Chapter 1: Matter and Measurement

Chemistry: The science that describes matter, its chemical and physical properties, the chemical and physical changes that matter undergoes, and the energy changes that accompany these processes.

Remember—chemistry is an experimental science based on observations. Theories have been developed by scientists to
(1) explain phenomena
(2) predict phenomena that have not yet happened.

How do we classify matter?
state of matter: solid liquid gas

solid liquid gas

Kinetic molecular theory—all matter consists of tiny particles (atoms, molecules, ions) in constant motion (above 0 K)
solids: particles packed close together & vibrate in place
liquids: particles are close but can move past each other (fluid)
gases: particles are far apart, move quickly, occupy full volume of container (fluid)

Chemistry is thought to be a complex subject to learn because one has to think at several different levels at once:

consider water,

H2O
symbolic view

macrophscopic particulate view
(observations, experimentation) molecular or atomic


pure substance: has unique set of properties and cannot be separated by physical means. A compound or element are pure substances.

(more later)

mixture: a conglomeration of pure substances.
- heterogeneous mixture: soil (rocks, clay, water, worms, bacteria, gases etc.)
- sugar + salt, sand + water
  you can easily see the uneven texture of material
- homogeneous mixture: solutions (salt in water, brass-alloy)
  can be separated by physical means into pure substances.
  for example: you can boil the water off a salt water solution, leaving the salt behind.

element: a pure substance composed of only 1 kind of small particle.
- 116 elements are known - see periodic table.

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Element</th>
<th>Symbol</th>
<th>Element</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ag</td>
<td>silver (argentum)</td>
<td>F</td>
<td>fluorine</td>
</tr>
<tr>
<td>Al</td>
<td>aluminum</td>
<td>Fe</td>
<td>iron (ferrum)</td>
</tr>
<tr>
<td>Au</td>
<td>gold (aurum)</td>
<td>H</td>
<td>hydrogen</td>
</tr>
<tr>
<td>B</td>
<td>boron</td>
<td>He</td>
<td>helium</td>
</tr>
<tr>
<td>Ba</td>
<td>barium</td>
<td>Hg</td>
<td>mercury (hydrargyrum)</td>
</tr>
<tr>
<td>Bi</td>
<td>bismuth</td>
<td>I</td>
<td>iodine</td>
</tr>
<tr>
<td>Br</td>
<td>bromine</td>
<td>K</td>
<td>potassium (kalium)</td>
</tr>
<tr>
<td>C</td>
<td>carbon</td>
<td>Kr</td>
<td>krypton</td>
</tr>
<tr>
<td>Ca</td>
<td>calcium</td>
<td>Li</td>
<td>lithium</td>
</tr>
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<td>Cd</td>
<td>cadmium</td>
<td>Mg</td>
<td>magnesium</td>
</tr>
<tr>
<td>Cl</td>
<td>chlorine</td>
<td>Mn</td>
<td>manganese</td>
</tr>
<tr>
<td>Co</td>
<td>cobalt</td>
<td>N</td>
<td>nitrogen</td>
</tr>
<tr>
<td>Cr</td>
<td>chromium</td>
<td>Na</td>
<td>sodium (natrium)</td>
</tr>
<tr>
<td>Cu</td>
<td>copper (cuprum)</td>
<td>Ne</td>
<td>neon</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Ni</td>
<td>nickel</td>
</tr>
<tr>
<td></td>
<td></td>
<td>O</td>
<td>oxygen</td>
</tr>
<tr>
<td></td>
<td></td>
<td>P</td>
<td>phosphorus</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Pb</td>
<td>lead (plumbum)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Pt</td>
<td>platinum</td>
</tr>
<tr>
<td></td>
<td></td>
<td>S</td>
<td>sulfur</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Sb</td>
<td>antimony (stibium)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Si</td>
<td>silicon</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Sn</td>
<td>tin (stannum)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Sr</td>
<td>strontium</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Ti</td>
<td>titanium</td>
</tr>
<tr>
<td></td>
<td></td>
<td>U</td>
<td>uranium</td>
</tr>
<tr>
<td></td>
<td></td>
<td>W</td>
<td>tungsten (Wolfram)</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Zn</td>
<td>zinc</td>
</tr>
</tbody>
</table>

atom is smallest part of an element that contains the same characteristics as the element.

