WILEY

Chapter 1
Scientific Measurements

Chemistry, 7th Edition
International Student Version
Brady/Jespersen/Hyslop

Matter and Its Classifications

Matter

- Anything that has mass and occupies space

Mass

- How much matter given object has
- Measure of object's momentum, or resistance to change in motion

Weight

- Force with which object is attracted by gravity

Example: Mass vs. Weight

on moon and on earth

- Weight on moon = $^1/_6$ weight on earth
- Astronaut mass regardless of location

Elements

- Substances that can't be decomposed into simpler materials by chemical reactions
- Substances composed of only one type of atom
- Simplest forms of matter that we can work with directly
- More complex substances composed of elements in various combinations

Compound

- Formed from two or more atoms of different elements
- Always combined in same fixed ratios by mass
- Can be broken down into elements by some chemical changes

e.g., Water decomposed to elemental hydrogen and oxygen

Pure Substance vs. Mixture

Pure substances

- Elements and compounds
- Composition always same regardless of source

Mixture

- Can have variable compositions
- Made up of two or more substances

e.g., CO_2 in water—varying amounts of “fizz” in soda

- Two broad categories of mixtures:
  - Heterogeneous
  - Homogeneous

Homogeneous Mixtures

- Same properties throughout sample
- **Solution**

  e.g.,
  - **Liquid solution**
    - Sugar in water
    - Coca-Cola (without ice)
  - **Gas solution**
    - Air
      - Contains nitrogen, oxygen, carbon dioxide and other gases
  - **Solid solution**
    - US 5¢ coin – Metal alloy
      - Contains copper and nickel metals

Heterogeneous Mixtures

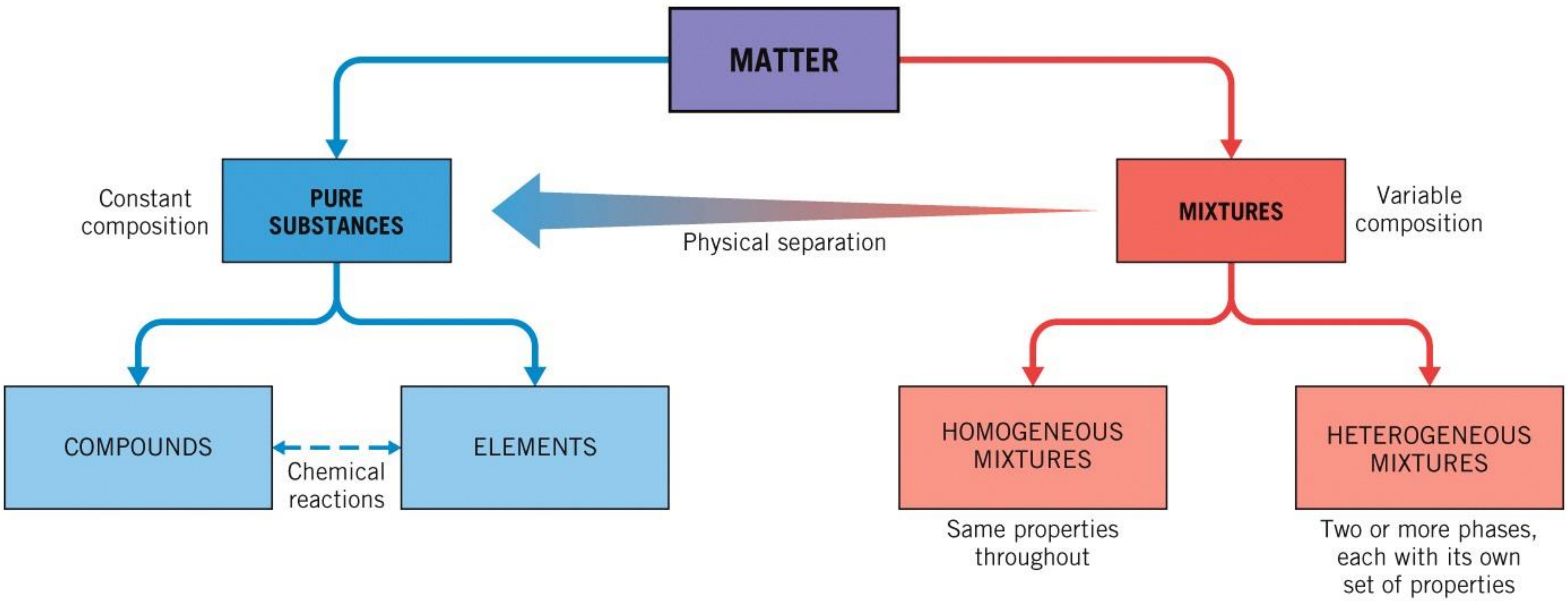
- Two or more regions of different properties
- Solution with multiple phases
- Separate layers

e.g.,

- Salad dressing
  - Oil and vinegar
- Ice and water
  - Same composition
  - Two different physical states

Andy Washnik

Relationship of Elements, Compounds, and Substances



Physical Change

- No new substances formed
- Substance may change state or the proportions

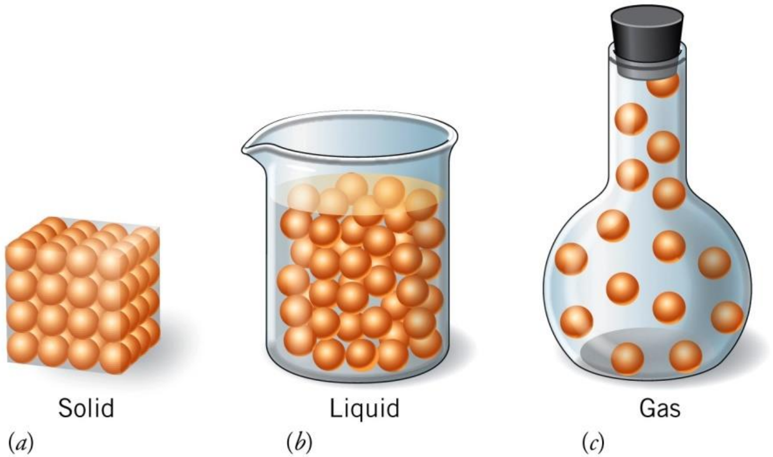
e.g., Ice melting

- Sugar or salt dissolving
- Stirring iron filings and sulfur together

States of Matter

Solids:

