2 Standards for Measurement

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Careful and accurate measurements of ingredients are important both when cooking and in the chemistry laboratory!

Foundations of College Chemistry, 14th Ed.

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Chapter Outline

Scientific Notation

Scientific Notation: A way to write very large or small numbers (measurements) in a compact form.

2.468×10^8

Number written from 1-10 — **Raised to a power (-/+ or fractional)**

Method for Writing a Number in Scientific Notation

1. **Move the decimal point in the original number so that it is located after the first nonzero digit.**
2. **Multiply this number by 10 raised to the number of places the decimal point was moved.**
3. **Exponent sign indicates which direction the decimal was moved.**

Scientific Notation Practice

Write 0.000423 in scientific notation.

Place the decimal between the 4 and 2.
4.23

The decimal was moved 4 places so the exponent should be a 4.

The decimal was moved to the right so the exponent should be negative.

$\boxed{4.23 \times 10^{-4}}$

Scientific Notation Practice

What is the correct scientific notation for the number 353,000 (to 3 significant figures)?

a. 35.3×10^4

b. 3.53×10^5

c. 0.353×10^6

d. 3.53×10^{-5}

e. 3.5×10^5

Measurement and Uncertainty

Measurement: A quantitative observation.

Examples: 1 cup, 3 eggs, 5 molecules, etc.

Measurements are expressed by

1. **a numerical value and**
2. **a unit of the measurement.**

Example: 50 kilometers

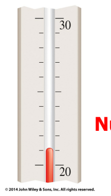
Numerical Value — **Unit**

A measurement always requires a unit.

Measurement and Uncertainty

Every measurement made with an instrument requires estimation.

Uncertainty exists in the last digit of the measurement because this portion of the numerical value is estimated.



21.2 °C

The other two digits are certain. These digits would not change in readings made by one person to another.

Numerical values obtained from measurements are never exact values.

Measurement and Uncertainty

Some degree of uncertainty exists in all measurements.

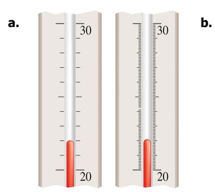
By convention, a measurement typically includes all certain digits plus one digit that is estimated.

Because of this level of uncertainty, any measurement is expressed by a limited number of digits.

These digits are called significant figures.

Measurement and Uncertainty

a. Recorded as 22.0 °C
(3 significant figures with uncertainty in the last digit)
b. Recorded as 22.11 °C
(4 significant figures with uncertainty in the last digit)



22.0 °C **22.11 °C**

Significant Figures

Because all measurements involve uncertainty, we must be careful to use the correct number of significant figures in calculations.

Rules for Counting Significant Figures

1. All nonzero digits are significant.
2. Some numbers have an infinite number of sig figs
Ex. 12 inches are always in 1 foot
Exact numbers have no uncertainty.
3. Zeroes are significant when:
a. They are in between non zero digits
Ex. 75.04 has 4 significant figures (7,5,0 and 4)

Significant Figures

Rules for Counting Significant Figures

3. Zeroes are significant when:
b. They are at the end of a number after a decimal point.
Ex. 32.410 has five significant figures (3,2,4,1 and 0)
4. Zeroes are not significant when:
a. They appear before the first nonzero digit.
Ex. 0.00321 has three significant figures (3,2 and 1)
b. They appear at the end of a number without a decimal point.
Ex. 6920 has three significant figures (6,9 and 2)

When in doubt if zeroes are significant, use scientific notation!

Let's Practice!

How many significant figures are in the following measurements?

3.2 inches	**2 significant figures**
25.0 grams	**3 significant figures**
103 people	**Exact number (∞ number of sig figs)**
0.003 kilometers	**1 significant figure**

Rounding Off Numbers

With a calculator, answers are often expressed with more digits than the proper number of significant figures.

These extra digits are omitted from the reported number, and the value of the last digit is determined by rounding off.

Rules for Rounding Off

If the first digit after the number that will be retained is:

1. < 5, the digit retained does not change.
   Ex. 53.2305 = 53.2 (other digits dropped)
   digit retained
2. > 5, the digit retained is increased by one.
   Ex. 11.789 = 11.8 (other digits dropped)
   digit rounded up to 8

Let's Practice!

