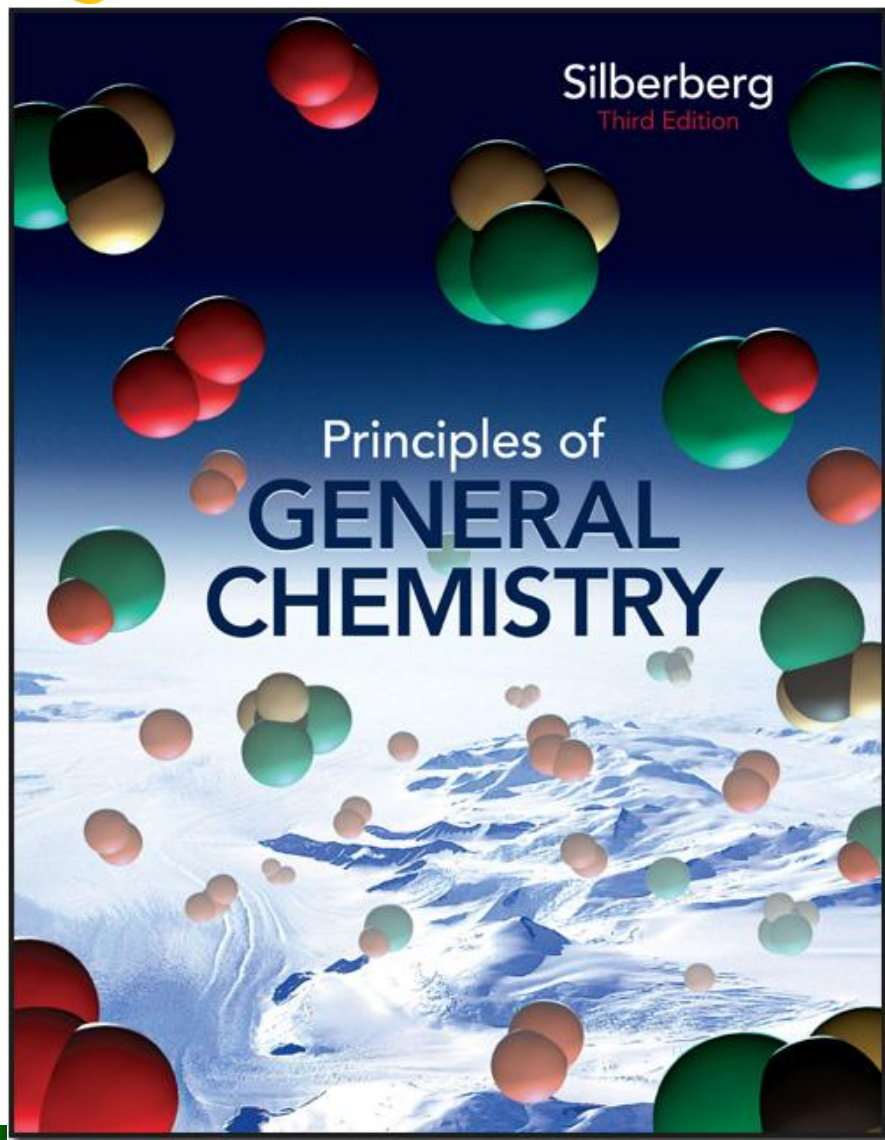




Dr. Diab Qadah

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CHEM 141
Chapter 1

Chapter 1

Keys to the Study of Chemistry



Chapter 1: Keys to the Study of Chemistry

1.1 Some Fundamental Definitions

1.2 The Scientific Approach: Developing a Model

1.3 Chemical Problem Solving

1.4 Measurement in Scientific Study

1.5 Uncertainty in Measurement: Significant Figures

Chemistry

Chemistry is the study of ***matter***,
its ***properties***,
the ***changes*** that matter undergoes,
and
the ***energy*** associated with these changes.



Definitions

Matter	anything that has both mass and volume - the “stuff” of the universe: books, planets, trees, etc.
Composition	the types and amounts of simpler substances that make up a sample of matter
Properties	the characteristics that give each substance a unique identity



Physical Properties

properties a substance shows by itself without interacting with another substance

- color, melting point, boiling point, density

Chemical Properties

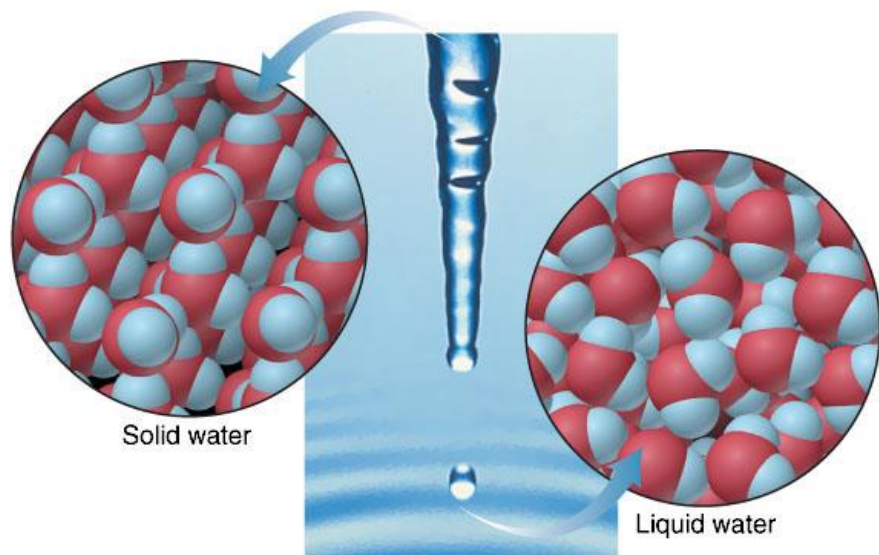
properties a substance shows as it interacts with, or transforms into, other substances

- flammability, corrosiveness



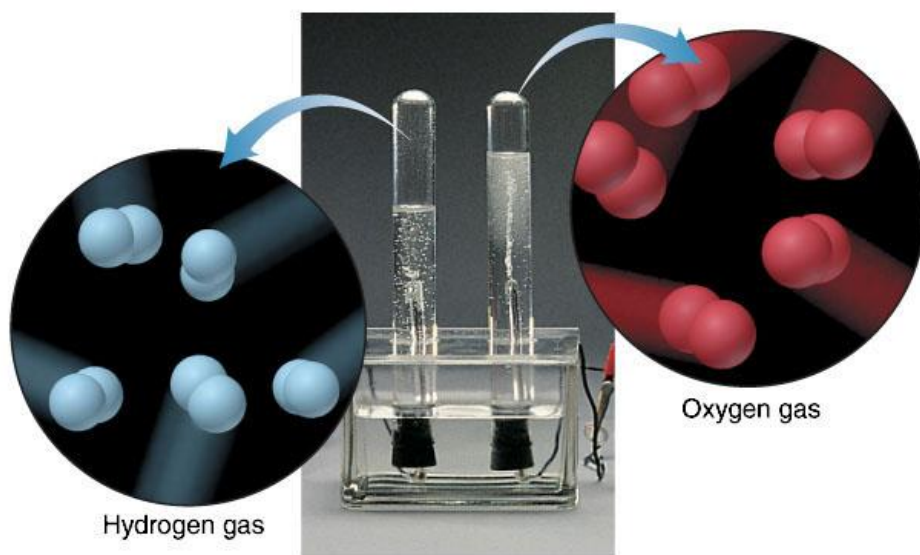
Figure 1.1 The distinction between physical and chemical change.

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A Physical change:

Solid form of water becomes liquid form; composition does *not* change because particles are the same.



B Chemical change:

Electric current decomposes water into different substances (hydrogen and oxygen); composition *does* change because particles are different.

