Unit 1 Includes the Following OpenStax Sections

- 1.1 Chemistry in Context
- 1.4 Measurements
- 1.5 Measurement Uncertainty, Accuracy, and Precision
- 1.6 Mathematical Treatment of Measurement Results

Figure 1.1

Chemical substances and processes are essential for our existence, providing sustenance, keeping us clean and healthy, fabricating electronic devices, enabling transportation, and much more. (credit “left”: modification of work by “vxla”/Flickr; credit “left middle”: modification of work by “the Italian voice”/Flickr; credit “right middle”: modification of work by Jason Trim; credit “right”: modification of work by “gosheshe”/Flickr)

Learning Objectives

- 1.1 Chemistry in Context
  - Outline the historical development of chemistry
  - Provide examples of the importance of chemistry in everyday life
  - Describe the scientific method
  - Differentiate among hypotheses, theories, and laws
  - Provide examples illustrating macroscopic, microscopic, and symbolic domains

Chemistry in Context

- Chemistry is the study of the composition, properties, and interactions of matter.

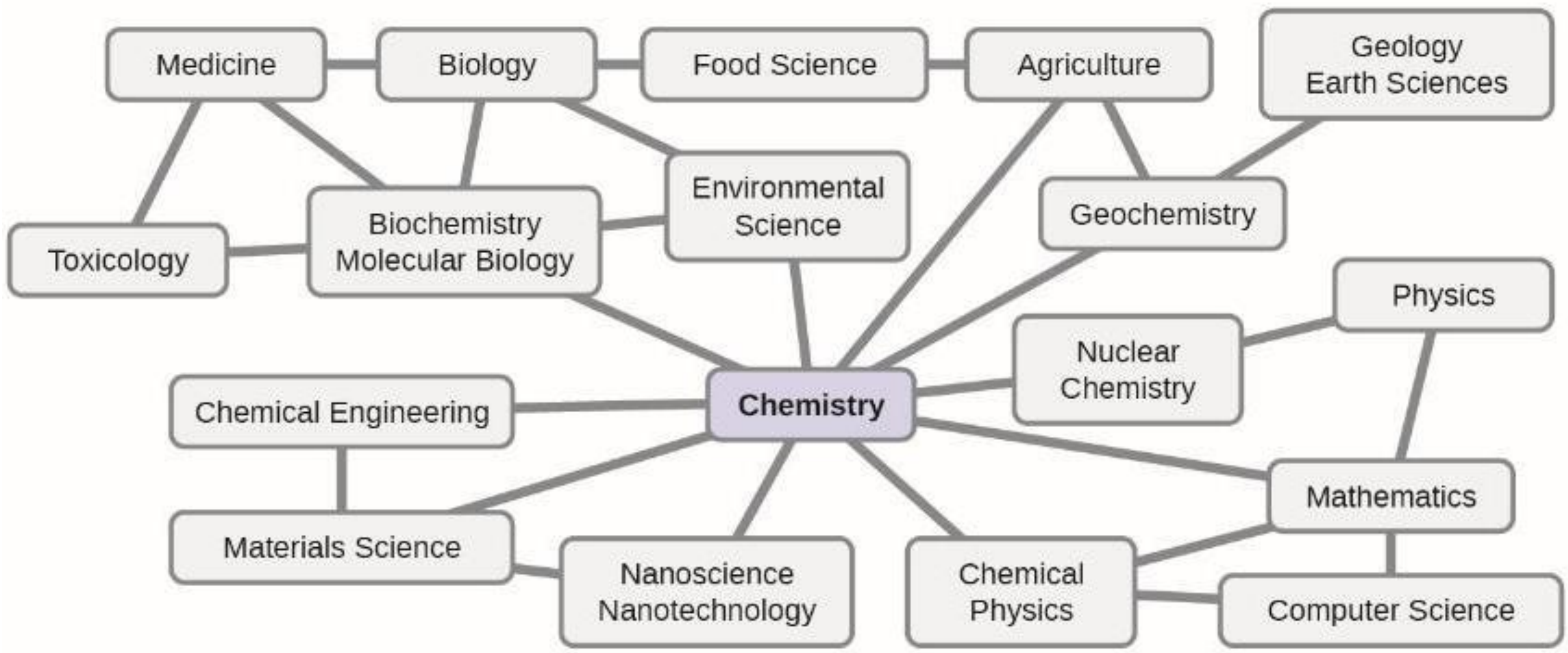
- Attempts to understand the behavior of matter extend back more than 2,500 years.
  - Greeks: Matter consists of four elements: earth, air, fire, and water.
  - Alchemists attempted to transform “base metals” into “noble metals.”

Figure 1.2

This portrayal shows an alchemist’s workshop circa 1580. Although alchemy made some useful contributions to how to manipulate matter, it was not scientific by modern standards. (credit: Chemical Heritage Foundation)

- Chemistry is interconnected to a vast array of other STEM disciplines.

Figure 1.3



Knowledge of chemistry is central to understanding a wide range of scientific disciplines. This diagram shows just some of the interrelationships between chemistry and other fields.

Chemistry and Everyday Life

- Examples of chemistry in everyday life:
    - Digesting food
    - Synthesizing polymers for clothing, cookware, and credit cards
    - Refining crude oil into gasoline and other products

- As you proceed through this course, you will discover:
    - Many different examples of changes in the composition and structure of matter.
    - How to classify these changes in matter and understand how they occur.
    - The changes in energy that accompany these changes in matter.

