CHEM 1411 General Chemistry *Chemistry: An Atoms First Approach* by Zumdahl[2]

R Measurement and Calculations in Chemistry

Chapter Objectives:

- Learn to use and manipulate units, and convert from one unit to another.
- Learn to use the appropriate number of significant figures in measurements and calculations.
- Learn the basic properties of matter and energy.

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What Is Chemistry?

- **<u>Chemistry</u> is the science that seeks to understand the composition, properties, and transformations of matter by studying the behavior of atoms and molecules.**
- Chemistry is subdivided into different specialized fields: *organic* chemistry, *inorganic* chemistry, *physical* chemistry, *bio*chemistry, *analytical,* chemistry, *environmental* chemistry, etc.
- We study chemistry to provide ourselves with a better understanding of the underlying workings of nature, to learn how to make new materials with useful properties that satisfy particular needs. Chemistry intersects with other important fields, such as biology, molecular biology and genetics, medicine, physics, etc.

Units of Measure

... there is poetry in science, but also a lot of bookkeeping.
Peter Medawar

Units

- Observations in science may be:
  - **qualitative** — a description which does not involve a number (e.g., "this coin is heavy," "the sky is blue").
  - **quantitative** — a **measurement**, which contains both a number and a unit (e.g., "this coin weighs 2.35 grams," "the frequency of the light from the sky has a wavelength of 421 nm").
- For a measurement to be meaningful, *both the number and the unit must be present* (usually).

The SI System

- In science, the most commonly used set of units are those of the **International System of Units** (the **SI System**, for Système International d'Unités).
- There are seven fundamental units in the SI system. The units for all other quantities (e.g., area, volume, energy) are derived from these base units.

Physical Quantity	Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Amount of substance	mole	mol
Temperature	kelvin	K
Electric current	ampere	A
Luminous intensity	candela	cd

Larger and Smaller Units

- In many instances, decimal multipliers are used in combination with the base or derived units, in cases where numbers are inconveniently large or small:
  - the diameter of a sodium atom:
    - long-hand: 0.000 000 000 372 m
    - scientific notation: 3.72 10^{-10} m
    - prefix units: 0.372 nm or 372 pm
  - the distance from the earth to the sun:
    - long-hand: 150,000,000,000 m
    - scientific notation: 1.5 10^{11} m
    - prefix units: 150 Gm

SI Prefix Multipliers — Large Units

	Factor	Prefix	Symbol	Example	Scale
	10^{24}	yotta	Y		volume of earth ~ 1 YL
	10^{21}	zetta	Z		radius of Milky Way ~ 1 Zm
	10^{18}	exa	E		age of universe ~ 0.4 Es
	10^{15}	peta	P		1 light-year ~ 9.5 Pm
	10^{12}	tera	T		distance from sun to Jupiter ~ 0.8 Tm
→	10^{9}	giga	G	1 Gm = 10^9 m	1 light-second ~ 0.3 Gm
→	10^{6}	mega	M	1 Mm = 10^6 m	1 Ms ~ 11.6 days
→	10^{3}	kilo	k	1 kg = 1000 g	
memorize!	10^{2}	hecto	h	1 hm = 100 m	
	10^{1}	deka	da	1 dag = 10 g	
	10^{0}	—	—	—	—

SI Prefix Multipliers — Small Units

	Factor	Prefix	Symbol	Example	Scale
	10^{0}	—	—	—	—
→	10^{-1}	deci	d	1 dm = 0.1 m	
→	10^{-2}	centi	c	1 cm = 0.01 m	
→	10^{-3}	milli	m	1 mg = 0.001 g	
→	10^{-6}	micro	μ	1 μm = 10^{-6} m	1 μL ~ a **very** tiny drop of water
→	10^{-9}	nano	n	1 ns = 10^{-9} s	radius of Cl atom ~ 0.1 nm
→	10^{-12}	pico	p	1 pm = 10^{-12} m	mass of bacterial cell ~ 1 pg
memorize!	10^{-15}	femto	f		radius of a proton ~ 1 fm
	10^{-18}	atto	a		time for light to cross an atom ~ 1 as
	10^{-21}	zepto	z		1 zmol ~ 600 atoms
	10^{-24}	yocto	y		1.7 yg ~ mass of a proton or neutron

TABLE R.2 › The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	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}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.

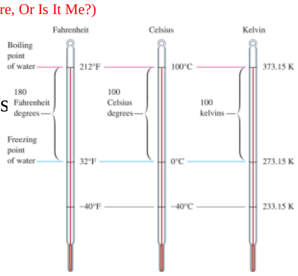
Mass, Length, Time, and Amount of Substance

- In the SI system, mass (the amount of matter in an object) is measured in **kilograms**, **kg** (1 kg = 2.205 U.S. lb). More commonly, the gram (g), milligram (mg,) and microgram (μg) are used in chemistry.
- In the SI system, length is measured in **meters**, **m** (1 m = 39.37 in) More commonly, the centimeter (cm, 2.54 cm = in), millimeter (mm), micrometer (μm), nanometer (nm) and picometer (pm) are used in chemistry.
- In the SI system, time is measured in **seconds**, **s**.
- The amount of substances is measured in **moles**, **mol**, which is the number of entities equal to the number of atoms in 12 grams of carbon-12 (Avogadro's number, 6.022×10^{23}) [more later].