compound: substance containing 2 or more elements in fixed ratio. The smallest part of compound with same characteristics is the
molecule - if the compound is held together by covalent bonds
formula unit - if the compound is held together by ionic bonds
(more later)

properties: characteristic that describe samples of matter.
- physical property: observed in the absence of change in composition
  - examples: color, density, hardness, melting point, boiling pt.
    electrical conductivity, mass, volume.
- divided into 2 kinds
  - extensive - depends on amount of substance present
    e.g. mass, volume
  - intensive - independent of amount of substance present
    e.g. color, temperature, density, melting point
- chemical property: a property that matter exhibits as it changes
  composition, e.g. magnesium metal can react with molecular oxygen.

physical change - occurs with no change in chemical composition,
but energy can be released or absorbed
(1) lead melting - process needs heat (endothermic)
\[
heat + \text{\text{solid}} \rightarrow \text{\text{liquid}}
\]
(2) water freezing - process releases heat (exothermic)
\[
\text{\text{liquid}} \rightarrow \text{\text{solid}} + heat
\]

chemical change: occurs when (1) one or more substances are used up
(2) one or more substances are produced
(3) energy is released or absorbed.

\[
\begin{align*}
\text{reactants} & : 2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g}) \\
\text{product} & : \text{balanced} \\
\text{Law of Conservation of Matter}
\end{align*}
\]
Units of Measurement

**International System of Units (SI)** adopted by National Bureau of Standards includes:

- **mass**
- **length**
- **time**
- **temperature**
- **amount of substance**
- **electric current**

<table>
<thead>
<tr>
<th>Unit</th>
<th>Symbol</th>
<th>Prefix</th>
<th>Factor</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>kilogram</td>
<td>kg</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>meter</td>
<td>m</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>second</td>
<td>s</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Kelvin</td>
<td>K</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>mole</td>
<td>mol</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ampere</td>
<td>A</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Common prefixes:**

- mega-: \( 10^6 \)  
  - example: \( Mg = 1 \times 10^6 \) grams
- kilo-: \( 10^3 \)  
  - example: \( kg = 1 \times 10^3 \) grams
- deci-: \( 10^{-1} \)  
  - example: \( dL = 0.1 \) liter = 100 mL
- centi-: \( 10^{-2} \)  
  - example: \( cm = 0.01 \) meter
- milli-: \( 10^{-3} \)  
  - example: \( mg = 1 \times 10^{-3} \) grams
- micro-: \( 10^{-6} \)  
  - example: \( \mu m = 1 \times 10^{-6} \) meter
- nano-: \( 10^{-9} \)  
  - example: \( ng = 1 \times 10^{-9} \) grams
- pico-: \( 10^{-12} \)  
  - example: \( pg = 1 \times 10^{-12} \) grams

**Temperature:** we will only be working with °C and K

- absolute zero = 0 K = -273.15 °C

**Volume:** 1 cubic centimeter = \( 1 \text{ cm} \times 1 \text{ cm} \times 1 \text{ cm} = 1 \text{ cm}^3 = 1 \text{ mL} \)

**Mass vs Weight:**

- **mass**: measure of quantity of matter a body contains (grams)
- **weight**: measure of gravitational attraction of earth for the body

weight = force = mass \( \times \) gravitational acceleration, g
length: measured in meters (cm, mm, etc.)
Volume: measured in liters (cmL)

**Significant Figures**

There are two kinds of numbers in the world:
- **Exact**: there are exactly 12 eggs in a dozen
- **Inexact**: any measurement:
  - I quickly measure the width of this paper: 220 mm, 2 sig fig.
  - If I am more precise: 216 mm, 3
  - If I am even more precise: 215.6 mm, 4

**Note**: **Precision vs. Accuracy**

- **Accuracy**: refers to how closely a measured value agrees with the correct value.
- **Precision**: refers to how closely individual measurements agree with each other.

