GENERAL CHEMISTRY

Principles and Modern Applications

TENTH EDITION

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Chapter 1:

Matter: Its Properties and Measurement

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CONTENTS

Matter: Its Properties and Measurement

- Chemistry and chemicals are integral part of life

A Hubble Space Telescope image of a cloud of hydrogen gas and dust (lower right half of the image) that is part of the Swan Nebula (M17). The colors correspond to light emitted by hydrogen (green), sulfur (red), and oxygen (blue). The chemical elements discussed in this text are those found on Earth and, presumably, throughout the universe.

- No chemicals added !!!!!
- Organically grown !!!!!
- All material objects (living or inanimate) are made up only of chemicals

- people have always practiced chemistry. Glazing pottery, ore smelting, tanning, fabric dyeing, making cheese, wine, beer, soap……

Chemistry is the central science. Scientific progress depends on the way scientists do their work, asking the right questions, designing the right experiments to supply the answers, and formulating plausible explanations of their findings.

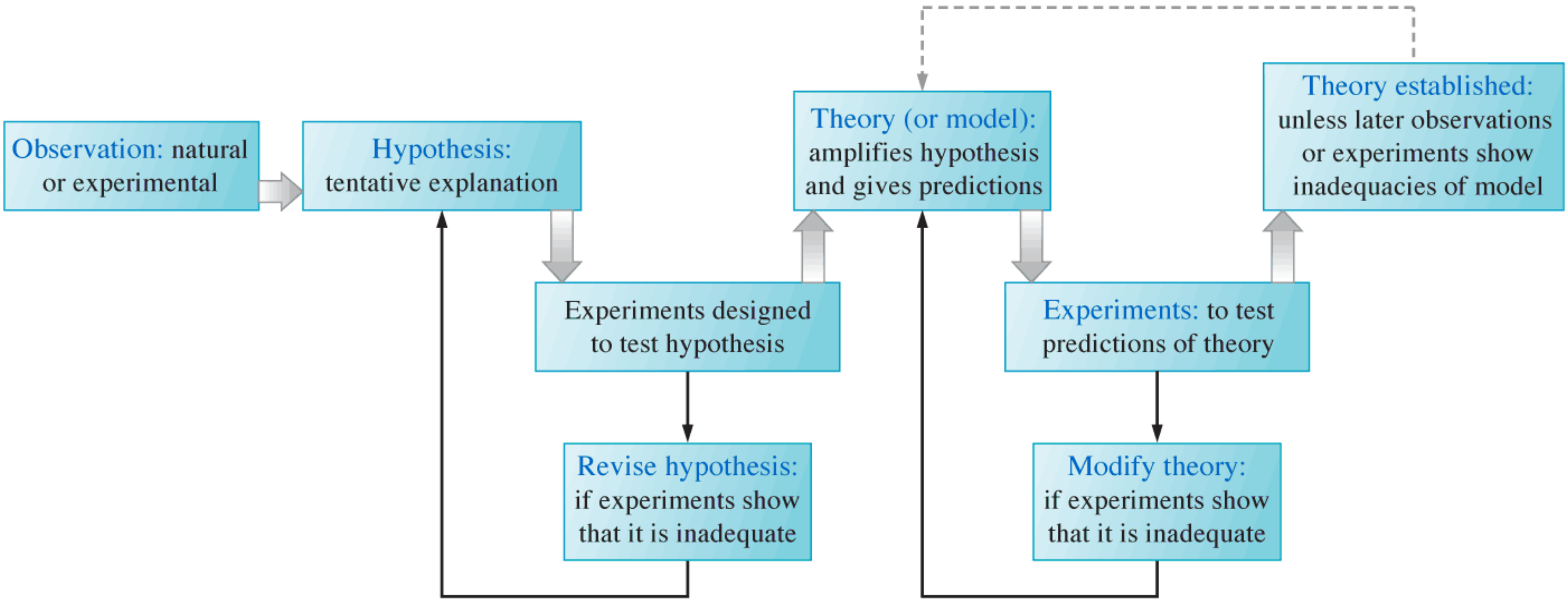
1-1 The Scientific Method

- The scientific method originated in the seventeenth century with such people as Galileo, Francis Bacon, Robert Boyle, and Isaac Newton.
- The key to the method is to make no initial assumptions, but rather to make careful observations of natural phenomena.
- When enough observations have been made so that a pattern begins to emerge, a generalization or natural law can be formulated describing the phenomenon.
  - Natural Laws are concise statements often in mathematical form, about natural phenomena
    e.g. Nicholoas Copernicus, 1473-1543 astronomical observations,
    e.g. Radioactive decay law.
  - The success of a natural law depends on its ability to explain observations and to predict new phenomena, however, Natural Law is not an absolute truth:
    e.g. Johannes Kepler/ half a century later / planets travel in elliptical orbits rather than circular stated by Copernicus.
  - To verify natural law, a scientist designs experiments to support their conclusions deduced from natural law.

1-1 The Scientific Method

- A **hypothesis is a tentative explanation of a natural law.**
- **If a hypothesis survives** testing by experiments, it is often referred to as a theory.
- In a broader sense, a **theory is a model or way of looking at nature that can be used to** explain natural laws and make further predictions about natural phenomena

The scientific method is the combination of observation, experimentation and the formulation of laws, hypotheses, and theories.



▲ FIGURE 1-1
The scientific method illustrated

1-1 The Scientific Method

Scientists need to be alert to unexpected observations: Chance and serendipidity

Many discoveries have been made by accident.

- 1839 Charles Goodyear ⇒ natural rubber

  / less brittle in cold // Less tacky in warm.

  Accidental spilling of a rubber-sulfur mixture on a hot stove

- X-rays, radioactivity, penicillin

Louis Pasteur (1822-1895)

" Chance favors the prepared mind"

- developer of germ theory
- Pasteurization
- rabies vaccination

Called the greatest physician of all time by some.

He was a chemist by training and profession.

1-2 Properties of Matter

Dictionary definitions of chemistry usually include the terms:

* *matter, composition, and properties, as in the statement that*

" chemistry is the science that deals with the composition and properties of matter.

Matter: Occupies space, has mass and inertia

Composition: Parts or components of a matter.

ex. H_2O, 11.19% H and 88.81% O

Properties: Qualities or attributes that we can use to distinguish one sample of matter from the other.

the properties of matter are generally grouped into two broad categories: physical and chemical.

Physical and Chemical properties and changes:

Physical prop:

- color, malleability, ductility
- no change in composition

  e.g. Water-ice

Chemical Prop:

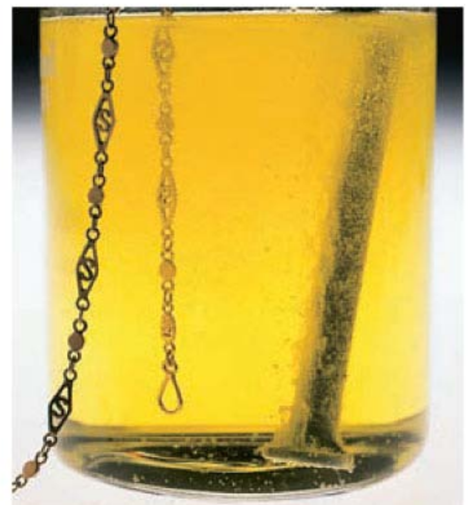
-one or more kinds of matter are converted to new kinds of matter with different compositions

-change in composition

e.g. burning a sheet of paper

Zinc and gold reaction with HCl

▶ FIGURE 1-2
Physical properties of sulfur and copper



▲ FIGURE 1-3
A chemical property of zinc and gold: reaction with hydrochloric acid

* Decomposition: A chemical change

* The decomposition of compounds into their constituent elements is a more difficult matter than the mere physical separation of mixtures.

