Unit 2: Matter and Energy

**Matter**  
*Introductory Definitions*

*matter*: anything having mass and volume

*mass*: the amount of matter in an object

*weight*: the pull of gravity on an object

*volume*: the space an object occupies

  units: L, dm$^3$, mL, cm$^3$

*state of matter*: solid, liquid, or gas  (plasma, neutron star)

*composition*: what the matter is made of

  copper: many Cu atoms

  water: many groups of 2 H’s and 1 O

*properties*: describe the matter

  -- what it looks like, smells like, etc.

  -- how it behaves

*atom*: a basic building block of matter

  ~100 diff. kinds
- **Elements** → contain only one type of atom

1. **monatomic** elements consist of unbonded, “like” atoms
   e.g., Fe, Al, Cu, He

2. **polyatomic** elements consist of several “like” atoms bonded together
   - **diatomic** elements: \( H_2 \), \( O_2 \), \( Br_2 \), \( F_2 \), \( I_2 \), \( N_2 \), \( Cl_2 \)
   - others: \( P_4 \), \( S_8 \)

**allotropes**: different forms of the same element in the same state of matter

**OXYGEN**
- oxygen gas
- ozone

**CARBON**
- elemental carbon
- graphite
- diamond
- buckyball

**molecule**: a neutral group of bonded atoms
<table>
<thead>
<tr>
<th>Description</th>
<th>Chemical Symbol</th>
<th>Model</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 oxygen atom</td>
<td>O</td>
<td></td>
</tr>
<tr>
<td>1 oxygen molecule</td>
<td>O₂</td>
<td></td>
</tr>
<tr>
<td>2 unbonded oxygen atoms</td>
<td>2 O</td>
<td></td>
</tr>
<tr>
<td>1 phosphorus atom</td>
<td>P</td>
<td></td>
</tr>
<tr>
<td>1 phosphorus molecule</td>
<td>P₄</td>
<td></td>
</tr>
<tr>
<td>4 unbonded phosphorus atoms</td>
<td>4 P</td>
<td></td>
</tr>
</tbody>
</table>

Elements may consist of either molecules or unbonded atoms.
- **Compounds**

...contain two or more different types of atoms
...have properties that are different from those of their constituent elements

Na (sodium): explodes in water \( \rightarrow \) table salt
Cl\(_2\) (chlorine): poisonous gas \( \rightarrow \) (NaCl)

Atoms can only be altered by *nuclear* means.
Molecules can be altered by *chemical* means.
(i.e., chemical reactions, chemical changes)

e.g., Dehydration of sugar

\[
C_{12}H_{22}O_{11}(s) \rightarrow 12 \text{ C}(s) + 11 \text{ H}_2\text{O}(g)
\]

Electrolysis of water

\[
2 \text{ H}_2\text{O}(l) \rightarrow 2 \text{ H}_2(g) + \text{ O}_2(g)
\]

In a chemical reaction, the atoms are rearranged.
**Compound Composition** → All samples of a given compound have the same composition.

Phosgene gas (COCl$_2$) is 12.1% carbon, 16.2% oxygen, and 71.7% chlorine by mass. Find # of g of each element in 254 g of COCl$_2$.

\[
\begin{align*}
X \text{ g C} &= 254 \text{ g (0.121)} = 30.7 \text{ g C} \\
X \text{ g O} &= 254 \text{ g (0.162)} = 41.1 \text{ g O} \\
X \text{ g Cl} &= 254 \text{ g (0.717)} = 182.1 \text{ g Cl}
\end{align*}
\]

A sample of butane (C$_4$H$_{10}$) contains 288 g carbon and 60 g hydrogen. Find…

A. total mass of sample

\[
288 \text{ g C} + 60 \text{ g H} = 348 \text{ g}
\]

B. % of each element in butane

\[
\begin{align*}
% \text{C} &= \frac{288 \text{ g C}}{348 \text{ g}} = 0.828 \\
% \text{H} &= \frac{60 \text{ g H}}{348 \text{ g}} = 0.172
\end{align*}
\]

\[
82.8\% \text{ C, 17.2\% H}
\]

C. how many g of C and H are in a 24.2 g sample

\[
\begin{align*}
X \text{ g C} &= 24.2 \text{ g (0.828)} = 20.0 \text{ g C} \\
X \text{ g H} &= 24.2 \text{ g (0.172)} = 4.2 \text{ g H}
\end{align*}
\]
A 550 g sample of chromium (III) oxide (Cr$_2$O$_3$) has 376 g Cr. How many grams of Cr and O are in a 212 g sample of Cr$_2$O$_3$?

\[
\% \text{ Cr} = \frac{376 \text{ g Cr}}{550 \text{ g}} = 68.4\% \text{ Cr and } 31.6\% \text{ O}
\]

\[
\begin{align*}
X \text{ g Cr} &= 212 \text{ g (0.684)} = 145 \text{ g Cr} \\
X \text{ g O} &= 212 \text{ g (0.316)} = 67 \text{ g O}
\end{align*}
\]

**Classifying Matter**

- **(Pure) Substances**
  - have a fixed composition
  - have fixed properties

<table>
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<tr>
<th>ELEMENTS</th>
<th>COMPOUNDS</th>
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<tr>
<td>e.g., Fe, N$_2$, S$_8$, U</td>
<td>e.g., H$_2$O, NaCl, HNO$_3$</td>
</tr>
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</table>

Pure substances have a chemical formula.
Mixtures
two or more substances mixed together
...have varying composition
...have varying properties
The substances are NOT chemically bonded, and they retain their individual properties.

