Chapter 1
What should I expect in lecture?

• Read book before you attend!

• Take notes
  – Evernote, pencil/paper, tablet PC, etc.

• Problem solving - *Participate!*
  – Bring calculator
  – Try problems (unsure, write “?”)
  – Contribute to discussion and working with peers
Chemical Classifications of Matter

Matter: anything that occupies space and has mass

1. Pure substances
2. Mixtures
Chemical Classifications of Matter

1. Pure substances

2. Mixtures
   - homogeneous mixtures
   - heterogeneous mixtures
Pure Substances

• uniform, or same, chemical composition throughout and from sample to sample
Pure substances

• two kinds of pure substances
  – elements

  – Compounds
Elements
Name the following elements (question 1.32; p.47)

(a) C

(b) Ca

(e) Cu

(f) Cl
Example 2

What are the symbols for the following elements (q. 1.37; p.47)

(a) iron

(b) lead

(c) silver

(d) gold
Pure Substances

• uniform, or same, chemical composition throughout and from sample to sample

• two kinds of pure substances
  – Elements
  – compounds
Compounds

**Compound**: a substance composed of two or more elements chemically combined in a definite proportion by mass.

- $\text{H}_2\text{O}$
- $\text{NaCl}$
- $\text{CO}_2$
Representations of Molecules

molecules composed of atoms

- balls represent atoms and colors distinguish different atoms

- An atom is the smallest unit of an element that retains the chemical properties of that element (more in ch. 2)
Representation of Oxygen Gas $(O_2)$

(Fig 1.11)
Molecular Representations

- **B**: H$_2$O
- **C**: H–O \( \text{H} \)
- **D**: balls & sticks
- **E**: space-filling

**formulas**

**line structure** (covalent bonding implied)

**balls & sticks**

**space-filling**
Chemical Classifications of Matter

1. Pure substances
   - elements
   - compounds

2. Mixtures
   - homogeneous mixtures
   - heterogeneous mixtures
States of Matter

Gas

Solid

Liquid

Solid (p 13 & 14)
States of Matter

We denote the states of matter symbolically:

Gases
- \( \text{CO}_2(g), \text{H}_2\text{O}(g) \)

Liquids
- \( \text{H}_2\text{O}(\ell), \text{Fe}(\ell) \)

Solids
- \( \text{CO}_2(s), \text{H}_2\text{O}(s) \)
Physical (Phase) Changes

Heat added

Solid → Melting → Liquid

Freezing

Liquid → Evaporation → Gas

Condensation

Heat removed
Phase changes are *physical changes* – they do not change the identity of the substance.
Physical Properties of Matter

*Physical* properties

– properties that can be observed without changing the composition of the substance

Four common physical properties are:

Mass and volume are also known as *extensive* properties (size dependent); density and temperature are known as *intensive* properties (size independent).
Mass

- _______________________________

- essentially the same physical quantity as weight; exception — weight is bound by gravity, mass is not

- Unit: grams (g)

  derived units: mg, kg, etc.

Metric base units and prefix factors will be dealt in *Math Toolbox 1.3,*
## Metric Prefixes

<table>
<thead>
<tr>
<th>Prefix*</th>
<th>Prefix Symbol</th>
<th>Meaning</th>
<th>Word</th>
<th>Multiple$^+$</th>
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<td><strong>Number</strong></td>
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<td>1,000,000,000</td>
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<td>G</td>
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<td>femto</td>
<td>f</td>
<td>0.000000000000001</td>
<td>quadrillionth</td>
<td>10^{-15}</td>
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</tbody>
</table>

*The prefixes most frequently used by chemists appear in bold type.*
Example 3

A slice of Swiss cheese contains 45 mg of sodium. What is this mass in units of grams? (q. 1.65; p.48)
Example 4

A package of Swiss cheese has a mass of 0.340 kg. What is this mass in units of grams? (q. 1.66; p. 48)
Volume

- amount of space occupied

Common units of volume

- $\text{cm}^3 \equiv \text{milliliter (mL)}$
- liters (L)
Example 5

If you drank 1.2 L of a sports drink, what volume did you consume in the following units? (q. 1.69; p.48)

(a) milliliters

(b) cubic centimeters

(c) cubic meters
Density

• ratio of the mass to the volume it occupies

• density = \( \frac{\text{mass}}{\text{volume}} \)

Units of density

• g/mL (solids and liquids)

• g/L (gases)