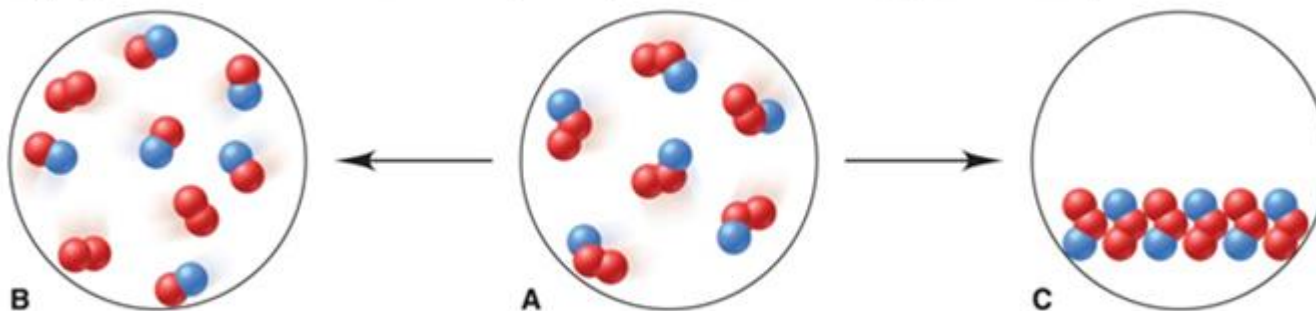


Sample Problem 1.1

Visualizing Change on the Atomic Scale

PROBLEM: The scenes below represent an atomic-scale view of substance A undergoing two different changes. Decide whether each scene shows a physical or a chemical change.

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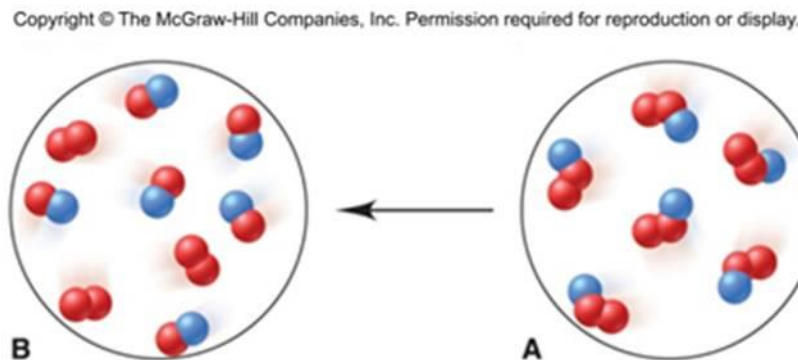


PLAN: We need to determine what change is taking place. The numbers and colors of the little spheres that represent each particle tell its “composition”. If the composition does not change, the change is physical, whereas a chemical change results in a change of composition.



Sample Problem 1.1

SOLUTION:



Each particle of substance A is composed of one blue and two red spheres.

Sample B is composed of two different types of particles – some have two red spheres while some have one red and one blue.

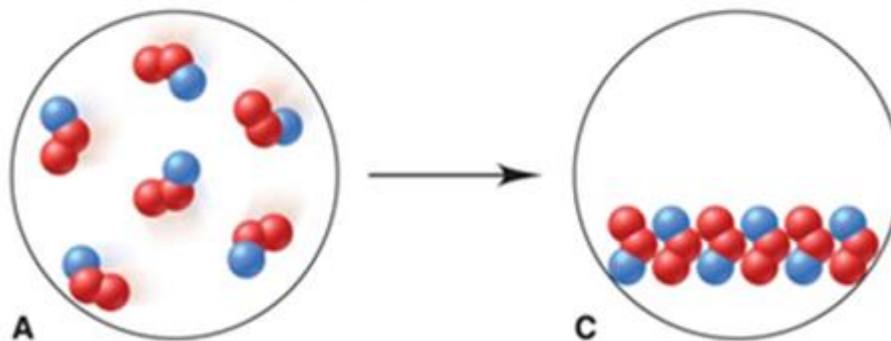
As A changes to B, the chemical composition has changed.

$A \rightarrow B$ is a **chemical change**.



Sample Problem 1.1

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Each particle of C is still composed of one blue and two red spheres, but the particles are closer together and are more organized. The composition remains unchanged, but the physical form is different.

A → C is a **physical change**.



The States of Matter

A **solid** has a fixed shape and volume. Solids may be hard or soft, rigid or flexible.

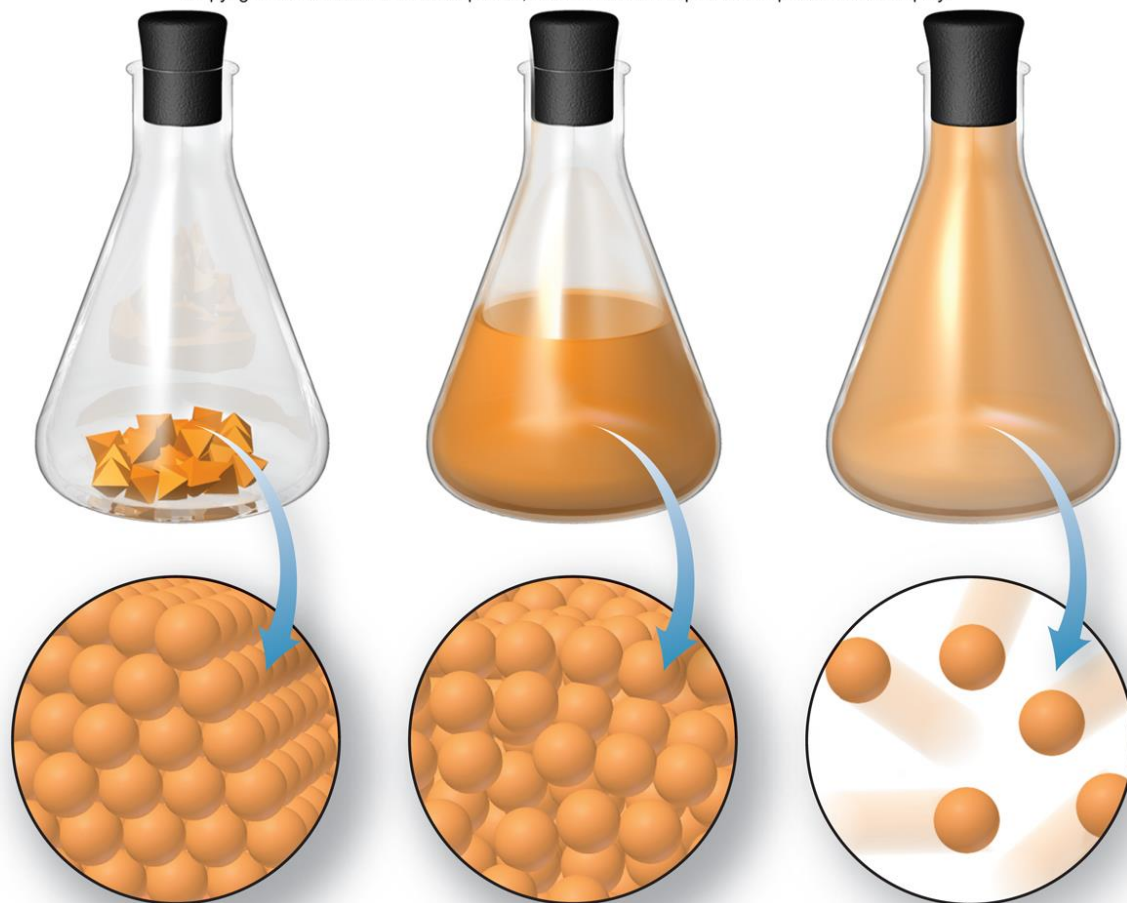
A **liquid** has a varying shape that conforms to the shape of the container, but a fixed volume. A liquid has an **upper surface**.

A **gas** has no fixed shape or volume and therefore does not have a surface.



Figure 1.2 The physical states of matter.

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Solid

Particles are close together and organized.

Liquid

Particles are close together but disorganized.

Gas

Particles are far apart and disorganized.



Temperature and Change of State

- A change of state is a ***physical*** change.
 - Physical form changes, composition does not.
- Changes in physical state are ***reversible***
 - by changing the temperature.
- A chemical change cannot simply be reversed by a change in temperature.



