GENERAL CHEMISTRY

Principles and Modern Applications TENTH EDITION

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Chapter 1: Matter—Its Properties and Measurement

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1-1 Properties of Matter

Matter is anything that occupies space and displays the properties of *mass* and inertia.

Composition refers to the parts or components of a sample of matter and their relative proportions

Properties are qualities that we can use to distinguish one sample of matter from others (physical and chemical).

Physical Properties and Physical Changes

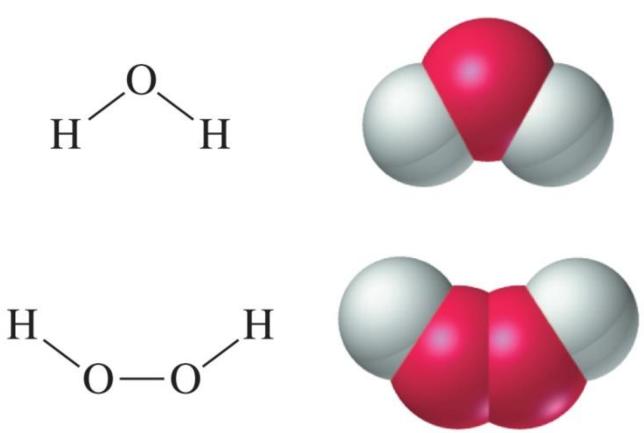
- A **physical property** is one that a sample of matter displays without changing its composition.
- In such a physical change, some of the physical properties of the sample may change, but its composition remains unchanged. When liquid water freezes into solid water (ice), it certainly looks different and, in many ways, it is different. Yet, the water remains 11.19% hydrogen and 88.81% oxygen by mass.

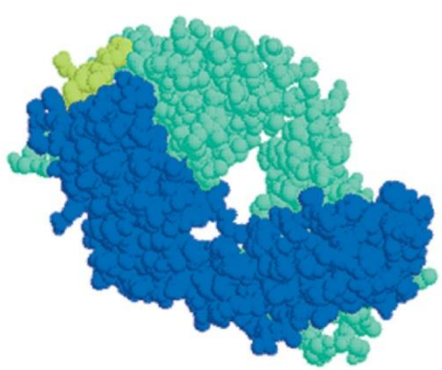
Chemical Properties and Chemical Changes

- In a chemical change, or chemical reaction, one or more kinds of matter are converted to new kinds of matter with different compositions. The key to identifying chemical change, then, comes in observing a change in composition.
- The burning of paper involves a chemical change.
- The ability of paper to burn is an example of a chemical property. A **chemical property** is the ability (or inability) of a sample of matter to undergo a change in composition under stated conditions.

1-2 Classification of Matter

- Matter is made up of very tiny units called **atoms**. Each different type of atom is the building block of a different **chemical element**.
- **Compounds** are comprised of two or more elements.
- **Molecules** are the smallest units of compounds.
- The composition and properties of an element or a compound are uniform throughout a given sample and from one sample to another.