- Fixed shape and volume
- Particles are close together
- Have restricted motion

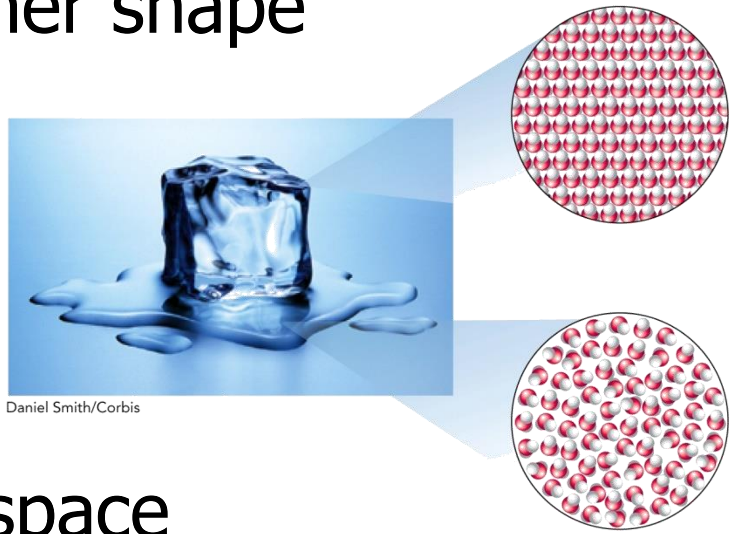


Liquids:

- Fixed volume, but take container shape
- Particles are close together
- Are able to flow

Gases:

- Expand to fill entire container
- Particles separated by lots of space



e.g., Ice, water, steam

Chemical Change

or Chemical Reaction

- Formation of new substance or compound
- Involves changing chemical makeup of substances
- New substance has different physical properties
- Can't be separated by physical means (physical properties)

e.g.,

- Fool's gold
- Compound containing sulfur and iron
  - No longer has same physical properties of free elements
  - Can't be separated using magnet

Learning Check:

- For each of the following, determine if it represents a chemical or physical change:

	Chemical	Physical
Magnesium burns when heated	X	
Magnesium metal tarnishes in air	X	
Magnesium metal melts at 922 K		X
Grape Kool-Aid lightens when water is added		X

Your Turn!

Which one of the following represents a physical change?

A. when treated with bleach, some dyed fabrics change color

B. grape juice left in an open unrefrigerated container turns sour

C. when heated strongly, sugar turns dark brown

D. in cold weather, water condenses on the inside surface of single pane windows

E. when ignited with a match in open air, paper burns

Physical Properties

Properties that can be observed without changing the chemical makeup of a substance

e.g.,

- Color
- Electrical conductivity
- Melting point and boiling point

Chemical Properties

- Describes how a substance undergoes a chemical reaction – its 'reactivity'
  - Involves changing chemical makeup of substances
  - New substance has different physical properties
  - Can't be separated by physical means

e.g.,

  - Fool's gold
  - Compound containing sulfur and iron
    - No longer has same physical properties of free elements
    - Can't be separated using magnet

Intensive vs. Extensive Properties

Intensive properties

- Independent of sample size
- Used to identify substances

e.g., Color
Density
Boiling point
Melting point
Chemical reactivity

Extensive properties

- Depend on sample size

e.g., volume and mass

Your Turn!

Which of the following is an extensive property?

A. Density

B. Melting point

C. Color

D. Temperature

E. Mass

Observations

- Fall into two categories

1. **Quantitative** observations
   - Numeric data
   - Measure with instrument

   e.g., Melting point, boiling point, volume, mass

2. **Qualitative** observations
   - Do not involve numerical information

   e.g., Color, rapid boiling, white solid forms

Measurements Include Units

1. **Measurements involve comparison**
   - Always measure relative to reference

     e.g., Foot, meter, kilogram
   - **Measurement = number + unit + error**

     e.g., Distance between 2 points = 25
     - What unit? inches, feet, yards, miles
     - Meaningless without units
2. **Measurements are inexact**
   - Measuring involves estimation
   - Always have **uncertainty**
   - The observer and instrument have inherent physical limitations

International System of Units (SI)

- Standard system of units used in scientific and engineering measurements
- Metric
- Seven Base Units

TABLE 1.2 The SI Base Units

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

SI Units

- Focus on first six in this book
- All physical quantities will have units derived from these seven SI base units

  e.g., Area
  - Derived from SI units based on definition of area
  - length × width = area
  - meter × meter = area

    $m \times m = m^2$
  - SI unit for area = square meters = m^2

Note: Units undergo same kinds of mathematical operations that numbers do

Learning Check

- What is the SI derived unit for velocity?

$$\text{Velocity } (v) = \frac{\text{distance}}{\text{time}}$$

$$\text{Velocity units} = \frac{\text{meters}}{\text{seconds}} = \frac{\text{m}}{\text{s}}$$

- What is the SI derived unit for volume of a cube?

$$\text{Volume } (V) = \text{length} \times \text{width} \times \text{height}$$

$$V = \text{meter} \times \text{meter} \times \text{meter}$$

$$\boldsymbol{V = \mathbf{m}^3}$$

Some Non-SI Metric Units Commonly Used in Chemistry

TABLE 1.3 Some Non-SI Metric Units Commonly Used in Chemistry

Measurement	Unit	Abbreviation	Value in SI Units
Length	angstrom	Å	1 Å = 0.1 nm = 10^{-10} m
Mass	atomic mass unit	u (amu)	1 u = 1.66054 × 10^{-27} kg (rounded to six digits)
	metric ton	t	1 t = 10^3 kg
Time	minute	min.	1 min. = 60 s
	hour	h	1 h = 60 min. = 3600 s
Temperature	degree Celsius	°C	$T_K = t_{°C} + 273.15$
Volume	liter	L	1 L = 1000 cm^3

Using Decimal Multipliers

- Use prefixes on SI base units when number is too large or too small for convenient usage
- Only commonly used are listed here
  - For more complete list see Table 1.5 in textbook
- Numerical values of multipliers can be interchanged with prefixes