* A chemical compound retains its identity during physical changes, but it can be decomposed into its constituent elements by *chemical changes.*

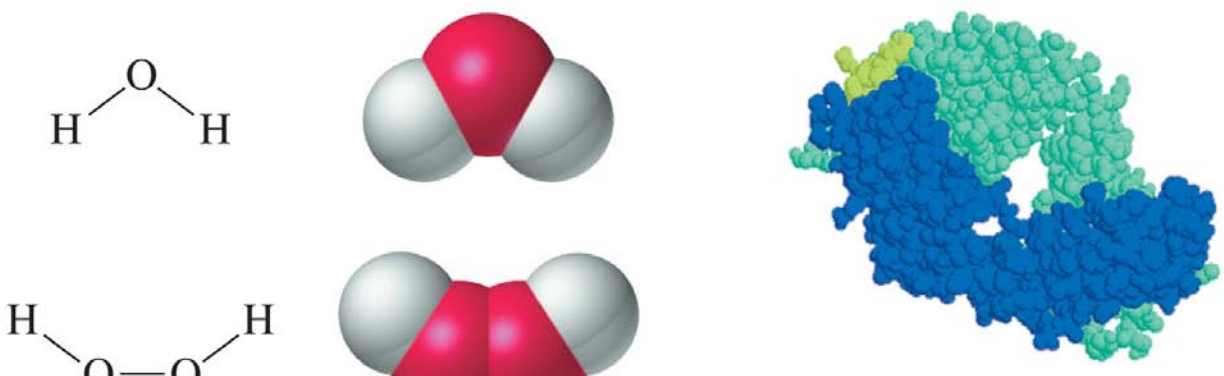
* ammonium dichromate, when heated, decomposes into the substances chromium(III) oxide, nitrogen, and water. This reaction, once used in movies to simulate a volcano.

▲

▲ FIGURE 1-6
A chemical change: decomposition of ammonium dichromate

1-3 Classification of Matter

- Matter is made of **atoms**. Each different type of atom is the building block of a different chemical **element**
- 114 **elements**. About 90% available from natural sources The remainder do not occur naturally and have been created only in laboratories.
- **Compounds** are comprised of atoms of two or more elements. Scientists have identified millions of different chemical compounds. In some cases, we can isolate a molecule of a compound
- **Molecules** are the smallest units of compounds having the same proportions of the constituent atoms as does the compound as a whole



Gamma globulin

(3) (4) (19996)

- The composition and properties of an element or a compound are uniform throughout a given sample and from one sample to another.
- Elements and compounds are called **substances.**
- **In the chemical sense, the term *substance*** should be used only for elements and compounds.
- A *mixture of substances* can vary in composition and properties from one sample to another. One that is uniform in composition and properties throughout is said to be a **homogeneous mixture or a *solution. E.g. solution of sucrose (cane sugar)*** in water, Ordinary air *Seawater ,* Gasoline is a homogeneous mixture
- In **heterogeneous mixtures sand and water, for example the components** separate into distinct regions. Thus, the composition and physical properties vary from one part of the mixture to another. E.g. Salad dressing, a slab of concrete, and the leaf of a plant are all heterogeneous. It is usually easy to distinguish heterogeneous from homogeneous mixtures. A scheme for classifying matter into elements and compounds and homogeneous and heterogeneous mixtures is summarized in Figure 1-4.

A classification scheme for matter

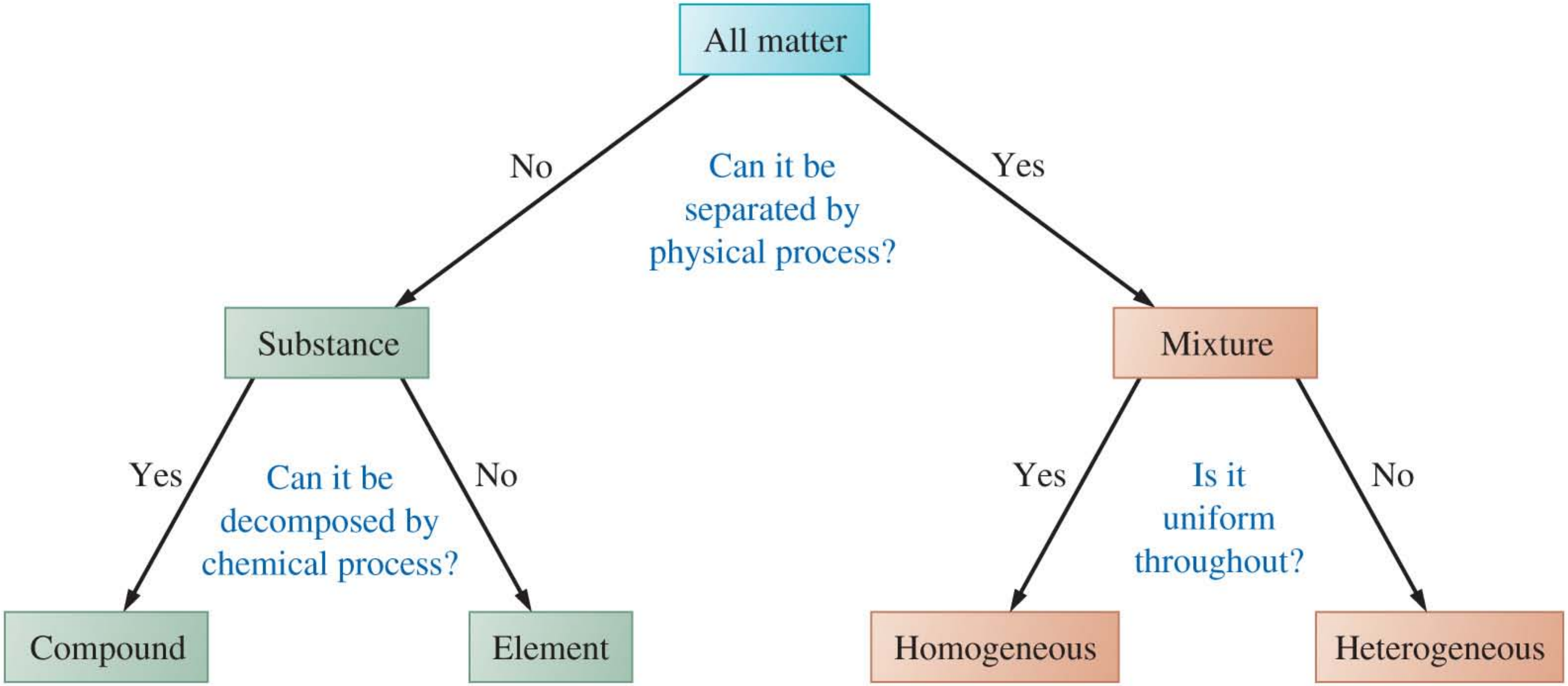


FIGURE 1-4 **Every sample of matter is either a single substance (an element or compound) or a mixture of substances. At the molecular level, an element consists of atoms of a single type and a compound consists of two or more different types of atoms, usually joined into molecules. In a homogeneous mixture, atoms or molecules are randomly mixed at the molecular level. In heterogeneous mixtures, the components are physically separated, as in a layer of octane molecules (a constituent of gasoline) floating on a layer of water molecules.**