The Scientific Method

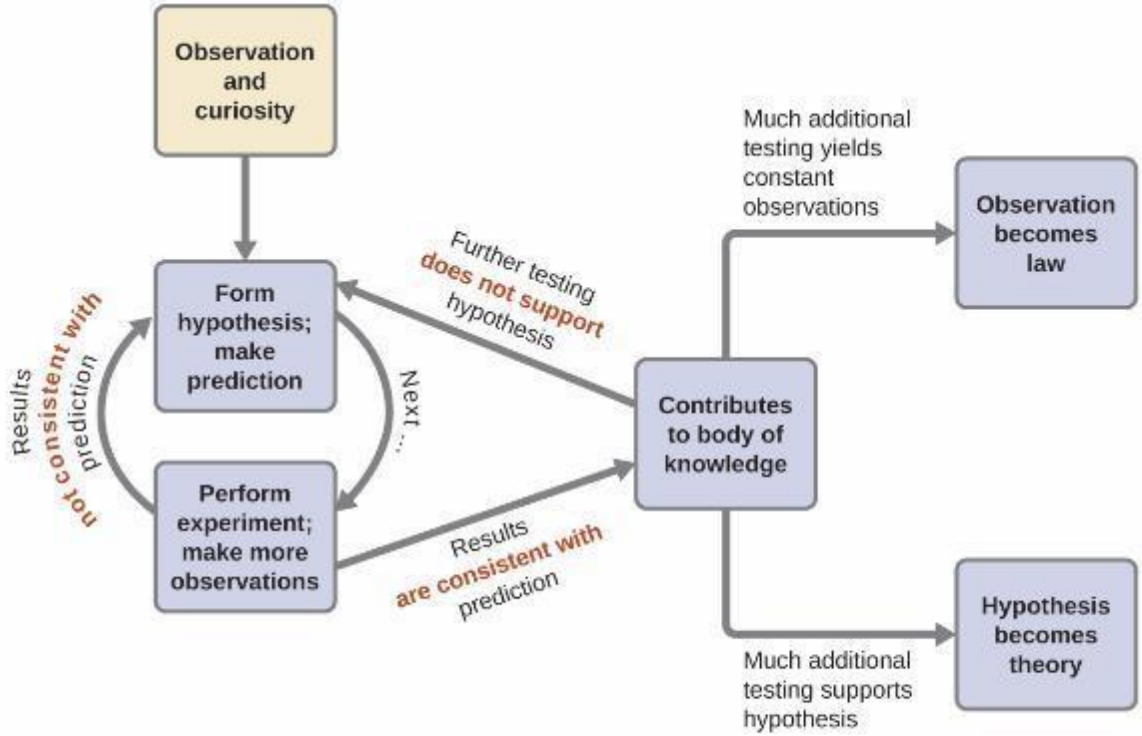
- Chemistry is a science based on observation and experimentation.
- Chemists often formulate a **hypothesis**: a tentative explanation of observations.
- The laws of science summarize a vast number of experimental observations, and describe or predict some facet of the natural world.
- **Theory:** A well-substantiated, comprehensive, testable explanation of a particular aspect of nature.

The Domains of Chemistry

Chemists study and describe the behavior of matter and energy in three different domains.

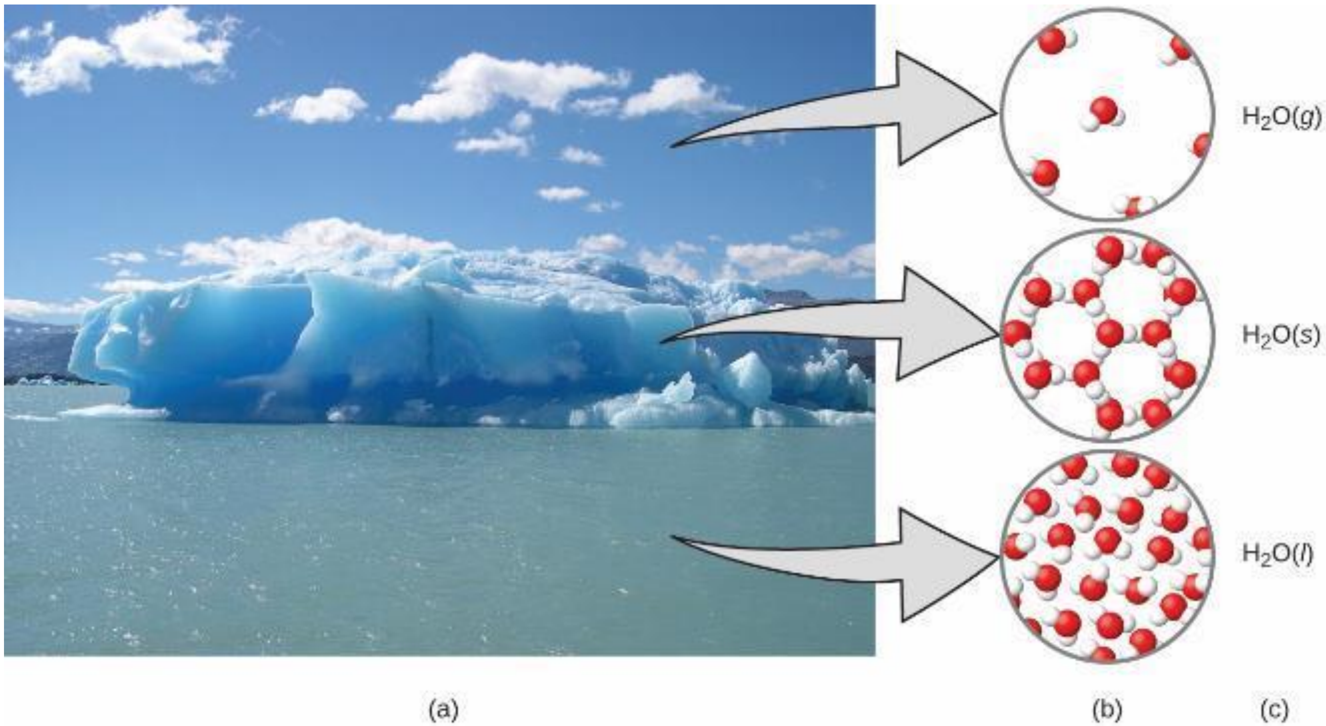
1) The macroscopic domain is familiar to us: It is the realm of everyday things that are large enough to be sensed directly by human sight or touch.
2) The microscopic domain of chemistry is almost always visited in the imagination. Micro also comes from Greek and means “small.” Some aspects of the microscopic domains are visible through a microscope.
3) The symbolic domain contains the specialized language used to represent components of the macroscopic and microscopic domains, such as chemical symbols.