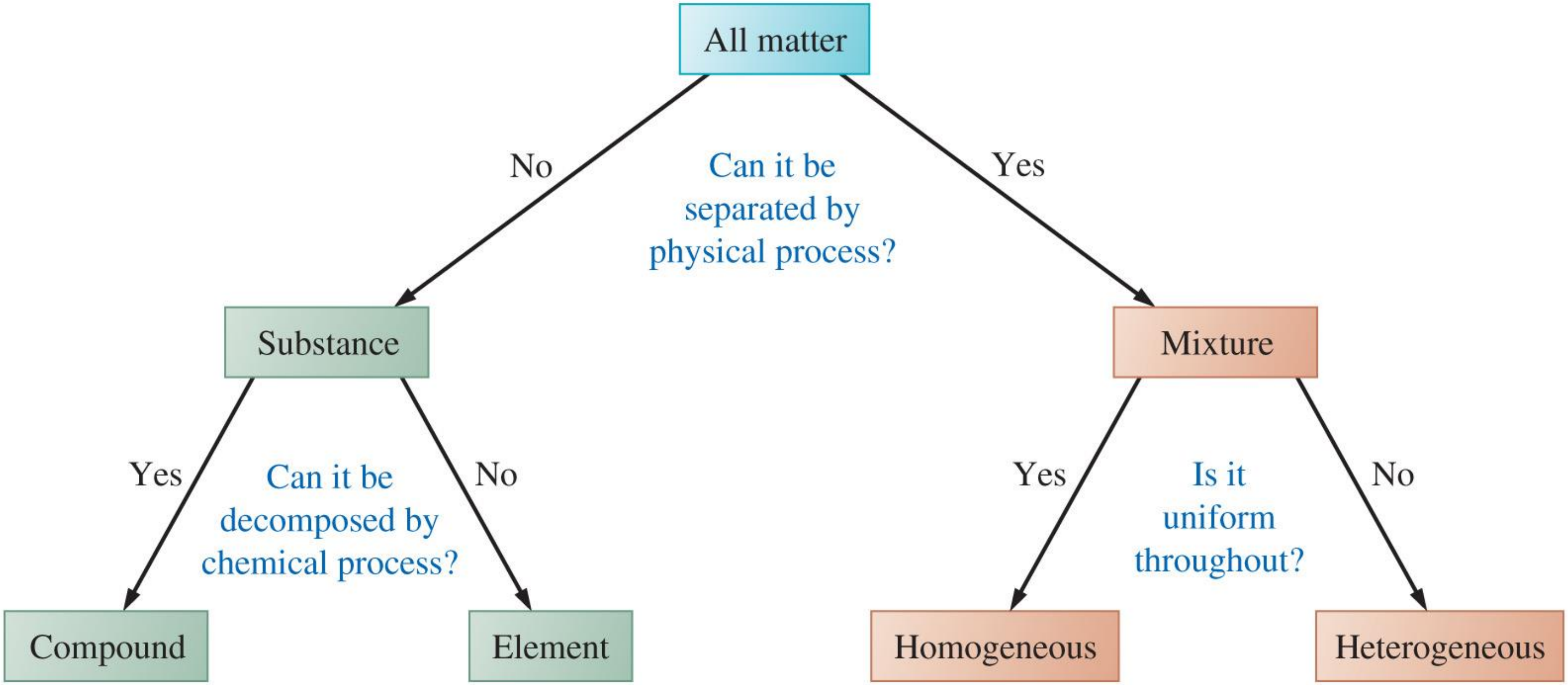


Gamma globulin

1-2 Classification of Matter

- 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**.
- In **heterogeneous mixtures**, the composition and physical properties vary from one part of the mixture to another.

Classification of Matter



1-3 The Measurement of Matter

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

The scientific system of measurement is called the *Système Internationale d Unités (International System of Units)* and is abbreviated **SI**. It is a modern version of the metric system, a system based on the unit of length called a meter (m).
There are seven fundamental quantities in the SI system.

SI is a decimal system. Quantities differing from the base unit by powers of ten are noted by the use of prefixes.

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)

