General Chemistry:
An Integrated Approach
Hill, Petrucci, 4th Edition

Chapter 1
Chemistry:
Matter and Measurement

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Chemistry: Principles and Applications

Chemical theory and applications are interwoven like the threads of a fine fabric.

Chlorine gas is an example of a chemical with many applications. Understanding the theory makes for better use of this chemical.

The applications of chemistry, much like the science itself, undergo **constant change**.

Getting Started: Some Key Ideas

CHEMISTRY

The study of the composition, structure, and properties of matter and of changes that occur in matter

Composition – what's in it?

E.g., water is 2 parts Hydrogen and 1 part Oxygen

Structure – how is it assembled?

E.g., crystals

Properties:

E.g., boiling point, density, flammability

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Key Ideas, cont.

Matter = anything that has mass and occupies space

Key Terms

- **Atoms** are the smallest distinctive units in a sample of matter.
- **Molecules** are larger units in which two or more atoms are joined together.
  - Examples: Water consists of molecules, each having two atoms of hydrogen and one of oxygen.
  - Oxygen *gas* consists of molecules, each having two atoms of oxygen.

Water molecule

Oxygen molecule

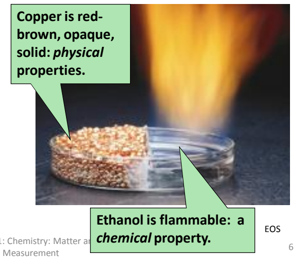
Properties of Matter

Physical property: characteristic displayed by a sample of matter without undergoing any change in its composition

e.g., color

Chemical property: characteristics displayed as a result of change in composition

e.g., flammability



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Physical and Chemical Changes

Physical Change: changes in appearance but not in composition
e.g., sublimation of ice in the winter

Chemical Change: changes resulting in altered composition and/or molecular structure
e.g., spoilage of foods

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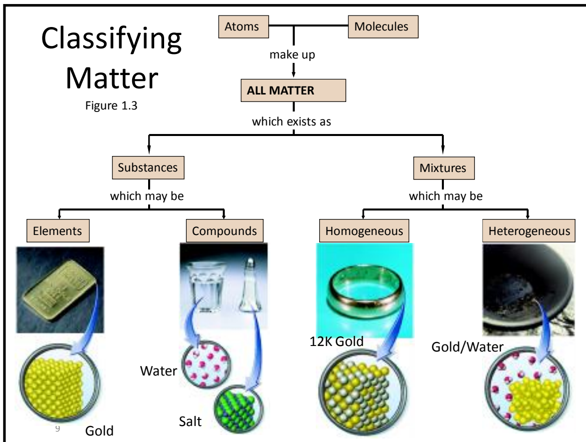
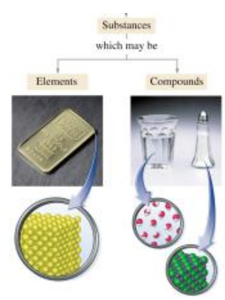
	Physical Properties
Property	**Example**
Qualitative	
Color	Sulfur is *yellow.*
Odor	Hydrogen sulfide *stinks.*
Solubility	Table salt *dissolves in water.*
Hardness	Diamond is *exceptionally hard.*
Electrical conductivity	Copper *conducts electricity.*
Quantitative	
Mass	A nickel has a mass of *5 grams.*
Temperature	Water for the bath is at *40 °C.*
Melting point	Lead melts at *327.5 °C.*
Density	At 20 °C, water has a density of *0.998 grams per milliliter.*
	Chemical Properties
Substance	**Typical Chemical Property**
Iron	*Rusts* (combines with oxygen to form iron oxide)
Carbon	*Undergoes combustion* (combines with oxygen to form carbon dioxide)
Silver	*Tarnishes* (combines with sulfur to form silver sulfide)
Sodium	*Reacts* violently *with water* to form hydrogen gas and a solution of sodium hydroxide.
Nitroglycerin	*Explodes* (decomposes, when detonated, to a mixture of gases)

Classifying Matter

Figure 1.3



Matter Classifications ...



- **Substance** – type of matter with fixed composition that does NOT vary from sample to sample
- **Element** – substance that cannot be broken down
- **Compound** – substance made up of atoms of two or more elements, with the different kinds of atoms combined in fixed proportions

Matter Classifications ...

- A **mixture** has no fixed composition; its composition may vary over a broad range
- A **solution** is a mixture that is **homogeneous**, which means that its composition and properties are the same throughout
- A **heterogeneous** mixture varies in composition and/or properties from one part of the mixture to another

Chemical Symbols

A one- or two-lettered designation derived from the name of the element

Most symbols are based on English names:

Hydrogen = H Neon = Ne Chromium = Cr

Note that the first letter is always capitalized and the second is lowercase

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Scientific Methods

- A **hypothesis** is a *tentative* explanation or prediction concerning some phenomenon
  - Tested via experiments
- A **theory** provides explanations of observed natural phenomena and predictions that can be tested by further experiments
- A **scientific law** is a summary of observed patterns in large collections of data, often expressed mathematically (model).