Temperature

(Is It Hot In Here, Or Is It Me?)

- In the SI system, temperature is measured in **kelvins**, **K**, but often the **Celsius degree**, C, is used instead.
- A kelvin is the same size as a Celsius degree, but with the zero point set at the coldest possible temperature, *absolute zero* (-273.15 C).
- In most mathematical formulas, K must be used instead of C.



$$K = °C + 273.15$$
$$°F = \tfrac{9}{5}°C + 32$$
$$°C = \tfrac{5}{9}(°F - 32)$$

Figure R.5

Derived Units

- From the SI base units, we can derive other units, such as those for area, volume, density, force, etc.

Quantity	Definition	Units
Area	Length × width	m^2
Volume	Length × width × height	m^3
Density	Mass / volume	kg/m^3, g/cm^3, g/mL
Speed	Distance / time	$m\ s^{-1}$
Acceleration	Change in speed / time	$m\ s^{-2}$
Frequency	Event / time	s^{-1}
Force	Mass × acceleration	$kg\ m\ s^{-2}$ (newton, N)
Pressure	Force / area	$kg\ m^{-1}\ s^{-2}$ (pascal, Pa)
Energy	Force × distance	$kg\ m^2\ s^{-2}$ (joule, J)

Volume

- **Volume**, the amount of space occupied by an object, is measured in **cubic meters** (**m**^{3}); more commonly, the **liter** (**L**), **cubic centimeter** (**cm**^{3}), and **milliliter** (**mL**, 1 mL = 1 cm^3) are used.

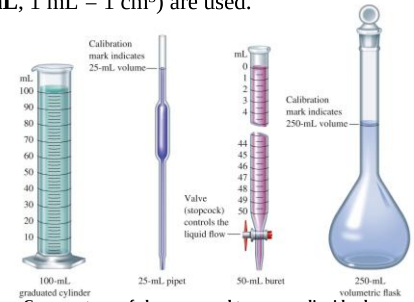


Figure R.2 **Common types of glassware used to measure liquid volume**

Density

- **Density**, the ratio of an object's mass (m) to its volume (V), is given by the formula:

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

  - Density has units of mass over volume: g/mL, g/L, lb/gal, kg/m^3, lb/ft^3, etc.
  - Because volume changes with temperature, density is *temperature-dependent*.

Density of water vs. temperature

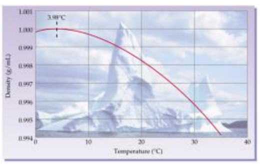


TABLE R.5 › Densities of Various Common Substances* at 20°C

Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

*At 1 atmosphere pressure.

Extensive and Intensive Properties

- **Extensive properties** *depend on the size of the sample* (mass, volume, length, etc.).
- **Intensive properties** are *independent of the size of the sample* (color, melting / boiling point, odor, etc.)
  - Despite the fact that the mass and volume of a sample are extensive properties, the density of a pure substance is an *intensive* property.

Examples: SI Prefixes

1. Complete the following conversion factors:

a. 1 km = _1000_ m

b. 1 cm = ____ m

c. 1.2×10^{-9} g = 1 ___g

d. 1 cm = _____ mm

e. 3.5 μg = 3.5×10__ g

f. 1,000,000 g = 1 ___g

g. 0.001 g = 1 ___g

Examples: Conversions with SI Prefixes

2. A bar of aluminum has a mass of 1210 g. What is its mass in kilograms (kg)?

$$1\,\text{kg} = ?\,\text{g}$$

$$1\,\text{kg} = 1000\text{g}$$

$$1210\text{g} \times \frac{1000\text{g}}{1\,\text{kg}} = \frac{\text{g}^2}{\text{kg}}$$ ✗

$$1210\text{g} \times \frac{1\,\text{kg}}{1000\text{g}} = \text{kg}$$ ✓

Examples: Metric Conversions

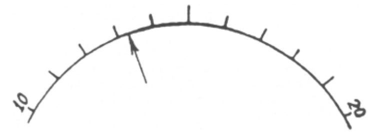
3. Convert 0.123 cm to mm.

Answer: 1.23 mm

Measurement, Significant Figures, and Scientific Notation

Uncertainty in Measurement

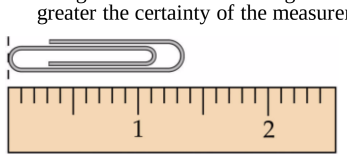
- For instance, on the scale below, the pointer is pointing between "13" and "14":



- We can "guess" that the pointer is about a third of the way to the 14, so we can estimate the reading as 13.3. Since the last digit is an estimation, anything further than that would be a wild guess.
- All measurements have some **degree of uncertainty** to them.