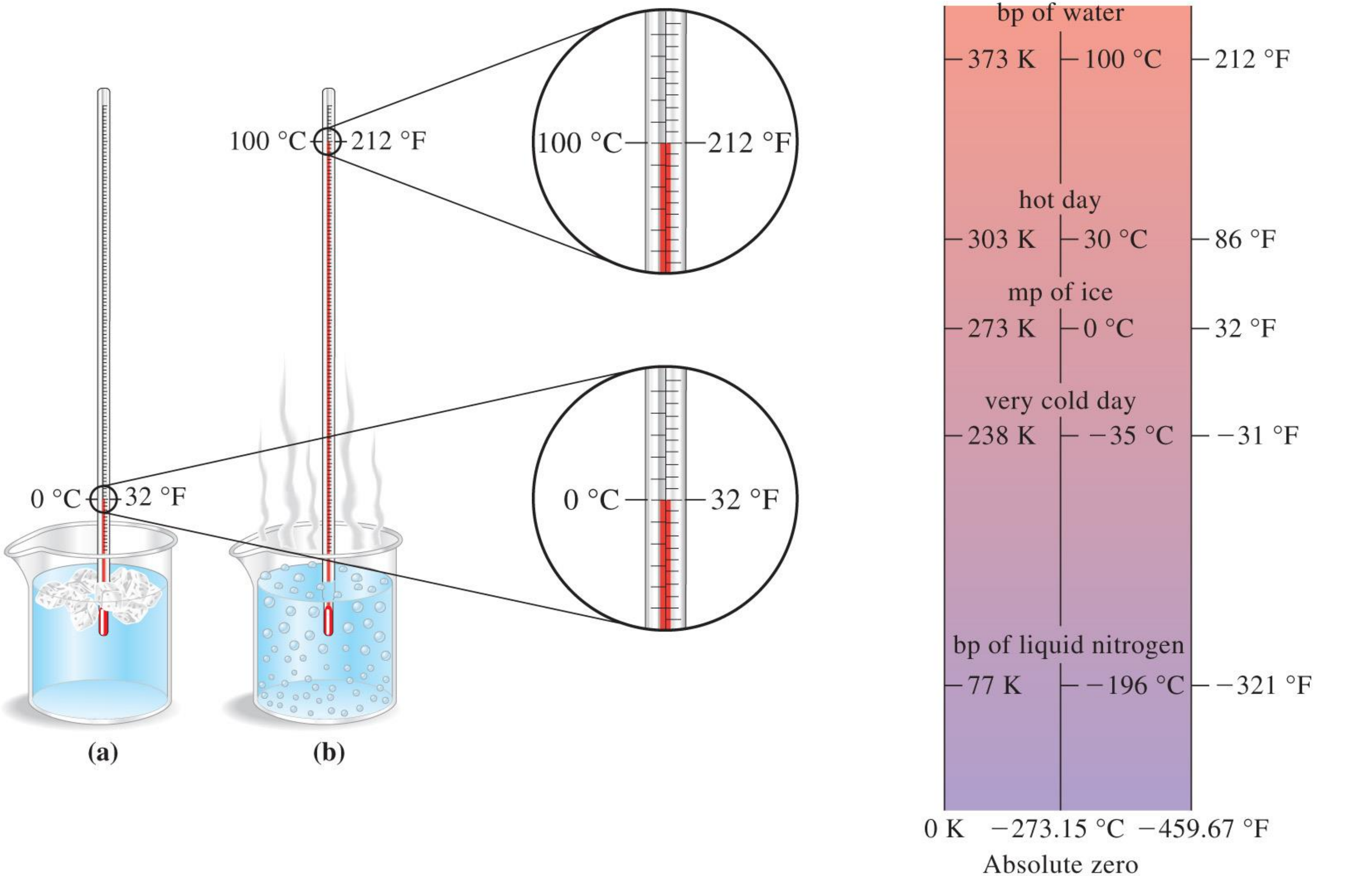
[a]The Greek letter μ (pronounced "mew").

Mass

- Mass is the **quantity** of matter in an object.
- In SI the standard of mass is 1 *kilogram* (kg).
- Weight is the force of gravity on an object