Figure 1.4



The scientific method follows a process similar to the one shown in this diagram. All the key components are shown, in roughly the right order. Scientific progress is seldom neat and clean: It requires open inquiry and the reworking of questions and ideas in response to findings.

Figure 1.5



(a) Moisture in the air, icebergs, and the ocean represent water in the macroscopic domain. (b) At the molecular level (microscopic domain), gas molecules are far apart and disorganized, solid water molecules are close together and organized, and liquid molecules are close together and disorganized. (c) The formula H_2O symbolizes water, and (*g*), (*s*), and (*l*) symbolize its phases. Note that clouds are actually comprised of either very small liquid water droplets or solid water crystals; gaseous water in our atmosphere is not visible to the naked eye, although it may be sensed as humidity. (credit a: modification of work by “Gorkaazk”/Wikimedia Commons)

Learning Objectives

- 1.4 Measurements
  - Explain the process of measurement
  - Identify the three basic parts of a quantity
  - Describe the properties and units of length, mass, volume, density, temperature, and time
  - Perform basic unit calculations and conversions in the metric and other unit systems

Measurements

- Measurements provide the information that is the basis of most of the hypotheses, theories, and laws in chemistry.

- Every measurement provides three kinds of information:

  1) The size or magnitude of the measurement: a **number**

  2) A standard of comparison for the measurement: a **unit**

  3) An indication of the uncertainty of the measurement.

Units

- Without units, a number can be meaningless or confusing!

- In chemistry, we use an updated version of the metric system known as the International System of Units, or SI units.
    - Used since 1964

Table 1.2 Base Units of the SI System

Property	Name of Unit	Symbol of Unit
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K
electric current	ampere	A
amount of substance	mole	mol
luminous Intensity	candela	cd

Table 1.3 Common Unit Prefixes

- Fractional or multiple SI units are named using a prefix and the name of the base unit.

Prefix	Symbol	Factor
femto	f	10^{-15}
pico	p	10^{-12}
nano	n	10^{-9}
micro	m	10^{-6}
milli	m	10^{-3}
centi	c	10^{-2}
deci	d	10^{-1}

