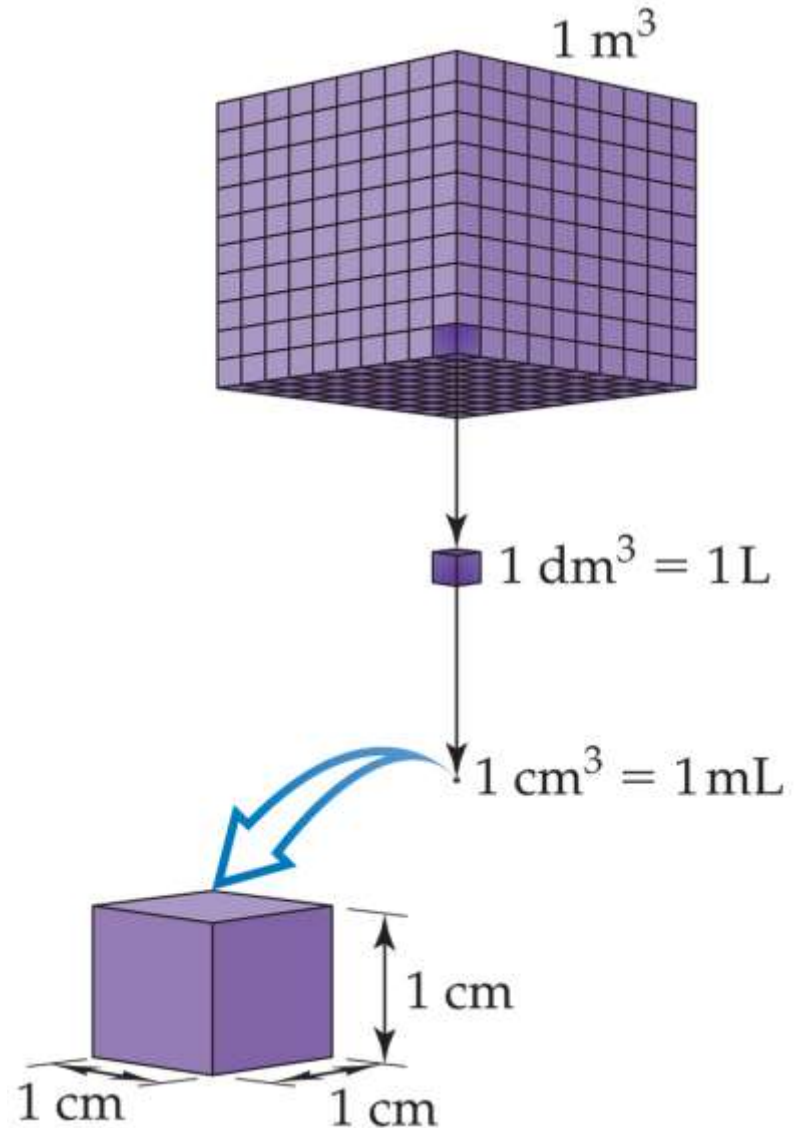
The volume occupied by a cube that is 1 m on each edge is a cubic meter,  $1 \text{ m}^3$

$$1 \text{ m}^3 = 1000 \text{ dm}^3$$

$$1 \text{ L} = 1 \text{ dm}^3$$

$$1 \text{ dm}^3 = 1000 \text{ cm}^3$$

$$1 \text{ cm}^3 = 1 \text{ mL}$$



## Sample Exercise 1.11 Converting Volume Units

Earth's oceans contain approximately  $1.36 \times 10^9 \text{ km}^3$  of water. Calculate the volume in liters.

