What is Chemistry?

**Definition:** Chemistry is the study of matter and its changes from one substance to another.

Chemistry is central to all sciences and overlaps with physics, biology, geology, and astronomy.

Learning about chemistry teaches you about the benefits and risks associated with chemicals and will help you to be an informed citizen and make intelligent choices concerning the world around you.

Chemistry teaches you to solve problems and communicate with others in an organized and logical manner.
Major Chemistry subdivisions:

- **Analytical (qualitative and quantitative) Chemistry:**
  - Qualitative Analytical Chemistry: What is a sample of matter composed of?
  - Quantitative Analytical Chemistry: Involves anything about the “stuff” that can be measured. How much “stuff” is in a sample of matter?

- **Biochemistry:**
  - The study of living systems (Biology + Chemistry)

- **Organic Chemistry:**
  - The study of properties and reactions of compounds that contain Carbon

- **Inorganic Chemistry:**
  - The study of properties and reactions of compounds that are not Carbon based

- **Physical Chemistry:**
  - The study of the physics involved with chemical changes.
Major Hints in being a successful Chemistry Student

• Study Everyday!! A 5 unit class requires 10 hours/week of study time MINIMUM to earn a “C”

• Don’t fall behind – you will never catch up!!!

• Always ask for help when you need it – don’t wait.

• Working in groups is very helpful IF you have the right group!!!
1.4 The Scientific Method: How Chemists Think

Ways to understand the world: Chemists use the scientific method—a way of learning that emphasizes observation and experimentation—to produce knowledge as the result of the senses.
The Scientific Method

• Observations involve measuring or observing some aspect of nature.
• Hypotheses are tentative interpretations of the observations.
• Laws summarize the results of a large number of observations.
• Theories are models that explain and give the underlying causes for observations and laws.
The Scientific Method

• Hypotheses, laws, and theories must be tested and validated by experiment.
• If they are not confirmed, they are revised and tested through further experimentation.
Scientific Method

Observation of an event or object

A question

A Hypothesis

Experimentation

Theory

New hypothesis

Further experimentation

Developement of new experimentation and theory

LAW
Theories are tested and validated by experiments.

• If a law, hypothesis, or theory is inconsistent with the findings of an experiment, it must be revised and new experiments must be conducted to test the revisions.

• Over time, poor theories are eliminated and good theories—those consistent with experiments—remain.
Example: The **atomic theory** of John Dalton (1766–1844)

- Dalton explained the law of conservation of mass by proposing that all matter was composed of small, indestructible particles called atoms.
- Dalton’s theory was a model of the physical world—it went beyond the laws and observations of the time to explain these laws and observations.
2.2 Scientific Notation: Writing Large and Small Numbers

- A number written in scientific notation has two parts.
- A **decimal part**: a number that is between 1 and 10.
- An **exponential part**: $10$ raised to an exponent, $n$. 

\[ 1.2 \times 10^{-10} \]

- decimal part
- exponential part

\[ \text{exponent (} n \text{)} \]
To convert a number to scientific notation

- If the decimal point is moved to the left, the exponent is positive.
- If the decimal is moved to the right, the exponent is negative.
Counting significant figures in a correctly reported measurement

1. All **nonzero** digits **are significant**.
2. **Interior zeros** (zeros between two numbers) **are significant**.
3. **Trailing zeros** (zeros to the right of a nonzero number) that fall after a decimal point **are significant**.
4. **Trailing zeros** that fall before a decimal point **are significant**.
5. **Leading zeros** (zeros to the left of the first nonzero number) **are NOT significant**. They only serve to locate the decimal point.
6. **Trailing zeros at the end** of a number, but before an implied decimal point, **are ambiguous** and should be avoided.
Exact Numbers

*Exact numbers have an unlimited number of significant figures.*

- Exact counting of discrete objects
- Integral numbers that are part of an equation
- Defined quantities
- *Some conversion factors are defined quantities while others are not.*
ACCURACY vs. PRECISION

• Accurate & precise

• inaccurate but precise

• inaccurate & imprecise
Figure 2.3, Page 15

Measuring temperature with various degrees of precision

(a) 30
(b) 30
(c) 30
Significant figures are based on the tools used to make the measurement. An imprecise tool will negate the precision of the other tools used. The following rules are used when measurements are used in calculations.
Rule 1. When the first digit after those you want to retain is 4 or less, that digit and all other to its right are dropped. The last digit retained is not changed. The following examples are rounded off to four digits:

74.693 = 74.69  
1.00629 = 1.006

Rule 2. When the first digit after those you want to retain is 5 or greater, that digit and all others to the right are dropped and the last digit retained is increased by one. These examples are rounded off to four digits:

1.026868 = 1.027  
18.02500 = 18.03

12.899 = 12.90
Many measurements in science involve either very large numbers or very small numbers (\#). Scientific notation is one method for communicating these types of numbers with minimal writing.