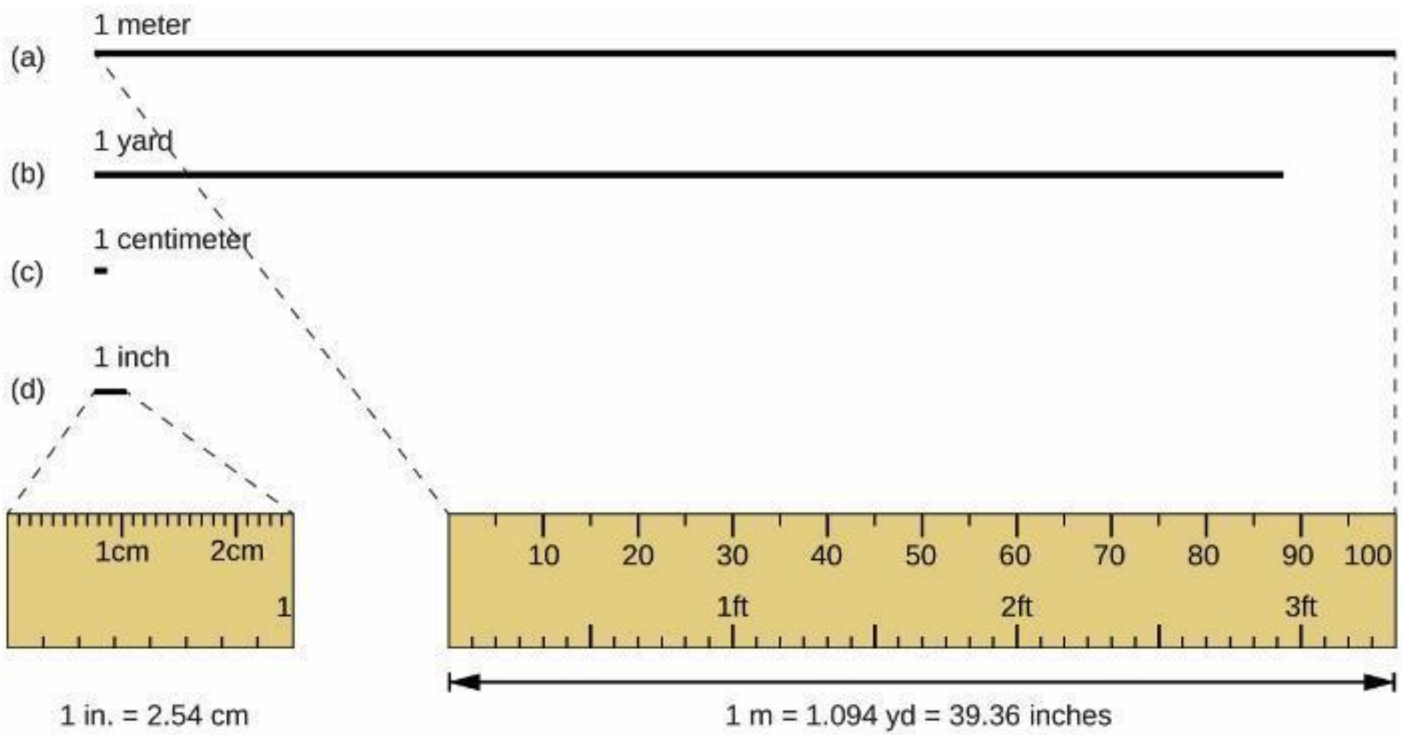
Table 1.3: Common Unit Prefixes

Prefix	Symbol	Factor
kilo	k	10^3
mega	M	10^6
giga	G	10^9
tera	T	10^{12}

Common SI Base Units: Length

- The SI unit of length is the meter (m).
- The meter was originally intended to be 1/10,000,000 of the distance from the North Pole to the equator.
- A meter is now defined as the distance light travels in a vacuum in 1/299,792,458 of 1 second.
- A meter is about 3 inches longer than 1 yard.

Figure 1.23



The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

Common SI Base Units: Mass

- The SI unit of mass is the kilogram (kg).
- A kilogram was originally defined as the mass of a liter of water.
- It is now defined by a certain cylinder of platinum-iridium alloy, which is kept in France.
- 1 kilogram is about 2.2 pounds.

Figure 1.24

This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology)

Common SI Base Units: Temperature

- The SI unit of temperature is the kelvin (K).
- No degree word nor symbol (°) is used with kelvin.
- The degree Celsius (°C) is also allowed in the SI system.
- Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places.
- Water freezes at 273.15 K (0 °C) and boils at 373.15 K (100 °C).

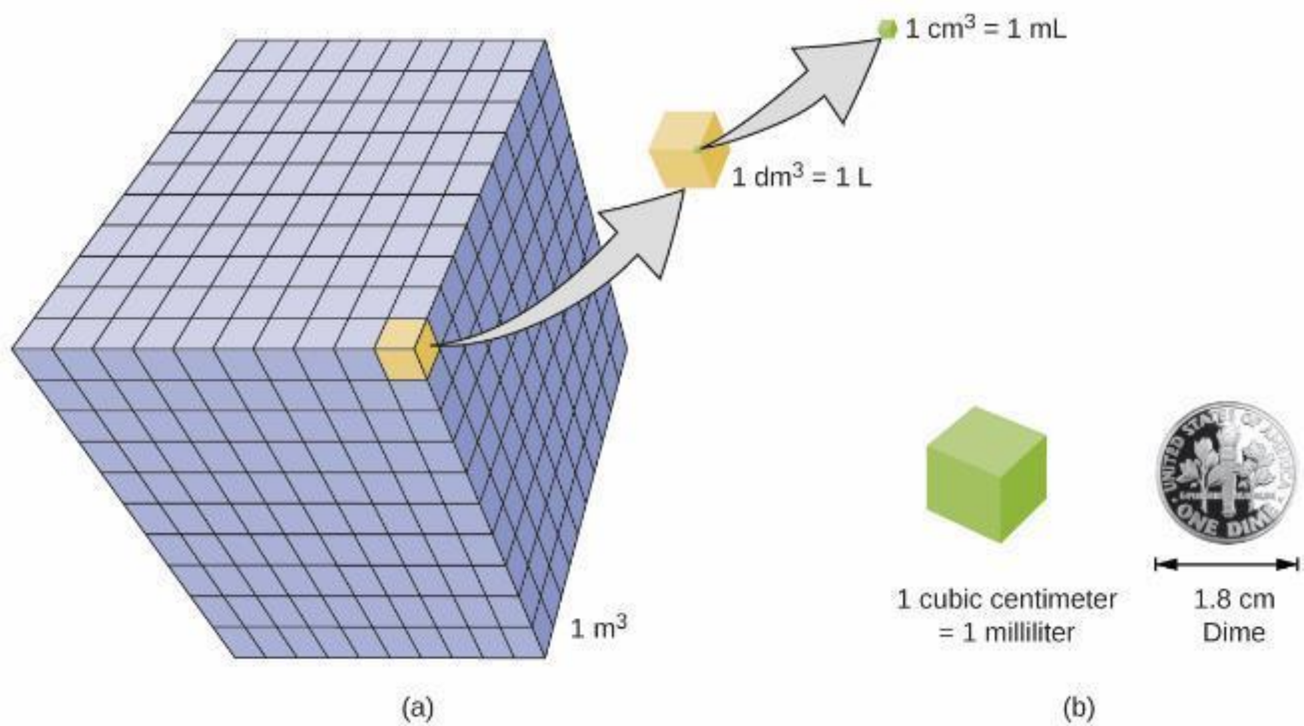
Common SI Base Units: Time

- The SI unit of time is the second (s).
- Smaller and larger time intervals can be expressed with the appropriate prefixes.
- Alternatively, hours, days, and years can be used.