See Table 1.4 for a listing of densities for common substances
# Density

<table>
<thead>
<tr>
<th>Substance</th>
<th>Physical State</th>
<th>Density (g/mL)*</th>
</tr>
</thead>
<tbody>
<tr>
<td>Helium</td>
<td>Gas</td>
<td>0.000178</td>
</tr>
<tr>
<td>Oxygen</td>
<td>Gas</td>
<td>0.00143</td>
</tr>
<tr>
<td>Cooking oil</td>
<td>Liquid</td>
<td>0.92</td>
</tr>
<tr>
<td>Water</td>
<td>Liquid</td>
<td>1.00</td>
</tr>
<tr>
<td>Mercury</td>
<td>Liquid</td>
<td>13.6</td>
</tr>
<tr>
<td>Gold</td>
<td>Solid</td>
<td>19.3</td>
</tr>
<tr>
<td>Copper</td>
<td>Solid</td>
<td>8.92</td>
</tr>
<tr>
<td>Zinc</td>
<td>Solid</td>
<td>7.14</td>
</tr>
<tr>
<td>Ice</td>
<td>Solid</td>
<td>0.92</td>
</tr>
</tbody>
</table>

*At room temperature and at normal atmospheric pressure, except gases at 0 degrees Celsius (°C) and water at 4°C.
Density

maple syrup  d=1.32 g/mL
shampoo      d=1.01 g/mL
water        d=1.00 g/mL
antifreeze   d=1.13 g/mL
corn oil     d=0.93 g/mL
dish detergent d=1.03 g/mL
Example 6

If the density of a sugar solution is 1.30 g/mL, what volume of this solution has a mass of 50.0 g? (q. 1.75; p.48)
Temperature

- measure of how hot or cold something is relative to a standard
- thermometer
- varying scales – Fahrenheit, Celsius, Kelvin
Temperature

Water boils: 212°F = 100°C = 373.15 K
Room temperature: 77°F = 25°C = 298.15 K
Water freezes: 32°F = 0°C = 273.15 K

Lowest possible temperature:
-460°F = -273.15°C = 0 K
Temperature

In chemistry, we use both the Kelvin and Celsius scales for temperature.

\[ T_K = T_C + 273.15 \]
The boiling point of liquid nitrogen ($N_2$) is 77 K. What is its boiling point in units of degrees Celsius? In degrees Fahrenheit? (q. 1.82; p.49)
Example 8

If the temperature in a room increases from 20.0° C to 30.0° C, what is the temperature change in units of Kelvin? (q. 1.122; p.51)
Temperature

Temperature _________in Kelvin is numerically equivalent to temperature _________in Celsius!
This is because the temperature increments on the Kelvin scale are the same as those on the Celsius scale.
Physical Changes

changing the physical properties of a substance without changing the chemical properties

\[ \text{H}_2\text{O}(s) \rightarrow \text{H}_2\text{O}(g) \]
\[ \text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(g) \]

Sublimation
Condensation
Vaporization
Deposition

Melting
Freezing
Chemical Changes

- one or more substances are converted into a new substance (or substances)

\[ \text{Na}_2\text{CO}_3(s) + 2\ \text{HCl(aq)} \rightarrow 2\ \text{NaCl(aq)} + \text{CO}_2(g) + \text{H}_2\text{O(ℓ)} \]

sodium carbonate    hydrochloric acid    sodium chloride    carbon dioxide    water
Chemical Properties

chemical change:

\[ \text{Na}_2\text{CO}_3(s) + 2 \text{HCl}(aq) \rightarrow 2 \text{NaCl}(aq) + \text{CO}_2(g) + \text{H}_2\text{O}(\ell) \]

sodium carbonate hydrochloric acid sodium chloride carbon dioxide water

physical change:

\[ \text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(g) \]

Both of the above changes produce bubbles – so observation alone may not be enough to distinguish between the two.
Physical or Chemical Change?

• Physical
  – A change in appearance is the result of a new set of conditions imposed on the same material
  – _________________________________________

• Chemical
  – A change in appearance is the result of the formation of a new material that has its own unique set of physical properties
  – _________________________________________
Example 6

Identify each of the following as a physical change or a chemical change. (q. 1.90; p.49)

(a) condensation of ethanol

(c) dissolving sugar in water

(d) burning a piece of paper

(e) combining sodium metal with water, producing sodium hydroxide and hydrogen gas
Energy in Chemical Reactions

- chemical reactions can absorb or release energy
- heat, light, electricity, etc.
- thermodynamics
What is Energy?

Energy is the capacity to transfer **heat** or do **work**.

work (mechanical) – force acting over some distance

– some chemical reactions can be used to produce work – some can not
Types of Energy

• kinetic energy
  – energy of motion
  \[
  \text{k.e.} = \frac{1}{2} mv^2
  \]
  \[
  \text{k.e.} \propto T
  \]

• potential energy
  – stored energy

Kinetic energy is converted to potential energy as the ball ascends.