Significant Figures

- The total number of digits in a measurement is called the number of **significant figures**.
- When reading a scale, the value you record should use all of the digits you are sure of, plus *one additional digit that you estimate*. This last estimated digit is the last significant figure in your reading. (On a digital readout, the last number on the screen is usually the last significant figure.)
- The greater the number of significant figures, the greater the certainty of the measurement.



Accuracy and Precision

- Whenever possible, a measurement must be performed more than once in order to improve the confidence with which the result is reported.
  - The **precision** of a series of measurements is a measure of how close all of the reported numbers are to each other (reproducibility).
  - The **accuracy** of the measurements is a measure of how close they are to the actual value.

(a) inaccurate imprecise — **(b)** inaccurate precise — **(c)** accurate precise

Random and Systematic Error

- In the following example, a lead block with a mass of 10.00 g has been measured by three students:

Student 1	Student 2	Student 3
10.49 g	9.78 g	10.03
9.79 g	9.82 g	9.99
9.92 g	9.75 g	10.03
10.31 g	9.80 g	9.98
Avg: 10.13 g	**Avg: 9.79 g**	**Avg: 10.01 g**
inaccurate imprecise	*inaccurate precise*	*accurate precise*

- The results in (1) are due to ***random error***. Random errors often average out in repeated trials.
- In (2), there is a ***systematic error*** — all of the measurements are off in one direction. Systematic errors do not average out in repeated trials, because the same error is being made every time.

Measured Numbers vs. Exact Numbers

- **Exact numbers** are relationships that are arrived at by counting discrete objects (3 eggs = 3.00000… eggs) or that are true by definition (12 inches = 1 foot, 60 s = 1 min, 5280 feet = 1 mile, 100 cm = 1 m, 2.54 cm = 1 inch, etc.).
  - There is *no uncertainty* in these numbers, and they have an *infinite* number of significant figures (i.e., they do not affect the number of significant figures in the result of a calculation).
- All **measured numbers** will have some limit to how precisely they are known, and there is a limit to the number of significant digits in the number.
  - *This must be taken into account when doing calculations with those numbers!*

Counting Significant Figures

Rules for Counting Significant Figures:

a. All nonzero digits are significant. (42 has 2 sf's.)

b. Leading zeros are not significant; they are there to locate the decimal point. (0.00123 g has three sf's.)

c. Zeros in the middle of a number (*middle zeros* or *captive zeros*) are significant. (4.803 cm has 4 sf's.)

d. Trailing zeros are significant if the number contains a decimal point. (55.220 K has five sf's; 50.0 mg has three sf's, 5.100 10^{-3} has four sf's.)

e. Trailing zeros are not significant if the number does not contain a decimal. (34,200 m has three sf's.)

- Because trailing zeros can be ambiguous, it is a good practice to avoid potential errors by reporting the number in scientific notation.

Counting Significant Figures

a. All nonzero digits are significant.

164.87	5 sf's
395	3 sf's

b. Leading zeros are not significant.

0.766	3 sf's
0.000033	2 sf's
00591.3	4 sf's

c. Middle zeros are always significant.

2.028	4 sf's
5107	4 sf's
0.00304	3 sf's

Counting Significant Figures

d. Trailing zeros are significant if the number contains a decimal point.

14.30	4 sf's
0.0030	2 sf's
500.	3 sf's
500.0	4 sf's

e. Trailing zeros are ambiguous if the number contains no decimal point. We usually assume that they are not significant.

2500	2 sf's
60	1 sf's

Manipulating Significant Figures

- The results of calculations are only as reliable as the least precise measurement.

$$\text{Mileage} = \frac{\text{Miles}}{\text{Gallons}} = \frac{278\,\text{mi}}{11.70\,\text{gal}} = 23.760683760684\ldots\ \text{mi/gal}$$

This is far more significant figures than either of the measurements

Rules for Calculating Numbers involving sig. figs.:

- During multiplication or division, the result has the same number of sf's as the factor with the fewest sf's.
- During addition or subtraction, the result has the same number of decimal places as the quantity with the fewest decimal places.
- The final answer is then rounded off appropriately.