$W \propto m \qquad W = g \cdot m$

Temperature



Kelvin from Celsius $T(\text{K}) = t(°\text{C}) + 273.15$

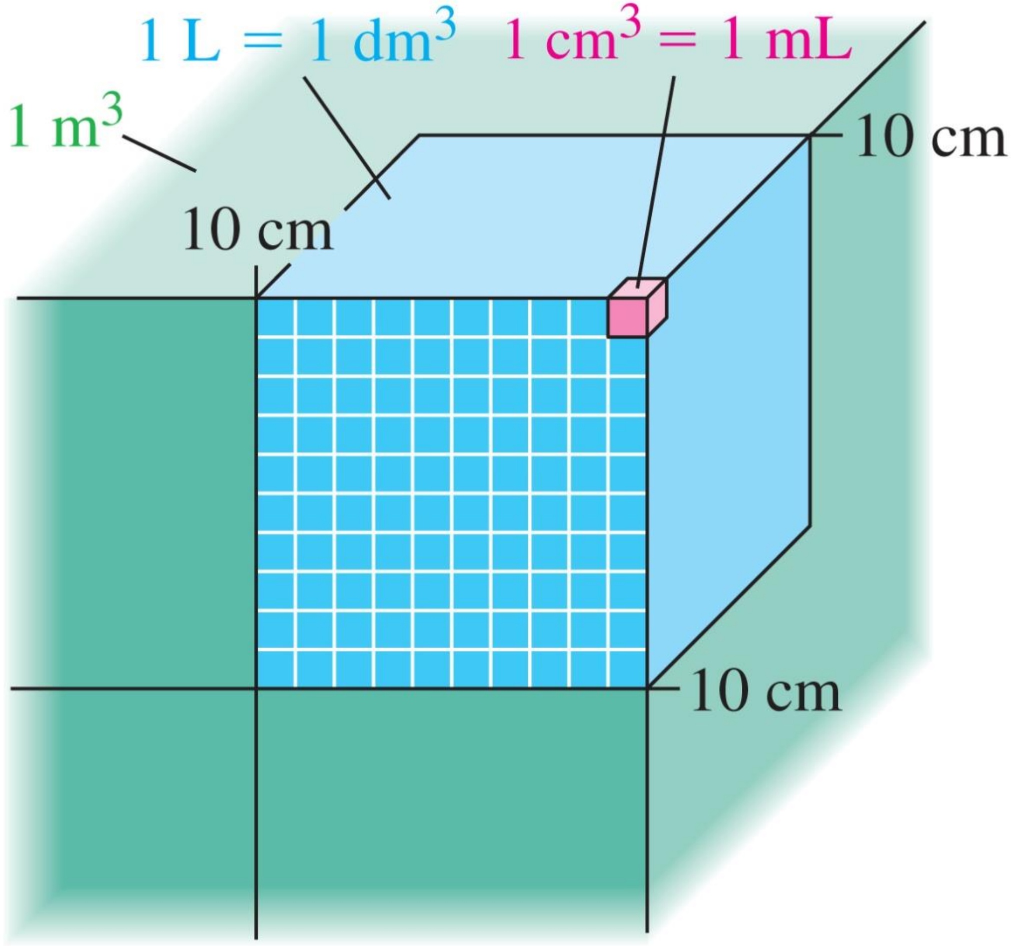
Fahrenheit from Celsius $t(°\text{F}) = \frac{9}{5}t(°\text{C}) + 32$

Celsius from Fahrenheit $t(°\text{C}) = \frac{5}{9}[t(°\text{F}) - 32]$

PRACTICE EXAMPLE A: A recipe in an American cookbook calls for roasting a cut of meat at 350 °F. What is this temperature on the Celsius scale?

$$t(°\text{C}) = \frac{5}{9}[350\text{-}32] = 177\ °\text{C}$$

Volume



Units

S.I. Units

Length	meter, m
Mass	Kilogram, kg
Time	second, s
Temperature	Kelvin, K
Quantity	Mole, 6.022×10^{23} mol^{-1}

Derived Quantities

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

Other Common 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

1-4 Density and Percent Composition

- **Density** is the ratio of mass to volume.

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

g/cm^3 or g/mL

Mass and volume are extensive properties

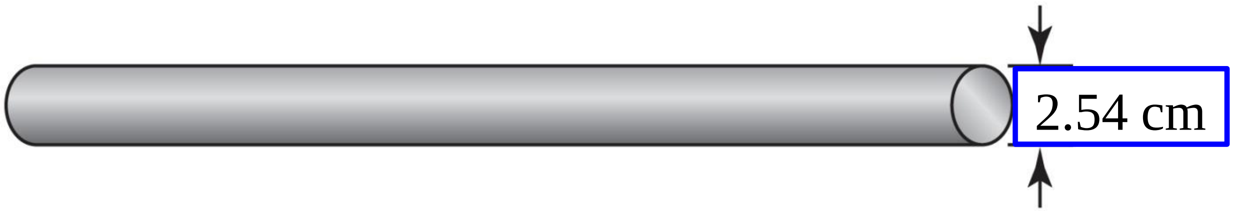
Density is an intensive property

- Solid densities: from about 0.2 g/cm^3 to 20 g/cm^3
- Liquid densities: from about 0.5 g/mL to 3–4 g/mL
- Gas densities: mostly in the range of a few grams per liter

EXAMPLE 1-2

Relating Mass, Volume and Density

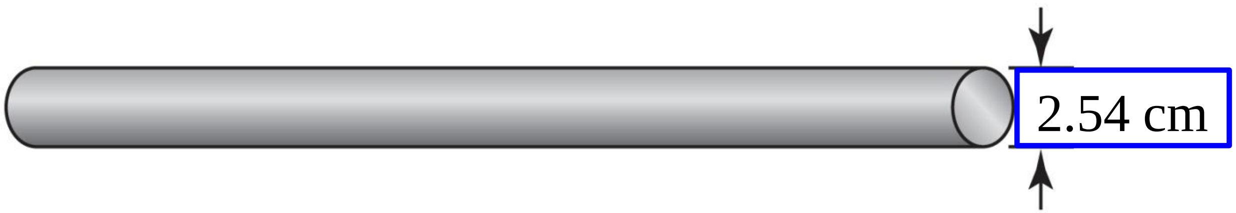
The stainless steel in the solid cylindrical rod pictured below has a density of 7.75 g/cm^3. If we want a 1.00 kg mass of this rod, how long a section must we cut off?



