General Chemistry 1
Chem 1402

Section 20: 8:00-9:15 AM Daily
Fribourgh Hall 102

Summer 2011, University of Arkansas - Little Rock

Text: Chemistry & Chemical Reactivity
7th ed., by Kotz, Treichel & Townsend
Instructor:

Dr. Robert E. Belford

Office: SCLB 277

Office Hours: M-F: 8:20-8:40

office phone: 569-8824
email: rebelford@ualr.edu

Chemistry WEB Page: http://www.ualr.edu/rebelford
The science dealing with the composition and properties of substances, and with the reactions by which substances are produced from or converted into other substances.
How Old is Modern Chemistry?

1776

John Dalton’s Elements 1808
History of Chemistry

**Applied Chemistry** - The application of chemistry based upon practical experience, dates back to prehistoric times.

**Modern Chemistry** - based upon the Philosophy of Experimental Verification, (the “**Scientific Method**”)
Prehistoric Chemistry

No distinction between Science and Technology

Combustion Reactions - the Camp Fire
Prehistoric Chemistry

Paleolithic Art
Prehistoric Chemistry

Mesolithic Salt Production

Sodium Chloride (NaCl), halite crystal (right), Dead Seas Brine “mushrooms” (above left).
Egyptian Era of Chemistry

Chemia - Alexandrian writings refer to chemia as the Egyptian Art or Magic, acquiring the name from a black soil in Egypt, which was used as a dye; here-in lies the etymological root origins of the word chemistry.
Classical Period of Chemistry

700 BC - 600 AD - The period of Greek Philosophy and Roman Practice.

Roman Salt Pans in Ostia

Painting by Andrea Locatelli
Classical Period of Chemistry

- Was not a true Science, used the forum of debate and argumentation, instead of experiments to resolve issues.
Scientific Approach to Problem Solving

1. Recognize Problem (Observation)
2. Propose Solutions (Hypothesis)
3. Test Hypothesis (Experiment)
The Scientific Method

• *A process* of studying natural phenomena that involves making observations, forming laws and theories, and testing theories by experimentation
Classical Versus Modern Chemistry

The Scientific Method

- Observation and Experiments
- Patterns & Trends
- Formulate & Tests Hypothesis
- Theory
- Laws
Matter and Measurement

What are the Characteristics of Matter?

1. Matter has Mass
2. Matter Occupies Space
3. Matter has Energy

What Is the Composition of Matter?

1. Matter is Composed of Elements
2. Matter is Composed of Compounds
**Elements** - Can Not Be Broken Down by Chemical Means - Represented by the Periodic Table (N, H, O ...)

**Compounds** - Can be broken down by chemical means into constituent elements (H₂O, CO, CO₂ ...)

**Elements and Compounds**
Symbols of the Elements

Elements with Non English Symbols

Sb - Antimony (Stibium)  Ag - Silver (Argentium)
Cu - Copper (Cuprum)      Na - Sodium (Natrium)
Au - Gold (Aurum)         Sn - Tin (Stannum)
Fe - Iron (Ferrum)         W- Tungsten (Wolfram)
Pb - Lead (Plumbum)        K - Potassium (Kalium)
Hg - Mercury (Hydrargyrum)
Symbols of the Elements

Tricky Elements

Mg – Magnesium
Mn – Manganese
Ra – Radium
Rn – Radon
(noble gas)

There is no such thing as manganesium!
Properties of Matter

1. **Physical Properties** - describe the physical state of matter, odor, color, volume, state, density, melting point, boiling pt, etc.

2. **Chemical Properties** - describe the atomic arrangement, composition and reactivity of matter

What are the Differences Between Physical and Chemical Changes?
Physical Changes - changes in the state of matter (melting, boiling…) do not change the identity of a substance (water can be a liquid, vapor, or ice; it is still water)

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

Chemical Changes - changes in the identity of a substance, decomposition of water into Hydrogen and Oxygen

\[ 2 \text{H}_2\text{O (l)} \rightarrow 2\text{H}_2(g) + \text{O}_2(g) \]
States of Matter

• What are the Physical States of Matter?
States of Matter

• **Solid** - Definite Shape and Volume

• **Liquid** - Indefinite Shape Definite Volume (Incompressible Fluid)

• **Gaseous** - Indefinite Shape and Volume (Compressible Fluid)

Can you name a 4th State?