Sample Problem 1.2

Distinguishing Between Physical and Chemical Change

PROBLEM: Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:

- (a) Frost forms as the temperature drops on a humid winter night.
- (b) A cornstalk grows from a seed that is watered and fertilized.
- (c) A match ignites to form ash and a mixture of gases.
- (d) Perspiration evaporates when you relax after jogging.
- (e) A silver fork tarnishes slowly in air.

PLAN: “Does the substance change composition or just change form?”



Sample Problem 1.2

SOLUTION:

(a) Frost forms as the temperature drops on a humid winter night.

physical change

(b) A cornstalk grows from a seed that is watered and fertilized.

chemical change

(c) A match ignites to form ash and a mixture of gases.

chemical change

(d) Perspiration evaporates when you relax after jogging.

physical change

(e) A silver fork tarnishes slowly in air.

chemical change



Energy in Chemistry

Energy is the ability to do work.

Potential Energy

is energy due to the ***position*** of an object.

Kinetic Energy

is energy due to the ***movement*** of an object.

Total Energy = Potential Energy + Kinetic Energy



Energy Changes

Lower energy states are **more stable** and are favored over higher energy states.

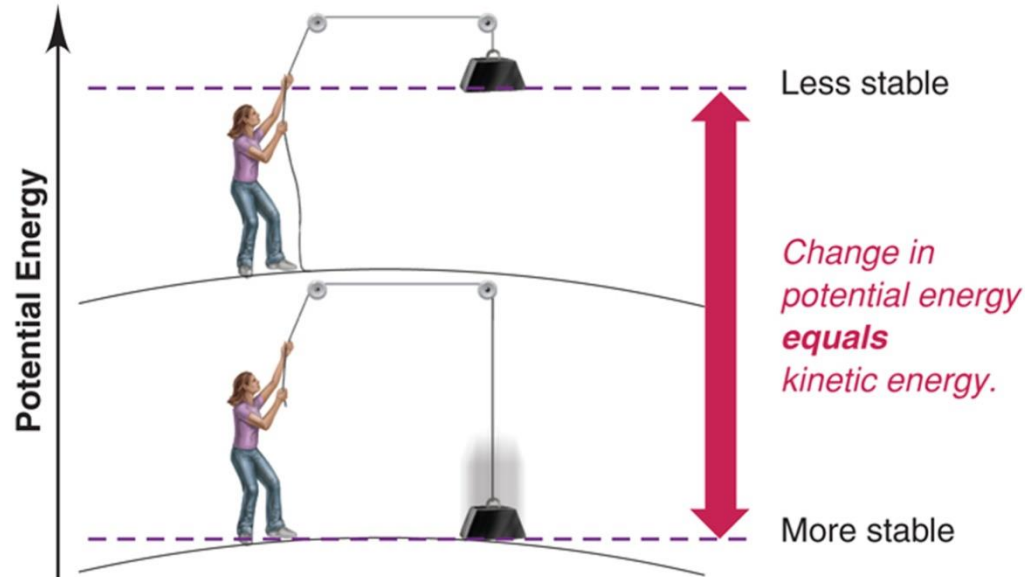
Energy is neither created nor destroyed

- it is **conserved**
- and can be **converted** from one form to another.



Figure 1.3A Potential energy is converted to kinetic energy.

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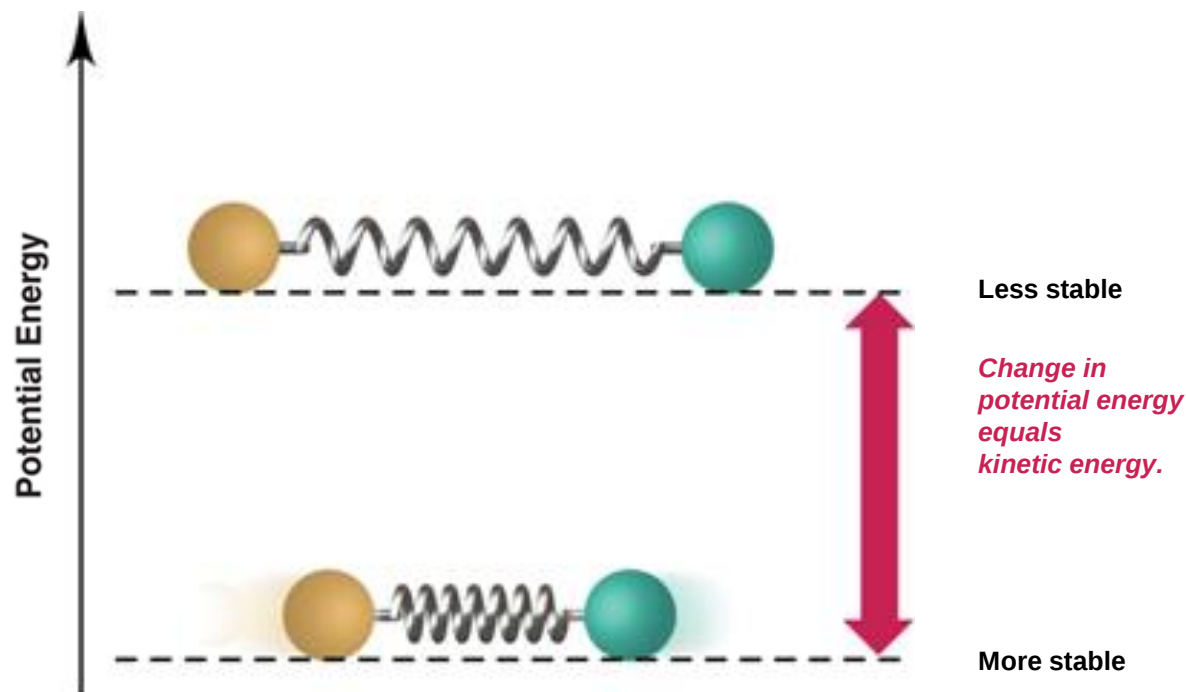
A A gravitational system. Potential energy is gained when a weight is lifted. It is converted to kinetic energy as the weight falls.

A gravitational system. The potential energy gained when a lifted weight is converted to kinetic energy as the weight falls.

A lower energy state is more stable.



Figure 1.3B Potential energy is converted to kinetic energy.



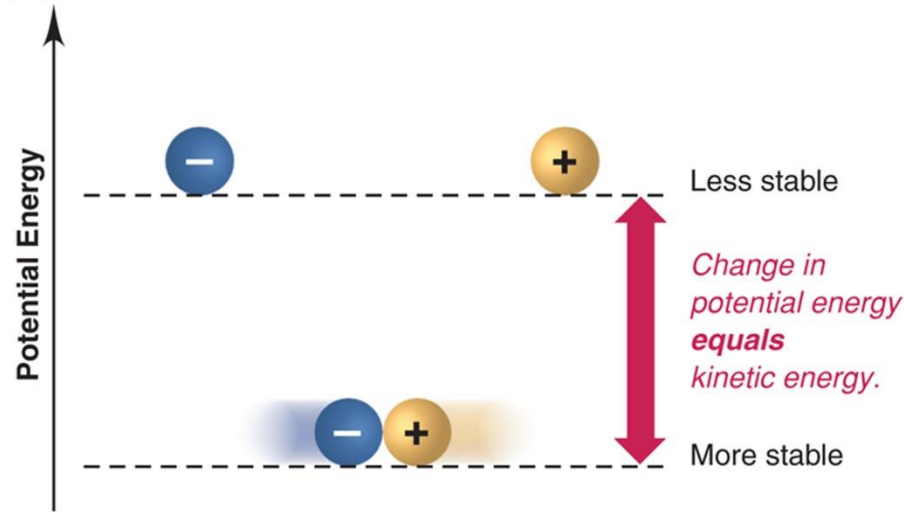
A system of two balls attached by a spring. The potential energy gained by a stretched spring is converted to kinetic energy when the moving balls are released.

Energy is conserved when it is transformed.



Figure 1.3C Potential energy is converted to kinetic energy.