Manipulating Significant Figures

3 sf's → 278 mi
4 sf's → 11.70 gal

$$\frac{278\ \text{mi}}{11.70\ \text{gal}} = 23.8\ \text{mi/gal}$$ ← 3 sf's

3.18 ← 2 decimal places
+ 0.01315 ← 5 decimal places
3.19315
3.19 ← 2 decimal places

Adding and Subtracting Measured Numbers

- **During addition or subtraction**, the answer can't have more digits to the right of the decimal point than any of the original numbers.

```
  6.14          0.002631
+ 0.0375      - 0.0014278
  6.1775        0.0012032
  6.18          0.001203
```

$5\bar{2}00$ + 63.4 = 5263.4 → **5260**

12.6198 − 12.5202 = **0.0996**

Bold results: **6.18**, **0.001203**

Multiplying and Dividing Measured Numbers

- **During multiplication or division**, the answer can't have more sf's than any of the original numbers.

$$\frac{5.77}{1.9} = 3.036842105 26\ldots \text{rounds off to } 3.0$$

$$\frac{(28.71)(00626)(128 54)}{(5.0)} = 46.20360016 8\ldots \text{rounds off to } 46$$

Combined Operations

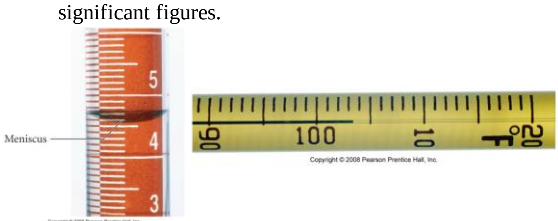
- **Combined Operations.** When operations involving both addition/subtraction and multiplication/division are performed, the order of operations is important when determining the number of significant figures in the final answer:

$$\frac{500.00}{275.0 - 225.00} = \frac{500.00}{50.0} = 10.0$$

$$10.00 \times 10.00 - 200.0/20.00 = 90.0$$

Examples: Significant Figures

1. (a) Report the volume on the graduated cylinder shown below (which is read at the bottom of the meniscus) to the correct number of significant figures. (b) Report the temperature on the thermometer shown below to the correct number of significant figures.



Answer: (a) 4.58 mL, (b) 103.4ºF

Examples: Significant Figures

2. How many significant figures are in each of the following?

a. 0.04450 m
b. 1000 m = 1 km
c. 0.00002 g
d. 5.0003 km
e. 1.000×10^{-3} mL
f. 10,000 m
g. 4080 kg
h. 1.4500 L
i. 5280 feet = 1 mile
j. 2.54 cm = 1 inch
k. 0.000304 s

Examples: Significant Figures

3. Perform the following calculations to the correct number of significant figures. (Example 1.6)

a. 1.10 × 0.5120 × 4.0015   3.4555 = **0.652**

b. 0.355 + 105.1 – 100.5820 = **4.9**

c. 4.562 × 3.99870   (452.6755 – 452.33) = **53**

d. (14.84 × 0.55) – 8.02 **0.1**

or 0.2 depending on when you round off

Examples: Significant Figures

4. Do the following calculations, rounding the answers off to the correct number of significant digits.

a. $12.43\,\text{miles}\left(\frac{5280\,\text{ft}}{1\,\text{mile}}\right)\left(\frac{12\,\text{in}}{1\,\text{ft}}\right)\left(\frac{2.54\,\text{cm}}{1\,\text{in}}\right) = ?$

b. $\frac{90.00\,\text{g}}{(.20\,\text{cm})(0.27\,\text{cm})(.05\,\text{cm})} = ?$

c. $8.233\text{g} + 35.5\text{g} + 75.8060\text{g} =$

d. $\frac{165.75\text{g}}{(2.20\,\text{mL} - 1.15\,\text{mL})} = ?$

Scientific Notation

- In science, we have the opportunity to work with numbers that are extremely large:

602,000,000,000,000,000,000,000

and numbers that are extremely small:

0.000000000000000000000000000911

- Numbers like this may be written more compactly using **scientific notation**:

$\mathbf{6.02\times10^{23}}$ $\mathbf{9.11\times10^{-28}}$

Scientific Notation

- To put a number into **scientific notation**, we move the decimal point behind the first significant figure, and multiply it by the appropriate power of ten.
- For a number larger than 1, move the decimal point to the left, behind the first nonzero digit, and use a *positive* power of ten to indicate how many places the decimal point was moved:

6.02	=	6.02×10^{0}
60.2	=	6.02×10^{1}
602	=	6.02×10^{2}
6020	=	6.02×10^{3}
6,020,000	=	6.02×10^{6}

Scientific Notation

- For a decimal number (smaller than 1), move the decimal point to the *right*, behind the first nonzero digit, and use a *negative* power of ten to indicate how many places the decimal point was moved:

0.602	=	6.02×10^{-1}
0.0602	=	6.02×10^{-2}
0.00602	=	6.02×10^{-3}
0.00000602	=	6.02×10^{-6}