**Plasma**
What is the Difference Between Homogeneous and Heterogeneous Matter?

1. **Homogeneous** - a pure substance, appears uniform throughout (milk, wine, water…) May be a Mixture or Pure Substance

2. **Heterogeneous** - a mixture, has parts which are obviously different
MATTER

Homogeneous

Physical Methods (Filtration)

Pure Substance

Physical Methods

Solutions

Distillation

Chromatography
Separation of Mixtures

1. **Heterogeneous Mixtures** - Filtration
   Separates particles based on mesh size

2. **Homogeneous Mixtures**
   - **Chromatography**: uses different affinities of solutes to a substrate for separation
   - **Distillation**: uses different boiling points to separate substances
Separation of Pentane & Octane

- Pentane, being lighter, has a lower boiling pt, and boils while the octane is still a liquid. It is than condensed by the cold water.
MATTER

Heterogeneous

Physical Methods

(Filtration)

Homogeneous

Physical Methods

Pure Substance

Solutions

Physical Methods

Distillation

Chromatography

Compounds

Chemical Methods

Elements
Physical Properties

Properties which can be observed and measured without changing the chemical composition of matter
Physical Properties

- Color
- State of Matte
- Melting Point
- Boiling Point
- Density
- Mass
- Solubility
- Electric Conductivity
- Malleability
- Ductility
- Viscosity
- Volume
Two Types of Physical Properties

1. **Intensive Properties** - are the same for all samples of a substance, can be used to identify substance, (color, boiling point, density…)

2. **Extensive Properties** - depend on the amount of a sample, can not be used to identify a substance, (volume, mass, length, shape…)
Is Density an Intensive or Extensive Property?

\[ \text{Density} = \frac{\text{Mass}}{\text{Volume}} \]
## Selected Densities

<table>
<thead>
<tr>
<th>Substance</th>
<th>State @ 20°C, 1 atm</th>
<th>Density (g/mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>gas</td>
<td>0.000084*</td>
</tr>
<tr>
<td>Oxygen</td>
<td>gas</td>
<td>0.00133*</td>
</tr>
<tr>
<td>Ethanol</td>
<td>liquid</td>
<td>0.789</td>
</tr>
<tr>
<td>Water</td>
<td>liquid</td>
<td>0.9982</td>
</tr>
<tr>
<td>Aluminum</td>
<td>solid</td>
<td>2.70</td>
</tr>
<tr>
<td>Iron</td>
<td>solid</td>
<td>7.87</td>
</tr>
<tr>
<td>Lead</td>
<td>solid</td>
<td>11.34</td>
</tr>
<tr>
<td>Mercury</td>
<td>liquid</td>
<td>13.6</td>
</tr>
<tr>
<td>Gold</td>
<td>solid</td>
<td>19.32</td>
</tr>
</tbody>
</table>

* Usually use units of g/L for gases
Temperature

-a property which determines if heat will be transferred between objects

In terms of temperature, in what direction is heat transferred?
Temperature Measurement

3 SCALES

Fahrenheit Scale (F)
Celsius Scale (C)
Kelvin Scale (K)
Temperature Measurement

<table>
<thead>
<tr>
<th>Kelvin</th>
<th>Celsius</th>
<th>Fahrenheit</th>
</tr>
</thead>
<tbody>
<tr>
<td>373 K</td>
<td>100°C</td>
<td>212°F</td>
</tr>
<tr>
<td>Δ=100K</td>
<td>Δ=100°C</td>
<td>Δ=180°F</td>
</tr>
<tr>
<td>273 K</td>
<td>0°C</td>
<td>32°F</td>
</tr>
<tr>
<td>0 K</td>
<td>-273°C</td>
<td>-460°F</td>
</tr>
</tbody>
</table>
Temperature Conversions

Given: $\Delta 100^\circ C = \Delta 180^\circ F$

dividing by 180 gives:

$$\Delta 1^\circ F = \Delta \left(\frac{1}{1.8}\right)^\circ C$$

and dividing by 100 gives:

$$\Delta 1^\circ C = \Delta (1.8)^\circ F$$

Note: These are Changes in Temperature, Not the Temperatures!
Temperature Conversions

\[ y = mx + b \]

\[ ^0F = \frac{212 - 32}{100 - 0} \]
\[ = \frac{180}{100} \]
\[ = \frac{9}{5} \]
\[ = 1.8 \]

\[ ^0C = \frac{^0F - 32}{1.8} \]
Temperature Conversions

At what Temperature do these scales converge?

