1.1 Use Atomic Theory of Matter to discuss chemical and physical changes. [Readings 1.1-1.3 Problems 5,
Atomic Theory

- All chemical substances are composed of atoms
- Atoms combine to form molecules
- All molecules of a substance are composed of the same atoms in the same proportions

Atoms and Molecules

Atom - smallest representative sample of an element
Molecule - neutral particle composed of two or more atoms combined in definite ratio of whole numbers

Representations of Molecules

Structural Model - shows arrangement of atoms
Ball and Stick Model - shows geometric arrangement
Space-filling Model - shows relative sizes of atoms
Physical vs. Chemical Changes

- Physical States: Solid, liquid, or gas
- Physical change
- Chemical reaction
- Results in the formation of an entirely different chemical substance

Physical Change

Chemical Change
Energy

- Chemical reaction either release or absorb energy.
  - Light
  - Heat
- Energy is the ability to do work
- Kinetic Energy
- Potential energy
- Chemical energy is a form of Potential Energy

1.2. Measurement and expressing measured and calculated quantities in Metric Units.

[Readings 1.4 Review Questions 24, 25, 26, 43, 44, & 45]

Fundamental Quantitative Units

<table>
<thead>
<tr>
<th>Quantity</th>
<th>Unit</th>
<th>Abbr.</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>Temperature</td>
<td>Kelvin</td>
<td>K</td>
</tr>
<tr>
<td>Electrical Current</td>
<td>ampere</td>
<td>A</td>
</tr>
<tr>
<td>Amount</td>
<td>mole</td>
<td>mol</td>
</tr>
<tr>
<td>Luminous Intensity</td>
<td>candela</td>
<td>cd</td>
</tr>
</tbody>
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<td>candela</td>
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</table>

Measurement - Mass

Standard kilogram

Mass of object on left pan balanced with known masses added to right

Measurement - comparison with a known
Mass - measure of amount of matter in a sample

English and Metric Units

<table>
<thead>
<tr>
<th>Quantity</th>
<th>English</th>
<th>Metric</th>
<th>Conversion</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass</td>
<td>pound</td>
<td>kilogram</td>
<td>1 lb = 453.6 g</td>
</tr>
<tr>
<td>Length</td>
<td>inch</td>
<td>centimeter</td>
<td>1 in = 2.54 cm</td>
</tr>
<tr>
<td>Volume</td>
<td>quart</td>
<td>liter</td>
<td>1 qt = 946.4 mL</td>
</tr>
</tbody>
</table>
Metric Units

Mass - kilogram
1 kg = 1000 g

Volume - liter
1 L = 1000 mL
1 L = 1 dm³
1 mL = 1 cm³

Measurement - Volume

Graduated Cylinder
Shown is 100 mL size
Scale shows mL levels

Buret
Shown in 50 mL size
Scale - 0.1 mL levels

Pipet
Used to dispense fixed volume when filled to etched line (arrow)

Volumetric Flask
Scale shown is 250 mL
Contains fixed volume when filled to line

Measurement - Temperature

<table>
<thead>
<tr>
<th>Fixed points</th>
<th>Freezing Pt. H₂O</th>
<th>Boiling Pt. H₂O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Celsius</td>
<td>0 °C</td>
<td>100 °C</td>
</tr>
<tr>
<td>Fahrenheit</td>
<td>32 °F</td>
<td>212 °F</td>
</tr>
<tr>
<td>Kelvin</td>
<td>273.16 K</td>
<td>373.16 K</td>
</tr>
</tbody>
</table>

Size of degree 1 K = 1 °C = 1.8 °F
Commonly Used Prefixes

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Meaning</th>
<th>Exp Not</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>mega</td>
<td>M</td>
<td>1,000,000</td>
<td>$10^6$</td>
<td>Mm</td>
</tr>
<tr>
<td>kilo-</td>
<td>k</td>
<td>1000</td>
<td>$10^3$</td>
<td>km</td>
</tr>
<tr>
<td>centi-</td>
<td>c</td>
<td>0.01</td>
<td>$10^{-2}$</td>
<td>cm</td>
</tr>
<tr>
<td>milli-</td>
<td>m</td>
<td>0.001</td>
<td>$10^{-3}$</td>
<td>mm</td>
</tr>
<tr>
<td>micro-</td>
<td>µ</td>
<td>0.000 001</td>
<td>$10^{-6}$</td>
<td>mm</td>
</tr>
<tr>
<td>nano-</td>
<td>n</td>
<td>0.000 000 001</td>
<td>$10^{-9}$</td>
<td>nm</td>
</tr>
</tbody>
</table>

Using Metric Prefixes

• Replace prefix with the exponential form
  • m = $10^{-3}$
  • 2.53 mm = 2.53 mm = 2.53 x $10^{-3}$ m
• Practice with the concept
  – 1.5 µm = ? m
  – 6.06 MHz = ? Hz
• 2.5 x $10^6$ scopes = ?