Derived SI Units: Volume and Density

- We can derive many units from the seven SI base units.
- **Volume:** The measure of the amount of space occupied by an object.
- The standard SI unit for volume is the cubic meter (m3), which is derived from the SI base unit of length.
- Other units for volume are the liter (L) and milliliter (mL).
  - 1 dm^3 = 1 L
  - 1 cm^3 = 1 mL

Figure 1.25



(a) The relative volumes are shown for cubes of 1 m^3, 1 dm^3 (1 L), and 1 cm^3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm^3 (1-mL) cube.

Density

- The density of a substance is the ratio of the mass of a sample of the substance to its volume.

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

- The standard SI unit for density is the kilogram per cubic meter (kg/m^3).

- Commonly used density units based on state of matter:

g/cm^3	(solids, liquids)
g/L	(gases)

Learning Objectives

- 1.5 Measurement Uncertainty, Accuracy, and Precision
  - Define accuracy and precision
  - Distinguish exact and uncertain numbers
  - Correctly represent uncertainty in quantities using significant figures
  - Apply proper rounding rules to computed quantities

Measurement Uncertainty, Accuracy, and Precision

- Counting is the only type of measurement that is free from uncertainty.

- The result of a counting measurement is an example of an **exact number**.

- The numbers for defined quantities are also exact.
  - 1 ft. is exactly 12 in.
  - 1 in. is exactly 2.54 cm
  - 1 g is exactly 0.001 kg

1.5 Measurement Uncertainty, Accuracy, and Precision

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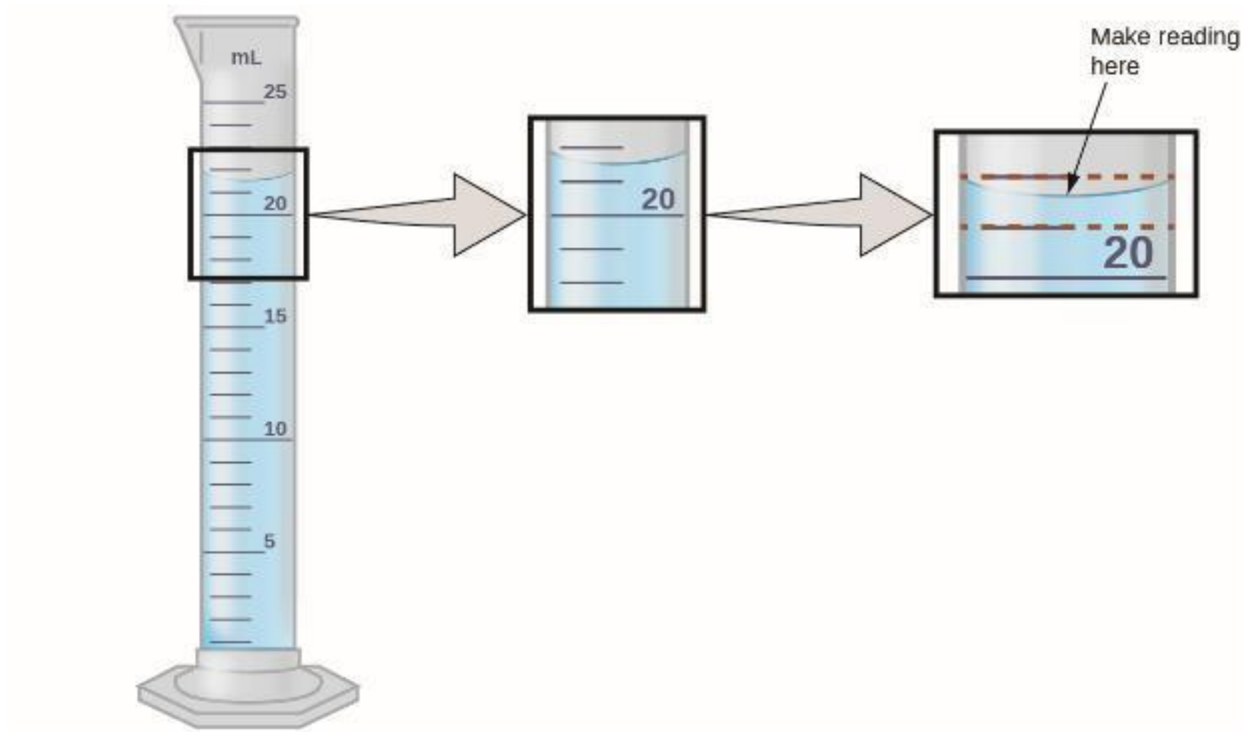
- Quantities derived from measurements other than counting are uncertain to varying extents.
- These numbers are *not* exact.
- There are always practical limitations of the measurement process used.
- A measured number must be reported in a way to indicate its uncertainty.
- In general, when recording a measurement, you are allowed to estimate one uncertain digit.

Figure 1.26



To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

Significant Figures

- On the previous slide, if one recorded the volume in the graduated cylinder to be 21.6 mL:
  - 2 and 1 are certain digits.
  - 6 is an estimate.
  - Someone else might perceive the volume to be 21.5 mL or 21.7mL.

- All the digits in a measurement, including the uncertain last digit, are called **significant figures** or significant digits.

- Frequently, we need to know the number of significant figures in a measurement reported by someone else.