-40°C = -40°F
Temperature Conversions

+40/-40 Method

1. Add 40 to number

2. If going from C to F, multiply by 1.8
   (the change is greater)

   If going from F to C, divide by 1.8
   (the change is smaller)

3. Subtract 40 from number
Temperature Conversions

0 K Is Called Absolute Zero and Is Thermodynamically the Coldest Possible Temperature

1. What Is Absolute 0 in Degree Celsius?

2. Use the +40/-40 Technique to Determine Absolute 0 in Degree Fahrenheit?
# Units of Measurement

**7 SI Base Units** - *Systeme International d’Units*

<table>
<thead>
<tr>
<th>Category</th>
<th>Unit</th>
<th>Symbol</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass</td>
<td>Kilogram</td>
<td>kg</td>
</tr>
<tr>
<td>Length</td>
<td>Meter</td>
<td>m</td>
</tr>
<tr>
<td>Time</td>
<td>Second</td>
<td>s</td>
</tr>
<tr>
<td>Quantity</td>
<td>Mole</td>
<td>mol</td>
</tr>
<tr>
<td>Temperature</td>
<td>Kelvin</td>
<td>K</td>
</tr>
<tr>
<td>Electric Current</td>
<td>Ampere</td>
<td>A</td>
</tr>
<tr>
<td>Light Intensity</td>
<td>Candela</td>
<td>cd</td>
</tr>
</tbody>
</table>
Measurements of Mass

Kilogram

Mass is the Quantity of matter present in an object. Weight refers to the force gravity pulls on an object with. An object on the earth or the moon have the same mass, but different weights.

\[ 1\text{kg} = \text{mass of a standard Pt-Ir alloy bar kept in a French Vault} \]

\[ 1\text{lb} = 453.59\text{g} \]
Measurement of Time

Second

Based on Cesium Beam Atomic Clock

Related to the frequency of radiation coming from the cesium 133 isotope

Measurement of Temperature

Kelvin

The Fraction $1/273.16$ of the temperature of water at the triple point

The triple pt. is the temperature at which water, ice and steam can coexist in equilibrium
Measurement of Quantity

Mole

Number of particles equal to the number of carbon-12 atoms in 12 grams of carbon-12
<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Factor</th>
<th>Prefix</th>
<th>Symbol</th>
<th>Factor</th>
</tr>
</thead>
<tbody>
<tr>
<td>Yotta-</td>
<td>Y</td>
<td>$10^{24}$</td>
<td>Centi-</td>
<td>c</td>
<td>$10^{-2}$</td>
</tr>
<tr>
<td>Zetta-</td>
<td>Z</td>
<td>$10^{21}$</td>
<td>Milli-</td>
<td>m</td>
<td>$10^{-3}$</td>
</tr>
<tr>
<td>Exa-</td>
<td>E</td>
<td>$10^{18}$</td>
<td>Micro-</td>
<td>$\mu$</td>
<td>$10^{-6}$</td>
</tr>
<tr>
<td>Peta-</td>
<td>P</td>
<td>$10^{15}$</td>
<td>Nano-</td>
<td>n</td>
<td>$10^{-9}$</td>
</tr>
<tr>
<td>Tera-</td>
<td>T</td>
<td>$10^{12}$</td>
<td>Pico-</td>
<td>p</td>
<td>$10^{-12}$</td>
</tr>
<tr>
<td>Giga-</td>
<td>G</td>
<td>$10^{9}$</td>
<td>Femto-</td>
<td>f</td>
<td>$10^{-15}$</td>
</tr>
<tr>
<td>Mega-</td>
<td>M</td>
<td>$10^{6}$</td>
<td>Atto-</td>
<td>a</td>
<td>$10^{-18}$</td>
</tr>
<tr>
<td>Kilo-</td>
<td>K</td>
<td>$10^{3}$</td>
<td>Zepto-</td>
<td>z</td>
<td>$10^{-21}$</td>
</tr>
<tr>
<td>Deci-</td>
<td>d</td>
<td>$10^{-1}$</td>
<td>Yocto-</td>
<td>y</td>
<td>$10^{-24}$</td>
</tr>
</tbody>
</table>
## Measurements of Length