  Example: 1 mL = 10^{-3} L
  - 1 km = 1000 m
  - 1 ng = 10^{-9} g
  - 1,130,000 m = 1.13×10^{6} m = **1.13 Mm**

Decimal Multipliers

TABLE 1.5 SI Prefixes—Their Meanings and Values[a]

Prefix	Meaning	Symbol	Prefix Value[b] (numerical)	Prefix Value[b] (power of ten)
exa		E		10^{18}
peta		P		10^{15}
tera		T		10^{12}
giga	billions of	G	1000000000	10^{9}
mega	millions of	M	1000000	10^{6}
kilo	thousands of	k	1000	10^{3}
hecto		h		10^{2}
deka		da		10^{1}
deci	tenths of	d	0.1	10^{-1}
centi	hundredths of	c	0.01	10^{-2}
milli	thousandths of	m	0.001	10^{-3}
micro	millionths of	μ	0.000001	10^{-6}
nano	billionths of	n	0.000000001	10^{-9}
pico	trillionths of	p	0.000000000001	10^{-12}
femto		f		10^{-15}
atto		a		10^{-18}

[a]Prefixes in red type are used most often.

[b]Numbers in these columns can be interchanged with the corresponding prefix.

Laboratory Measurements

- **Four common**
  1. Length
  2. Volume
  3. Mass
  4. Temperature

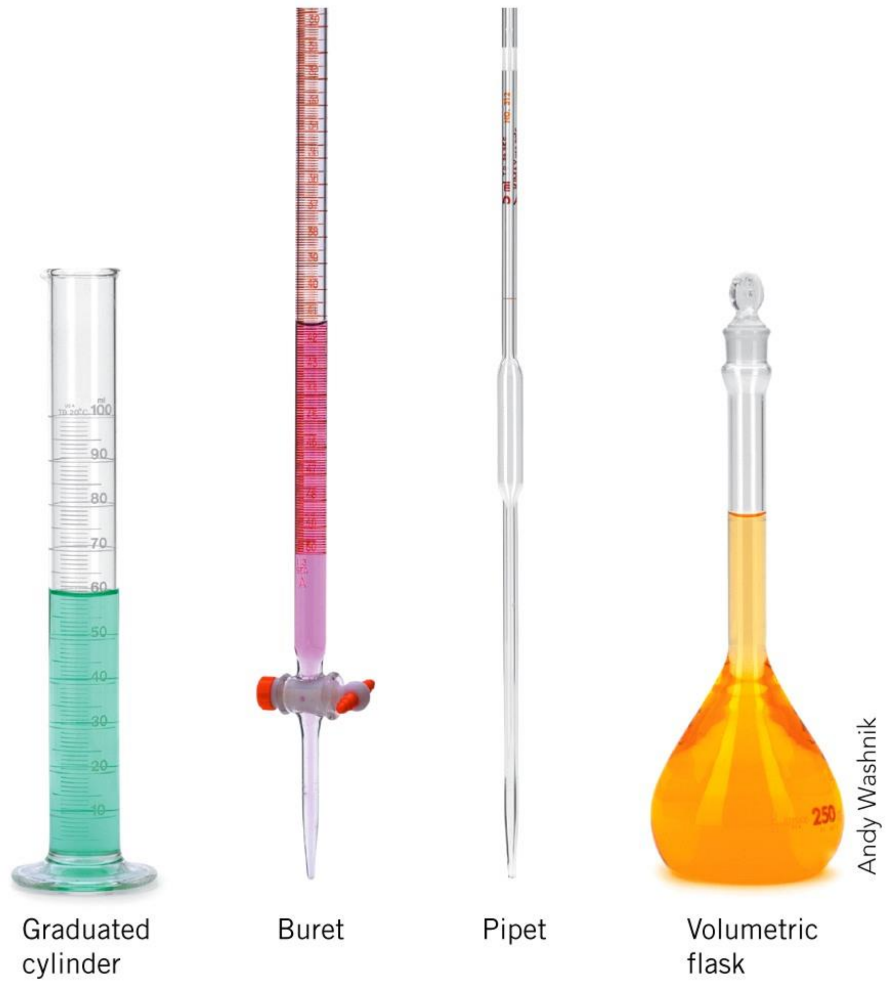
Laboratory Measurements

1. Length

- SI Unit is meter (m)
- Meter too large for most laboratory measurements
- Commonly use
  - Centimeter (cm)
    - 1 cm = 10^{-2} m = 0.01 m
  - Millimeter (mm)
    - 1 mm = 10^{-3} m = 0.001 m