Arithmetic with Scientific Notation

$$(6.71\times10^{2})(2.21\times10^{3}) = (6.71\times2.21)(10^{2+3})$$
$$= 14.8\times10^{5} = 1.48\times10^{6}$$

$$\frac{7.143\times10^{8}}{2.55\times10^{3}} = (7.143\div2.55)(10^{8-3}) = 2.80\times10^{5}$$

$$(1.6\times10^{5})^{3} = (1.6)^{3}\times(10^{5\times3}) = 4.1\times10^{15}$$

$$\sqrt{6.32\times10^{8}} = \sqrt{6.32}\times(10^{8/2}) = 2.51\times10^{4}$$

Examples: Scientific Notation

5. Write the following in scientific notation:

a. 360.2 ____________________

b. 0.000000208 ____________________

c. 5.61 ____________________

d. 60200000000 ____________________

e. 224.5×10^{-14} ____________________

f. 0.02245×10^{8} ____________________

Examples: Scientific Notation

6. Write the following in long hand notation:

a. 1.046×10^{-4} ____________________

b. 2.5×10^{5} ____________________

c. 1638×10^{-3} ____________________

d. 0.00224×10^{-4} ____________________

e. 0.0224×10^{7} ____________________

f. 500×10^{0} ____________________

Dimensional Analysis

Units

- Every number that you measure in the laboratory and most of the numbers that you calculate have not only a **numerical value**, but also a set of **units** associated with them. A number without units is (usually) meaningless; for a calculation to be regarded as correct, **the correct units must be included in the final answer**.
- Units can be manipulated just like numbers:

$$\begin{array}{r} 6\text{ cm} \\ \underline{-\ 2\text{ cm}} \\ 4\text{ cm} \end{array} \qquad 8.0\,\text{ft}\times 2.0\,\text{ft} = 16\,\text{ft}^2 \qquad \frac{24\,\text{mL}}{4.0\,\text{mL}} = 6.0\,(\text{no units})$$

$$\frac{24\,\text{g}}{6.0\,\text{mL}} = 4.0\,\text{g/mL} \qquad 5.0\,\text{ft}\times 3.0\,\text{lb} = 15\,\text{ft}\cdot\text{lb}$$

$$0.00165\,\text{s}^{-1} = 0.00165/\text{s} = 0.00165\,\text{Hz}$$

Manipulating Units

- **Dimensional analysis** (also known as the **factor-label method**) is a way of analyzing the setup of a problem by first manipulating the units in the same way you would manipulate the numbers in the calculations.
  - If your final units are correct, there is a good chance the problem has been set up correctly.
  - If you end up with incorrect units (for instance, units of time when you're measuring distance), or units which are clearly nonsense (for instance, cm•inches), the problem has been set up incorrectly.

Examples: Unit Conversions

1. **Convert 5.0 lb to kg (kilograms)**
   - The conversion factor is 2.2 lb = 1 kg.
   - Do you multiply 5.0 by 2.2, divide 5.0 by 2.2, or divide 2.2 by 5.0?

$$5.0\,\text{lb}\left(\frac{2.2\,\text{lb}}{1\,\text{kg}}\right) = 11\,\text{lb}^2/\text{kg}$$

$$\frac{2.2\,\text{lb}}{1\,\text{kg}}\left(\frac{1}{5.0\,\text{lb}}\right) = 0.44\,\text{kg}^{-1}$$

$$5.0\,\text{lb}\left(\frac{1\,\text{kg}}{2.2\,\text{lb}}\right) = 2.3\,\text{kg}$$

Examples: Unit Conversions

2. Convert 58 cm^3 to gallons.

1 cm^3 = 1 mL
1 L = 1000 mL
1 L = 1.057 qt
1 gal = 4 qt

Answer: 0.015 gal

Examples: Unit Conversions

3. What is the volume in liters of a sample of acetone having a mass of 925 g? (The density of acetone is 0.788 g/mL)

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

$$V = \frac{d}{m} \qquad V = \frac{\text{g/mL}}{\text{g}} = \text{mL}^{-1}$$

$$V = \frac{m}{d} \qquad V = \frac{\text{g}}{\text{g/mL}} = \text{mL}$$

Answer: 1.17 L

Examples: Unit Conversions

4. What is the mass in grams of a sample of acetone that has a volume of 1.180 L? The density of acetone is 0.788 g/mL.

Answer: 930. g

Examples: Unit Conversions

5. The radius of a copper atom is 0.1280 nanometers (nm). What is its radius in picometers (pm)?

Answer: 128.0 pm

Examples: Unit Conversions

6. Convert 105 ms to s.

Answer: 0.105 s

Examples: Unit Conversions

7. How many square centimeters (cm^2) are there in 2.00 square meters (m^2)?

$$1\,\text{m} = 100\,\text{cm}$$

$$2.00\,\text{m}^2 \times \frac{100\,\text{cm}}{1\,\text{m}} = 200\,\text{m} \bullet \text{cm} \quad ✗$$

$$2.00\,\text{m}^2 \times \left(\frac{100\,\text{cm}}{1\,\text{m}}\right)^2 = 20000\,\text{cm}^2 \quad ✓$$

Examples: Unit Conversions

8. Convert 75 mi/hr to ft/s.

In a case like this, since you're converting two sets of units, it's helpful to split it up into a numerator and denominator:

$$75\,\text{mi/hr} = \frac{75\,\text{mi}}{1\,\text{hr}}$$

Answer: 110 ft/s

Examples: Unit Conversions

9. To leave the surface of the Earth, an object must attain an *escape velocity* of 11,200 m/s. What is this speed in units of miles per hour?