Separating Mixtures: a physical process

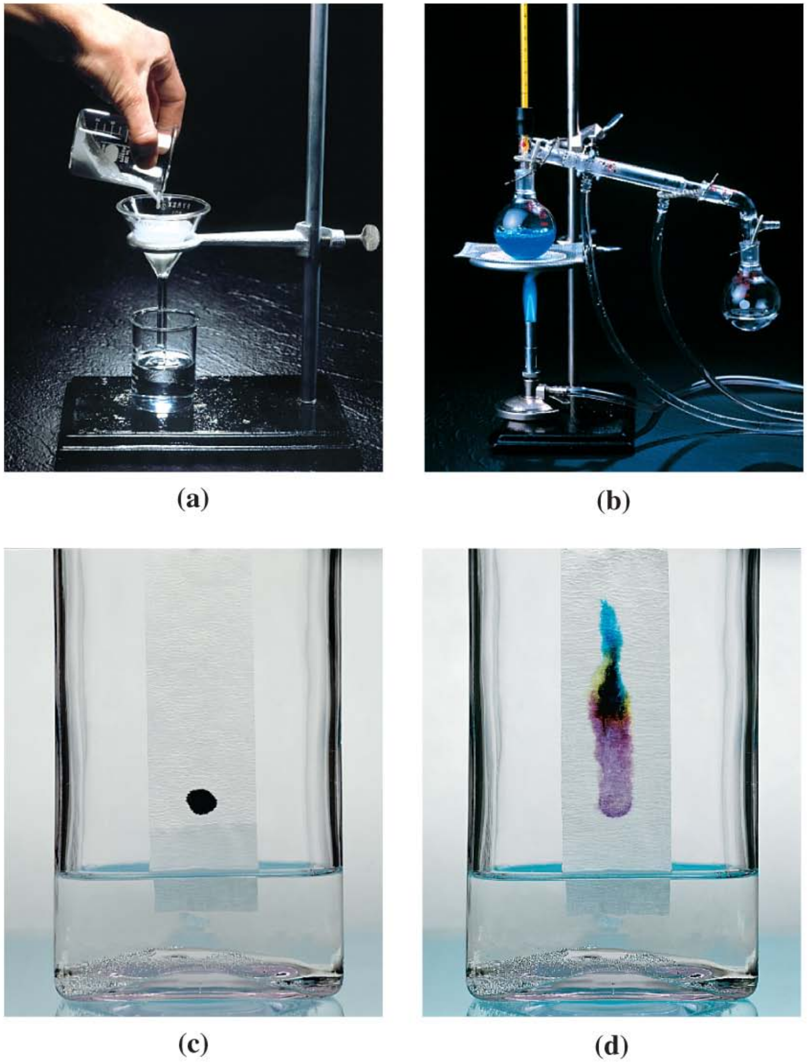


FIGURE 1-5

-Homogeneous mixtures
- Sieving
- Filtration

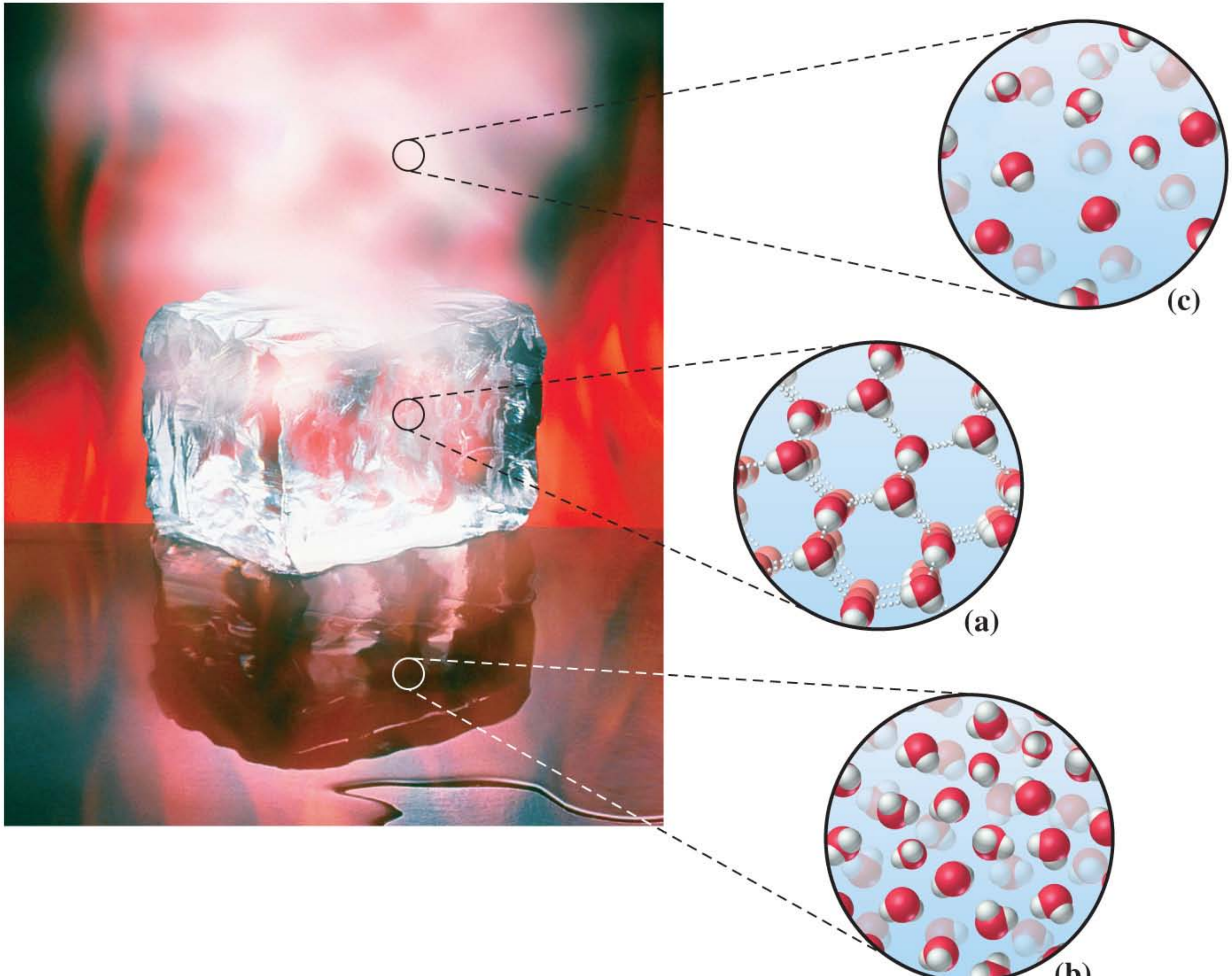
- Heterogeneous mixtures
- Distillation
- Chromatography