Using Metric Prefixes

• Replace prefix with the exponential form
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• Practice with the concept
  – 1.5 µm = ? m
  – 6.06 MHz = ? Hz
• 2.5 x $10^6$ scopes = a bad chemistry joke!
1.2a. Express measured and calculated quantities in exponential form.  
[Readings App. A.1]

Exponential Notation (Appendix A.1)

Express very large or very small numbers as a number between 1 and 10 times a power of 10.
Example: the number of atoms of carbon in a 12g sample is $600,000,000,000,000,000,000,000$
In exponential notation: $6 \times 10^{23}$
The weight of one carbon atom is $\approx 0.0000000000000000000002g$
In exponential notation: $2 \times 10^{-23}g$

Exponentials

- $10^x$ \(\Rightarrow\) move the decimal point to the right
- $10^y$ \(\Rightarrow\) move the decimal point to the left
- $2.05 \times 10^2 = ??$
- $2.05 \times 10^{-2} = 0.0205$
1. Exponentials

- $10^x$: move the decimal point to the right
- $10^{-y}$: move the decimal point to the left
- $2.05 \times 10^2 = 205$

1.3. Express calculated quantities in the proper number of significant figures.
[Readings 1.5 Review Quest. 53, 54, 55, & 56]

Accuracy and Precision

Accuracy
The closeness of a measurement to the true value

Precision
How reproducible measurements are; shown by the number of significant figures reported in a physical quantity
**Measurement - Mass**

- Balance showing mass to milligram value 20.988 g
- Balance showing mass to tenth of milligram 0.0000g

**Precision and Significant Figures**

- Significant Figures - number of digits of a measurement with one uncertain digit
- Upper Scale 11.3 m
- Lower Scale 11.3m
- Three significant figures
- Four significant figures
- (Uncertain digits are shown in yellow)

**Systematic Error**

- Systematic Error - constant error in one direction from true value. Caused here by mistake in labeling scale or ruler.
1.4. Express calculated quantities in the proper number of significant figures.
[Readings 1.5 Review Quest. 53, 54, 55, & 56]

Calculated Values & Sig Figs
When #s are combined, uncertainty is combined.
Example: A 25 cm$^3$ sample of Al has a density of 2.71g/mL. What is the sample’s weight?

$$m = D \cdot V$$

$$2.71\, ? \times 25.\, ?$$

$$????$$

$$1355\, ?$$

$$542\, ?$$

$$67???? \, 67.????g$$

Rules for Sig Figs
Addition and Subtraction:
Result has number of decimal places as least precise measurement.
Example:

$$150.301\, ?$$
$$-\, 106.81\, ?$$

$$43.49?? \, 43.49$$
Rules for Significant Figures

• 1. All digits 1 through 9 are significant
  – 1562 has four significant digits
• 2. Leading zeros are not significant
  – 0.00015 has only two sig figs
• 3. Confined zeros are significant
  – 3.015 has four sig figs
• 4. Trailing zeros are significant
  – 35.000 has five sig figs

Rules for Sig Figs

Multiplication & Division:
Result has number of sig figs as least precise measurement.
Example:
\[ 15.63 \div 5.5 = 2.84181818182 \]
4 sig figs 2 sig figs
\[ \therefore 2 \text{ sig figs} = 2.8 \]

Rounding Off

In a series of calculations (m&d), round at the final step.
To round, consider only the first unknown #.
If number to remove is less than five, drop.
If # to remove is greater than five, round up.
Examples:
round to 2 sig figs: 2.54, 2.548, 2.55
round to correct # of sig figs:
\[ 2.654 \div 5.5 = ? \]
\[ (35.662 - 34.82) \div 4.668 = ? \]
1.5. Use the Factor Label Method to convert quantities among different unit systems. [Readings 1.6 & 1.7 Problems 63, 64, 65, 67, 69, 74, & 75]

The Factor Label Method

- Also called Dimensional Analysis
- Use in Unit Conversion
- Goal Quantity = (given) x (conversion factor)
- Conversion Factors (Factors equal to 1)