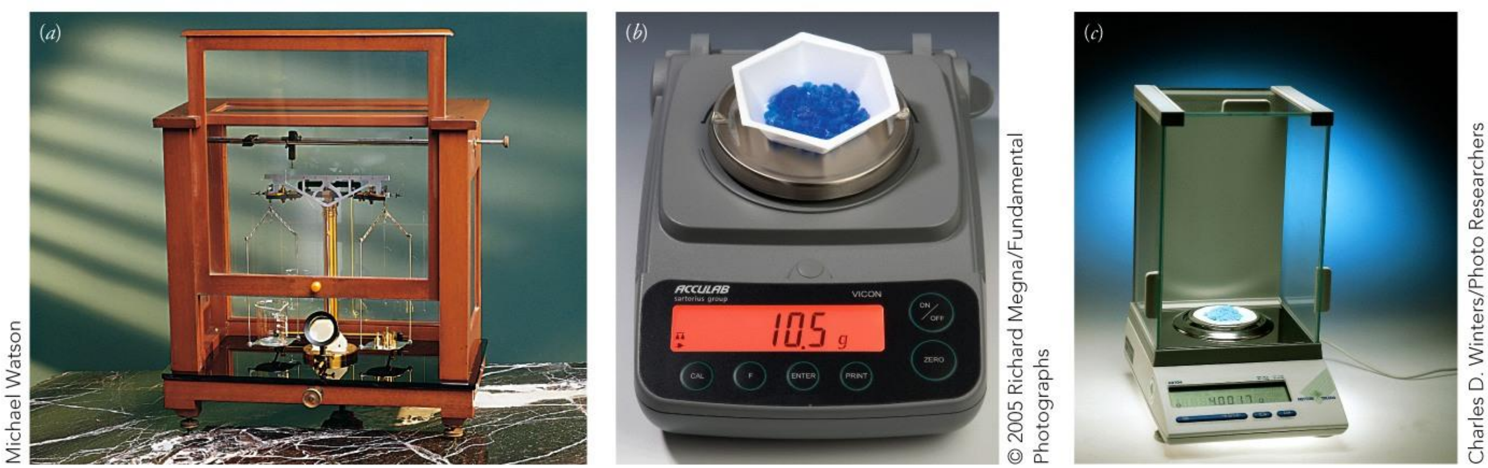
2. Volume

- Dimensions of $(\text{length})^3$
- SI unit for Volume = m^3
- Most laboratory measurements use *V* in liters (L)
  - 1 L = 1 dm^3 (exactly)
- Chemistry glassware marked in L or mL
  - 1 L = 1000 mL
- What is a mL?
  - 1 mL = 1 cm^3



3. Mass

- SI unit is kilogram (kg)
  - Frequently use grams (g) in laboratory as more realistic size
- 1 kg = 1000 g   1 g = 0.1000 kg = $\frac{1}{1000}$ g
- Mass is measured by comparing weight of sample with weights of known standard masses
- Instrument used = balance

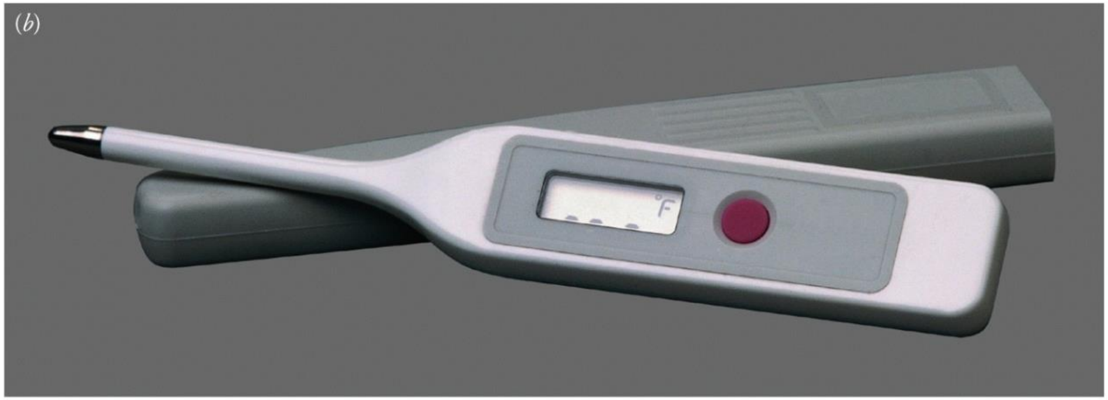


4. Temperature

- Measured with thermometer
- Three common scales

A. Fahrenheit scale

- Common in US
- Water freezes at 32 °F and boils at 212 °F
- 180 degree units between melting and boiling points of water

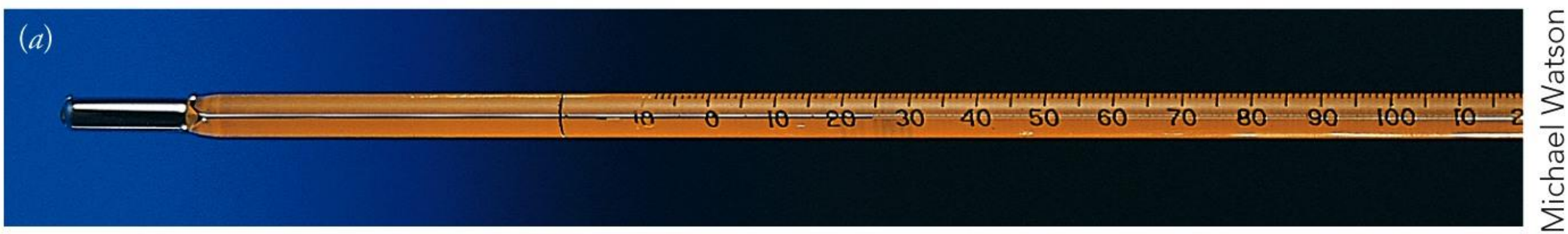


Corbis Images

4. Temperature

B. Celsius scale

- Most common for use in science
- Water freezes at 0 °C
- Water boils at 100 °C
- 100 degree units between melting and boiling points of water