- Chromatography is a separation science based on the differing abilities of compounds to adhere to the surfaces of various solid substances, e.g. paper, starch

States of the Matter: Macroscopic and microscopic views

- Eye

- atomic and molecular level



1-4 The Measurement of Matter

Chemistry is a *quantitative science, which means that in many cases we can* measure a property of a substance and compare it with a standard having a known value of the property. We express the measurement as the product of a *number and a unit.*

TABLE 1.1 SI Base Quantities

Physical Quantity	Unit	Symbol
Length	meter[a]	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Amount of substance[b]	mole	mol
Electric current[c]	ampere	A
Luminous intensity[d]	candela	cd

[a]The official spelling of this unit is "metre," but we will use the American spelling.
[b]The mole is introduced in Section 2-7.
[c]Electric current is described in Appendix B and in Chapter 20.
[d]Luminous intensity is not discussed in this text.

TABLE 1.2 SI Prefixes

Multiple	Prefix
10^{18}	exa (E)
10^{15}	peta (P)
10^{12}	tera (T)
10^{9}	giga (G)
10^{6}	mega (M)
10^{3}	kilo (k)
10^{2}	hecto (h)
10^{1}	deka (da)
10^{-1}	deci (d)
10^{-2}	centi (c)
10^{-3}	milli (m)
10^{-6}	micro (μ)[a]
10^{-9}	nano (n)
10^{-12}	pico (p)
10^{-15}	femto (f)
10^{-18}	atto (a)
10^{-21}	zepto (z)
10^{-24}	yocto (y)

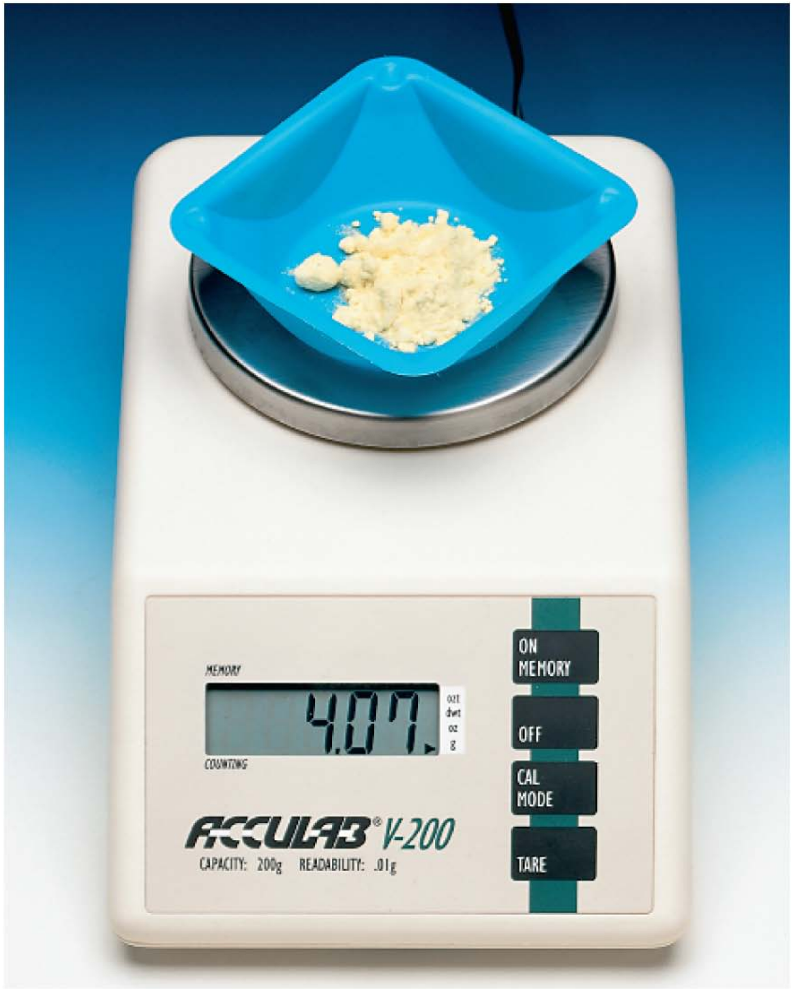
[a]The Greek letter μ (pronounced “mew”).

Mass:

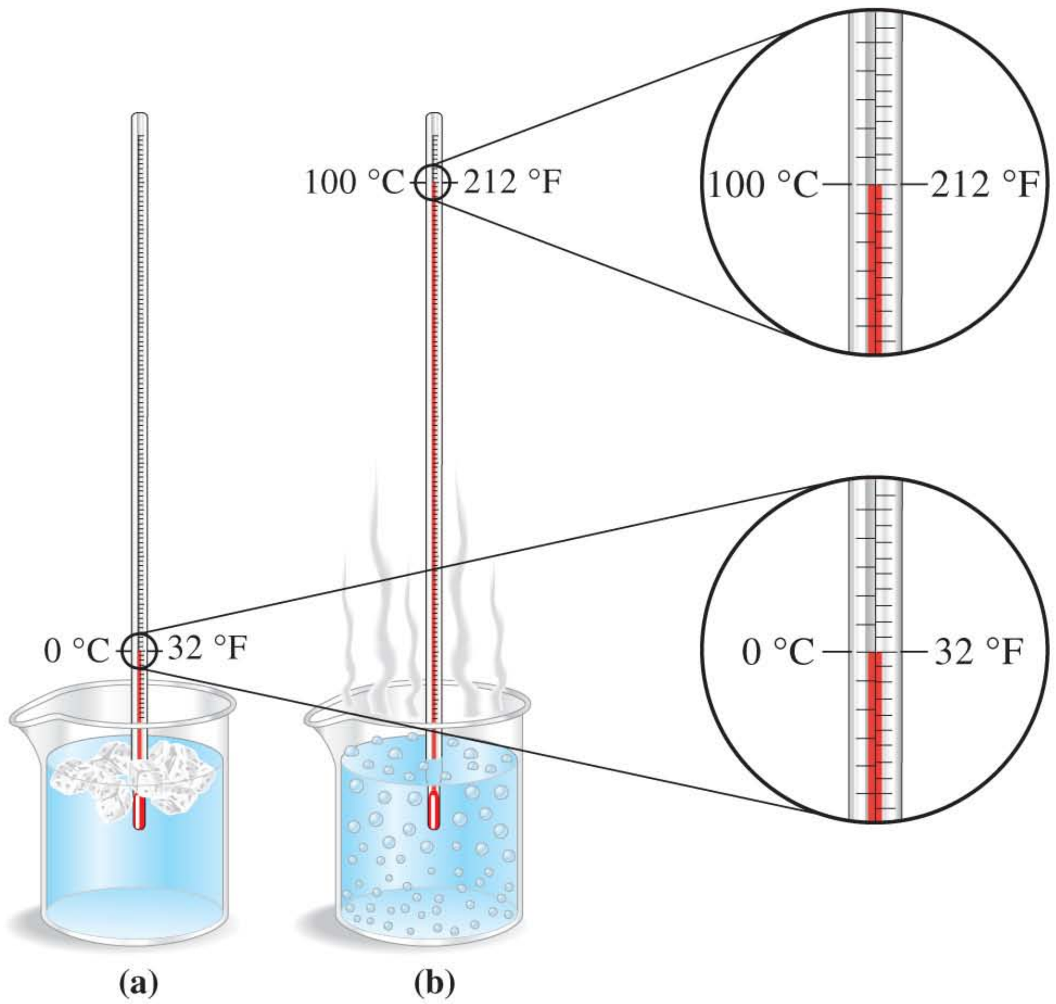
Mass is the **quantity** of matter in an object.