### Solution

This problem involves conversion of  $\text{km}^3$  to L. From the back inside cover of the text we find  $1 \text{ L} = 10^{-3} \text{ m}^3$ , but there is no relationship listed involving  $\text{km}^3$ . From our knowledge of metric prefixes, however, we have  $1 \text{ km} = 10^3 \text{ m}$  and we can use this relationship between lengths to write the desired conversion factor between volumes:

$$\left(\frac{10^3 \text{ m}}{1 \text{ km}}\right)^3 = \frac{10^9 \text{ m}^3}{1 \text{ km}^3}$$

Thus, converting from  $\text{km}^3$  to  $\text{m}^3$  to L, we have

$$\text{Volume in liters} = (1.36 \times 10^9 \text{ km}^3) \left(\frac{10^9 \text{ m}^3}{1 \text{ km}^3}\right) \left(\frac{1 \text{ L}}{10^{-3} \text{ m}^3}\right) = 1.36 \times 10^{21} \text{ L}$$

# Density

Density is the amount of mass in a unit volume of the substance.

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

The densities of solids and liquids are commonly expressed in units of (g/mL) (not SI unit).

**SI unit of density is kg/m<sup>3</sup>**

Because most substances change volume when they are heated or cooled, densities are temperature dependent.



- The density of water is 1.00 g/mL (mass equal volume).
  - Densities are temperature dependent (because most substances change volume when they are heated or cooled. So, the temperature should be specified).
  - The density and weight are sometimes confused.
- For example; iron has density more than air, but 1kg of air has the same mass as 1kg of iron, but iron occupies a smaller volume, which giving it a higher density.

**TABLE 1.6 ■ Densities of Some Selected Substances at 25 °C**

Substance	Density (g/cm <sup>3</sup> )
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32

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

**(a)** Calculate the density of mercury if  $1.00 \times 10^2$  g occupies a volume of  $7.36 \text{ cm}^3$ . **(b)** Calculate the volume of 65.0 g of the liquid methanol (wood alcohol) if its density is  $0.791 \text{ g/mL}$ . **(c)** What is the mass in grams of a cube of gold (density =  $19.32 \text{ g/cm}^3$ ) if the length of the cube is  $2.00 \text{ cm}$ ?

### 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

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

# Other Derived SI Units

$$\text{Distance} = L = \text{m}$$

$$\text{Area} = L \times L = \text{m}^2$$

$$\text{Volume} = L \times L \times L = \text{m}^3$$

$$\text{Force} = m a = \text{kg} (\text{m/s}^2) = \text{kg m s}^{-2}$$

$$\text{Energy} = \frac{1}{2} m v^2 = \text{kg} (\text{m/s})^2 = \text{kg m}^2 \text{s}^{-2} = \text{J}$$

$$\text{Pressure} = F / A = \text{kg m s}^{-2} / \text{m}^2 = \text{kg m}^{-1} \text{s}^{-2} = \text{Pa}$$



Q & A



At what temperature  $^{\circ}\text{F}$  and  $^{\circ}\text{C}$  are the same?

Answer: -40

If you have equal masses of the following metals, which will occupy the largest volume?

- Au, density =  $19.3 \text{ g/cm}^3$
- Pb, density =  $11.3 \text{ g/cm}^3$
- Ag, density =  $10.5 \text{ g/cm}^3$
- Cu, density =  $8.92 \text{ g/cm}^3$
- Al, density =  $2.70 \text{ g/cm}^3$



Estimate room temperature ( $\sim 72^{\circ}\text{F}$ ) in  $^{\circ}\text{C}$ .

- $\sim 15^{\circ}\text{C}$
- $\sim 22^{\circ}\text{C}$
- $\sim 27^{\circ}\text{C}$
- $\sim 32^{\circ}\text{C}$
- $\sim 37^{\circ}\text{C}$

# Which represents the largest volume?

- 0.25 L
- $2.5 \times 10^2$  mL
- $2.5 \times 10^6$   $\mu$ L
- $2.5 \times 10^8$  nL
- $2.5 \times 10^{10}$  pL

**(a)** Calculate the density of a 374.5-g sample of copper if it has a volume of 41.8 cm<sup>3</sup>. **(b)** A student needs 15.0 g of ethanol for an experiment. If the density of ethanol is 0.789 g/mL, how many milliliters of ethanol are needed? **(c)** What is the mass, in grams, of 25.0 mL of mercury (density = 13.6 g/mL)?

***Answers: (a) 8.96 g/cm<sup>3</sup>, (b) 19.0 mL, (c) 340 g***