4. Temperature

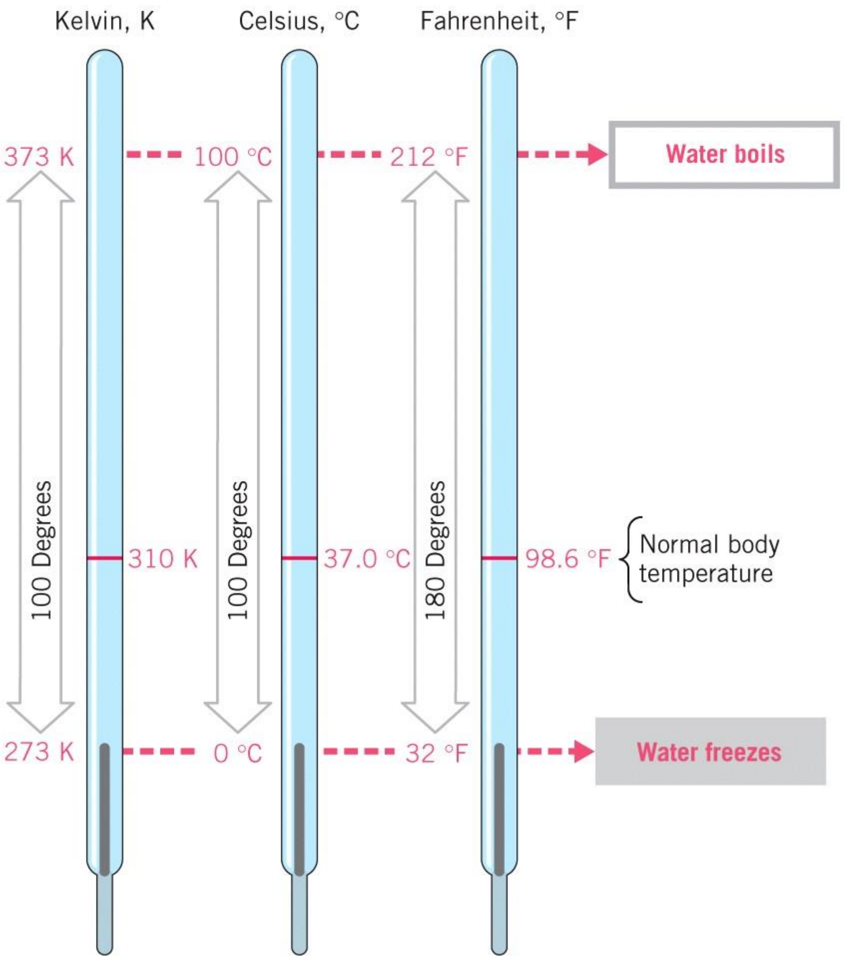
C. Kelvin scale

- SI unit of temperature is **kelvin (K)**
  - **Note:** No degree symbol in front of K
- Water freezes at 273.15 K and boils at 373.15 K
  - 100 degree units between melting and boiling points
- Only difference between Kelvin and Celsius scale is zero point

Absolute Zero

- Zero point on Kelvin scale
- Corresponds to nature's lowest possible temperature

Temperature Conversions



- **How to convert between °F and °C?**

$$\boxed{t_F = \left(\frac{9\ ^\circ F}{5\ ^\circ C}\right) t_C + 32^\circ F}$$

Example: 100 ° C = ? °F

$$t_F = \left(\frac{9\ ^\circ F}{5\ ^\circ C}\right) 100\ ^\circ C + 32^\circ F$$

$$t_F = 212\ ^\circ F$$

Temperature Conversions

- Common laboratory thermometers are marked in Celsius scale
- Must convert to Kelvin scale

$$\boxed{T_K = (t_C + 273.15\ ^\circ\text{C})\frac{1\ \text{K}}{1\ ^\circ\text{C}}}$$

- Amounts to adding 273.15 to Celsius temperature

Example: What is the Kelvin temperature of a solution at 25 °C?

$$T_K = (25\ ^\circ\text{C} + 273.15\ ^\circ\text{C})\frac{1\ \text{K}}{1\ ^\circ\text{C}}$$ **= 298 K**

Learning Check: T Conversions

1. Convert 121 °F to the Celsius scale.

$$t_F = \left(\frac{9\ ^\circ F}{5\ ^\circ C}\right) t_C + 32\ ^\circ F \qquad t_C = \left(t_F - 32\ ^\circ F\right)\left(\frac{5\ ^\circ C}{9\ ^\circ F}\right)$$

$$t_C = \left(121\ ^oF - 32\ ^\circ F\right)\left(\frac{5\ ^\circ C}{9\ ^\circ F}\right)$$ **= 49 ° C**

2. Convert 121 °F to the Kelvin scale.

- We already have in °C so…

$$T_K = (t_C + 273.15\ ^\circ C)\frac{1\ K}{1\ ^\circ C} = (49 + 273.15\ ^\circ C)\frac{1\ K}{1\ ^\circ C}$$

T_K = 332 K

Learning Check: T Conversions

3. Convert 77 K to the Celsius scale.

$$T_K = (t_C + 273.15\ °C)\frac{1\ K}{1\ °C} \qquad t_C = (T_K - 273.15\ K)\frac{1\ °C}{1\ K}$$

$$t_C = (77\ K - 273.15\ K)\frac{1\ °C}{1\ K} = -196\ °C$$

4. Convert 77 K to the Fahrenheit scale.

- We already have in °C so

$$t_F = \left(\frac{9\ °F}{5\ °C}\right)(-196\ °C) + 32\ °F = -321\ °F$$

Your Turn!

In a recent accident some drums of uranium hexafluoride were lost in the English Channel. The melting point of uranium hexafluoride is 64.53 ˚ C. What is the melting point of uranium hexafluoride on the Fahrenheit scale?

A. 67.85 °F
B. 96.53 °F
C. 116.2 °F
D. 337.5 °F
E. 148.2 °F

$$t_F = \left(\frac{9\ °F}{5\ °C}\right) t_C + 32\ °F$$

$$t_F = \left(\frac{9\ °F}{5\ °C}\right) 64.53\ °C + 32\ °F$$

Your Turn!

On an absolute temperature scale, 100 ° F is *not* double 50 ° F (i.e., not 'twice as hot'). What temperature, in ° F, would really be double 50 ° F (hint: convert 50 ° F to K, double the Kelvin temperature, then convert back to ° F)?

A. 560 ° F
B. 25 ° F
C. 200 ° F
D. 283.15 ° F
E. 566 ° F

$$t_C = (50\ °F - 32\ °F)\left(\frac{5\ °C}{9\ °F}\right) = 10\ °C$$

$$10\ ^{o}C + 273.15\ K = 283.15\ K$$

Doubled is 566.3 K

$$566.3\ K - 273.15\ K = 293.15\ ^{o}C$$

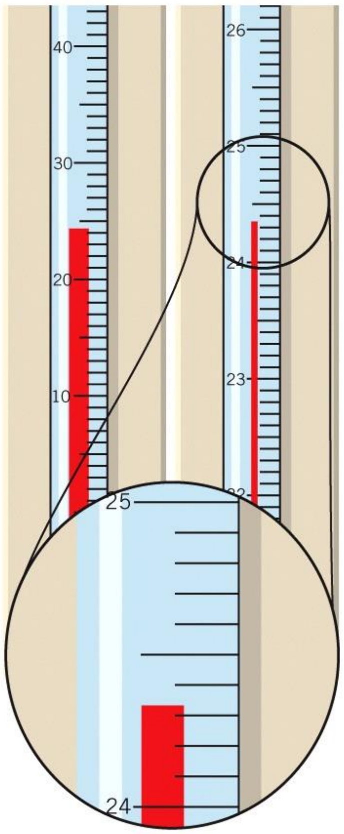
$$t_F = \left(\frac{9\ °F}{5\ °C}\right)293.15\ °C + 32\ °F = 560\ °F$$