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In any measurement, the number of significant figures is important; the number of significant figures is the number of digits believed to be correct by the person doing the measuring. It includes 1 estimated digit.

Let's look at another example — volume measurement.

Volumes of liquids are measured in the laboratory in many different ways:

- **Beaker**
  - 80 mL
  - 60 mL
  - 40 mL
  - 20 mL
  - 50 ± 2 mL (1 of 10 mL)
  - Digit containing error

- **Graduated Cylinder**
  - 40 mL
  - 25 mL
  - Meniscus
  - 37.3 ± 0.2 mL
  - (1 of smallest division = 1 mL)

- **Buret**
  - 20 mL
  - 19 mL
  - 21 mL
  - 20.27 ± 0.02 mL
  - (Smallest division = 0.1 mL)

**Rule of Thumb**: read measurements to 1/5 or 0.2 of the smallest division.
Volume obtained from the buret has 4 significant figures - the last digit is a best estimate, whereas the volume measured from a beaker only has 2 significant figures. If a person needed only a rough estimate of volume the beaker volume is satisfactory, otherwise one would use the graduated cylinder or better yet the buret.

Rules for significant figures:

(1) Leading zeros are never significant
imbedded zeros are always significant
trailing zeros are significant only if the decimal point is specified.

(2) Addition or subtraction: last digit retained is set by the first doubtful digit.

(3) Multiplication or division: answer contains no more sig. fig. than the least accurately known number.

Examples:

<table>
<thead>
<tr>
<th>Number of Sig. Fg.</th>
<th>Scientific Notation</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td>6.82 x 10^3</td>
</tr>
<tr>
<td>4</td>
<td>1.072 (x10^3)</td>
</tr>
<tr>
<td>1</td>
<td>3 x 10^2</td>
</tr>
<tr>
<td>3</td>
<td>3.00 x 10^2</td>
</tr>
<tr>
<td>4</td>
<td>3.000 x 10^2</td>
</tr>
</tbody>
</table>

Addition: \(4.7832\) (same for subtraction) + \(1.234\) \(= 8.0172\) \(\downarrow 8.04\)

Multiplication: \(2.8723 \times 1.6\) \(= 4.59568\) \(\downarrow 4.6\)

Dimensional Analysis (factor-label or unit factor method)

This is a very powerful mathematical tool for solving problems using unit factors (a factor in which numerator & denominator have different units but represent equivalent amounts :: it is equal to unity)
Rounding Off:

When rounding off numbers to a certain number of significant figures - do so to the nearest value.

\[ 2.3467 \times 10^4 \] to 3 significant figures \[ 2.35 \times 10^4 \]
\[ 1.612 \times 10^3 \] to 2 significant figures \[ 1.6 \times 10^3 \]

What happens if there is a 5. For example

\[ 2.35 \times 10^2 \] to 2 significant figures

\[ 2.4 \times 10^2 \]
\[ 2.3 \times 10^2 \]

How do we round this number.

There is an arbitrary rule: if the number before the 5 is odd, round up.
if the number before the 5 is even, round down.

\[ 2.35 \times 10^2 \] to 2 significant figures \[ 2.4 \times 10^2 \]
\[ 2.45 \times 10^2 \] to 2 significant figures \[ 2.4 \times 10^2 \]

Of course,

\[ 2.451 \times 10^2 \] to 2 significant figures \[ 2.5 \times 10^2 \]
since \[ 2.451 \times 10^2 \] is closer to \[ 2.5 \times 10^2 \] than \[ 2.4 \times 10^2 \]
I suggest using dimensional analysis for converting units - but using it only as a last resort for problem solving, unless you are very skillful at it. It is useful as a check for correct units.