Answer: 25,000 mi/hr (or 2.50×10^4 mi/hr)

Examples: Unit Conversions

10. Convert 37ºC to ºF and Kelvin.

Answer: 98.6ºC (*what's wrong with this number?*)

Examples: Unit Conversions

11. A small hole in the heat shield of the space shuttle requires a 32.70 cm^2 patch. If the patching material costs NASA \$2.75/in^2, what is the cost of the patch?

Answer: \$13.94

Examples: Unit Conversions

12. A runner wants to run 10.0 km. She knows that her running pace is 7.5 miles per hour. How many minutes must she run?

Answer: 50. min.

Examples: Unit Conversions

13. A car is advertised as having a gas mileage of 15 km/L. What is this in miles per gallon?

Answer: 35 mi/gal

Matter and Energy

1

The States of Matter

- **Matter** is anything that occupies space and has mass. Matter is classified by its *state* and by its *composition*:
- **Solids** have a fixed shape and volume that does not conform to the container shape.
  – The atoms or molecules vibrate, but don't move past each other, making solids *rigid* (more or less) and *incompressible*.
  – In *crystalline solids*, the atoms and molecules are arranged with some kind of long-range, repeating order, as in diamonds, ice, or salt
  – In *amorphous solids*, there is no long-range order, as in charcoal, glass, plastics.

2

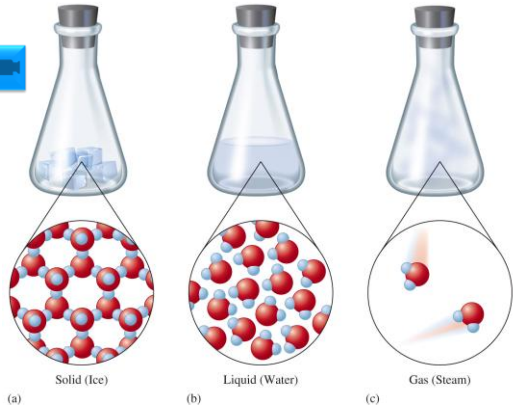
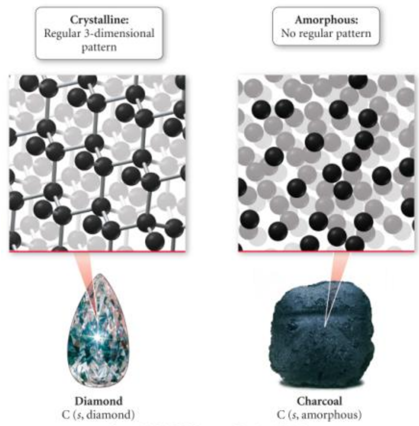


Figure R.7

3

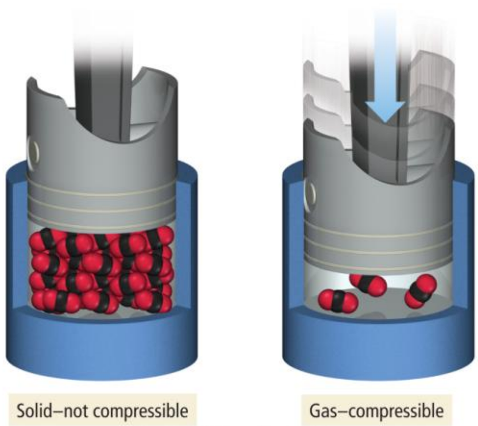


4

The States of Matter

- **Liquids** have fixed volumes that conform to the container shape (i.e., they form *surfaces*).
  – The particles are packed closely, as in solids, but are free to move past each other, making them *fluid* and *incompressible* (more or less).
  – e.g., liquid water
- **Gases** have no fixed shape or volume; they conform to the container shape, but fill the entire volume (i.e., there is no surface).
  – Gas particles are widely separated, making them both *fluid* and *compressible*.
  – e.g., steam

5



6

The Composition of Matter

- Matter can be classified as either **pure substances**, which have **fixed compositions**, or **mixtures**, which have **variable compositions**.
  - Pure substances (elements and compounds) are unique materials with their own chemical and physical properties, and are composed of only one type of atom or molecule.
  - Mixtures are simply random combinations of two or more different types of atoms of molecules, and *retain the properties of the individual substances*. They can therefore be separated (although sometimes with difficulty) by physical means (such as boiling, distillation, melting, crystallizing, magnetism, etc.).