**GENERIC FORMAT:** \# . \# \#... x 10^\#

A negative exponent represents a number less than 1 and a positive exponent represents a number greater than 1.

6.75 x 10^{-3} is the same as 0.00675
6.75 x 10^{3} is the same as 6750
Give the following in scientific notation (or write it out) with the appropriate significant figures.

1. \(528900300000\) = \(5.289003 \times 10^{11}\)
2. \(0.000000000003400\) = \(3.400 \times 10^{-12}\)
3. \(0.23\) = \(2.3 \times 10^{-1}\)
4. \(5.678 \times 10^{-7}\) = \(0.0000005678\)
5. \(9.8 \times 10^{4}\) = \(98000\)
Dimensional Analysis

Dimensional Analysis (also call unit analysis) is one method for solving math problems that involve measurements. The basic concept is to use the units associated with the measurement when determining the next step necessary to solve the problem. **Always start with the given measurement** then immediately follow the measurement with a set of parentheses.

Keep in mind, try to ask yourself the following questions in order to help yourself determine what to do next.

1. **Do I want that unit?**
   If not, get rid of it by dividing by it if the unit is in the numerator, (if the unit is in the denominator, then multiply).

2. **What do I want?**
   Place the unit of interest in the opposite position in the parentheses.

**Numerator**  
**Denominator**
Calculate the number of weeks in 672 hours.

4 weeks

How many miles will a car travel in 3.00 hours at an average speed of 62.0 miles per hour?

186 miles

If you went to the bank to obtain 325 dollars in nickels, how many nickels would you receive?

6500 nickels
MEASUREMENTS

- There are different types of measurements that can be made in the lab for length, mass, volume, temperature, area, time, heat and pressure.

<table>
<thead>
<tr>
<th>Unit</th>
<th>Metric</th>
<th>English</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td>Meter (m)</td>
<td>Inches (in) or Feet (ft)</td>
</tr>
<tr>
<td>Mass</td>
<td>Gram (g)</td>
<td>Pounds (lb)</td>
</tr>
<tr>
<td>Volume</td>
<td>Liters (L)</td>
<td>Gallon (gal)</td>
</tr>
<tr>
<td>Temperature</td>
<td>Celsius (°C) and Kelvin (K)</td>
<td>Fahrenheit (°F)</td>
</tr>
<tr>
<td>Area</td>
<td>Square meters (m²)</td>
<td>Square feet (ft²)</td>
</tr>
<tr>
<td>Time</td>
<td>Seconds (s)</td>
<td>Minutes (min) or Hours (hr)</td>
</tr>
<tr>
<td>Heat</td>
<td>Calories (cal) or Joules (J)</td>
<td>British Thermal Units (BTU)</td>
</tr>
<tr>
<td>Pressure</td>
<td>Atmospheres (atm), Torr, or mmHg</td>
<td>Pounds/sq in (lb/in²)</td>
</tr>
</tbody>
</table>
### Prefix Multipliers

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Multiplier</th>
</tr>
</thead>
<tbody>
<tr>
<td>tera-</td>
<td>T</td>
<td>1,000,000,000,000,000</td>
</tr>
<tr>
<td>giga-</td>
<td>G</td>
<td>1,000,000,000</td>
</tr>
<tr>
<td>mega-</td>
<td>M</td>
<td>1,000,000</td>
</tr>
<tr>
<td>kilo-</td>
<td>k</td>
<td>1,000</td>
</tr>
<tr>
<td>deci-</td>
<td>d</td>
<td>0.1</td>
</tr>
<tr>
<td>centi-</td>
<td>c</td>
<td>0.01</td>
</tr>
<tr>
<td>milli-</td>
<td>m</td>
<td>0.001</td>
</tr>
<tr>
<td>micro-</td>
<td>μ</td>
<td>0.000001</td>
</tr>
<tr>
<td>nano-</td>
<td>n</td>
<td>0.0000000001</td>
</tr>
<tr>
<td>pico-</td>
<td>p</td>
<td>0.0000000000001</td>
</tr>
<tr>
<td>femto-</td>
<td>f</td>
<td>0.0000000000000001</td>
</tr>
</tbody>
</table>
MEASUREMENTS

Significant Figures & Calculations

• Adding/subtracting

  The result should be rounded to the same number of **decimal places** as the measurement with the **least number of decimal places**.