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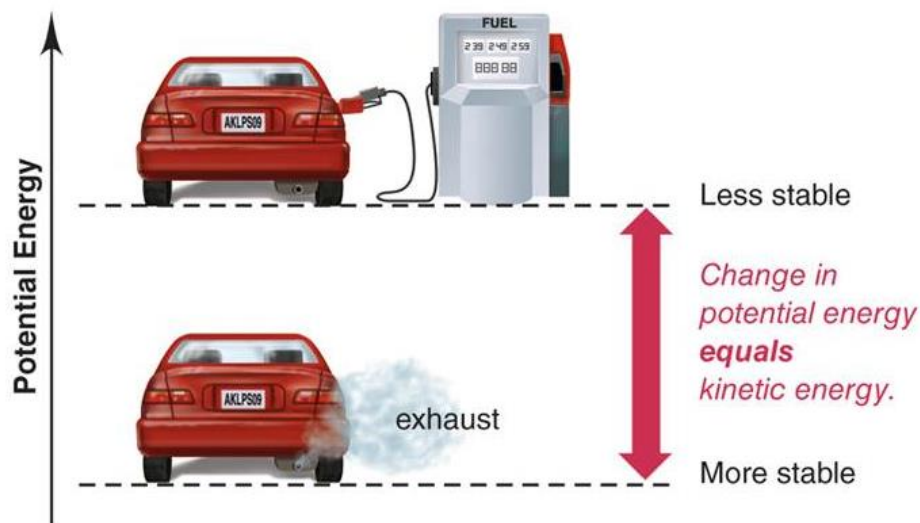


C A system of oppositely charged particles. Potential energy is gained when the charges are separated. It is converted to kinetic energy as the attraction pulls the charges together.

A system of oppositely charged particles. The potential energy gained when the charges are separated is converted to kinetic energy as the attraction pulls these charges together.



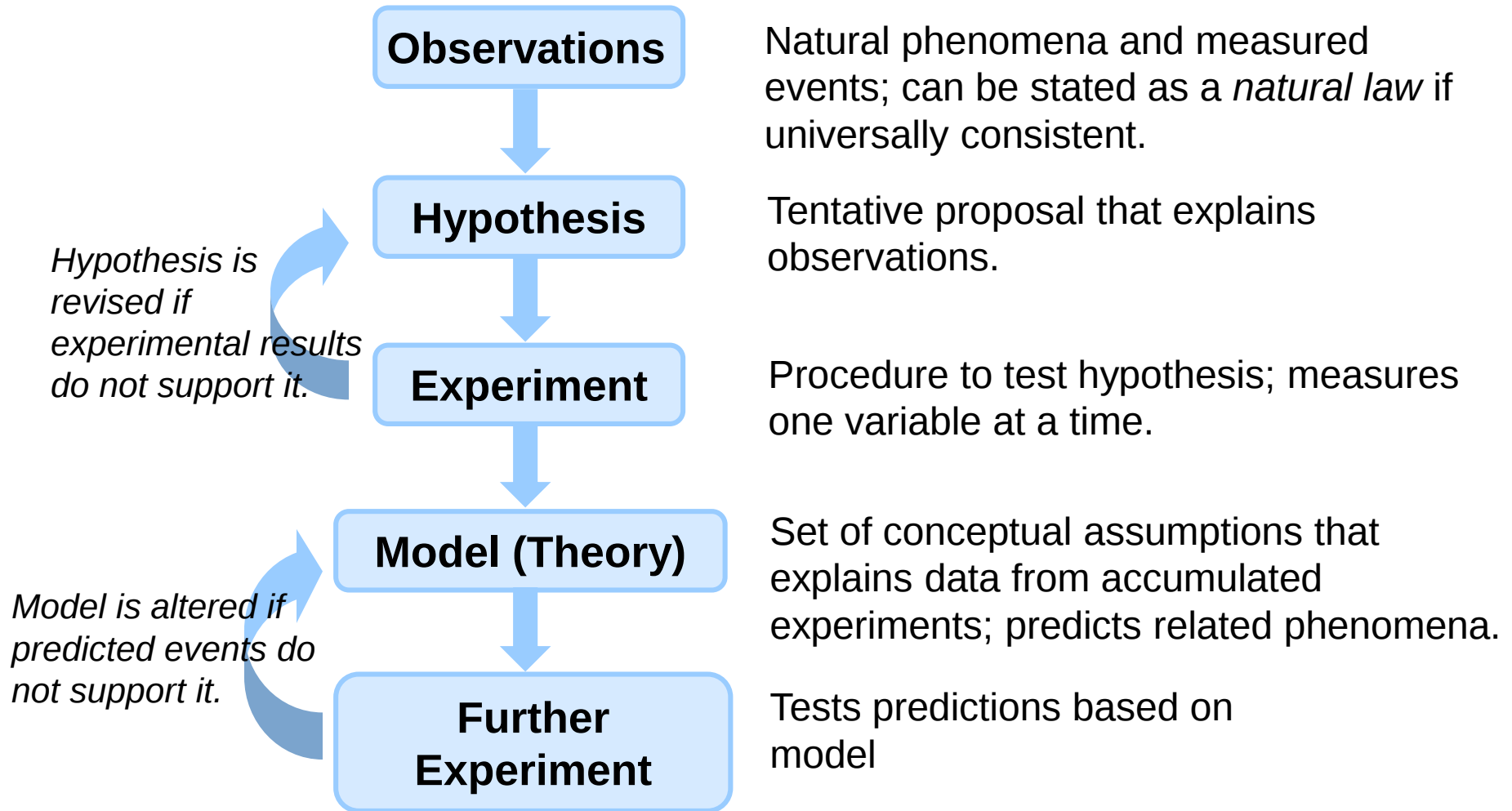
Figure 1.3D Potential energy is converted to kinetic energy.



D A system of fuel and exhaust. A fuel is higher in chemical potential energy than the exhaust. As the fuel burns, some of its potential energy is converted to the kinetic energy of the moving car.

A system of fuel and exhaust. A fuel is higher in chemical potential energy than the exhaust. As the fuel burns, some of its potential energy is converted to the kinetic energy of the moving car.

Figure 1.4 **The scientific approach to understanding nature.**



Chemical Problem Solving

- All measured quantities consist of
 - a ***number*** and a ***unit***.
- Units are manipulated like numbers:
 - $3 \text{ ft} \times 4 \text{ ft} = 12 \text{ ft}^2$

$$\text{– } \frac{350 \text{ mi}}{7 \text{ h}} = \frac{50 \text{ mi}}{1 \text{ h}} \quad \text{or } 50 \text{ mi} \cdot \text{h}^{-1}$$



Conversion Factors

A ***conversion factor*** is a ratio of equivalent quantities used to express a quantity in different units.

The relationship $1 \text{ mi} = 5280 \text{ ft}$ gives us the conversion factor:

$$\frac{1 \text{ mi}}{5280 \text{ ft}} = \frac{\cancel{5280 \text{ ft}}}{\cancel{5280 \text{ ft}}} = 1$$



A conversion factor is chosen and set up so that all units cancel except those required for the answer.

PROBLEM: The height of the Angel Falls is 3212 ft. Express this quantity in miles (mi) if 1 mi = 5280 ft.

PLAN: Set up the conversion factor so that ft will cancel and the answer will be in mi.

SOLUTION:

$$3212 \text{ ft} \times \frac{1 \text{ mi}}{5280 \text{ ft}} = 0.6083 \text{ mi}$$



Systematic Approach to Solving Chemistry Problems

- State Problem
- Plan
 - Clarify the known and unknown.
 - Suggest steps from known to unknown.
 - Prepare a visual summary of steps that includes conversion factors, equations, known variables.
- Solution
- Check
- Comment
- Follow-up Problem

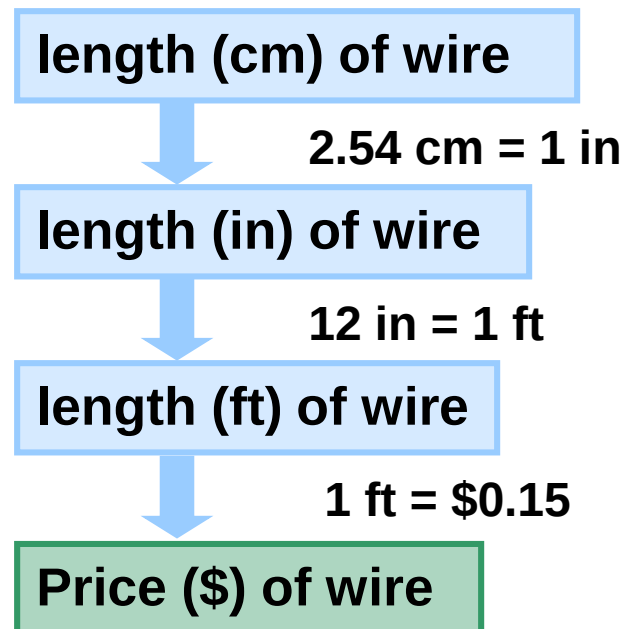


Sample Problem 1.3

Converting Units of Length

PROBLEM: To wire your stereo equipment, you need 325 centimeters (cm) of speaker wire that sells for \$0.15/ft. What is the price of the wire?