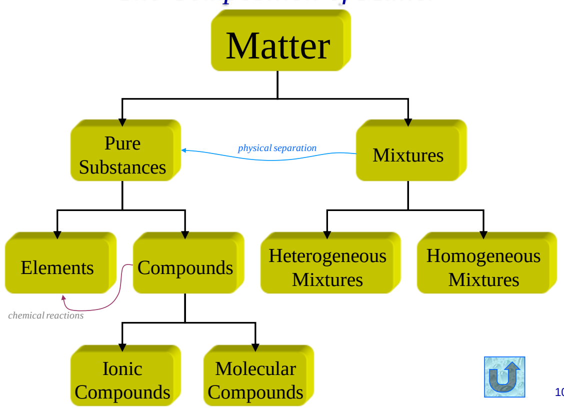
The Composition of Matter — Pure Substances

- **Pure Substances**:
  - An **element** is the simplest type of matter with unique physical and chemical properties; elements consist of only one kind of atom.
  - A **compound** is a pure substance that is composed of atoms or two or more different elements. There are two major types of chemical compounds: **ionic compounds** and **molecular compounds**.
  - Compounds cannot be broken down by physical means, but can be broken down (although sometimes with difficulty) by chemical reactions.

The Composition of Matter — Mixtures

- Matter usually occurs as **mixtures**. A **mixture** is a blend of two or more substances added together in some random proportion without chemically changing the individual substances themselves.
  - **Heterogeneous mixtures** are those in which the mixing is not uniform and which therefore have regions of different compositions — i.e., there are observable boundaries between the components (e.g., ice-water, salad dressing, milk, dust in air).
  - **Homogeneous mixtures** (or **solutions**) are those in which the mixing *is* uniform and which therefore have a constant composition throughout; there are no observable boundaries because the substances are intermingled on the molecular level (e.g., salt water, sugar water, air).

The Composition of Matter



The Composition of Matter

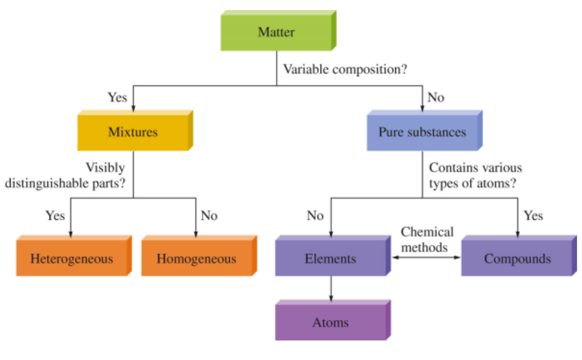
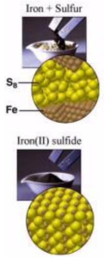
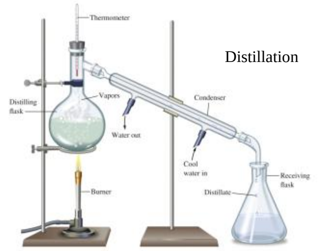


Figure R.10

Separation of Mixtures

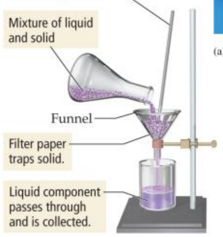
- Since the components of a mixture are different substances, with at least some physical properties that are unique to each compound, mixtures can be separated by physical means into their components by techniques such as filtration, distillation, chromatography, etc.



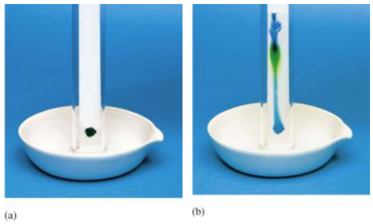


Distillation

Separation of Mixtures



Filtration



Chromatography

Physical and Chemical Changes

- A **physical change** occurs when a substance alters its physical form, but not its composition — the atoms and molecules in the sample retain their identities during a physical change (e.g., ice melting into liquid water, liquid water boiling to steam.)

$$H_2O(s) \rightarrow H_2O(l)$$

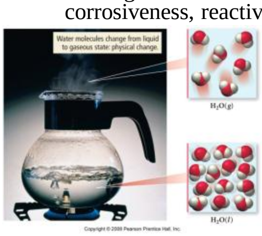
- A **chemical change** is one that alters the composition of matter — the atoms in the sample rearrange their connections in a ***chemical reaction***, transforming the substance into a different substance (e.g., the rusting of iron, the formation of water from hydrogen and oxygen, etc.)

$$4Fe(s) + 3O_2(g) \longrightarrow 2Fe_2O_3(s)$$

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$

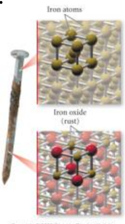
Physical and Chemical Properties

- **Physical properties** are properties that do not involve a change in a substance's chemical makeup (e.g., melting and boiling points, color, density, odor, solubility, etc.).
- **Chemical properties** are properties that do involve a change in chemical makeup (e.g., flammability, corrosiveness, reactivity with acids, etc.).