Examples:
(1) How many centimeters are in 20.0 inches? (1 inch = 2.54 cm)

\[ ? \text{ cm} = 20.0 \text{ inches} \times \left( \frac{2.54 \text{ cm}}{1 \text{ inch}} \right) = 50.8 \text{ cm} \]

unit factor

(2) Express 6.2 gallons in milliliters
( 1 gallon = 4 qt ; 1 qt = 0.946 L ; 1 L = 1000 mL)

\[ ? \text{ mL} = 6.2 \text{ gallons} \times \frac{4 \text{ quarts}}{1 \text{ gal}} \times \frac{0.946 \text{ L}}{1 \text{ qt}} \times \frac{1000 \text{ mL}}{1 \text{ L}} \]
\[ = 2.3 \times 10^4 \text{ mL} \quad \left(4 \text{ quarts are exact numbers}\right) \]

(3) Express \(6.12 \times 10^3\) cm² in ft².

\[ ? \text{ ft}^2 = 6.12 \times 10^3 \text{ cm}^2 \times \frac{(1 \text{ inch})^2}{(2.54 \text{ cm})^2} \times \frac{(1 \text{ ft})^2}{(12 \text{ inch})^2} \]
\[ = 6.59 \text{ ft}^2 \]

Density and Specific Gravity

Density is defined as mass per unit volume

\[ 1 \text{ mL} = 1 \text{ cm}^3 \]

\[ \text{density} = \frac{\text{mass}}{\text{volume}} \]

Liquido y solido: unido: \( \frac{g}{\text{mL}} \) or \( \frac{g}{\text{cm}^3} \)

Gas: unidades: \( \frac{g}{L} \quad (g \cdot L^{-1}) \)

Example: Given a substance with mass of 472 g with a volume of 97.3 cm³, what is its density?

\[ D \left( \frac{g}{\text{cm}^3} \right) = \frac{472 \text{ g}}{97.3 \text{ cm}^3} = 4.85 \text{ g/cm}^3 \]

Specific gravity has no units. It is a ratio:

\[ \text{Density of substance} = \frac{\text{density of substance}}{\text{density of water}} \]

Numerically, density = specific gravity

\[ \text{density of substance} = 1.00 \text{ g/mL} \]
Example: The density of water is dependent both on temperature and phase.

<table>
<thead>
<tr>
<th>Phase</th>
<th>Ice $\rightarrow$ Water</th>
</tr>
</thead>
<tbody>
<tr>
<td>$T (\degree C)$</td>
<td>0°</td>
</tr>
<tr>
<td>Density (g/mL)</td>
<td>0.917</td>
</tr>
</tbody>
</table>

![Density vs. Temperature Graph](image)

$\star$ Notes on Plasma: another form of the gaseous state.

**Plasma**

At several thousand degrees gas molecules are traveling with such high energies that on collision they drive off electrons. This process may continue to strip off each of the electrons of an atom until at very high temperatures all of the electrons may be removed. This process produces a high concentration of high-energy charged particles that have become known as plasma. This is believed to be the condition of much of the gaseous material in the sun, stars, and interstellar space. Electrons may also be driven off in other processes involving the absorption of a great deal of energy. Today the term plasma is used for a high concentration of these high-energy particles that have lost one or more electrons.

At very high temperatures, 10,000,000$^\circ$C and higher, the plasma particles have gained so much translational energy that on collision their nuclei approach within $10^{-12}$ cm of each other where the powerful nuclear attractive forces take effect and cause nuclear fusion processes to occur. These are some of the main reactions that produce the energy and the light elements in the sun and stars. The fusion process will be discussed in Chap. 26.

The study of plasma is one of the new and very important fields of research today.