PLAN: We know the length (in cm) of wire and cost per length (\$/ft). We have to convert cm to inches and inches to feet. Then we can find the cost for the length in feet.



Sample Problem 1.3

SOLUTION:

Length (in) = length (cm) x conversion factor

$$= 325 \text{ cm} \times \frac{1 \text{ in}}{2.54 \text{ cm}} = 128 \text{ in}$$

Length (ft) = length (in) x conversion factor

$$= 128 \text{ in} \times \frac{1 \text{ ft}}{12 \text{ in}} = 10.7 \text{ ft}$$

Price (\$) = length (ft) x conversion factor

$$= 10.7 \text{ ft} \times \frac{\$ 0.15}{1 \text{ ft}} = \boxed{\$ 1.60}$$



Table 1.1 SI Base Units

Physical Quantity (Dimension)	Unit Name	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



Table 1.2 Common Decimal Prefixes Used with SI Units

Prefix*	Prefix Symbol	Word	Conventional Notation	Exponential Notation
tera	T	trillion	1,000,000,000,000	1×10^{12}
giga	G	billion	1,000,000,000	1×10^9
mega	M	million	1,000,000	1×10^6
kilo	k	thousand	1,000	1×10^3
hecto	h	hundred	100	1×10^2
deka	da	ten	10	1×10^1
—	—	one	1	1×10^0
deci	d	tenth	0.1	1×10^{-1}
centi	c	hundredth	0.01	1×10^{-2}
milli	m	thousandth	0.001	1×10^{-3}
micro	μ	millionth	0.000001	1×10^{-6}
nano	n	billionth	0.000000001	1×10^{-9}
pico	p	trillionth	0.0000000000001	1×10^{-12}
femto	f	quadrillionth	0.0000000000000001	1×10^{-15}

*The prefixes most frequently used by chemists appear in bold type.



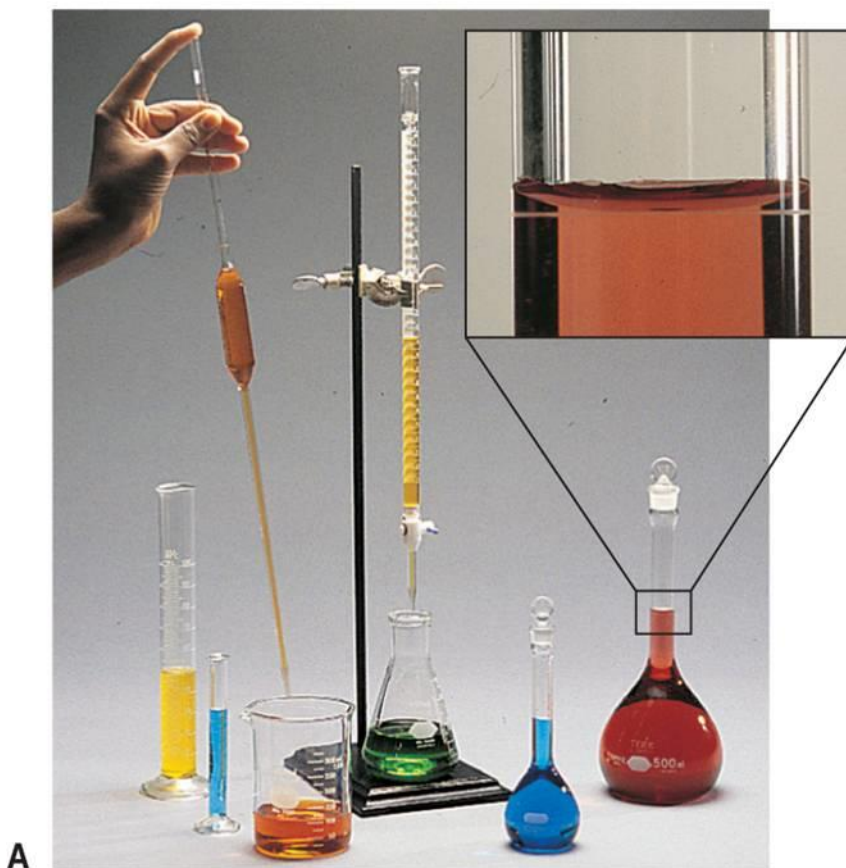
Table 1.3 Common SI-English Equivalent Quantities

Quantity	SI to English Equivalent	English to SI Equivalent
Length	1 km = 0.6214 mile	1 mi = 1.609 km
	1 m = 1.094 yd	1 yd = 0.9144 m
	1 m = 39.37 inches	1 ft = 0.3048 m
	1 cm = 0.3937 inch	1 in = 2.54 cm
Volume	1 cubic meter (m ³) = 35.31 ft ³	1 ft ³ = 0.02832 m ³
	1 dm ³ = 0.2642 gal	1 gal = 3.785 dm³
	1 dm ³ = 1.057 qt	1 qt = 0.9464 dm ³
	1 cm ³ = 0.03381 fluid ounce	1 qt = 946.4 cm ³
		1 fluid ounce = 29.57 cm ³
Mass	1 kg = 2.205 lb	1 lb = 0.4536 kg
	1 g = 0.03527 ounce (oz)	1 oz = 28.35 g



Figure 1.5 Common laboratory volumetric glassware.

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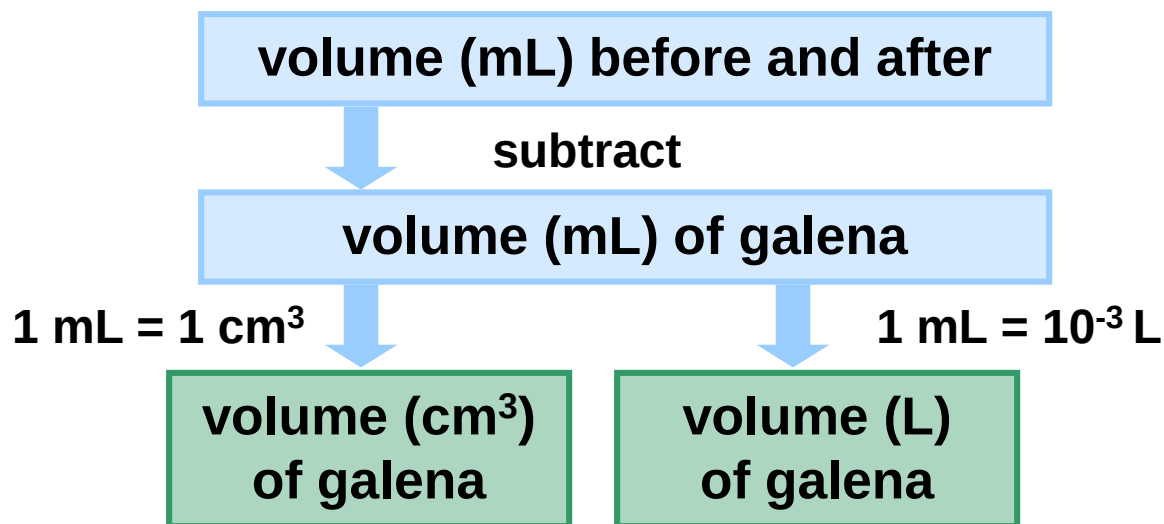


Sample Problem 1.4

Converting Units of Volume

PROBLEM: A graduated cylinder contains 19.9 mL of water. When a small piece of galena, an ore of lead, is added, it sinks and the volume increases to 24.5 mL. What is the volume of the piece of galena in cm^3 and in L?