Calculate the Volume needed $$V = \frac{m}{d} = m \times \frac{1}{d}$$

Calculate the Length $$V = h\,\pi r^2 \qquad h = \frac{V}{\pi r^2}$$

EXAMPLE 1-2



Calculate the Volume needed $$V = \frac{m}{d} = m \times \frac{1}{d}$$

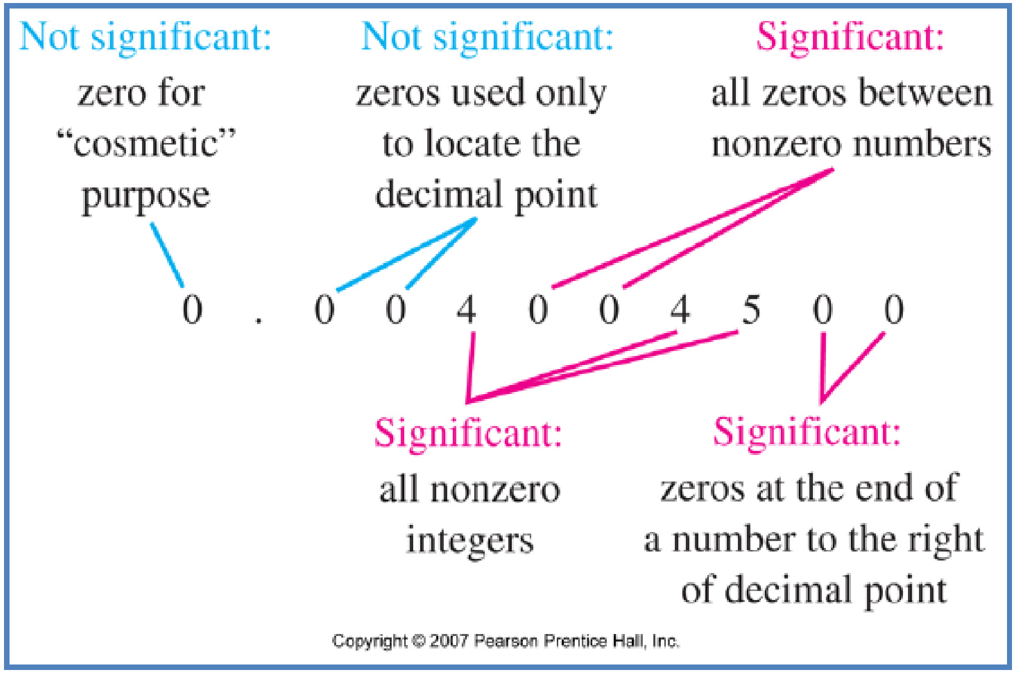
$$V = (1.00\,\cancel{kg})(1000\,\cancel{g}/\cancel{kg}) \times \frac{1\ cm^3}{7.75\ \cancel{g}} = 129\ cm^3$$

Calculate the Length $$V = h\,\pi r^2 \qquad h = \frac{V}{\pi r^2}$$

$$h = \frac{V}{\pi r^2} = \frac{129\ cm^{\cancel{3}}}{(3.14)\{(0.5)(2.54\ cm)\}^{\cancel{2}}} = 25.5\ cm$$

1-5 Significant Figures

- The use of appropriate significant figures is important. Because the correct application of the significant numbers rule avoids needlessly suggesting a high degree of precision that comes from calculations and does not guarantee precision in real experiments.



- 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.

Example

Number	Significant figure	Number	Significant figure
0,0050 L	2	1,34000 x 10 ^{7} nm	6
18,00 g	4	5600 ng	2
0,00012 kg	2	87000 L	2
83,000 L	5	78 002,3 ng	6
0,006002 g	4	0,000007800 g	4

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.