Round off the following numbers to the given number of significant figures.

79.137 (four)	79.14
0.04345 (three)	0.0435
136.2 (three)	136
0.1790 (two)	0.18

Significant Figures in Calculations

The results of a calculation are only as precise as the least precise measurement.

Calculations Involving Multiplication or Division

The significant figures of the answer are based on the measurement with the least number of significant figures.

Example

79.2 x 1.1 = 87.12

The answer should contain two significant figures, as 1.1 contains only two significant figures.

79.2 x 1.1 = 87

Let's Practice!

Round the following calculation to the correct number of significant figures.

$$\frac{(12.18)(5.2)}{13} = 4.872$$

a. 4.9
b. 4.87
c. 4.8
d. 4.872
e. 5.0

The answer is rounded to 2 sig figs.
(5.2 and 13 each contain only 2 sig. figures)

Significant Figures in Calculations

The results of a calculation are only as precise as the least precise measurement.

Calculations Involving Addition or Subtraction

The significant figures of the answer are based on the precision of the least precise measurement.

Example Add 136.23, 79, and 31.7.

$$\begin{array}{r} 136.23 \\ 79 \\ \underline{31.7} \\ 246.93 \end{array}$$

The least precise number is 79, so the answer should be rounded to 247.

Let's Practice!

Round the following calculation to the correct number of significant figures.

142.57 - 13.0

a. 129.57
b. 129.6
c. 130
d. 129.5
e. 129

$$\begin{array}{r} 142.57 \\ \underline{-\ 13.0} \\ 129.57 \end{array}$$

The answer is rounded to the tenths place.

Let's Practice!

Round the following calculation to the correct number of significant figures.

$$\frac{12.18 - 5.2}{10.1}$$

a. 0.69109
b. 0.70
c. 0.693
d. 0.69

$$\begin{array}{r} 12.18 \\ -\ 5.2 \\ \hline 6.98 \end{array}$$

The numerator must be rounded to the tenths place.

$$\frac{7.0}{10.1} = 0.693069$$

Final answer is now rounded to 2 significant figures.

Let's Practice!

How many significant figures should the answer to the following calculation contain?

1.6 + 23 – 0.005

a. 1
b. 2
c. 3
d. 4

$$\begin{array}{r} 1.6 \\ 23 \\ 0.005 \\ \hline 24.595 \end{array}$$

Round to least precise number (23).
Round to the ones place (25).

The Metric System

Metric or International System (SI):
Standard system of measurements for mass, length, time and other physical quantities.
Based on standard units that change based on factors of 10.
Prefixes are used to indicate multiples of 10.
This makes the metric system a decimal system.

Quantity	Unit Name	Abbreviation
Length	Meter	m
Mass	Kilogram	kg
Temperature	Kelvin	K
Time	Second	s
Amount of Substance	Mole	mol
Electric current	Ampere	A

The Metric System

Common Prefixes and Numerical Values for SI Units

Prefix	Symbol	Numerical Value	Power of 10
Mega	M	1,000,000	10^{6}
Kilo	k	1,000	10^{3}
—	—	1	10^{0}
Deci	d	0.1	10^{-1}
Centi	c	0.01	10^{-2}
Milli	m	0.001	10^{-3}
Micro	μ	0.000001	10^{-6}
Nano	n	0.000000001	10^{-9}

Measurements of Length

Meter (m): standard unit of length of the metric system.

Definition: the distance light travels in a vacuum during 1/299,792,458 of a second.

Common Length Relationships:

1 meter (m) = 100 centimeters (cm)
= 1000 millimeters (mm)
1 kilometer (km) = 1000 meters

Relationship Between the Metric and English System:

1 inch (in.) = 2.54 cm

Dimensional Analysis: A Problem Solving Method

Dimensional analysis: converts one unit of measure to another by using conversion factors.

Conversion factor: A ratio of equivalent quantities.

Example:
1 km = 1000 m

Conversion factor: $\frac{1\ \text{km}}{1000\ \text{m}}$ or $\frac{1000\ \text{m}}{1\ \text{km}}$

Conversion factors can always be written two ways.
Both ratios are equivalent quantities and will equal 1.