PLAN: The volume of the galena is equal to the difference in the volume of the water before and after the addition.



Sample Problem 1.4

SOLUTION:

$$(24.5 - 19.9) \text{ mL} = \text{volume of galena} = 4.6 \text{ mL}$$

$$4.6 \text{ mL} \times \frac{1 \text{ cm}^3}{1 \text{ mL}} = 4.6 \text{ cm}^3$$

$$4.6 \text{ mL} \times \frac{10^{-3} \text{ L}}{1 \text{ mL}} = 4.6 \times 10^{-3} \text{ L}$$

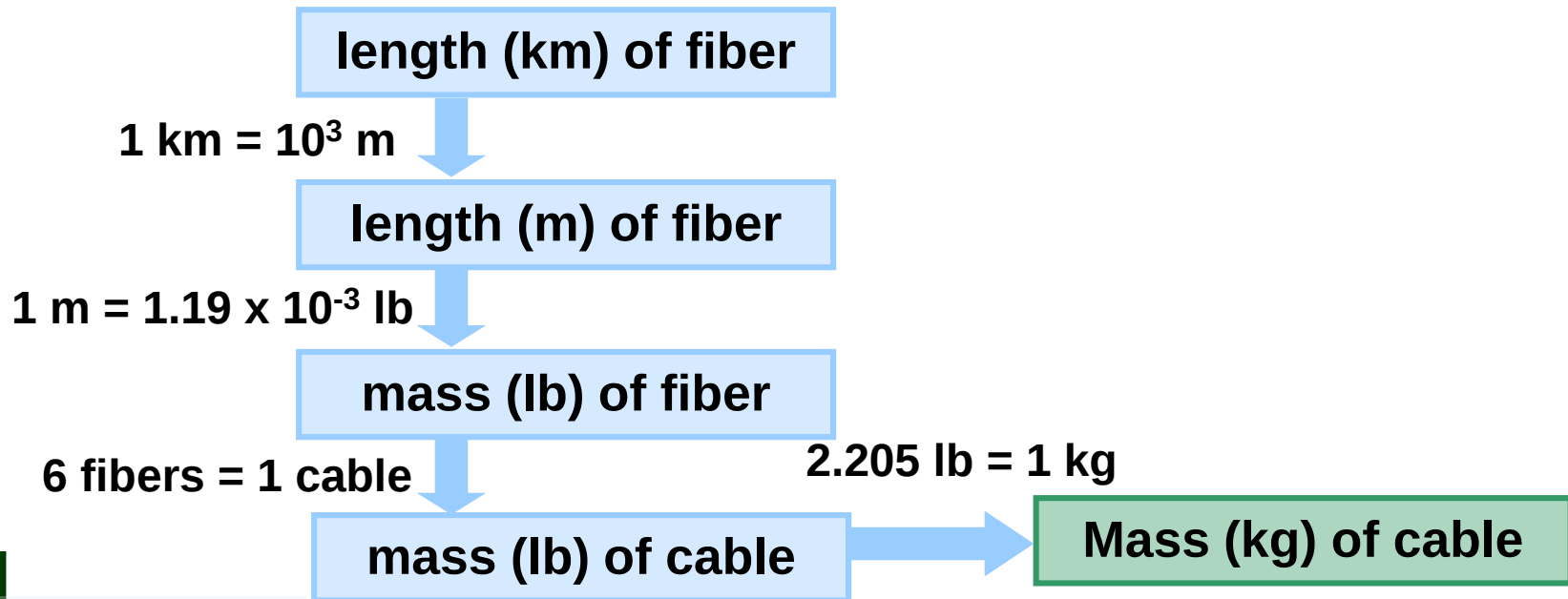


Sample Problem 1.5

Converting Units of Mass

PROBLEM: Many international computer communications are carried out by optical fibers in cables laid along the ocean floor. If one strand of optical fiber weighs 1.19×10^{-3} lb/m, what is the mass (in kg) of a cable made of six strands of optical fiber, each long enough to link New York and Paris (8.94×10^3 km)?

PLAN: The sequence of steps may vary but essentially we need to find the length of the entire cable and convert it to mass.



Sample Problem 1.5

SOLUTION:

$$8.84 \times 10^3 \cancel{\text{km}} \times \frac{10^3 \text{ m}}{1 \cancel{\text{km}}} = 8.84 \times 10^6 \text{ m}$$

$$8.84 \times 10^6 \cancel{\text{m}} \times \frac{1.19 \times 10^{-3} \text{ lb}}{1 \cancel{\text{m}}} = 1.05 \times 10^4 \text{ lb}$$

$$\frac{1.05 \times 10^4 \text{ lb}}{1 \cancel{\text{fiber}}} \times \frac{6 \cancel{\text{fibers}}}{1 \text{ cable}} = 6.30 \times 10^4 \text{ lb/cable}$$

$$\frac{6.30 \times 10^4 \cancel{\text{lb}}}{1 \text{ cable}} \times \frac{1 \text{ kg}}{2.205 \cancel{\text{lb}}} = \boxed{2.86 \times 10^4 \text{ kg/cable}}$$



Density

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

At a given temperature and pressure, the density of a substance is a characteristic physical property and has a specific value.



Table 1.4 Densities of Some Common Substances*

Substance	Physical State	Density (g/cm ³)
Hydrogen	gas	0.0000899
Oxygen	gas	0.00133
Grain alcohol	liquid	0.789
Water	liquid	0.998
Table salt	solid	2.16
Aluminum	solid	2.70
Lead	solid	11.3
Gold	solid	19.3

*At room temperature (20°C) and normal atmospheric pressure (1 atm).

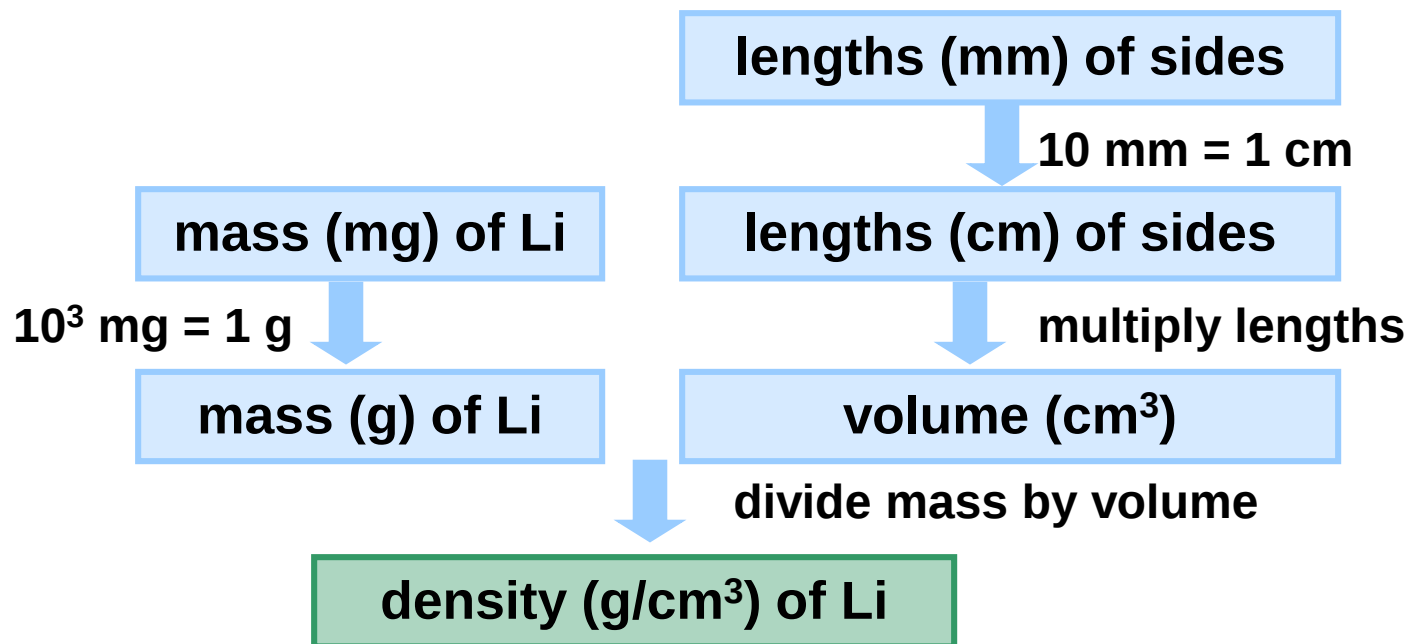


Sample Problem 1.6

Calculating Density from Mass and Length

PROBLEM: Lithium, a soft, gray solid with the lowest density of any metal, is a key component of advanced batteries. A slab of lithium weighs 1.49×10^3 mg and has sides that are 20.9 mm by 11.1 mm by 11.9 mm. Find the density of lithium in g/cm^3 .