Uncertainties in Measurements

- Measurements all inexact
  - Contain uncertainties or errors
- Sources of errors
  - Limitations of reading instrument
- Ways to minimize errors
  - Take series of measurements
  - Data clusters around central value
  - Calculate average or mean values
  - Report average value

Limits in Reading Instruments

- Consider two Celsius thermometers
- **Left** thermometer has markings every 1°C
  - T between 24 °C and 25 °C
  - About 3/10 of way between marks
  - Can estimate to 0.1 °C = uncertainty
  - T = 24.3 ± 0.1 °C
- **Right** thermometer has markings every 0.1 ° C
  - T reading between 24.3 °C and 24.4 °C
  - Can estimate 0.01 °C
  - T = 24.32 ± 0.01 °C



Limits in Reading Instruments

- Finer graduations in markings
  - Means smaller uncertainties in measurements
- Reliability of data
  - Indicated by number of digits used to represent it
- **What about digital displays?**
  - Mass of beaker = 65.23 g on digital balance
  - Still has uncertainty
  - Assume ½ in last readable digit
  - Record as 65.230 ± 0.005 g

Significant Figures

- Scientific convention:
  - **All digits in measurement up to and including first estimated digit are significant.**
- Number of certain digits plus first uncertain digit
- Digits in measurement from first non-zero number on left to first estimated digit on right

Rules for Significant Figures

1. All non-zero numbers are significant.

   e.g., 3.456 has **4** sig. figs.

2. Zeros between non-zero numbers are significant.

   e.g., 20,089 or 2.0089×10^4 has **5** sig. figs

3. Trailing zeros always count as significant if number has decimal point

   e.g., 500. or 5.00×10^2 has **3** sig. figs

Rules for Significant Figures

4. Final zeros on number without decimal point are NOT significant

   e.g., 104,956,000 or 1.04956×10^8

   has **6** sig. figs.

5. Final zeros to right of decimal point are significant

   e.g., 3.00 has **3** sig. figs.

6. Leading zeros, to left of first nonzero digit, are never counted as significant

   e.g., 0.00012 or 1.2×10^{-4} has **2** sig. figs.

Learning Check

How many significant figures does each of the following numbers have?

		Scientific Notation	# of Sig. Figs.
1.	413.97	4.1397×10^2	**5**
2.	0.0006	6×10^{-4}	**1**
3.	5.120063	5.120063	**7**
4.	161,000	1.61×10^5	**3**
5.	3600.	3.600×10^3	**4**

Your Turn!

How many significant figures are in 19.0000?

A. 2
B. 3
C. 4
D. 5
E. 6

Your Turn!

How many significant figures are in 0.0005650850?

A. 7
B. 8
C. 9
D. 10
E. 11

- Could be rewritten as 5.650850×10^{-4}

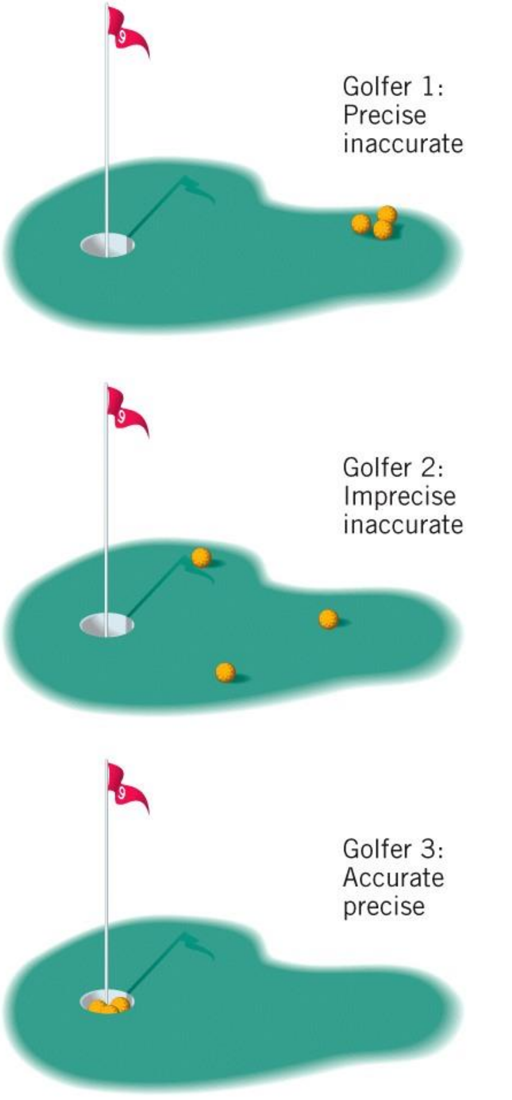
Accuracy and Precision

Accuracy

- How close measurement is to true or accepted true value
- Measuring device must be calibrated with standard reference to give correct value

Precision

- How well set of repeated measurements of same quantity agree with each other
- More significant figures equals more precise measurement



Rounding to Correct Digit

1. If digit to be dropped is **greater than 5**, last remaining digit is **rounded up**.

   e.g., 3.67<u>7</u> is rounded up to 3.68

2. If number to be dropped is **less than 5**, last remaining digit **stays the same**.

   e.g., 6.63<u>2</u> is rounded to 6.63

3. If number to be dropped is exactly 5, then if digit to left of 5 is

   a. Even, it **remains the same.**

   e.g., 6.6<u>5</u> is rounded to 6.6

   b. Odd, it **rounds up.**

   e.g., 6.3<u>5</u> is rounded to 6.4

Scientific Notation

- Clearest way to present number of significant figures unambiguously
  - Report number between 1 and 10 followed by correct power of 10
  - Indicates only significant digits

e.g., 75,000 people attend a concert

- If a rough estimate
  - Uncertainty $\pm$ 1000 people
  - 7.5×10^4
- If number estimated from aerial photograph
  - Uncertainty $\pm$100 people
  - 7.50×10^4

Learning Check

Round each of the following to three significant figures. Use scientific notation where needed.

