**General Chemistry 1**  
Chem 1402  
Section 210: 8:00-9:15 AM Daily  
Fribourgh Hall 102  
Spring 2006, University of Arkansas - Little Rock  

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

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**Scientific Approach to Problem Solving**  
1. Recognize Problem (Observation)  
2. Propose Solutions (Hypothesis)  
3. Test Hypothesis (Experiment)

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**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

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**Symbols of the Elements**

**Elements with Non English Symbols**

<table>
<thead>
<tr>
<th>Element</th>
<th>English Name</th>
<th>Latin Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sb</td>
<td>Antimony</td>
<td>Stibium</td>
</tr>
<tr>
<td>Ag</td>
<td>Silver</td>
<td>Argentium</td>
</tr>
<tr>
<td>Cu</td>
<td>Copper</td>
<td>Cuprum</td>
</tr>
<tr>
<td>Au</td>
<td>Gold</td>
<td>Aurum</td>
</tr>
<tr>
<td>Fe</td>
<td>Iron</td>
<td>Ferrum</td>
</tr>
<tr>
<td>Pb</td>
<td>Lead</td>
<td>Plumbum</td>
</tr>
<tr>
<td>K</td>
<td>Potassium</td>
<td>Kalium</td>
</tr>
<tr>
<td>Hg</td>
<td>Mercury</td>
<td>Hydrargyrum</td>
</tr>
</tbody>
</table>
Symbols of the Elements

Tricky Elements
Mg – Magnesium  Ra – Radium
Mn – Manganese  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

- Solid - Definite Shape and Volume
- Liquid - Indefinite Shape Definite Volume (Incompressible Fluid)
- Gas – Indefinite Shape and Volume (Compressible Fluid)

Can you name a 4th State?

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

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 pts to separate substances
MATTER

Heterogeneous

Physical Methods (Filtration)

Homogeneous

Physical Methods

Distillation Chromatography

Pure Substance

Solutions

Compounds

Chemical Methods

Elements

Physical Properties

• Color
• Solubility
• State of Matte
• Electric Conductivity
• Melting Point
• Malleability
• Boiling Point
• Ductility
• Density
• Viscosity
• Mass
• 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…)

Selected Densities

<table>
<thead>
<tr>
<th>Substance</th>
<th>State @ 20°C, 1atm</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 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: Δ100°C = Δ180°F

dividing by 180 gives:

\[ \Delta 1^\circ F = \Delta (1/1.8) ^\circ C \]

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

\[ T(\degree C) = \frac{9}{5} T(\degree F) - 32 \]

\[ T(\degree F) = \frac{5}{9} (T(\degree C) + 32) \]

+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

<table>
<thead>
<tr>
<th>7 SI Base Units - Systeme International d’Units</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass</td>
</tr>
<tr>
<td>Kilogram</td>
</tr>
<tr>
<td>Length</td>
</tr>
<tr>
<td>Meter</td>
</tr>
<tr>
<td>Time</td>
</tr>
<tr>
<td>Second</td>
</tr>
<tr>
<td>Quantity</td>
</tr>
<tr>
<td>Mole</td>
</tr>
<tr>
<td>Temperature</td>
</tr>
<tr>
<td>Kelvin</td>
</tr>
<tr>
<td>Electric Current</td>
</tr>
<tr>
<td>Ampere</td>
</tr>
<tr>
<td>Light Intensity</td>
</tr>
<tr>
<td>Candela</td>
</tr>
</tbody>
</table>

Selected SI Prefixes

- Yotta- \( 10^{24} \)
- Zetta- \( 10^{21} \)
- Exa- \( 10^{18} \)
- Peta- \( 10^{15} \)
- Tera- \( 10^{12} \)
- Giga- \( 10^9 \)
- Mega- \( 10^6 \)
- Kilo- \( 10^3 \)
- Deci- \( 10^{-1} \)

- Yotta- \( 10^{-2} \)
- Zetta- \( 10^{-3} \)
- Exa- \( 10^{-6} \)
- Peta- \( 10^{-12} \)
- Tera- \( 10^{-15} \)
- Giga- \( 10^{-18} \)
- Mega- \( 10^{-21} \)
- Kilo- \( 10^{-24} \)

Measurements of Length

- kilometer \( 10^3 \) m
- meter \( 1 \) m
- decimeter \( 10^{-1} \) m
- centimeter \( 10^{-2} \) m
- millimeter \( 10^{-3} \) m
- micrometer \( 10^{-6} \) m
- nanometer \( 10^{-9} \) m
- Angstrom \( 10^{-10} \) m
**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*

\[ 1N = 1Kg \cdot m^2/sec^2 \]

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**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.

\[ 1ml = 1cm^3 = 1cc \]

\[ 1L = 1dm^3 = 1000cm^3 \]

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**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

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**How can we represent the accuracy of a Measurement?**

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

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)

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**How can we represent the precision of a Measurement?**

Average Deviation:

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

- \( M_i \) = Measured value of the \( i^{th} \) measurement
- \( M_{\text{ave}} \) = Average measured value
- \( n \) = Number of measurements

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

Standard Deviation (σ):

\[ \sigma = \sqrt{\frac{\sum (M_i - \bar{M})^2}{n}} \]

Estimated Standard Deviation (s):

\[ s = \sqrt{\frac{\sum (M_i - \bar{M})^2}{n-1}} \]

- Use s unless you have a very large number of measurements

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

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
2. 2300
3. 32.00
4. 34.483

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

- \( 1 = 1 \times 10^0 \)
- \( 2 = 2 \times 10^0 \)
- \( 10 = 1 \times 10^1 \)
- \( 20 = 2 \times 10^1 \)
- \( 0.1 = 1/10 = 10^{-1} \)
- \( 0.2 = 2 \times 10^{-1} \)
- \( 100 = 1 \times 10^2 \)
- \( 200 = 2 \times 10^2 \)
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 $345\times10^4$ it will be marked wrong even though they are equal. It must be written as $3.45\times10^6$

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) $13.7325 - 14.21$

b) $(1.1)(2.62)(13.5278) \approx 2.650$

determine sig figs for the following calculations

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, therefor:

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

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

Conversion Factors

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

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?