PLAN: Density is expressed in g/cm^3 so we need the mass in g and the volume in cm^3 .



Sample Problem 1.6

SOLUTION:

$$1.49 \times 10^3 \text{ mg} \times \frac{1 \text{ g}}{10^3 \text{ mg}} = 1.49 \text{ g}$$

$$20.9 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 2.09 \text{ cm}$$

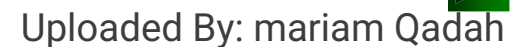
Similarly the other sides will be 1.11 cm and 1.19 cm, respectively.

$$\text{Volume} = 2.09 \text{ cm} \times 1.11 \text{ cm} \times 1.19 \text{ cm} = 2.76 \text{ cm}^3$$

$$\text{density of Li} = \frac{1.49 \text{ g}}{2.76 \text{ cm}^3} = \boxed{0.540 \text{ g/cm}^3}$$



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Temperature Scales

Kelvin (K) - The “absolute temperature scale” begins at absolute zero and has only positive values. Note that the kelvin is not used with the degree sign ($^{\circ}$).

Celsius ($^{\circ}\text{C}$) - The Celsius scale is based on the freezing and boiling points of water. This is the temperature scale used most commonly around the world. The Celsius and Kelvin scales use the **same size degree** although their starting points differ.

Fahrenheit ($^{\circ}\text{F}$) – The Fahrenheit scale is commonly used in the US. The Fahrenheit scale has a different degree size and different zero points than both the Celsius and Kelvin scales.



Temperature Conversions

$$T \text{ (in K)} = T \text{ (in } ^\circ\text{C)} + 273.15$$

$$T \text{ (in } ^\circ\text{C)} = T \text{ (in K)} - 273.15$$

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

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



Sample Problem 1.7

Converting Units of Temperature

PROBLEM: A child has a body temperature of 38.7°C , and normal body temperature is 98.6°F . Does the child have a fever? What is the child's temperature in kelvins?

PLAN: We have to convert $^{\circ}\text{C}$ to $^{\circ}\text{F}$ to find out if the child has a fever. We can then use the $^{\circ}\text{C}$ to Kelvin relationship to find the temperature in Kelvin.

SOLUTION:

Converting from $^{\circ}\text{C}$ to $^{\circ}\text{F}$ $\frac{9}{5}(38.7^{\circ}\text{C}) + 32 = \mathbf{101.7^{\circ}\text{F}}$

Yes, the child has a fever.

Converting from $^{\circ}\text{C}$ to K $38.7^{\circ}\text{C} + 273.15 = \mathbf{311.8\text{ K}}$



Significant Figures

Every measurement includes some ***uncertainty***. The ***rightmost*** digit of any quantity is always ***estimated***.

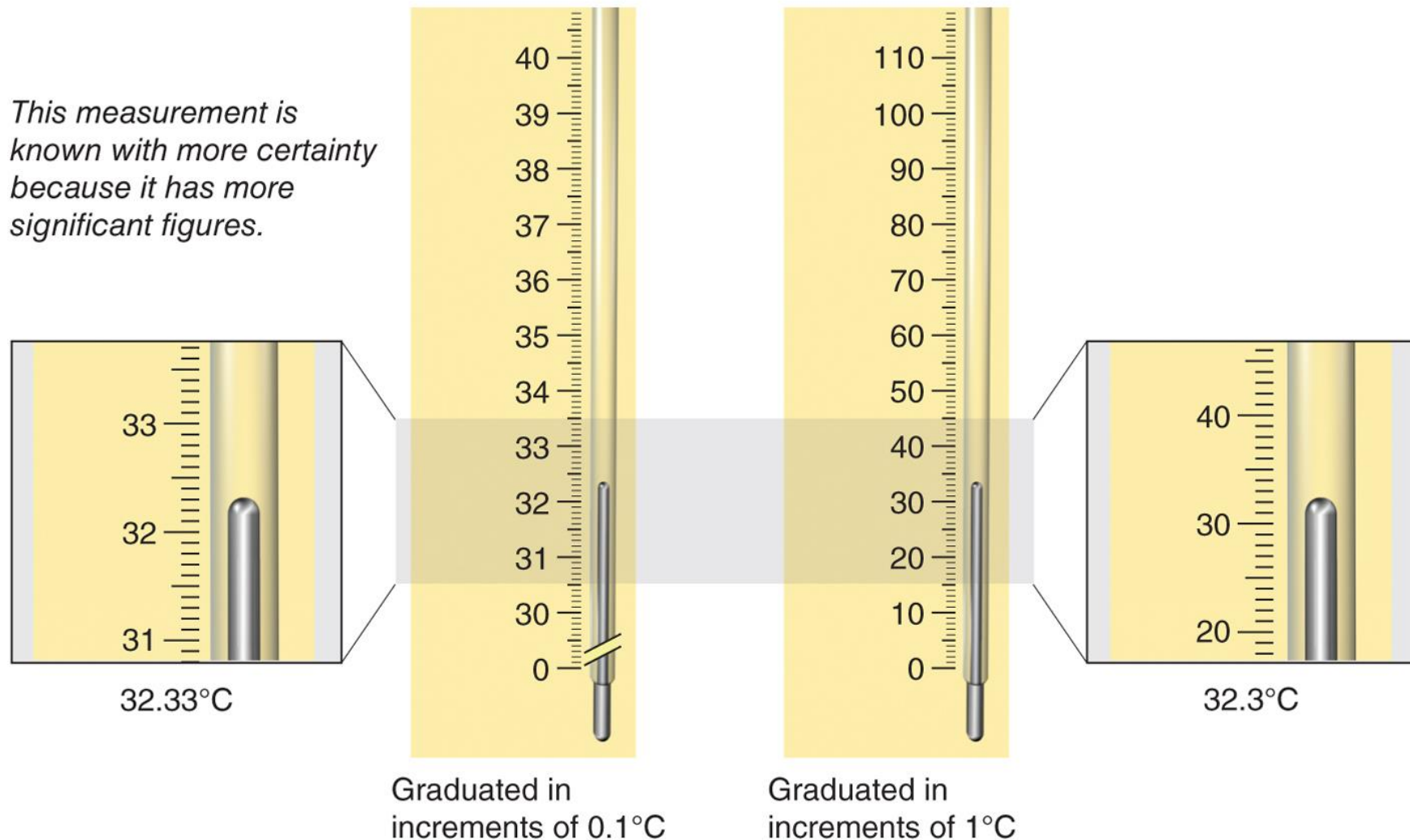
The recorded digits, both certain and uncertain, are called ***significant figures***.

The greater the number of significant figures in a quantity, the greater its certainty.



Figure 1.7 The number of significant figures in a measurement.

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Determining Which Digits are Significant

All digits are significant

- *except zeros that are used only to position the decimal point.*

- Make sure the measured quantity has a decimal point.
- Start at the left and move right until you reach the first nonzero digit.
- Count that digit and every digit to its right as significant.



- Zeros that end a number are significant
 - whether they occur before or after the decimal point
 - as long as a decimal point is present.
- 1.030 mL has 4 significant figures.
- 5300. L has 4 significant figures.
- If no decimal point is present
 - zeros at the end of the number are **not** significant.
- 5300 L has only 2 significant figures.