2.54 cm = 1 in
The Factor Label Method

• The Factor Label Method
• Use in Unit Conversion
• Goal Quantity = (given) x (conversion factor)
• Conversion Factors (Factors equal to 1)

\[
\frac{2.54 \text{ cm}}{1 \text{ in}} = \frac{1 \text{ in}}{1 \text{ in}}
\]
The Factor Label Method

- Single-step conversions
- Convert 5.6 lbs to kilograms

\[ \text{kg} = \text{lbs} \times \frac{1 \text{kg}}{1 \text{lbs}} = \]

\[ \text{kg} = \text{lbs} \times \frac{1 \text{kg}}{1 \text{lbs}} = \]
The Factor Label Method

- Single-step conversions
- Convert 5.6 lbs to kilograms

\[ \text{kg} = \frac{5.6 \text{lbs}}{2.2 \text{lbs}} \times 1 \text{kg} = 2.5 \text{kg} \]

The Factor Label Method

- Single-step conversions
- Convert 5.6 lbs to kilograms

\[ \text{kg} = \frac{5.6 \text{lbs}}{2.2 \text{lbs}} \times 1 \text{kg} = 2.5 \text{kg} \]

The Factor Label Method

- Multi-step conversions
- Convert 18.3 ft to meters
The Factor Label Method

- Multi-step conversions
- Convert 18.3 ft to meters

\[ m = \frac{ft}{in} \times \frac{in}{ft} \times \frac{cm}{in} \times \frac{m}{cm} \]

\[ m = 18.3 \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{1 \text{ m}}{100 \text{ cm}} \]
The Factor Label Method

- Multi-step conversions
- Convert 18.3 ft to meters

\[
m = 18.3 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{1 \text{ m}}{100 \text{ cm}} = 5.58 \text{ m}
\]

Length

How many feet are in 18,221 mm?

1 in = 2.54 cm

Mass (Weight)

How many oz. are in a 1.00 kg mass?

1 lbs = 454 g
16 oz = 1.0 lb
1.6. Calculate and use derived quantities from measurements (e.g. density and specific gravity).
[Readings 1.7, Problems 77, 79, 81, 83, & 89]

Properties of Matter

• Physical Properties
  • changes of physical properties do not alter the chemical makeup of a substance
  • e.g. boiling water does not change the chemical makeup of water
  • A physical change is a change which does not alter the chemical properties of the substance

Properties of Matter

• Chemical Properties
  • A chemical change results in the complete change in the chemical makeup of a substance
  • A chemical change is irreversible by physical means
Density

- An intensive property
- Every substance has its own characteristic density
- Density relates a substance’s mass to its volume

What is the density of Al?

Experiment: Six pieces of Al

<table>
<thead>
<tr>
<th></th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
</tr>
</thead>
<tbody>
<tr>
<td>weight</td>
<td>5.0g</td>
<td>10g</td>
<td>15g</td>
<td>20g</td>
<td>25g</td>
<td>30g</td>
</tr>
<tr>
<td>volume</td>
<td>1.9mL</td>
<td>3.7mL</td>
<td>5.6mL</td>
<td>7.4mL</td>
<td>9.2mL</td>
<td>11mL</td>
</tr>
<tr>
<td>D=w/V</td>
<td>2.63</td>
<td>2.70</td>
<td>2.68</td>
<td>2.70</td>
<td>2.72</td>
<td>2.73</td>
</tr>
<tr>
<td>g/mL</td>
<td>g/mL</td>
<td>g/mL</td>
<td>g/mL</td>
<td>g/mL</td>
<td>g/mL</td>
<td>g/mL</td>
</tr>
</tbody>
</table>

Average 2.69 g/mL = 2.7 g/mL

Graphical Solution

\[ y = 2.742749 x + -2.364439 \times 10^{-1} \]
Sample Problem

What is the weight in lbs. of a rectangular solid Al object 2.0cm by 40mm by 10cm?

\[(2.0\text{cm} \cdot 40\text{mm} \cdot 10\text{cm})(2.7\text{g/mL})(1\text{mL/cm}^3)\]
\[(1\text{cm/10mm})(1\text{lbs/454g})=\]
\[= 0.48\text{lbs}\]

Specific Gravity

- Relates the density of a substance to that of water

\[
\text{specific gravity} = \frac{d_{\text{substance}}}{d_{\text{water}}}
\]
Specific Gravity

- At room temperature, water has a density of 8.34 lb/gal.
- The specific gravity of ethyl acetate is 0.902. What is its density in lb/gal and in g/ml? (1 gal = 3.7854 L)