• Multiplying/dividing

  The result should contain the same number of **significant figures** as the measurement with the **least number of significant figures**.
1. All nonzero numbers are significant figures.

2. Zero’s follow the rules below.
   - Zero’s between numbers are significant.
     - 30.09 has 4 SF
   - Zero’s that precede are NOT significant.
     - 0.000034 has 2 SF
   - Zero’s at the end of decimals are significant.
     - 0.009000 has 4 SF
   - Zero’s at the end without decimals are either.
     - 4050 has either 4 SF or 3 SF
Adding & Subtracting

\[
\begin{array}{ccc}
345.678 & + & 0.07283 & - & 1587 \\
12.67 & - & 0.0162789 & - & 120 \\
358.348 & - & 0.0565511 & = & 1467 \\
358.35 & - & 0.05655 & = & 1470 \text{ or } 1.47 \times 10^3
\end{array}
\]

All Answers are Incorrect!!!

Multiplication & Division

\[
\begin{array}{ccc}
(12.034)(3.98) = & 47.89532 & 47.9 \text{ is correct} \\
98.657 & 43 = & 2.294348837 & 2.3 \text{ is correct} \\
(13.59)(6.3) = & 7.13475 & 7.1 \text{ is correct}
\end{array}
\]
### Table 3.3 Metric–USCS and USCS–USCS Conversion Factors

**Length**
- 1 in. $\equiv 2.54$ cm (definition)
- 1 ft $\equiv 12$ in.
- 1 yd $\equiv 3$ ft
- 1 mi $\equiv 5280$ ft

**Mass**
- 1 lb $\equiv 453.59$ g
- 1 kg $\equiv 2.205$ lb
- 1 lb $\equiv 16$ oz

**Volume**
- 1 L $\equiv 1.057$ qt
- 1 gal $\equiv 3.785$ L
- 1 ft$^3$ $\equiv 28.32$ L
- 1 qt $\equiv 16$ fl oz
- 1 gal $\equiv 4$ qt

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CONVERSIONS

Convert the following:

1. The Hope diamond is the world’s largest blue diamond. It weighs 44.4 carats. If 1 carat is defined as 200 mg, calculate the mass of the diamond in grams and ounces. (28.375 grams = 1 ounce)

   0.313 oz, 8.88 g

2. The Sears tower in Chicago is 1454 feet tall. How high is this in meters? (39.4 inches = 1 meter)

   443.2 m
CONVERSIONS

Convert the following:

1. **7.00 in\(^3\) → mL**

There is no direct conversion from in\(^3\) to mL so now you will have to develop a multi-step process that will start with in\(^3\) and end with mL. If you memorize that 1 mL = 1 cm\(^3\), this problem becomes easy. All you need to look up is the relationship between in and cm.

1 in = 2.54 cm  
1 mL = 1 cm\(^3\)

\[
7.00 \text{ in}^3 \times \frac{2.54 \text{ cm}^3}{1 \text{ in}} \times \frac{1 \text{ mL}}{1 \text{ cm}^3} = ?
\]

Place the conversion inside the parenthesis and the cube on the outside. Then cube the number inside the parenthesis.

\[
\begin{align*}
7.00 \text{ in}^3 & \times \frac{16.387 \text{ cm}^3}{1 \text{ in}^3} \times \frac{1 \text{ mL}}{1 \text{ cm}^3} = 115 \text{ mL}
\end{align*}
\]
CONVERSIONS & WORD PROBLEMS

Now it is time to apply these techniques to word problems. Nothing changes but it helps if you separate the words from the numbers. Therefore your first step should be to make a list.

1. How many miles will a car drive on 23.0 L of gasoline if the car averages 59.0 km/gal?

\[
\text{mi} = ? \quad 23.0 \text{ L} \quad 59.0 \text{ km / gal}
\]

Note that \(\text{mi} \) & \(\text{km}\) are units for \textit{length} and \(\text{L} \) & \(\text{gal}\) are units for \textit{volume}. Looking at the units you should notice that you will need to convert \(\text{km}\) to \(\text{mi}\) and \(\text{L}\) to \(\text{gal}\) so list the conversion factors you will use. You can only convert units of the same measurement type (You can not directly convert \(\text{km}\) to \(\text{gal}\), unless there is an additional stipulation connecting the two units like the 59.0 km/gal).