1. 37.459 — **37.5 or 3.75×10^1**
2. 5431978 — **5.43×10^6**
3. 132.7789003 — **133 or 1.33×10^2**
4. 0.00087564 — **8.77×10^{-4}**
5. 7.665 — **7.66**

Your Turn!

Round 0.00564458 to four significant figures and express using scientific notation.

A. 5.64×10^{-2}
B. 5.000×10^{-3}
C. 5.645×10^{-4}
D. 0.56446
E. 5.645×10^{-3}

Significant Figures in Calculations

Multiplication and Division

- Number of significant figures in answer = number of significant figures in **least precise** measurement

e.g., $10.54 \times 31.4 \times 16.987$ **= 5620 = 5.62×10^3**

4 sig. figs. × 3 sig. figs. × 5 sig. figs. = 3 sig. figs.

e.g., $5.896 \div 0.008$ **= 700 = 7×10^2**

4 sig. figs. ÷ 1 sig. fig. = 1 sig. fig.

Your Turn!

Give the value of the following calculation to the correct number of significant figures.

$$\left(\frac{635.4 \times 0.0045}{2.3589}\right)$$

A. 1.21213

B. 1.212

C. 1.212132774

D. 1.2

E. 1

Significant Figures in Calculations

Addition and Subtraction

- Answer has same number of decimal places as quantity with **fewest number** of decimal places.

e.g.,	12.9753	4 decimal places
	319.5	1 decimal place
	+ 4.398	3 decimal places
	336.9	1 decimal place

e.g.,	397	0 decimal places
	– 273.15	2 decimal places
	124	0 decimal place

Learning Check

For each calculation, give the answer to the correct number of significant figures.

1. 10.0 g + 1.03 g + 0.243 g = **11.3 g** or **1.13×10^{1} g**

2. 19.556 ° C – 19.552 ° C = **0.004 ° C** or **4×10^{-3} ° C**

3. 327.5 m × 4.52 m = **1.48×10^{3} m**

4. 15.985 g ÷ 24.12 mL = **0.6627 g/mL** or **6.627×10^{-1} g/mL**

Your Turn!

When the expression,

412.272 + 0.00031 – 1.00797 + 0.000024 + 12.8

is evaluated, the result should be expressed as:

A. 424.06

B. 424.064364

C. 424.1

D. 424.064

E. 424

Your Turn!

When the expression,

152.6736 - 0.0028 – 125.00797 + 56.795 + 12.853

is evaluated, the result should be expressed to the:

A. tenths place

B. hundredths place

C. thousandths place

D. ten-thousandths place

E. hundred-thousandths place

Learning Check

For the following calculation, give the answer to the correct number of significant figures.

1. $$\frac{(71.359\text{ m} - 71.357\text{ m})}{(3.2\text{ s} \times 3.67\text{ s})} = \frac{(0.00\underline{2}\text{ m})}{(1\underline{1}.744\text{ s}^2)}$$

$$= \mathbf{2 \times 10^{-4}\ m/s^2}$$

2. $$\frac{(13.674\text{ cm} \times 4.35\text{ cm} \times 0.35\text{ cm})}{(856\text{ s} + 1531.1\text{ s})}$$

$$= \frac{(2\underline{0}.818665\text{ cm}^3)}{(238\underline{7}.1\text{ s})} = \mathbf{0.0087\ cm^3/s}$$

Exact Numbers

- Numbers that come from definitions
  - 12 in. = 1 ft
  - 60 s = 1 min
- Numbers that come from direct count
  - Number of people in small room
- Have no uncertainty
- Assume they have infinite number of significant figures
- Do not affect number of significant figures in multiplication or division

Your Turn!

Express the answer to the following with the correct number of significant figures:

$$36.00\ \text{in} \times \left(\frac{\text{ft.}}{12\ \text{in.}}\right) =$$

A. 3 ft.

B. 3.0 ft

C. 3.00 ft.

D. 3.000 ft.

E. 3.00000 ft.

36.00 has four sig. figs.
The conversion is exact.
So the answer has four sig. figs.

Your Turn!

For the following calculation, give the answer to the correct number of significant figures.

$$\frac{(14.5\ \text{cm} \times 12.334\ \text{cm})}{(2.223\ \text{cm} - 1.04\ \text{cm})}$$

A. 179 cm^2

B. 1.18 cm

C. 151.2 cm

D. 151 cm

E. 178.843 cm^2

$$\frac{(17\underline{8}.843\ \text{cm}^2)}{(1.1\underline{8}3\ \text{cm})}$$

Dimensional Analysis

- Also called the Factor Label Method
- Not all calculations use specific equation
- Use units (dimensions) to analyze problem

Conversion Factor

- Fraction formed from valid equality or equivalence between units
- Used to switch from one system of measurement and units to another

$$\text{Given Quantity} \times \text{Conversion Factor} = \text{Desired Quantity}$$

Conversion Factors

Example: How to convert a person's height from 68.0 in to cm?

- Start with fact

  2.54 cm = 1 in. (exact)

- Dividing both sides by 1 in. or 2.54 cm gives 1

$$\frac{2.54\ \text{cm}}{1\ \text{in.}} = \frac{1\ \cancel{\text{in.}}}{1\ \cancel{\text{in.}}} = \mathbf{1} \qquad \frac{2.54\ \cancel{\text{cm}}}{2.54\ \cancel{\text{cm}}} = \frac{1\ \text{in.}}{2.54\ \text{cm}} = \mathbf{1}$$

- Cancel units
- Leave ratio that equals 1
- Use fact that units behave as numbers do in mathematical operations

Dimensional Analysis

- Now multiply original number by conversion factor that cancels old units and leaves new

Given Quantity × Conversion Factor = Desired Quantity

$$68.0\ \cancel{\text{in.}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in.}}} = \mathbf{173\ cm}$$

- Dimensional analysis can tell us when we have done wrong arithmetic

$$68.0\ \text{in.} \times \frac{1\ \text{in.}}{2.54\ \text{cm}} = \mathbf{26.8\ in^2/cm}$$

  - Units not correct

Using Dimensional Analysis

Example: Convert 0.097 m to mm.