Significant Figures

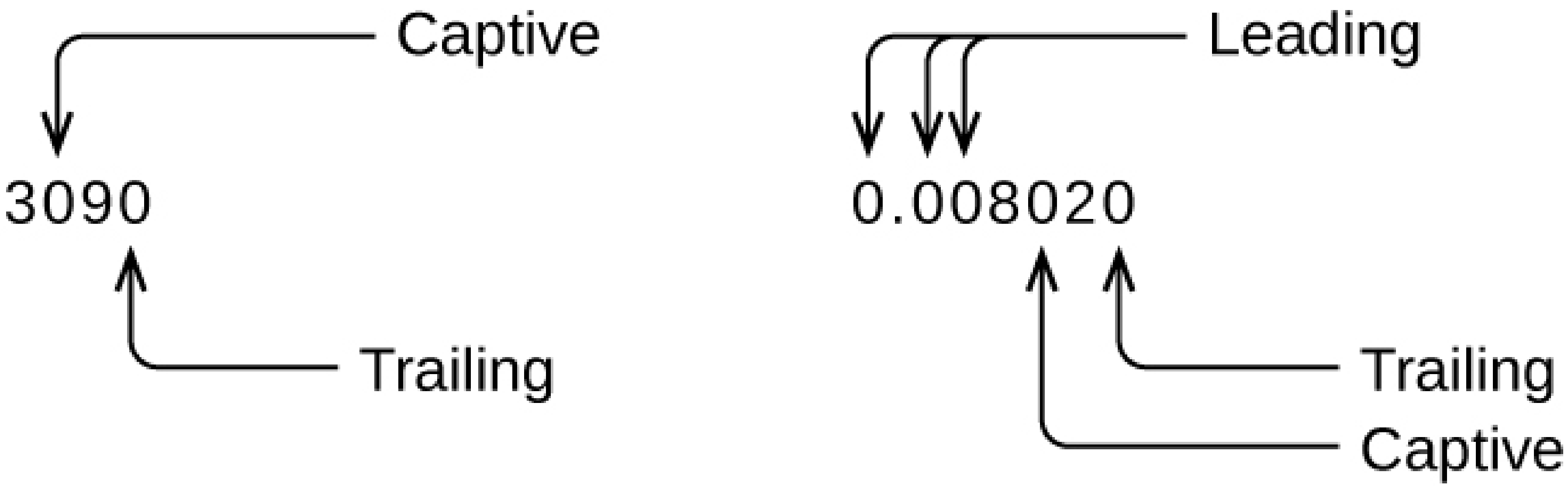
These numbers are always *significant*.

- Nonzero digits
- Captive zeros
- Trailing zeroes
  - When they are to the right of the decimal place
  - When in scientific notation

These numbers are always *not significant*.

- Leading zeros
- Trailing zeros
  - When they are to the left of the decimal place

Significant Figures



Significant Figures in Calculations

Results calculated from measured numbers are at least as uncertain as the measurement itself.

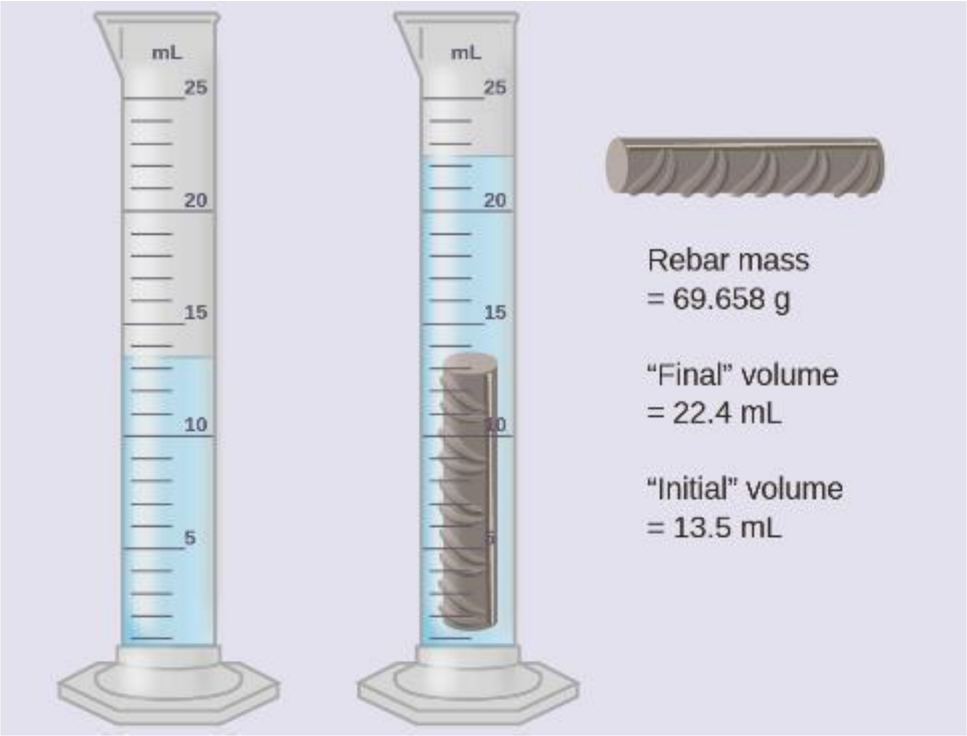
1) When we add or subtract numbers, we should round the result to the same number of decimal places as the number with the least number of decimal places (the least precise value in terms of addition and subtraction).
2) When we multiply or divide numbers, we should round the result to the same number of digits as the number with the least number of significant figures (the least precise value in terms of multiplication and division).
3) If the digit to be dropped (the one immediately to the right of the digit to be retained) is less than 5, we “round down” and leave the retained digit unchanged; if it is more than 5, we “round up” and increase the retained digit by 1; if the dropped digit is 5, we round up or down, whichever yields an even value for the retained digit.