\[
1 \text{ mi} = 1.61 \text{ km} \quad 1 \text{ L} = 1.0567 \text{ qt} \quad 4 \text{ qt} = 1 \text{ gal}
\]

Always start with the single unit measurement:

\[
23.0 \text{ L} \quad 1.0567 \text{ qt} \quad 1 \text{ gal} \quad 59.0 \text{ km} \quad 1 \text{ mi} = 223 \text{ mi}
\]

\[
\begin{align*}
1 \text{ L} & \quad 4 \text{ qt} & \quad 1 \text{ gal} & \quad 1.61 \text{ km} \\
\end{align*}
\]
MEASUREMENTS

TEMPERATURE

• A physical property of matter that determines the direction of heat flow.
  – Heat in = endothermic process
  – Heat out = exothermic process

• Temperature is measured with a thermometer.

• Temperature: \( F = (1.8 \ C) + 32 \)

  \[
  C = \frac{F - 32}{1.8}
  \]

  \[ K = C + 273.15 \]
Normal body temperature is 37°C or 98.6°F.
# Measurements - Temperature

## Fill in the blanks

<table>
<thead>
<tr>
<th>Celsius</th>
<th>Fahrenheit</th>
<th>Kelvin</th>
</tr>
</thead>
<tbody>
<tr>
<td>69</td>
<td>156</td>
<td>342</td>
</tr>
<tr>
<td>-34</td>
<td>-29</td>
<td>239</td>
</tr>
<tr>
<td>-162</td>
<td>-260</td>
<td>111</td>
</tr>
</tbody>
</table>
Introduction to Density

- Density is the measurement of the mass of an object per unit volume of that object.
  
  \[ d = \frac{m}{V} \]

- Density is usually measured in g/mL, g/cm\(^3\) for solids or liquids and g/L for gases.

- Volume may be measured in the lab using a graduated cylinder or calculated using:
  
  - **Volume** = \( \text{length} \times \text{width} \times \text{height} \) if a box or
  
  \[ V = \pi r^2 h \] if a cylinder.

- Remember 1 mL = 1 cm\(^3\).
How to measure the density of a solid in the laboratory.

• Obtain a clean graduated cylinder.

• Fill the graduated cylinder with enough water to cover the object. Record the volume.

• Carefully place the object into the water filled graduated cylinder.

• Record the new water level.

• The volume of the object is the $V_{\text{final}} - V_{\text{initial}}$. 
1. An irregular shaped object has a mass of 50.0 g was placed in a graduated cylinder with an initial volume of 25.0 mL. After placing the object in the cylinder the volume increased to 77.5 mL. What is its density?

2. Calculate the mass of 12.0 cm x 12.0 cm x 0.150 cm piece of copper which has a density of 9.00 g/cm³

3. Calculate the volume occupied by 15.4 grams of nickel which has a density of 8.91 g/mL
Example: A cylindrical piece of metal is 2.03 inches high and has a diameter 17.0 mm wide and weighs 31.599 g. Identify the metal. (Table 3.4 p. 77)

\[ r = \frac{1}{2}d = \frac{1.70 \text{ cm}}{2} = 0.85 \text{ cm} \]
\[ h = 2.03 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 5.1562 \text{ cm} \]

\[ V = \pi r^2 h = \pi (0.85 \text{ cm})^2 (5.1562 \text{ cm}) = 11.7 \text{ cm}^3 \]

\[ D = \frac{m}{V} = \frac{31.599 \text{ g}}{11.7 \text{ mL}} = 2.70 \text{ g/mL} \]

The metal is aluminum!
Density as a Conversion Factor

- Table 2.4 provides a list of the densities of some common substances.
- This is useful when solving homework problems.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Density (g/cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Charcoal, oak</td>
<td>0.57</td>
</tr>
<tr>
<td>Ethanol</td>
<td>0.789</td>
</tr>
<tr>
<td>Ice</td>
<td>0.92</td>
</tr>
<tr>
<td>Water</td>
<td>1.0</td>
</tr>
<tr>
<td>Glass</td>
<td>2.6</td>
</tr>
<tr>
<td>Aluminum</td>
<td>2.7</td>
</tr>
<tr>
<td>Titanium</td>
<td>4.50</td>
</tr>
<tr>
<td>Iron</td>
<td>7.86</td>
</tr>
<tr>
<td>Copper</td>
<td>8.96</td>
</tr>
<tr>
<td>Lead</td>
<td>11.4</td>
</tr>
<tr>
<td>Gold</td>
<td>19.3</td>
</tr>
<tr>
<td>Platinum</td>
<td>21.4</td>
</tr>
</tbody>
</table>

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