Potential energy is converted to kinetic energy as the ball descends.

Some of the kinetic energy of the ball is transferred to the ground.

Kinetic energy is transferred from the server’s hand to the volleyball.
Example 7

Which of the following samples of carbon monoxide (CO) is at a higher temperature?
Math Toolbox 1.1
Many numbers you will encounter in chemistry are either very large or very small. We need a simple way to deal with these numbers without having to write them out in the usual way.

For example:

a nanometer is 0.000000001 meters
or there are 1,000,000,000 nanometers in a meter

This is very cumbersome to write and to talk about. Let’s look at a more condensed way to express numbers like this.
Scientific Notation

In scientific notation, we will express numbers in the following way

\[ C \times 10^n \]

Where \( C \) is called the coefficient and \( n \) is the exponent. The coefficient is a single digit between 1 and 9.

For numbers smaller than 1, \( n \) will be negative.

For numbers larger than 1, \( n \) will be positive.

Let’s look at some examples.
Scientific Notation Example

#1: Express 845 in scientific notation

*Let’s think about this:* 
C \times 10^n where C is a number between 1 and 9, and n is positive or negative.

If the number given is >1, then n is positive.  
If the number given is <1 then n is negative!

#2 Express 10,300 in scientific notation.

#3 Express 0.0235 in scientific notation.
Convert each of the following values to scientific notation.

(a) 29,500

(b) 0.000082

(c) 650,000,000

(d) 0.01000
Need help?

If your answers are always off by a factor of 10, please ask us how to use your calculator!
Convert each of the following values from scientific notation to decimal form.

(a) $1.86 \times 10^{-5}$

(b) $1 \times 10^{7}$

(c) $4.53 \times 10^{5}$

(d) $6.1 \times 10^{-3}$
Rules of Exponents

1. When multiplying – add exponents
   \[ 10^{-4} \times 10^6 = 10^{(-4+6)} = 10^2 \]

2. When dividing – subtract exponents
   \[ 10^{-4} \div 10^6 = 10^{(-4-6)} = 10^{-10} \]

3. When raising to a power – multiply
   \[ (10^{-4})^2 = 10^{(-4\times2)} = 10^{-8} \]

4. When extracting a root – multiply
   \[ (10^{-4})^{1/2} = 10^{(-4\times1/2)} = 10^{-2} \]

Let’s look at some examples.
Scientific Notation Example 3 (q 1.5, p 46)

Perform each of the following operations without a calculator and report answers in scientific notation.

(a) \((2.0 \times 10^{-2}) \times (2.5 \times 10^{4})\)

(b) \((4.0 \times 10^{-2}) \div (2.0 \times 10^{3})\)

(c) \((3.0 \times 10^{4})^{2}\)

(d) \((4.3 \times 10^{2}) + (6.90 \times 10^{3})\)
Significant Figures (Math toolbox 1.2)

• Why do we worry about this stuff?

• Measurements define how accurate our numbers (the values) are!
  – Do NOT believe the 6 digits on your calculator.
  – If you measured your height in a doctor’s office you would not report it to the nearest 0.001 inches!
Significant Figures (p. 39)

- All digits that are not zero are significant.
- Sandwiched zeros are significant.
- Zeros between decimal point and first non-zero digit are not significant.
- Numbers >1, all zeros to right of decimal point are significant.
Significant Figures

• Numbers without decimal points, zero may or may not be significant. Use scientific notation to clarify.
Significant figures examples

• How many significant figures?
  A. 456 g
  B. 15,300 g
  C. 0.052 g
  D. 45.360 g
  E. 0.5 mL
Multiplication or division

• Your answer contains the smallest number of significant figures as determined by your original numbers.

D = 3.2167 g/ 1.2cm³ = 2.7 g/cm³
Addition or subtraction

• Answer contains smallest number of significant figures to the right of the decimal point, as determined by the original numbers.

• $82.326\text{ cm} + 4.1\text{ cm} = 86.4\text{ cm}$

To nearest 0.001

To nearest 0.1
Units and Conversions
Dimensional Analysis

• This is a way to set up calculations to get correct units.
• Making conversion factors:
  \(1 \text{ m} = 100 \text{ cm}\)
• From this equation we can get two conversion factors:

\[
\frac{100 \text{ cm}}{1 \text{ m}} \quad \text{or} \quad \frac{1 \text{ m}}{100 \text{ cm}}
\]
Conversions!

Convert 525 cm to m.

① How? Is there a relationship between m and cm?
② Yes, you know that there are 100 cm in 1 m.
③ Make a conversion factor!
Conversion examples!

A. Convert 1.25 g to mg.

B. Convert 544 mL to L.

C. Convert 1.2 g/mL to g/cm³