Dimensional Analysis: A Problem Solving Method

Any unit can be converted to another unit by multiplying the quantity by a conversion factor.

Unit_1 x conversion factor = Unit_2

Example

$$2\ \text{m} \times \frac{1\ \text{km}}{1000\ \text{m}} = 0.002\ \text{km}$$

Units are treated like numbers and can cancel.

A conversion factor must cancel the original unit and leave behind only the new (desired) unit.

The original unit must be in the *denominator* and new unit must be in the *numerator* to cancel correctly.

Dimensional Analysis: A Problem Solving Method

Many chemical principles or problems are illustrated mathematically.

A systematic method to solve these types of numerical problems is key.

Our approach: the dimensional analysis method

Create solution maps to solve problems.

Overall outline for a calculation/conversion progressing from known to desired quantities.

Dimensional Analysis: A Problem Solving Practice

Convert 215 centimeters to meters.

Solution Map:

known quantity cm ⟶ m **desired quantity**

$$215\ \cancel{\text{cm}} \times \frac{1\ \text{m}}{100\ \cancel{\text{cm}}} = 2.15\ \text{m}$$

Convert 125 meters to kilometers.

Solution Map:

known quantity m ⟶ km **desired quantity**

$$125\ \cancel{\text{m}} \times \frac{1\ \text{km}}{1000\ \cancel{\text{m}}} = 0.125\ \text{km}$$

Let's Practice!

How many micrometers are in 0.03 meters?

a. 30,000
b. 300,000
c. 300
d. 3000

Solution Map:

known quantity m ⟶ μm **desired quantity**

$$0.03\ \cancel{\text{m}} \times \frac{1{,}000{,}000\ \mu\text{m}}{1\ \cancel{\text{m}}} = 30{,}000\ \mu\text{m}$$

Dimensional Analysis: A Problem Solving Method

Some problems require a series of conversions to get to the desired unit.

Each arrow in the solution map corresponds to the use of a conversion factor.

Example

Convert from days to seconds.

Solution Map:

days ⟶ hours ⟶ minutes ⟶ seconds

$$1\ \cancel{\text{day}} \times \frac{24\ \cancel{\text{hours}}}{1\ \cancel{\text{day}}} \times \frac{60\ \cancel{\text{minutes}}}{1\ \cancel{\text{hour}}} \times \frac{60\ \text{seconds}}{1\ \cancel{\text{minute}}} = 8.64 \times 10^4\ \text{sec}$$

Dimensional Analysis: A Problem Solving Practice

Metric to English Conversions

How many feet are in 250 centimeters?

Solution Map:

cm ⟶ inches ⟶ ft

$$250\ \cancel{\text{cm}} \times \frac{1\ \cancel{\text{inch}}}{2.54\ \cancel{\text{cm}}} \times \frac{1\ \text{foot}}{12\ \cancel{\text{inches}}} = 8.20\ \text{ft}$$

Let's Practice!

Metric to English Conversions

How many meters are in 5 yards?

a. 9.14
b. 457
c. 45.7
d. 4.57

Solution Map:

yards ⟶ feet ⟶ inches ⟶ cm ⟶ m

$$5\ \cancel{\text{yards}} \times \frac{3\ \cancel{\text{feet}}}{1\ \cancel{\text{yard}}} \times \frac{12\ \cancel{\text{inches}}}{1\ \cancel{\text{foot}}} \times \frac{2.54\ \cancel{\text{cm}}}{1\ \cancel{\text{inch}}} \times \frac{1\ \text{m}}{100\ \cancel{\text{cm}}} = 4.57\ \text{m}$$

Let's Practice!

Metric to English Conversions

How many cm^3 are in a box that measures 2.20 x 4.00 x 6.00 inches?

Solution Map:

(in ⟶ cm)^3

$$2.20\ \text{in} \times 4.00\ \text{in} \times 6.00\ \text{in} = 52.8\ \text{in}^3$$

$$52.8\ \cancel{\text{in}^3} \times \left(\frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}}\right)^3 = 865\ \text{cm}^3$$

Measurement of Mass

Mass: amount of matter in an object

Mass is measured on a balance.