Models

Models are tangible items or representative pictures of atoms and molecules

Scientific Measurements and Units

All measurements contain two essential pieces of information:

a number (the quantitative piece)

a unit (the qualitative piece)

The number 60 is somewhat meaningless without units. Consider this for one's wages:

$ per week

$ per hour which is preferable?!

Scientific Measurements and Units

Scientists worldwide use common measurement units called the International System of Units (SI)

Physical Quantity	Name of Unit	Symbol of Unit
Length	Meter*	m
Mass	Kilogram	kg
Time	Second	s
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

Common SI Prefixes

Multiple	Prefix
10^{12}	*tera* (T)
10^{9}	*giga* (G)
10^{6}	*mega* (M)
10^{3}	*kilo* (k)
10^{2}	*hecto* (h)
10^{1}	*deca* (da)
10^{-1}	*deci* (d)
10^{-2}	*centi* (c)
10^{-3}	*milli* (m)
10^{-6}	*micro* (μ)*
10^{-9}	*nano* (n)
10^{-12}	*pico* (p)

Examples:

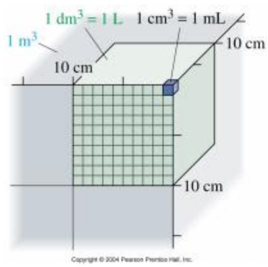
Gigahertz (GHz)
Megabytes (MB)
Terawatts (TW)

Length and Area

- The base unit of length is the **meter**, a little longer than a yard.
- Common derived units include:
  - kilometer (km; 1000 m), about 2/3 of a mile.
  - centimeter (cm; 0.01 m) and millimeter (mm; 0.001 m)
  - A contact lens is about 1 cm in diameter and 1 mm thick.
- The derived unit of area is the square meter (m^2) – an area one meter on a side.

Volume

- The derived unit of volume (space taken up by an object) is the cubic meter (m^3).
- A very common unit of volume, not SI but still used, is the **liter** (L).
- The milliliter (mL; 0.001 L) is also used, as is the cubic centimeter (cm^3).
- 1 mL = 1 cm^3.
- There are about five mL in one teaspoon.



Measuring Volume

Volume units typically use the Liters base unit

Very convenient for measuring the volume of irregularly shaped containers

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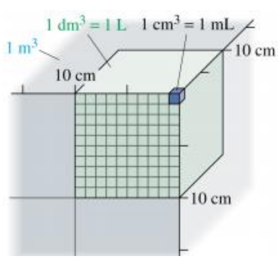
Measuring Volume, cont.



Regularly shaped objects can use a variant of the volume unit … cubic distance units

e.g., m^3, cm^3, etc.

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Example 1.1

Convert the unit of each of the following measurements to a unit that replaces the power of ten by a prefix.

(a) 9.56×10^{-3} m **(b)** 1.07×10^{3} g

Example 1.2

Use exponential notation to express each of the following measurements in terms of an SI base unit.

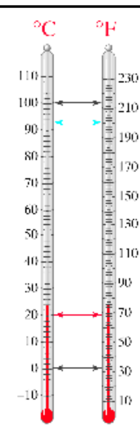
(a) 1.42 cm **(b)** 645 µs

Mass and Time

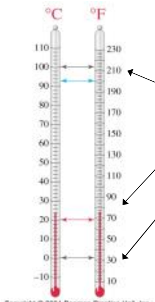
- **Mass** is the quantity of matter in an object; *weight* is a force.
- The base unit of mass is the **kilogram** (kg; 1000 g); it already has a prefix.
- A 1-L bottle of soft drink weighs about a kilogram.
- Commonly used mass units include the gram and the milligram (mg; 0.001 g).
- The SI base unit of time is the **second** (s).
- Smaller units of time include the millisecond (ms), microsecond (µs), and nanosecond (ns).
- Larger units of time usually are expressed in the nontraditional units of minutes, hours, days, and years.

Temperature

- Temperature is the property that tells us the direction that heat will flow.
- The base unit of temperature is the **kelvin** (**K**). **Note that a change of 1 ºC = 1 K**
- We often use the **Celsius** scale (℃) for scientific work.
  - On the Celsius scale, 0 ℃ is the freezing point of water, and 100 ℃ is the boiling point.
- The **Fahrenheit** scale (℉) is most commonly encountered in the U.S.
  - On the Fahrenheit scale, freezing and boiling water are 32 ℉ and 212 ℉, respectively.
- $T_F = 1.8T_{°C} + 32$
- $T_C = (T_F - 32)/1.8$



Comparing Temperature Scales



Some Celsius/Fahrenheit equivalents

b.p. of water 100/212

"Room-temperature" 20/68

f.p. of water 0/32

The SI unit for temperature is kelvins (K).