Sample Problem 1.8

Determining the Number of Significant Figures

PROBLEM: For each of the following quantities, underline the zeros that are significant figures (sf), and determine the number of significant figures in each quantity. For **(d)** to **(f)**, express each in exponential notation first.

(a) 0.0030 L

(b) 0.1044 g

(c) 53,069 mL

(d) 0.00004715 m

(e) 57,600. s

(f) 0.0000007160 cm³

PLAN: We determine the number of significant figures by counting digits, paying particular attention to the position of zeros in relation to the decimal point, and underline zeros that are significant.



Sample Problem 1.8

SOLUTION:

(a) 0.0030 L has **2 sf** (b) 0.1044 g has **4 sf**

(c) 53,069 mL has **5 sf**

(d) 0.00004715 m = **4.715x10⁻⁵ m** has **4 sf**

(e) 57,600. s = **5.7600x10⁴ s** has **5 sf**

(f) 0.0000007160 cm³ = **7.160x10⁻⁷ cm³** has **4 sf**



Rules for Significant Figures in Calculations

1. **For multiplication and division.** The answer contains the same number of significant figures as there are in the measurement with the fewest significant figures.

Multiply the following numbers:

$$9.2 \text{ cm} \times 6.8 \text{ cm} \times 0.3744 \text{ cm} = 23.4225 \text{ cm}^3 = \mathbf{23 \text{ cm}^3}$$



Rules for Significant Figures in Calculations

2. **For addition and subtraction.** The answer has the *same number of decimal places as there are in the measurement with the fewest decimal places.*

Example: adding two volumes

$$\begin{array}{r} 83.5 \text{ mL} \\ + 23.28 \text{ mL} \\ \hline 106.78 \text{ mL} = \mathbf{106.8 \text{ mL}} \end{array}$$

Example: subtracting two volumes

$$\begin{array}{r} 865.9 \text{ mL} \\ - 2.8121 \text{ mL} \\ \hline 863.0879 \text{ mL} = \mathbf{863.1 \text{ mL}} \end{array}$$

Rules for Rounding Off Numbers

1. If the digit removed is ***more than 5***, the preceding number increases by 1.

5.379 rounds to 5.38 if 3 significant figures are retained.

2. If the digit removed is ***less than 5***, the preceding number is unchanged.

0.2413 rounds to 0.241 if 3 significant figures are retained.



3. If the digit removed ***is 5 followed by zeros or with no following digits***, the preceding number increases by 1 if it is odd and remains unchanged if it is even.

17.75 rounds to 17.8, but 17.65 rounds to 17.6.

If the ***5 is followed by other nonzero digits***, rule 1 is followed:

17.6500 rounds to 17.6, but 17.6513 rounds to 17.7

4. Be sure to carry two or more additional significant figures through a multistep calculation and round off the ***final*** answer ***only***.



Exact Numbers

Exact numbers have no uncertainty associated with them.

Numbers may be exact by definition:

$$1000 \text{ mg} = 1 \text{ g}$$

$$60 \text{ min} = 1 \text{ hr}$$

$$2.54 \text{ cm} = 1 \text{ in}$$

Numbers may be exact by count:
exactly 26 letters in the alphabet

Exact numbers do not limit the number of significant digits in a calculation.



Sample Problem 1.9

Significant Figures and Rounding

PROBLEM: Perform the following calculations and round each answer to the correct number of significant figures:

$$(a) \quad \frac{16.3521 \text{ cm}^2 - 1.448 \text{ cm}^2}{7.085 \text{ cm}}$$

$$(b) \quad \frac{4.80 \times 10^4 \text{ mg} \left(\frac{1 \text{ g}}{1000 \text{ mg}} \right)}{11.55 \text{ cm}^3}$$

PLAN: We use the rules for rounding presented in the text: **(a)** We subtract before we divide. **(b)** We note that the unit conversion involves an exact number.



Sample Problem 1.9

SOLUTION:

$$(a) \quad \frac{16.3521 \text{ cm}^2 - 1.448 \text{ cm}^2}{7.085 \text{ cm}} = \frac{14.904 \text{ cm}^2}{7.085 \text{ cm}} = 2.104 \text{ cm}$$

$$(b) \quad \frac{4.80 \times 10^4 \text{ mg} \left(\frac{1 \text{ g}}{1000 \text{ mg}} \right)}{11.55 \text{ cm}^3} = \frac{48.0 \text{ g}}{11.55 \text{ cm}^3} = 4.16 \text{ g/cm}^3$$



Precision, Accuracy, and Error

Precision refers to how close the measurements in a series are to each other.

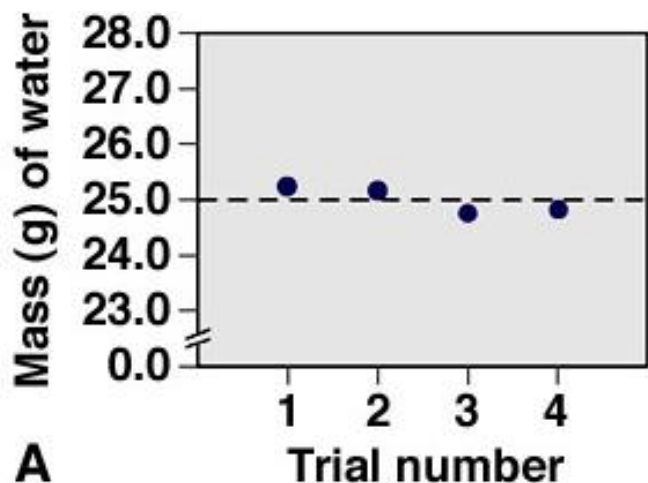
Accuracy refers to how close each measurement is to the actual value.

Systematic error produces values that are either **all** higher or **all** lower than the actual value.
This error is part of the experimental system.

Random error produces values that are both higher and lower than the actual value.



Figure 1.9 Precision and accuracy in a laboratory calibration.



precise and accurate

precise but not accurate

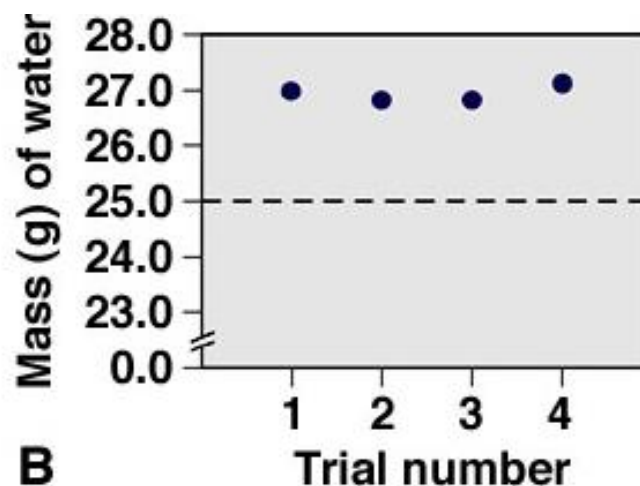
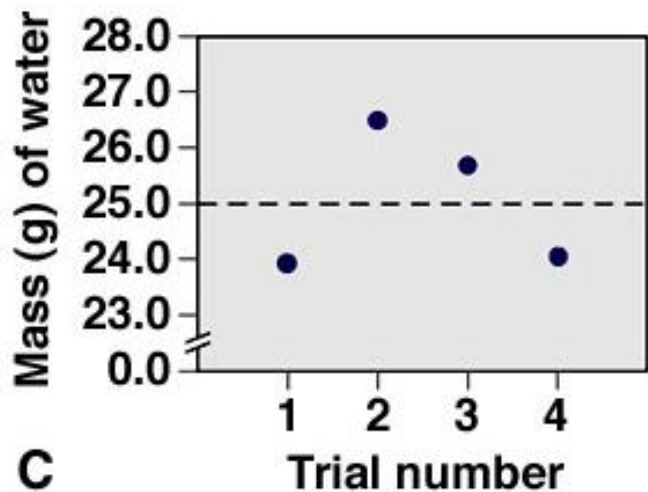


Figure 1.9
continued

Precision and accuracy in the laboratory.



random error

systematic error