- Relationship is 1 mm = 1 × 10^{-3} m
- Can make two conversion factors

$$\frac{1\ \text{mm}}{1 \times 10^{-3}\ \text{m}} \qquad\qquad \frac{1 \times 10^{-3}\ \text{m}}{1\ \text{mm}}$$

- Since going from **m** to **mm** use one on left.

$$0.097\ \cancel{\text{m}} \times \frac{1\ \text{mm}}{1 \times 10^{-3}\ \cancel{\text{m}}}\ \mathbf{= 97\ cm}$$

Learning Check

Example: Convert 3.5 m^3 to cm^3.

- Start with basic equality 1 cm = 0.01 m
- Now cube both sides
  - Units and numbers
  - $(1\ cm)^3 = (0.01\ m)^3$
  - $1\ cm^3 = 1 \times 10^{-6}\ m^3$
- Can make two conversion factors

$$\frac{1\ cm^3}{1 \times 10^{-6}\ m^3} \quad \text{or} \quad \frac{1 \times 10^{-6}\ m^3}{1\ cm^3}$$

$$3.5\ \cancel{m^3} \times \frac{1\ cm^3}{1 \times 10^{-6}\ \cancel{m^3}} = \mathbf{3.5 \times 10^6\ cm^3}$$

Non-metric to Metric Units

Convert speed of light from 3.00×10^8 m/s to mi/hr

- Use dimensional analysis
- 1 min = 60 s   60 min = 1 hr
- 1 km = 1000 m   1 mi = 1.609 km

$$\frac{3.00 \times 10^8 \text{ m}}{\cancel{\text{s}}} \times \frac{60 \cancel{\text{ s}}}{1 \cancel{\text{ min}}} \times \frac{60 \cancel{\text{ min}}}{1 \text{ hr}} = 1.08 \times 10^{12} \text{ m/hr}$$

$$\frac{1.08 \times 10^{12} \cancel{\text{ m}}}{\text{hr}} \times \frac{1 \cancel{\text{ km}}}{1000 \cancel{\text{ m}}} \times \frac{1 \text{ mi}}{1.609 \cancel{\text{ km}}} =$$

6.71×10^8 mi/hr

Your Turn!

The Honda Insight hybrid electric car has a gas mileage rating of 56 miles to the gallon. What is this rating expressed in units of kilometers per liter?

1 gal = 3.784 L   1 mile = 1.609 km

A. 1.3×10^2 km L^{-1}

B. 24 km L^{-1}

C. 15 km L^{-1}

D. 3.4×10^2 km L^{-1}

E. 9.2 km L^{-1}

$$56\,\frac{\cancel{\text{mi}}}{\cancel{\text{gal}}} \times \frac{1\,\cancel{\text{gal}}}{3.784\ \text{L}} \times \frac{1.609\ \text{km}}{1\,\cancel{\text{mi}}}$$

Your Turn!

The volume of a basketball is 433.5 in^3. Convert this to mm^3.

A. 1.101×10^{-2} mm^3

B. 7.104×10^{6} mm^3

C. 7.104×10^{4} mm^3

D. 1.101×10^{4} mm^3

E. 1.101×10^{6} mm^3

$$433.5\,\cancel{\text{in}^3} \times \left(\frac{2.54\,\cancel{\text{cm}}}{\cancel{\text{in}}}\right)^{\cancel{3}} \times \left(\frac{\cancel{\text{m}}}{100\,\cancel{\text{cm}}}\right)^{\cancel{3}} \times \left(\frac{1000\,\text{mm}}{\cancel{\text{m}}}\right)^{\cancel{3}}$$

Density

- Ratio of object's mass to its volume

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad d = \frac{m}{V}$$

- Intensive property (size independent)
  - Determined by taking ratio of two extensive properties (size dependent)
  - Frequently ratio of two size dependent properties leads to size independent property
  - Sample size cancels
- Units
  - **g/mL** or **g/cm^3**

Learning Check

- A student weighs a piece of gold that has a volume of 11.02 cm^3 of gold. She finds the mass to be 212 g. What is the density of gold?

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

$$d = \frac{212\ \text{g}}{11.02\ \text{cm}^3} = \mathbf{19.3\ g/cm^3}$$

Another student has a piece of gold with a volume of 1.00 cm^3. What does it weigh? **19.3 g**

What if it were 2.00 cm^3 in volume? **38.6 g**

Density

- Most substances expand slightly when heated
  - Same mass
  - Larger volume
  - Less dense
- Density decreases slightly as temperature increases
- Liquids and solids
  - Change is very small
  - Can ignore except in very precise calculations
- Density useful to transfer between mass and volume of substance

Learning Check

1. Glass has a density of 2.2 g/cm^3. What is the volume occupied by 35 g of glass?

$$V = \frac{m}{d} = \frac{35\text{ g}}{2.2\text{ g/cm}^3} = \textbf{16. g}$$

2. What is the mass of 400. cm^3 of glass?

$$m = d \times V = 2.2\text{ g/cm}^3 \times 400.\text{ cm}^3 = \mathbf{8.8 \times 10^2\ cm^3}$$

Your Turn!

Titanium is a metal used to make artificial joints. It has a density of 4.54 g/cm^3. What volume will a titanium hip joint occupy if its mass is 205 g?

A. 9.31×10^{2} cm^3

B. 4.51×10^{1} cm^3

C. 2.21×10^{-2} cm^3

D. 1.07×10^{-3} cm^3

E. 2.20×10^{-1} cm^3

$$V = \frac{205\ \text{g}}{4.54\ \text{g/cm}^3} =$$

Your Turn!

A sample of zinc metal (density = 7.14 g cm^{-3}) was submerged in a graduated cylinder containing water. The water level rose from 162.5 cm^3 to 186.0 cm^3 when the sample was submerged. How many grams did the sample weigh?

A. 1.16×10^3 g

B. 1.33×10^3 g

C. 23.5 g

D. 1.68×10^2 g

E. 3.29 g

$$\text{mass} = \text{density} \times \text{volume}$$

$$\text{volume} = (186.0\ \text{cm}^3 - 162.5\ \text{cm}^3)$$

$$= 23.5\ \text{cm}^3$$

$$\text{mass} = 7.14\ \text{g cm}^{-3} \times 23.5\ \text{cm}^3$$