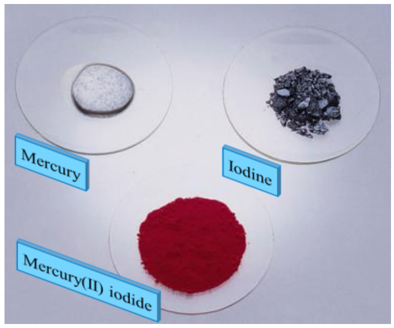
Boiling of water
a **physical** change

Rusting of iron
a **chemical** change



Chemical Changes

- In a chemical change, the products have a different composition, and therefore different properties, from the starting materials.



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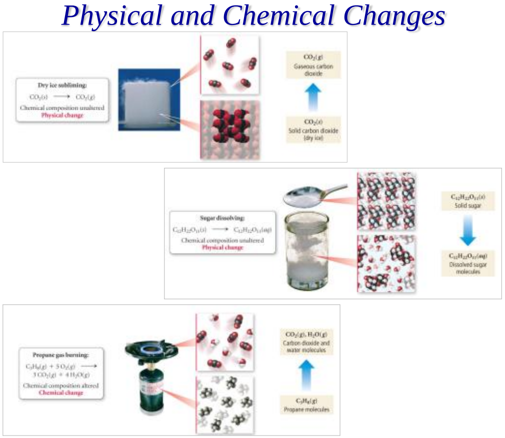
Physical and Chemical Changes



Physical and Chemical Properties of Copper



- **Copper — Physical Properties**
  - reddish brown, metallic luster
  - it is malleable (easily formed into thin sheets) and ductile (easily drawn into wires)
  - good conductor of heat and electricity
  - can be mixed with zinc to form brass or with tin to form bronze
  - density = 8.95 g/cm^3
  - melting point = 1083 C
  - boiling point = 2570 C
- **Copper — Chemical Properties**
  - slowly forms a green carbonate in moist air
  - reacts with nitric acid and sulfuric acid
  - forms a deep blue solution in aqueous ammonia



Examples: Physical and Chemical Changes

1. Which of the following processes are physical changes, and which are chemical changes? (sim. to Example 1.1)

   a. the evaporation of rubbing alcohol

   b. the burning of lamp oil

   c. the bleaching of hair with hydrogen peroxide

   d. the forming of frost on a cold night

   e. the beating of a copper wire into a sheet

   f. a nickel dissolving in acid to produce H_2 gas

   g. dry ice evaporating without melting

   h. the burning of a log in a fireplace

Energy

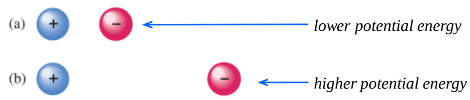
- **Energy** is defined as the ability to do work or produce heat.
  - **Work** is done when a force is exerted through a distance ($w = Fd$).
  - **Heat** is the energy that flows from one object to another because of a temperature difference.
- Energy may be converted from one form to another, but it is neither created nor destroyed (the *law of conservation of energy*).
- The total energy possessed by an object is the sum of its **kinetic energy** (energy of motion) and **potential energy** (stored energy due to position).
- Energy is measured in **Joules** (**J**) (kg m^2 s^{-2}) or **calories** (**cal**).

Potential and Kinetic Energy

- In general, systems tend to move from situations of high potential energy (less stable) to situations having lower potential energy (more stable).
- Gravitational potential energy decreases as the weight falls, while the kinetic energy of the weight increases as it falls faster and faster, but the sum of the kinetic and potential energies is always the same.
- When the weight hits the ground, the energy in converted into thermal energy, raising the temperature of the ground.

Energy

- **Opposite charges attract each other.** A positive and a negative charge have a *higher* potential energy when they are *far away* from each other, and a *lower* potential energy when they are *close together*.



- **Like charges repel each other.** A positive and a negative charge have a *lower* potential energy when they are *far away* from each other, and a *higher* potential energy when they are *close together*.

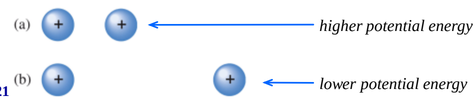
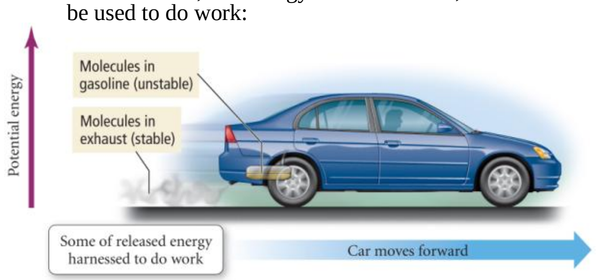


Figure 1.21

Chemical Energy

- The **chemical potential energy** of a substance results from the relative positions and the attractions and repulsions among all its particles. Under some circumstances, this energy can be released, and can be used to do work:



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