Note that a change of 1 ºC = 1 K

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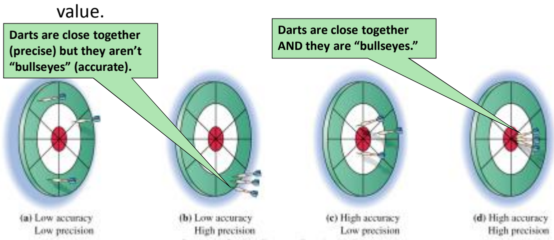
Example 1.3

At home you like to keep the thermostat at 72 °F. While traveling in Canada, you find the room thermostat calibrated in degrees Celsius. To what Celsius temperature would you need to set the thermostat to get the same temperature you enjoy at home?

$$T_{\circ C} = \frac{T_{\circ F} - 32}{1.8} = \frac{72 - 32}{1.8} = 22.2^{\circ}C$$

Precision and Accuracy in Measurements

- **Precision** – how closely repeated measurements approach one another.
- **Accuracy** – closeness of measurement to "true" (accepted) value.



Significant Figures

- All digits in a number that are known with certainty plus the first uncertain digit
- The more significant digits obtained, the better the precision of a measurement
- The concept of significant figures applies only to measurements
- *Exact values* have an unlimited number of significant figures

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Significant Figures in *Data*

- **Data** = measurements. (Results = calculations)
- All nonzero digits in data are significant.
- Zeroes may or may not be significant.
- To determine the number of significant figures in a measurement:
  - Begin counting with the first nonzero digit.
  - Stop at the end of the number.
- Problem: Zeroes in numbers without a decimal (100 mL, 5000 g) may or may not be significant.
  - To avoid ambiguity, such numbers are often written in scientific notation:
  - 1000 mL (?? sig fig) 1.00×10^3 mL (3 sig fig)

Rules for Zeros in Significant Figures

Zeros at the end of a number are significant if they are to the *right* of the decimal point

e.g., 0.1002300 1023.00

Zeros at the end of a number may or may not be significant if the number is written *without* a decimal point

e.g., 1000. compared to 1000

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Significant Figures in Calculations

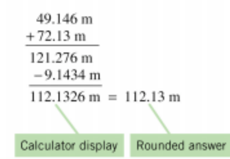
Multiplication and Division: the reported results should have no more significant figures than the factor with the fewest significant figures

1.827 m × 0.762 m = ?

0.762 has 3 sigfigs so the reported answer is **1.39 m^2**

Significant Figures in Calculations

Addition and Subtraction: the reported results should have the same number of decimal places as the number with the fewest decimal places



NOTE - Be cautious of round-off errors in multi-step problems. Wait until calculating the final answer before rounding.

Example 1.4

Calculate the area, in square meters, of the poster board whose dimensions are given in Table 1.5. Report the correct number of significant figures in your answer.

Example 1.5

For a laboratory experiment, a teacher wants to divide all of a 453.6-g sample of sulfur equally among the 21 members of her class. How many grams of sulfur should each student receive?

Example 1.6

Perform the following calculation, and round off the answer to the correct number of significant figures.

49.146 m + 72.13 m – 9.1434 m = ?

A Problem-Solving Method

Chemistry problems usually require calculations, and yield *quantitative* (numerical) answers

The unit-conversion method is useful for solving most chemistry problems – the focus here is on “unit equivalents”

For example,
1 inch = 2.54 cm

Other Equivalents and Conversion Factors

Metric	Common
Mass	
1 kg	= 2.205 lb
453.6 g	= 1 lb
28.35 g	= 1 ounce (oz)
Length	
1 m	= 39.37 in.
1 km	= 0.6214 mi
2.54 cm	= 1 in.*
Volume	
1 L	= 1.057 qt
3.785 L	= 1 gal
29.57 mL	= 1 fluid ounce (fl oz)

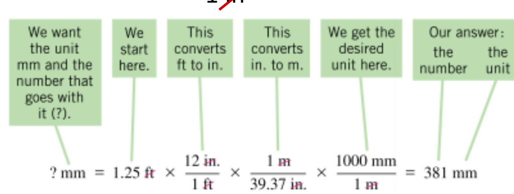
A conversion factor is the fractional expression of the equivalents

$$\frac{1\,inch}{2.54\,cm}\ or\ \frac{2.54\,cm}{1\,inch}$$

Two Examples

How many cm are in 26 inches?

$$26\ \text{in} \times \frac{2.54\ \text{cm}}{1\ \text{in}} = 66\ \text{cm}$$



Density(密度)

Density is the ratio of mass per unit volume of a substance

$$d = \frac{mass}{Volume} = \frac{m}{V}$$

common units are...

$$d = \frac{g}{cm^3} \text{ or } \frac{g}{mL} \text{ or } \frac{g}{L}$$

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