Two types of mixtures...

homogeneous: (or solution)
particles are microscopic;
sample has same composition and properties throughout;
evenly mixed
e.g., salt water
Kool Aid

heterogeneous:
different composition and properties in the same sample;
unevenly mixed
e.g., tossed salad
raisin bran

alloy: a homogeneous mixture of metals
e.g., bronze (Cu + Sn)
brass (Cu + Zn)
pewter (Pb + Sn)
suspension: settles over time
e.g., liquid meds,
muddy water
Contrast…

24K GOLD  
- pure gold
- 24/24 atoms are gold
- element

14K GOLD  
- mixture
- 14/24 atoms are gold
- homogeneous mixture

Chart for Classifying Matter

MATTER

PURE SUBSTANCE  MIXTURE

ELEMENT  HOMOGENEOUS

COMPOUND  HETEROGEOUS
A sample of bronze contains 68 g copper and 7 g tin.

A. Find total mass of sample.

\[ 68 \text{ g Cu} + 7 \text{ g Sn} = 75 \text{ g} \]

B. Find % Cu and % Sn.

\[ \% \text{ Cu} = \frac{68 \text{ g Cu}}{75 \text{ g}} = 90.7\% \text{ Cu and 9.3\% Sn} \]

C. How many grams of each element does a 346 g sample of bronze contain?

Best answer: don’t know.

(Bronze is a mixture and isn’t necessarily always 90.7% Cu and 9.3% Sn.)

However, assuming these % are correct…

\[ X \text{ g Cu} = 346 \text{ g (0.907)} = 314 \text{ g Cu} \]
\[ X \text{ g Sn} = 346 \text{ g (0.093)} = 32 \text{ g Sn} \]
- **Separating Mixtures**
  ...involves physical means, or physical changes

1. **sorting**: by color, shape, texture, etc.
2. **filter**: particle size is different
3. **magnet**: one substance must contain iron
4. **chromatography**: some substances dissolve more easily than others
5. **density**: “sink vs. float”
   perhaps use a **centrifuge**
   **decant**: to pour off the liquid
6. **distillation**: different boiling points

No chemical reactions are needed;
substances are NOT bonded.
**Density** → how tightly packed the particles are

\[
\text{Density} = \frac{\text{mass}}{\text{volume}} \quad \Rightarrow \quad D = \frac{m}{V}
\]

Typical units: \( \text{g/cm}^3 \) for solids, \( \text{g/mL} \) for fluids

To find volume, use…
1. a formula
2. water displacement method

** Density of water = 1.0 g/mL = 1.0 g/cm\(^3\)**

Things that are “less dense” float in things that are “more dense.”

The density of a liquid or solid is nearly constant, no matter what the sample’s mass.
Galilean Thermometer Problem

On a cold morning, a teacher walks into a cold classroom and notices that all bulbs in the Galilean thermometer are huddled in a group. Where are the bulbs, at the top of the thermometer or at the bottom?

1. Bulbs have essentially fixed masses and volumes. Therefore, each bulb has a relatively fixed density.
2. The surrounding liquid has a fixed mass, but its volume is extremely temperature-dependent.
3. The density of the liquid can be written as...

\[ D_{liq} = \frac{m_{liq}}{V_{liq}} \]

...so...

...if the liquid is cold: ...but if it’s hot:

\[ \frac{m_{liq}}{V_{liq}} = D_{liq} \]

\[ \frac{m_{liq}}{V_{liq}} = \frac{m_{liq}}{V_{liq}} = D_{liq} \]

On a cold morning, where are the bulbs?

AT THE TOP
**Density Calculations**

1. A sample of lead (Pb) has mass 22.7 g and volume 2.0 cm\(^3\). Find sample’s density.

\[
D = \frac{m}{V} = \frac{22.7 \text{ g}}{2.0 \text{ cm}^3} = 11.35 \text{ g/cm}^3
\]

2. Another sample of lead occupies 16.2 cm\(^3\) of space. Find sample’s mass.

\[
m = D \cdot V = 11.35 \cdot \frac{\text{g}}{\text{cm}^3} \cdot 6.2 \text{ cm}^3 = 184 \text{ g}
\]

3. A 119.5 g solid cylinder has radius 1.8 cm and height 1.5 cm. Find sample’s density.

\[
V = \pi r^2 h = 3.14 \cdot (1.8 \text{ cm})^2 \cdot (1.5 \text{ cm}) = 15.3 \text{ cm}^3
\]

\[
D = \frac{m}{V} = \frac{119.5 \text{ g}}{15.3 \text{ cm}^3} = 7.81 \text{ g/cm}^3
\]

4. A 153 g rectangular solid has edge lengths 8.2 cm, 5.1 cm, and 4.7 cm. Will this object sink in water?

\[
V = LW \cdot H = 8.2 \text{ cm} \cdot (5.1 \text{ cm}) \cdot (4.7 \text{ cm}) = 197 \text{ cm}^3
\]

\[
D = \frac{m}{V} = \frac{153 \text{ g}}{197 \text{ cm}^3} = 0.78 \text{ g/cm}^3 \rightarrow \text{NO; IT FLOATS}
\]


Properties of Matter

**CHEMICAL** properties tell how a substance reacts with other substances.
**PHYSICAL** properties can be observed without chemically changing the substance.

