Chapter 2: Measurement and Calculations

I. The Scientific Method

A. Definition: a logical approach to solving problems through observations, data, hypotheses, testing and formulating theories or conclusions.

B. Basic parts of the Scientific Method

1. Observations:
   a. They often involve measurements and collection of data.
   b. Observations are based on using the 5 senses to acquire information.
      i. Qualitative observations: descriptive in nature (no numbers)
         a) Examples: Color, odor, texture, shape
      ii. Quantitative observations involve numbers
         a) Examples: length, mass, liquid volume

2. Hypotheses
   a. These are testable statements.
   b. Hypotheses serve as the basis for experimental design.
   c. Often these are in the “if-then” form.

3. Experiments
   a. Experiments are done to test a hypothesis.
   b. Through many tests a model is often developed that is a best fit explanation of the data. Models can be visual, verbal, or mathematical.

4. Theory:
   a. A broad generalization that explains a body of facts or phenomena.

5. General order / steps of the scientific method:
   Define the problem
   Collect information
   Form a hypothesis
   Do an experiment/ analyze the data
   Form a conclusion

Hypothesis
A testable statement. A basis for making predictions and carrying out experiments.

Model
Usually proposed after a hypothesis is shown to be correct. A model that explains many phenomena may become part of a theory.

Theory
A broad generalization that explains a body of facts. It is supported by data. May contain a model.
II. Units of Measurement

A. Measurements are __quantitative__ information

1. They containing both a __number__ and a __unit__

   Example: 163 grams

   163 = number, grams = unit

B. The System International (SI)

Used by scientists to insure a uniform __standard__ of measurement for everyone

1. There are seven base units (bold units are commonly used in chemistry)

<table>
<thead>
<tr>
<th>Quantity</th>
<th>unit abbreviation</th>
<th>unit</th>
</tr>
</thead>
<tbody>
<tr>
<td>a. Length</td>
<td>m</td>
<td>meter</td>
</tr>
<tr>
<td>b. Mass</td>
<td>Kg</td>
<td>kilogram</td>
</tr>
<tr>
<td>c. Time</td>
<td>s</td>
<td>second</td>
</tr>
<tr>
<td>d. Temp.</td>
<td>K</td>
<td>Kelvin</td>
</tr>
<tr>
<td>e. Amount</td>
<td>mol</td>
<td>mole</td>
</tr>
</tbody>
</table>

2. All other units of measurement are __derived__ (calculated) from the above units

   a. Density is a derived unit

      Density = mass / length³

3. All base SI units use __prefixes__ to represent larger or smaller quantities.

   a. ml = 1/1000 of a liter
   b. cm = 1/100 of a meter
   c. kl = 1000 liters


C. Mass vs. Weight:

1. Mass: a measure of the amount of substance. __Gravity__ does not affect this

2. Weight: a measure of __force__ due to __gravity__.

D. Density: The ratio of mass to __volume__.

1. Density = mass/volume (d = m/v)

2. Density is a physical property of all matter at a given temperature and __pressure__

3. Density is a great tool to identify __unknown__ substances.

4. Practice problems:

   a. A sample of aluminum has a mass of 8.4g and a volume of 4.2 cm³. Calculate the density. \[ \text{Density} = \frac{8.4}{4.2} = 2 \text{ g/cm}^3 \]

   b. A diamond has a density of 3.26 g/ml. What is the mass of a diamond that has a volume of 0.351 ml? \[ M = D \cdot V = 3.26 \times 0.351 = \]

   c. **What is the volume** of a sample of liquid mercury that has a mass of 76.2 g, given that the density of mercury is 13.6 g/ml? \[ V = \frac{M}{D} = \frac{76.2}{13.6} = \]
III. Scientific Measurements

A. Accuracy vs Precision

1. **Accuracy** indicates how close a measurement is to the __accepted__ or actual or correct value.

2. **Precision** indicates how __reproducible__ a measurement is.
   
   a. Label the targets below as accurate, precise, both or neither. (Goal is the center).

B. Percent Error

1. Calculated to determine how close an experimental value is to the accepted or known value.

2. \[ \%\text{error} = \left( \frac{\text{experimental value} - \text{accepted value}}{\text{accepted value}} \right) \times 100 \]

3. Practice Problem:
   
   a. A student measures the mass and volume of a substance and calculates its density as 1.40 g/ml. The correct value is 1.30 g/ml. What is the percent error?
   
   \[ \frac{(1.4 - 1.3)}{1.3} \times 100 \]

   b. What is the percent error for a mass measurement of 17.7 g, given that the correct value is 21.2 g?
   
   \[ \frac{17.7 - 21.2}{21.2} \times 100 \]

C. Significant Figures

1. Note that each tool used to measure is *limited* to how accurate it is, or its level of certainty in the measurement.

2. Significant figures are measure of the all the __known__ digits (able to be read) plus one more estimated digit.

   a. The final digit in any measurement is _less_ certain and is estimated.

   b. Sig. Figs. indicate the limitation of the tool used for measuring
D. Rules for determining significant figures

1. If there are no zeros in the number, all digits are significant.
   Example: 2.87 has ___3___ significant digits.

2. Zeros between digits are significant
   Example: 30.41 has ___4___ significant digits

3. Zeros to the left or in front of digits are NOT significant
   Example: 0.00085 has ___2___ significant digits

4. Zeros at the end of a number AND to the right of a decimal are significant
   Example: 56.000 has ___5___ significant figures

5. Zeros at the end of a number or to the left of a digit WITHOUT a decimal are NOT significant
   Example: 2100. Has __________significant figures.

Examples:
28.6g = ________sig figs. 0.04604 L = ________sig figs
3440. cm = ________sig figs 804.05 g = ________ sig figs
910m = ________ sig figs .0067000kg = ________sig figs

6. For addition or subtraction the answer should be limited by the least number of decimal places held, not the total number of sig. figs.

7. For multiplication or division the answer should be limited to the least number of sig. figs. given in the problem or measured in the lab.

8. ROUND CORRECTLY!

9. Conversion factors are considered EXACT and DO NOT limit the number of sig. figs.

E. Scientific Notation

1. This allows very __large__ and very __small__ numbers to be written concisely

2. Numbers are in the form of a number between 1.0 and 9.99999... x 10 raised to any power.
   a. Examples: Convert these numbers to scientific notation:
      i. 0.00012mm
      ii. 560000cm
      iii. 0.0000000005730
   b. Convert these numbers to a decimal form:
      i. 7.050 x 10^3g
      ii. 4.00005 x 10^7mg
      iii. 7.84 x 10^-6nm
      iv. 8.4700 x 10^-3g