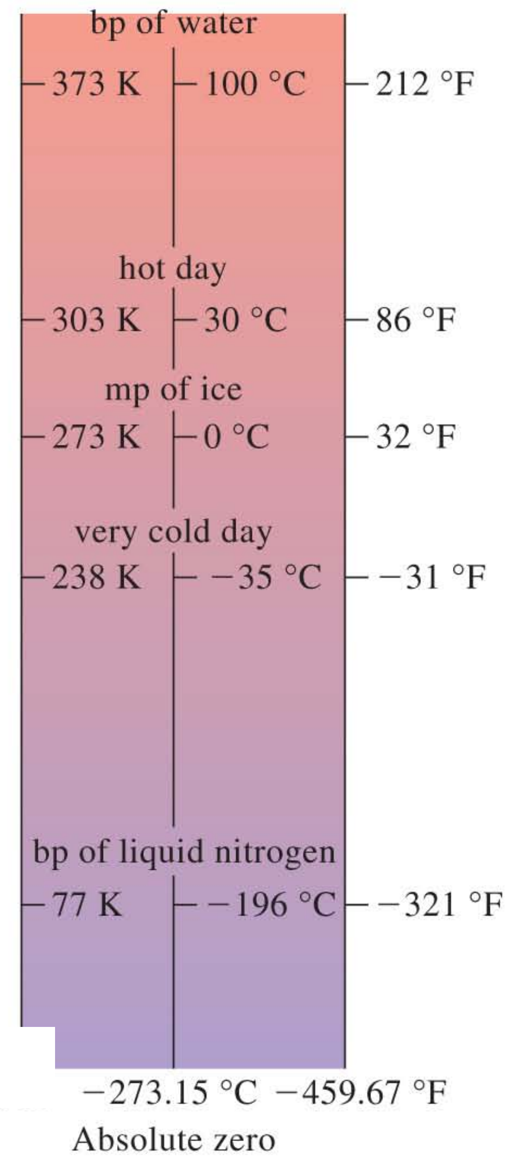
Weight is the force of gravity on an object

W α m $\quad$ $W = g \times m$

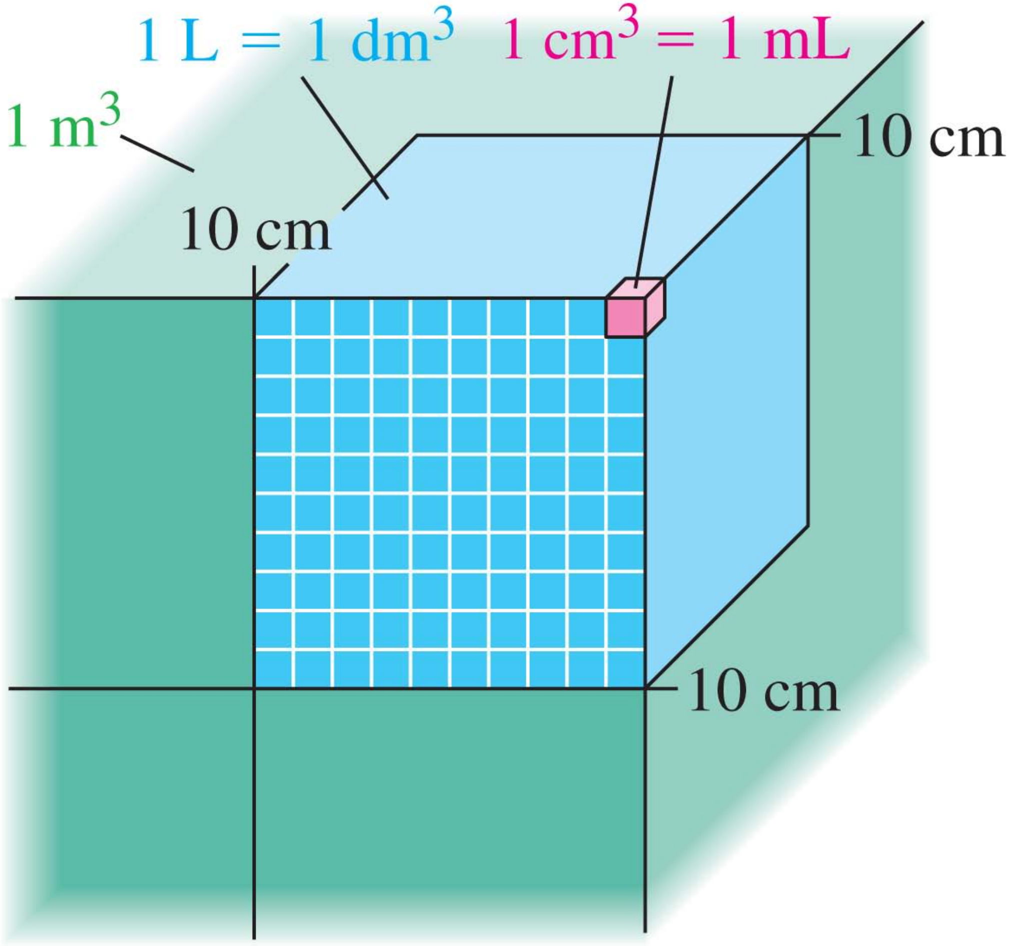


Temperature: To establish a temperature scale, we arbitrarily set certain fixed points and temperature increments called degrees