**EXTENSIVE** properties depend on the amount of substance present.
**INTENSIVE** properties do not depend on the amount of substance.

Examples:
- electrical conductivity: \[ \text{P} / \text{I} \]
- reactivity with water: \[ \text{C} / \text{I} \]
- heat content (calories): \[ \text{C} / \text{E} \]
- ductile: can be drawn (pulled) into wire: \[ \text{P} / \text{I} \]
- malleable: can be hammered into shape: \[ \text{P} / \text{I} \]
- brittle: \[ \text{P} / \text{I} \]
- magnetism: \[ \text{P} / \text{I} \]
States of Matter

SOLID

LIQUID

GAS

Changes in State

Energy put into system.

>SOLID | LIQUID | GAS

- sublimation
- melting
- boiling

SOLID ← freezing

LIQUID ← condensation

GAS ← deposition

Energy removed from system.
Energy → the ability to do work

potential energy: stored energy
-- stored in bonds between atoms
e.g., in food, gasoline, batteries

kinetic energy: energy of motion \[ KE = \frac{1}{2} mv^2 \]
-- “hot” gas particles move faster, have more KE

Law of Conservation of Energy: \[ E_{after} = E_{before} \]

\[ 2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O} + \text{energy} \]

\[ \text{C}_2\text{H}_2 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} \]

---

Diagram showing energy, potential energy (PE) of reactants and products, and kinetic energy (KE) of products.
- **Energy Changes**

  **Endothermic Change**
  - System absorbs heat
  - -- Beaker feels cold

  **Exothermic Change**
  - System releases heat
  - -- Beaker feels hot

  - Water boiling
  - Paper burning
  - Steam condensing
  - CO$_2$ subliming
  - Water freezing
  - Ice melting

  ![Energy Levels Diagram](image)
The Mole

Atoms are so small, it is impossible to count them by the dozens, thousands, or even millions. To count atoms, we use the concept of the mole.

1 mole of atoms = \(602,000,000,000,000,000,000,000\) atoms

That is, 1 mole of atoms = \(6.02 \times 10^{23}\) atoms

- How Big is \(6.02 \times 10^{23}\)?

6.02 \(\times\) \(10^{23}\) marbles would cover the entire Earth (including the oceans) to a depth of 2 miles.

6.02 \(\times\) \(10^{23}\) $1$ bills stacked face-to-face would stretch from the Sun to Pluto and back 7.5 million times. It takes light 9,500 years to travel that far.

For any element on the Periodic Table, one mole of that element (i.e., \(6.02 \times 10^{23}\) atoms of that element) has a mass in grams equal to the decimal number on the Table for that element.
Island Diagram Problems

1. How many moles is $3.79 \times 10^{25}$ atoms of zinc?

\[ X \text{ mol Zn} = 3.79 \times 10^{25} \text{ at. Zn} \left( \frac{1 \text{ mol Zn}}{6.02 \times 10^{23} \text{ at. Zn}} \right) = 63.0 \text{ mol Zn} \]

2. How many atoms is 0.68 moles of zinc?

\[ X \text{ at. Zn} = 0.68 \text{ mol Zn} \left( \frac{6.02 \times 10^{23} \text{ at. Zn}}{1 \text{ mol Zn}} \right) = 4.1 \times 10^{23} \text{ at. Zn} \]

3. How many grams is 5.69 moles of uranium?

\[ X \text{ g U} = 5.69 \text{ mol U} \left( \frac{238 \text{ g U}}{1 \text{ mol U}} \right) = 1.35 \times 10^3 \text{ g U} \]

4. How many grams is $2.65 \times 10^{23}$ atoms of neon?

\[ X \text{ g Ne} = 2.65 \times 10^{23} \text{ at. Ne} \left( \frac{1 \text{ mol Ne}}{6.02 \times 10^{23} \text{ at. Ne}} \right) \left( \frac{20.2 \text{ g Ne}}{1 \text{ mol Ne}} \right) = 8.9 \text{ g Ne} \]

5. How many atoms is 421 g of promethium?

\[ X \text{ at. Pm} = 421 \text{ g Pm} \left( \frac{1 \text{ mol Pm}}{145 \text{ g Pm}} \right) \left( \frac{6.02 \times 10^{23} \text{ at. Pm}}{1 \text{ mol Pm}} \right) = 1.75 \times 10^{24} \text{ at. Pm} \]