Significant Figures in Calculations

The following examples illustrate the application of this rule in rounding a few different numbers to three significant figures:

- 0.028675 rounds “up” to 0.0287 (the dropped digit, 7, is greater than 5)
- 18.3384 rounds “down” to 18.3 (the dropped digit, 3, is less than 5)
- 6.8752 rounds “up” to 6.88 (the dropped digit is 5, and the retained digit is even)
- 92.85 rounds “down” to 92.8 (the dropped digit is 5, and the retained digit is even)

Example 1.7



mL
25
20
15
10
5
mL
25
20
15
10
5
Rebar mass
= 69.658 g
"Final" volume
= 22.4 mL
"Initial" volume
= 13.5 mL

Example 1.7

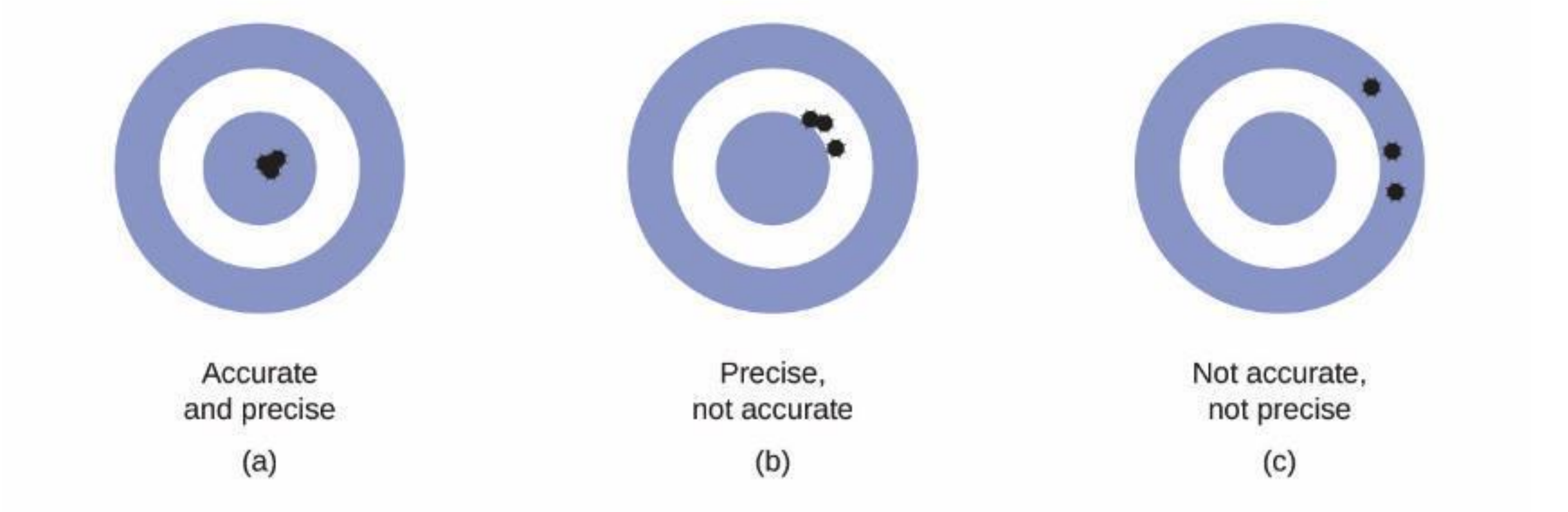
mL
25
20
15
10
5
mL
25
20
15
10
Mass
= 51.842 g
"Final" volume
= 19.8 mL
"Initial" volume
= 17.1 mL

Accuracy and Precision

- A measurement is said to be **precise** if it yields very similar results when repeated in the same manner.

- A measurement is considered **accurate** if it yields a result that is very close to the true or accepted value.

Figure 1.27



(a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

Table 1.5 Volume (ML) of Cough Medicine Delivered by 10-oz (296 ML) Dispensers

Dispenser #1	Dispenser # 2	Dispenser # 3
283.3	298.3	296.1
284.1	294.2	295.9
283.9	296.0	296.1
284.0	297.8	296.0
284.1	293.9	296.1

- Dispenser #1 is precise, but not accurate.
- Dispenser #2 is more accurate, but less precise.
- Dispenser #3 is both accurate and precise.

Learning Objectives

- 1.6 Mathematical Treatment of Measurement Results
  - Explain the dimensional analysis (factor label) approach to mathematical calculations involving quantities
  - Use dimensional analysis to carry out unit conversions for a given property and computations involving two or more properties

Mathematical Treatment of Measurement Results

- A quantity of interest may not be easy (or even possible) to measure directly but instead must be calculated from other directly measured properties and appropriate mathematical relationships.

- The mathematical approach we will be using is known as **dimensional analysis**.

- Dimensional analysis is based on the premise that the units of quantities must be subjected to the same mathematical operations as their associated numbers.

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Conversion Factors and Dimensional Analysis

- A conversion factor or unit conversion factor is a ratio of two equivalent quantities expressed with different measurement units.

- Example: The lengths 2.54 centimeters and 1 inch are equivalent.

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

$$\frac{2.54\ \text{cm}}{1\ \text{in.}} \text{ or } \frac{1\ \text{in.}}{2.54\ \text{cm}}$$

Table 1.6 Common Conversion Factors

Length	Volume	Mass
1 m = 1.0936 yd.	1 L = 1.0567 qt.	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt. = 0.94635 L	1 lb = 453.59 g
1 km = 0.62137 mi	1 ft^3 = 28.317 L	1 (avoirdupois) oz = 28.349 g
1 mi = 1609.3 m	1 tbsp = 14.1787 mL	1 (troy) oz = 31.103 g

Conversion Factors and Dimensional Analysis

- We must use the form of the conversion factors that results in the original unit canceling out, leaving only the sought unit.