Volume



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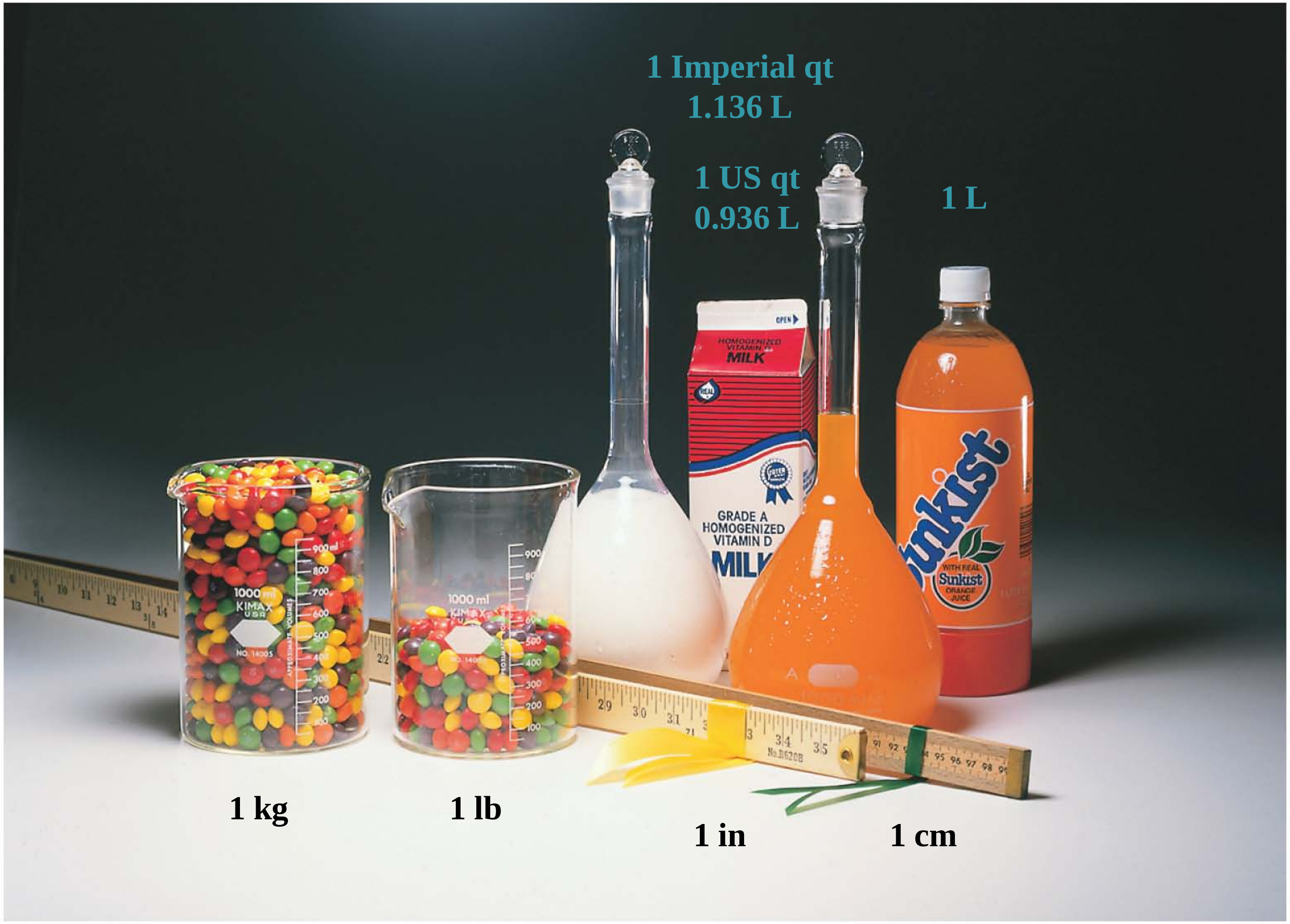
SI and non-SI Units Compared



SI Units

Length	meter, m
Mass	Kilogram, kg
Time	second, s
Temperature	Kelvin, K
Quantity	Mole, 6.022×10^{23}

mol^{-1}

Non-SI Units

Length	Angstrom, Å, 10^{-8} cm
Volume	Liter, L, 10^{-3} m^3
Energy	Calorie, cal, 4.184 J
Pressure	

1 Atm = 1.064×10^{2} kPa

1 Atm = 760 mm Hg

Derived Units

Force	Newton, kg m s^{-2}
Pressure	Pascal, kg m^{-1} s^{-2}
Energy	Joule, kg m^{2} s^{-2}

1-5 Density and Percent Composition

Density (d) = mass (m) / Volume (V) (g/mL) or (g/cm^3)

- Mass and volume are **extensive** properties. An **extensive property**

is *dependent* on the quantity of matter observed.

- Density is an **intensive** property. An **intensive property** is i*ndependent* of the amount of matter observed.

  - density of pure water at 25 ° C has a unique value, whether the sample fills a small beaker (small mass/small volume) or a swimming pool (large mass/large volume)

- **Density is a function of temperature.**

  density of water at 4 ºC = 1000 g/1000 mL, 1.000 g/mL.

  20 ºC= 0.9982 g/mL

Density in Conversion Pathways

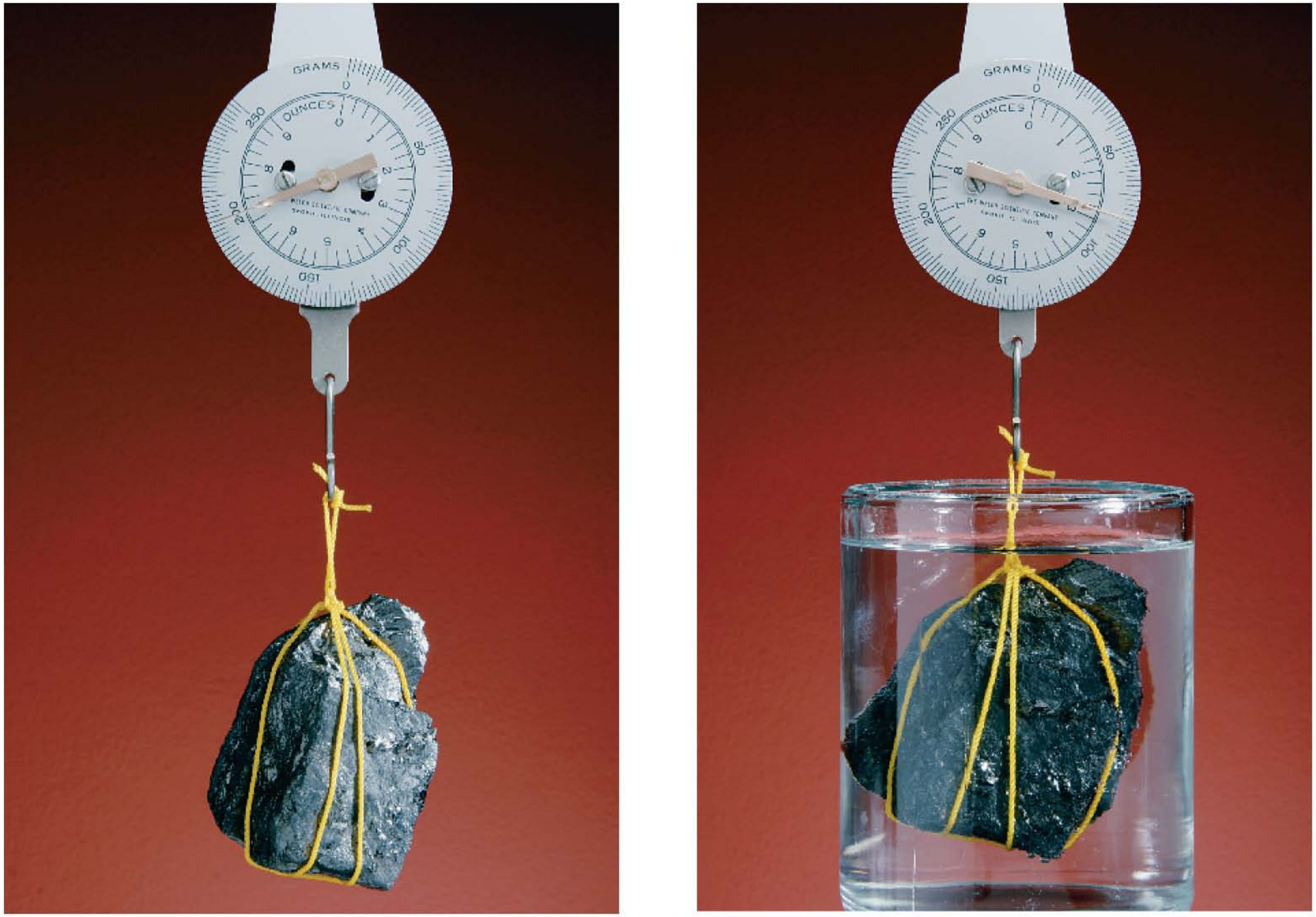
What is the mass of a cube of osmium that is 1.25 inches on each side?