Weight: effect of gravity on an object.

Weight is measured on a scale, which measures force against a spring.

Mass is independent of location, but weight is not.

Mass is the standard measurement of the metric system.

The SI unit of mass is the kilogram. (The gram is too small a unit of mass to be the standard unit.)

Measurement of Mass

1 kilogram (kg) is the mass of a Pt-Ir cylinder standard.

Metric to English Conversions

1 kg = 2.2015 pounds (lbs)

Metric Units of Mass

Prefix	Symbol	Gram Equivalent	Exponential Equivalent
kilogram	kg	1000 g	10^3 g
gram	g	1 g	10^0 g
decigram	dg	0.1 g	10^{-1} g
centigram	cg	0.01 g	10^{-2} g
milligram	mg	0.001 g	10^{-3} g
microgram	μg	0.000001 g	10^{-6} g

Let's Practice!

Convert 343 grams to kilograms.

Solution Map:

g ⟶ kg

Use the new conversion factor:

$$\frac{1\ \text{kg}}{1000\ \text{g}} \quad \text{or} \quad \frac{1000\ \text{g}}{1\ \text{kg}}$$

$$343\ \cancel{\text{g}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 0.343\ \text{kg}$$

Let's Practice!

How many centigrams are in 0.12 kilograms?

a. 120
b. 1.2 x 10^4
c. 1200
d. 1.2

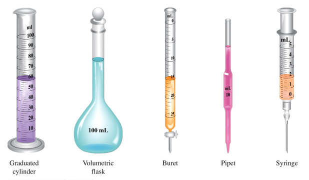
Solution Map:

kg ⟶ g ⟶ cg

$$0.12\ \cancel{\text{kg}} \times \frac{1000\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{100\ \text{cg}}{1\ \cancel{\text{g}}} = 1.2 \times 10^4\ \text{cg}$$

Measurement of Volume

Volume: the amount of space occupied by matter.
The SI unit of volume is the cubic meter (m^3)
The metric volume more typically used is the liter (L) or milliliter (mL).
A liter is a cubic decimeter of water (1 kg) at 4 °C.



Volume can be measured with several laboratory devices.

Measurement of Volume

Common Volume Relationships

1 L = 1000 mL = 1000 cm^3
1 mL = 1 cm^3
1 L = 1.057 quarts (qt)

Volume Problem

Convert 0.345 liters to milliliters.

Solution Map:

$$\text{L} \longrightarrow \text{mL}$$

$$0.345\,\cancel{\text{L}} \times \frac{1000\text{ mL}}{1\,\cancel{\text{L}}} = 345\text{ mL}$$

Let's Practice!

How many milliliters are in a cube with sides measuring 13.1 inches each?

a. 3690
b. 3.69
c. 369
d. 3.69 x 10^4

Solution Map:

$$\text{in.} \longrightarrow \text{cm} \longrightarrow \text{cm}^3 \longrightarrow \text{mL}$$

Convert from inches to cm:

$$13.1\,\cancel{\text{in.}} \times \frac{2.54\text{ cm}}{1\,\cancel{\text{in.}}} = 33.3\text{ cm}$$

Determine the volume of the cube:

$$\text{Volume} = (33.3\text{ cm}) \times (33.3\text{ cm}) \times (33.3\text{ cm}) = 3.69 \times 10^4\text{ cm}^3$$

Convert to the proper units:

$$3.69 \times 10^4\,\cancel{\text{cm}^3} \times \frac{1\text{ mL}}{1\,\cancel{\text{cm}^3}} = 3.69 \times 10^4\text{ mL}$$

Measurement of Temperature

Thermal energy: **A form of energy involving the motion of small particles of matter.**

Temperature: **measure of the intensity of thermal energy of a system (i.e. how hot or cold).**

Heat: **flow of energy due to a temperature difference.**

Heat flows from regions of higher to lower temperature.

The SI unit of temperature is the Kelvin (K).

Temperature is measured using a thermometer.

Different Temperature Scales

Temperature can be expressed in 3 commonly used scales. Celsius (°C), Fahrenheit (°F), and Kelvin (K).