- Example: Convert 34 inches to centimeters.

$$34\ \text{in.} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in.}}} = 86\ \text{cm}$$

Conversion of Temperature Units

- Temperature refers to the hotness or coldness of a substance.

- **Celsius scale**
  - Water freezes at 0 °C.
  - Water boils at 100 °C.

- **Fahrenheit scale**
  - Water freezes at 32 °F.
  - Water boils at 212 °F.

- 100 °C covers the same temperature interval as 180 °F.

Conversion of Temperature Units

- The SI unit of temperature is the kelvin (K).
- Unlike the Celsius and Fahrenheit scales, the kelvin scale is an absolute temperature scale.
- Zero kelvin corresponds to the lowest temperature that can theoretically be achieved.
- **Kelvin scale**
  - Water freezes at 273.15 K.
  - Water boils at 373.15 K.
- 100 °C covers the same temperature interval as 100 K.

Mathematical Relationships Between Temperature Scales

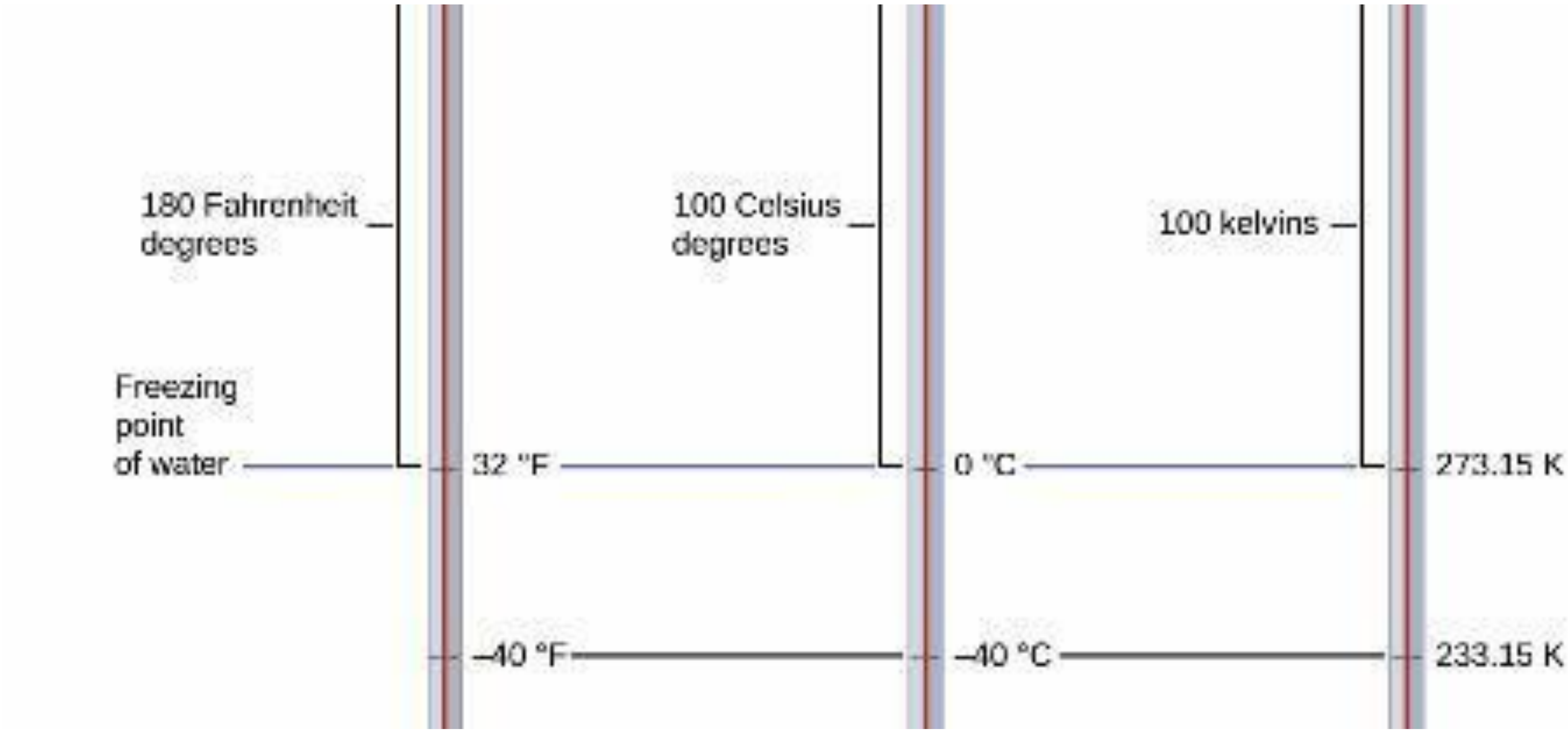
Fahrenheit and Celsius

$$T_{^\circ\mathrm{F}} = \left(\frac{9\ ^\circ\mathrm{F}}{5\ ^\circ\mathrm{C}} \times T_{^\circ\mathrm{C}} \right) + 32\ ^\circ\mathrm{C}$$

Kelvin and Celsius

$$T_{\mathrm{K}} = T_{^\circ\mathrm{C}} + 273.15$$

Figure 1.28



The Fahrenheit, Celsius, and kelvin temperature scales are compared.

Exercise 55