Have volume, need density = $22.48\ g/cm^3$

(converts in. to cm) (converts cm to cm^3) (converts cm^3 to g osmium)

$$? \text{ g osmium} = \left[1.25\ \text{in.} \times \frac{2.54\ \text{cm}}{1\ \text{in.}}\right]^3 \times \frac{22.48\ \text{g osmium}}{1\ \text{cm}^3} = 719\ \text{g osmium}$$

Measuring Volume of an Irregular Object



1-6 Uncertainties in Scientific Measurements

All measurements are subject to error.

Systematic errors: Built-in / inherent errors

- have a definite value and an assignable cause, and are of the same magnitude for replicate measurements made in the same way.
- Systematic errors lead to bias in measurement results. Bias can be negative or positive in sign.

  e.g. thermometer constantly measures 2° C low.
- must be avoided by carefully calibrating a method against a known sample or result

Random errors: results either too high or too low

- *are observed* by scatter in the data and can be dealt with effectively by taking the average of many measurements
- Limitations in experimenters ability to read a scale.

Accuracy & Precision:

- Accuracy: is the closeness of a measured value X_i to the true or accepted value X_t
- Precision: Reproducibility of a measurement.

Reproducibility	~ 0.1 g	~ 0.0001 g
Precision	low	high

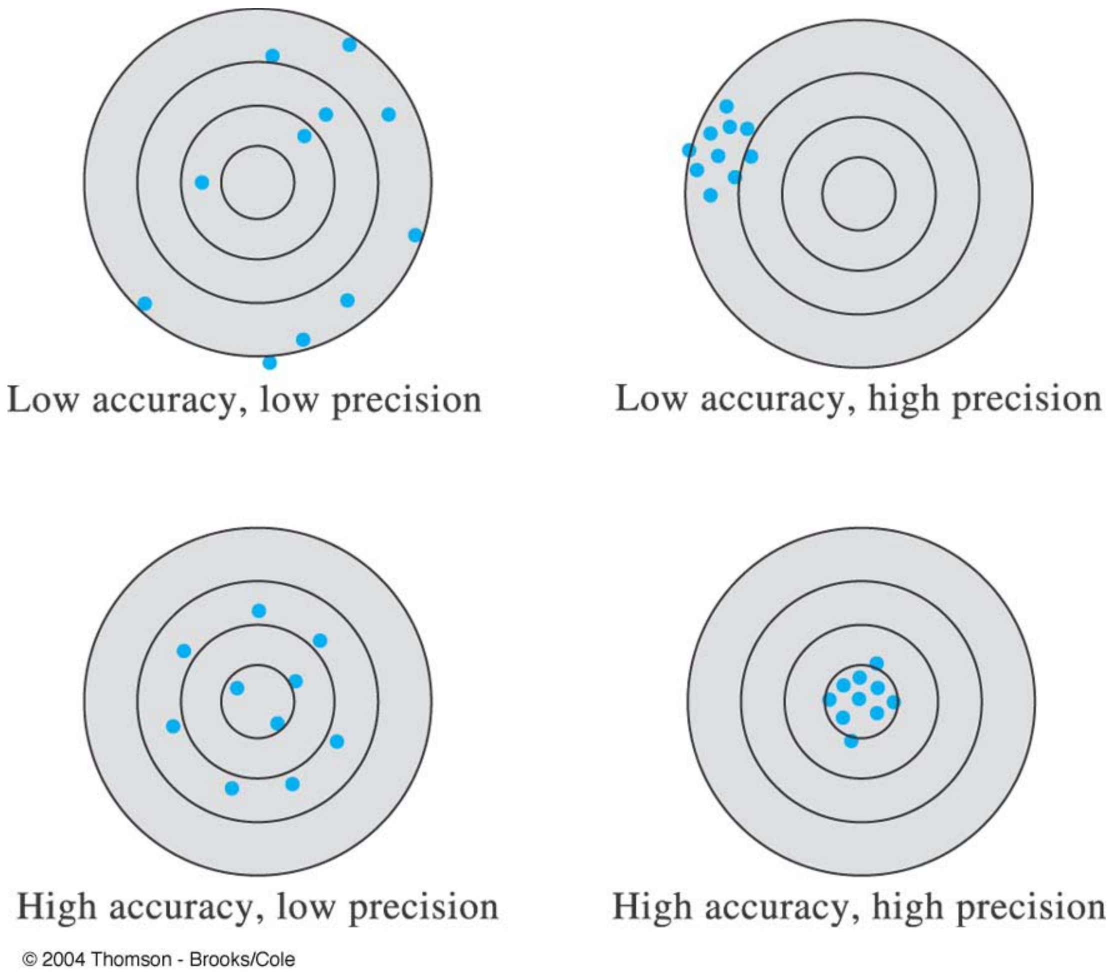
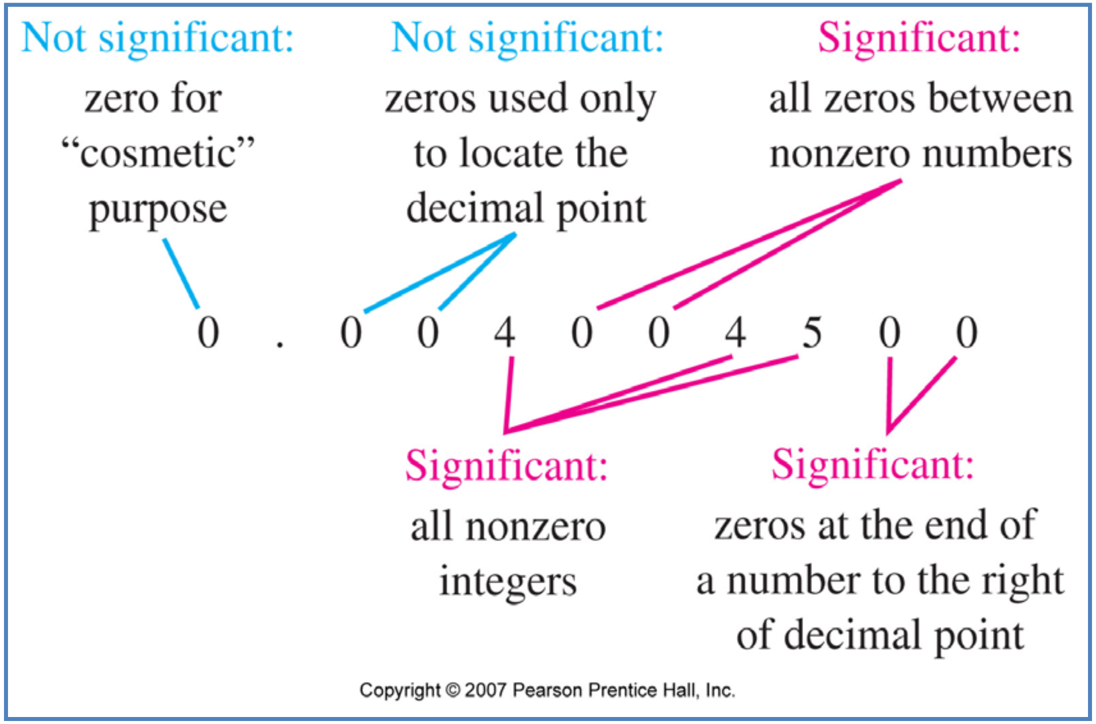