<table>
<thead>
<tr>
<th>Unit</th>
<th>Symbol</th>
<th>Conversion</th>
</tr>
</thead>
<tbody>
<tr>
<td>kilometer</td>
<td>km</td>
<td>$10^3 \text{ m}$</td>
</tr>
<tr>
<td>meter</td>
<td>m</td>
<td>$1 \text{ m}$</td>
</tr>
<tr>
<td>decimeter</td>
<td>dm</td>
<td>$10^{-1} \text{ m}$</td>
</tr>
<tr>
<td>centimeter</td>
<td>cm</td>
<td>$10^{-2} \text{ m}$</td>
</tr>
<tr>
<td>millimeter</td>
<td>mm</td>
<td>$10^{-3} \text{ m}$</td>
</tr>
<tr>
<td>micrometer</td>
<td>μm</td>
<td>$10^{-6} \text{ m}$</td>
</tr>
<tr>
<td>nanometer</td>
<td>nm</td>
<td>$10^{-9} \text{ m}$</td>
</tr>
<tr>
<td>Angstrum</td>
<td>Å</td>
<td>$10^{-10} \text{ m}$</td>
</tr>
</tbody>
</table>
Measurements of Mass

Mass is the Quantity of matter present in an object. Weight refers to the force gravity pulls on a mass with. An object on the earth or the moon would have the same mass, but different weights.

\[ 1 \text{kg} = 1000 \text{g} \]
\[ 1 \text{g} = 1000 \text{mg} \]
\[ 1 \text{lb} = 453.59 \text{g} \]
Derived SI Units

Units of Measurement Derived From the Fundamental SI Units

All measurable quantities can be measured in terms of the 7 SI units

*Force – Newton*

$1N = 1Kg\cdot m^2/sec^2$
Liter: Derived Unit of Volume

Volume is the space matter occupies, which can be described in terms of the 3 dimensions of the Cartesian coordinate system.

1ml = 1cm³ = 1cc

1L = 1dm³ = 1000cm³
Uncertainty in Measurement

**Exact Numbers** - Counted Quantities

**Inexact Numbers** - Measured Quantities

- Values Depend on Scale
- Report 1st Uncertain Value
- Guess the Value Between the Smallest Units of the Scale
- Different Measurements Will Give Different Values
Uncertainty in Measurement

**Accuracy** - How Close a Measured Value Is to the True Value.

**Precision** - How Close Successive Measured Values Are to Each Other

**Significant Figures** - First Uncertain and All Certain Digits of a Measured Number
How can we Represent the Accuracy of a Measurement?

% Error

\[
\% E = \left| \frac{\text{Measured Value} - \text{Theoretical Value}}{\text{Theoretical Value}} \right| (100)
\]

Where the Theoretical Value is the Accepted Value

- Note the text does not use absolute values
- Can you think of an advantage to using absolute values?

(The average \%E for random error does not go to zero)
How can we Represent the Precision of a Measurement?

Average Deviation:

$$Av. \text{ Dev} = \sum_{i=1}^{n} \frac{|M_i - M_{ave}|}{n}$$

$M_i$ = Measured Value of $i^{th}$ Measurement
$M_{ave}$ = Average Measured Value
$n$ = Number of Measurements
How can we Represent the Precision of a Measurement?

Standard Deviation (σ):

\[
\sigma = \sqrt{\frac{\sum_{i=1}^{n} (M_i - M_{\text{ave}})^2}{n}}
\]

Estimated Standard Deviation (s):

\[
s = \sqrt{\frac{\sum_{i=1}^{n} (M_i - M_{\text{ave}})^2}{n-1}}
\]

- Use s unless you have a very large number of measurements
How do we Express The Uncertainty of a Measured Number When We Write It?

**Significant Figures** - Report First Uncertain and All Certain Digits of a Measured Number
Uncertainty in Measurement

Read the following measurement to the correct number of significant figures.

2.84 or 2.85, maybe 2.83

Note: there is no deviation for the certain digits (2.8). The deviation for successive measurements comes from the uncertain digit. So deviation is a function of the scale.
Uncertainty in Measurement

Read the following measurement to the correct number of significant figures.