Celsius and Fahrenheit are both measured in degrees, but the scales are different.

H_2O	°C	°F	K
Freezing Point	0 °C	32 °F	273.15 K
Boiling Point	100 °C	212 °F	373.15 K

The Fahrenheit scale has a range of 180° between freezing and boiling.

The lowest temperature possible on the Kelvin scale is absolute zero (-273.15 °C).

Converting Between Temperature Scales

Mathematical Relationships Between Temperature Scales

$$\boxed{\begin{array}{c}\text{K} = {}^\circ\text{C} + 273.15 \\ {}^\circ\text{F} = 9/5({}^\circ\text{C}) + 32\end{array}}$$

Temperature Problem

Convert 723 °C to temperature in both K and °F.

Solution Map:

$${}^\circ\text{C} \longrightarrow \text{K}$$

$$\text{K} = 723 + 273.15 = 996\text{ K}$$

$${}^\circ\text{C} \longrightarrow {}^\circ\text{F}$$

$${}^\circ\text{F} = 9/5(723) + 32 = 1333\ {}^\circ\text{F}$$

Let's Practice!

What is the temperature if 98.6 °F is converted to °C?

a. 37
b. 371
c. 210
d. 175

Solution Map:

°F $\longrightarrow$ °C

98.6 = 9/5 (°C) + 32

98.6 - 32 = 9/5 (°C)

66.6 = 9/5 (°C)

°C = (5/9)(66.6) = 37 °C

Density

Density (d): the ratio of the mass of a substance to the volume occupied by that mass.

$$\boxed{d = \frac{\text{mass}}{\text{volume}}}$$

Density is a physical property of a substance.
The units of density are generally expressed as g/mL or g/cm^3 for solids and liquids and g/L for gases.
The volume of a liquid changes as a function of temp, so density must be specified for a given temperature.
Ex. The density of H_2O at 4 °C is 1.0 g/mL while the density is 0.97 g/mL at 80 °C.

Density: Specific Gravity

Specific gravity (sp gr): ratio of the density of a substance to the density of another substance (usually H_2O at 4 °C).

Specific gravity is unit-less (in the ratio all units cancel).

Let's Practice!

Calculate the density of a substance if 323 g occupy a volume of 53.0 mL.

Solution:

$$\boxed{d = \frac{\text{mass}}{\text{volume}}}$$

$$\frac{323\text{ g}}{53.0\text{ mL}} = 6.09\text{ g/mL}$$

Let's Practice!

The density of gold is 19.3 g/mL.
What is the volume of 25.0 g of gold?

Solution Map:

Use density as a conversion factor!

g Au $\longrightarrow$ mL Au

$$25.0\cancel{\text{ g}} \times \frac{1\text{ mL}}{19.3\cancel{\text{ g}}} = 1.30\text{ mL}$$

Let's Practice!

What is the mass of 1.50 mL of ethyl alcohol?
(d = 0.789 g/mL at 4 °C)

a. 1.90 g
b. 1.18 g
c. 0.526 g
d. 2.32 g
e. 1.50 g

Solution Map:

mL $\longrightarrow$ g

$$1.50\cancel{\text{ mL}} \times \frac{0.789\text{ g}}{1\cancel{\text{ mL}}} = 1.18\text{ g}$$

Learning Objectives

2.1 Scientific Notation

Write decimal numbers in scientific notation.

2.2 Measurement and Uncertainty

Explain the significance of uncertainty in measurements in chemistry and how significant figures are used to indicate a measurement.

2.3 Significant Figures

Determine the number of significant figures in a given measurement and round measurements to a specific number of significant figures.

Learning Objectives

2.4 Significant Figures in Calculations

Apply the rules for significant figures in calculations involving addition, subtraction, multiplication, and division.

2.5 The Metric System

Name the units for mass, length, and volume in the metric system and convert from one unit to another.

2.6 Dimensional Analysis: A Problem Solving Method

Use dimensional analysis to solve problems involving unit conversions.

Learning Objectives

2.7 Measurement of Temperature

Convert measurements among the Fahrenheit, Celsius and Kelvin temperature scales.

2.8 Density

Solve problems involving density.