Figure 5-2 Illustration of accuracy and precision using the pattern of darts on a dartboard. Note that we can have very precise results (upper right) with a mean that is not accurate and an accurate mean (lower left) with data points that are imprecise.

1-7 Significant Figures

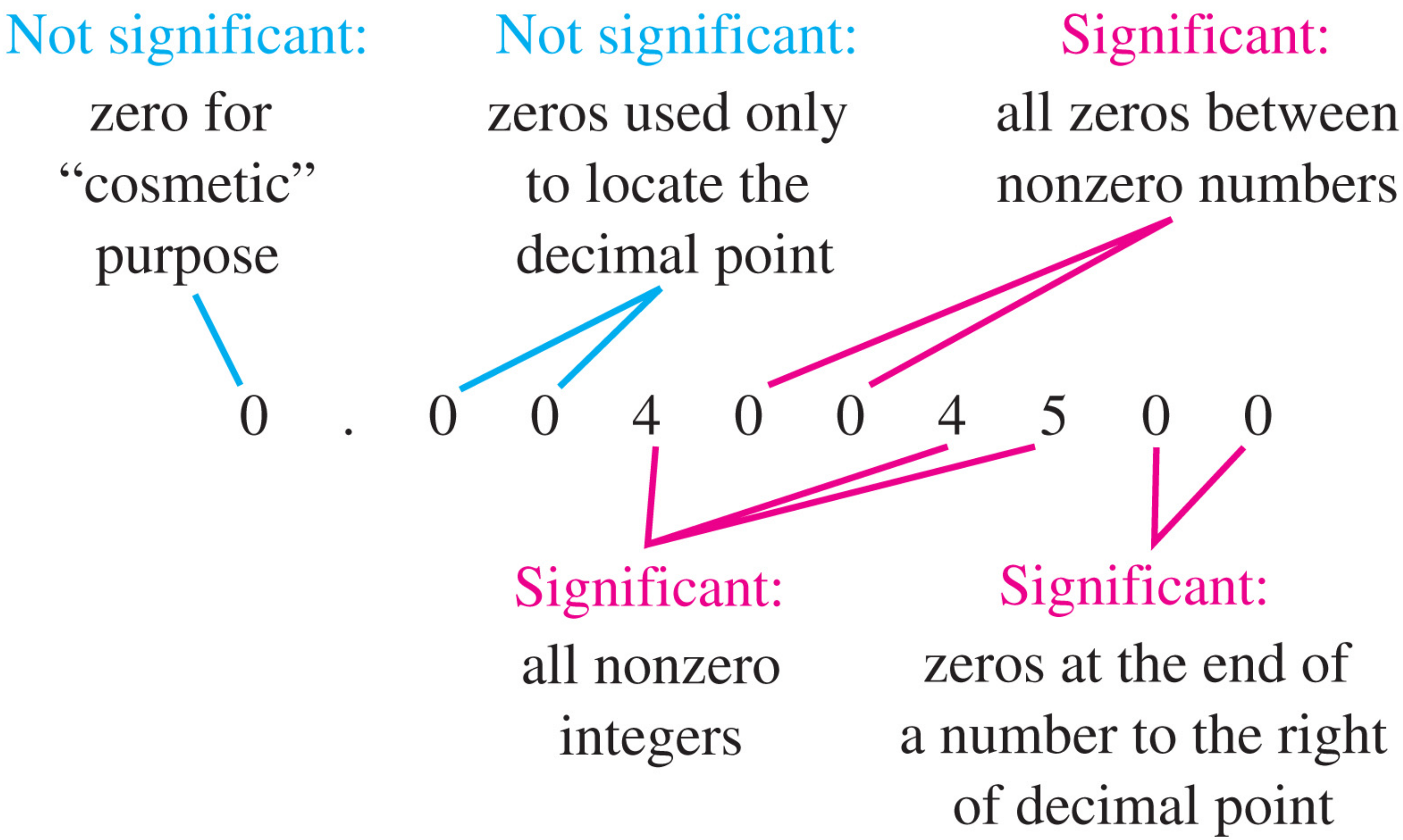


- All nonzero digits are *significant*.
- Zeros are also *significant*, but with two important *exceptions for quantities less than one*. Any zeros
  1- preceding the decimal point, or
  2- following the decimal point and preceding the first nonzero digit, are *not significant.*

-The case of terminal zeros that precede the decimal point in quantities *greater than one is ambiguous*.

*The quantity 7500 m is an example of an ambiguous case.

1-7 Significant Figures



▲ Figure 1-11

Determining the number of significant figures in a quantity

Significant Figures in numerical calculations:

Division and Multiplication:

The result of multiplication or division may contain only as many significant figures as the least precisely known quantity in the calculation.

14.79 cm * 12.11 cm * 5.05 cm = 904 cm3

(4 sig. fig.) (4 sig. fig.) (3 sig. fig.)   (3 sig. fig.)

Addition or subtraction :

The result of addition or subtraction must be expressed with the same number of digits beyond the decimal point as the quantity carrying the smallest number of such digits.

15.02 g + 9986.0 g + 3.518 g = 10,004.5 ~~3 8~~ g

Rounding Off Numerical Results

- increase the final digit by one unit if the digit dropped is
  - 5, 6, 7, 8, or 9 and
- leave the final digit unchanged if the digit dropped is
  - 0, 1, 2, 3, or 4.

* To three significant figures, 15.44 rounds off to 15.4,
and 15.45 rounds off to 15.5.
15.55 rounds to 15.6, and
17.65 rounds to 17.7.