27.5 or 27.6 maybe 27.8

Note: Read from the bottom of the meniscus
Representing Significant Figures

1. Non Zeros are always significant

2. Leading Zeros are never significant.

3. Captive zeros are always significant

4. Trailing zeros are only significant if the number has a decimal point
Predict the number of sig figs for the following numbers

1. 0.0053  1. 2
2. 2300   2. 2
3. 32.00  3. 4
4. 34.483 4. 5
Scientific Notation

Scientific Notation - Convention of Expressing Any Base 10 Number As a Product of a Number Between One and 9, multiplied by 10 to the Power of Some Exponent

\[ 1 = 1 \times 10^0 \]
\[ 10 = 1 \times 10^1 \]
\[ 0.1 = \frac{1}{10} = 10^{-1} \]
\[ 100 = 1 \times 10^2 \]
\[ 2 = 2 \times 10^0 \]
\[ 20 = 2 \times 10^1 \]
\[ 0.2 = 2 \times 10^{-1} \]
\[ 200 = 2 \times 10^2 \]
Scientific Notation

Advantages of Scientific Notation:

• Allows Awkwardly Large and Small Numbers to Be Expressed in Terms of Compact and Easily Written Numbers

• Allows Accurate Representation of the Number of Significant Figures in a Number, That Is a Measurement’s Precision, the “Certainty” of Our Measurements
Scientific Notation

Note: Points will be deducted for improper use of Scientific Notation

If you report 3,450,000 as $345 \times 10^4$ it will be marked wrong even though they are equal. It must be written as $3.45 \times 10^6$
Sig Figs in Calculations

1. Addition and Subtraction
   - Result is limited to precision of least precise measurement (determined by the largest uncertain digit)

2. Multiplication and Division
   - Result is limited to the number of significant figures of the value with the least number of significant figures
Determine Sig Figs for the following Calculations

a) \[ 13.7325 - 14.21 = -0.4775 = -0.48 \]

b) \[ \frac{(1.1)(2.62)(13.5278)}{2.650} = 14.712121 = 15 \]
Determine Sig Figs for the following Calculations

c). $4.8 \times 10^2 + 9.9968 \times 10^5$

Convert to Common Power
(I like to bring them to the largest value)

$0.0048 \times 10^5$

$+9.9968 \times 10^5$

$\frac{10.0016}{10.0016} = 1.00016 \times 10^6$
Rounding off Numbers

Often your calculator will give answers with more numbers than are significant, how do we deal with this?

1. If digit to be removed is less than 5, preceding digit stays the same (Round Down).

2. If digit to be removed is greater than or equal to 5, preceding digit is increased by 1 (Round Up).

NOTE: during calculations, use all digits and round off at the end, according to preceding rules.
Dimensional Analysis

-The Incorporation of Units Into Algebraic Solutions

1 foot & 12 inches are identical lengths, therefore:

\[
\frac{1 \text{ ft}}{1 \text{ in}} = \frac{12 \text{ in}}{1 \text{ in}} = 1 = \frac{1 \text{ ft}}{1 \text{ ft}} = \frac{12 \text{ in}}{1 \text{ ft}}
\]

Equivalence Statement
Conversion Factors

\[
\frac{1\text{ft}}{12\text{in}} = \frac{12\text{in}}{12\text{in}} = 1 = \frac{1\text{ft}}{1\text{ft}} = \frac{12\text{in}}{1\text{ft}}
\]

To convert from:

- ft to in, multiply by 12in/ft
- in to ft, multiply by 1ft/12in
Tricks

1. Algebraically cancel units in calculations
2. Start calculations with given quantities
3. Visualize answer in desired quantities
Important

Always Include Units In Calculations

Check All Solutions for Proper Dimensions

-Answers Without Units Will Be Considered Wrong
Solve the Following Problem

Give the volume in liters of a box which is 2.4 yards by 2.4 inches by 2.4 feet in size

\[
2.4 \, \text{yd} \times (2.4 \, \text{ft}) \times (2.4 \, \text{in}) \left( \frac{3 \, \text{ft}}{\text{yd}} \right) \left( \frac{12 \, \text{in}}{\text{ft}} \right)^2 \left( \frac{2.54 \, \text{cm}}{\text{in}} \right)^3 \left( \frac{\text{mL}}{\text{cm}^3} \right) \left( \frac{1 \, \text{L}}{1000 \, \text{mL}} \right) = 98 \, \text{L}
\]
Solve the Following Problem

What is the value of a gold bar with dimensions of 1.5cm x 2.5cm x 2.0cm if gold sells for $300/oz and has a density of 19.32g/ml?

\[
1.5cm(2.5cm)(2.0cm) \left( \frac{1mL}{cm^3} \right) \left( \frac{19.32g}{mL} \right) \left( \frac{16oz}{453.59g} \right) \left( \frac{$300}{oz} \right) = $1500
\]