THE ATOMIC THEORY

- The smallest particle of matter that still retains its properties is called an atom.

- In the fifth century B.C., the Greek philosopher Democritus proposed that matter is composed of a finite number of discrete particles, named atomos (meaning uncuttable or indivisible).

- In 1808, John Dalton, built on ideas of Democritus, and formulated a precise definition of the building blocks of matter.

- Dalton’s model represented the atom as a featureless ball of uniform density.

- This model is referred to as the “soccer ball” model.
Dalton’s Atomic Theory consisted of three parts:

- Each element is composed of tiny indestructible particles called atoms.
- All atoms of a given element are similar to one another, but different from atoms of other elements.
- Atoms combine in simple, whole-number ratios to form compounds.
In 1897, J.J. Thomson performed experiments with a cathode ray tube. Negatively charged particles from cathode were pulled towards positively charged plate, anode, and allowed to pass through and be detected on a fluorescent screen.

In absence of a magnetic field, the cathode rays were not deflected.

In presence of a magnetic and electric fields, the cathode rays were deflected towards the positive plate, indicating a negative charge.

The cathode rays were later named electrons.

Based on these findings, Thomson proposed an atomic model composed of negatively charged electrons embedded in a uniform positively charged sphere.

This model is called the “plum pudding” model.
In 1910, Ernest Rutherford carried out a number of experiments to further probe the nature of the atom. In these experiments he bombarded a thin sheet of gold foil with α-particles (large, positively charged) emitted from a radioactive source.

The majority of the particles were observed to pass through undeflected or slightly deflected. Some of the particles were observed to be deflected at large angles. Few of the particles were observed to be turned back towards the direction they came from. Thomson’s model of the atom could not account for these results.

Based on these observations, Rutherford proposed a model of the atom consisting of a small, massive positive center (nucleus), surrounded by electrons in mostly empty space.

This model is called the nuclear model.
The Structure of the Atom

- In 1932, James Chadwick discovered the existence of a second nuclear particle. This neutral particle was named neutron.

Current Model of the Atom:

- The atom is an electrically neutral spherical entity.
- It is composed of a positively charged center surrounded by negatively charged electrons.
- The electrons (e⁻) move rapidly through the atomic volume, held by the attractive forces to the nucleus.
- The atomic nucleus consists of positively charged protons (p⁺) and neutrally charged neutrons (n⁰).

- The modern atom consists of 3 subatomic particles:

<table>
<thead>
<tr>
<th>TABLE 4.1 Subatomic Particles</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass (kg)</td>
</tr>
<tr>
<td>----------</td>
</tr>
<tr>
<td>proton</td>
</tr>
<tr>
<td>neutron</td>
</tr>
<tr>
<td>electron</td>
</tr>
</tbody>
</table>

Mass Relationships in the Atom:

- The number of protons in an atom determines its identity, and is called atomic number (Z).
- In a neutral atom, the number of protons (+) are equal to the number of electrons (−).
- Almost all the mass of the atom rests in the nucleus.
- The number of protons and neutrons in an atom is called the mass number (A).
- The number of neutrons in an atom can be determined by the difference between mass number (A) and atomic number (Z):

\[
\text{# of neutrons} = \text{mass number} - \text{atomic number} \\
\text{# of } n^0 = A - Z
\]
The number of **protons** in an atom identify it as a particular element, and is called **atomic number**.

Elements are arranged on the periodic table according to their atomic number.

Over time some elements have been named for planets, mythological figures, minerals, colors, geographic locations and famous people. Some examples are shown below:

<table>
<thead>
<tr>
<th>Element</th>
<th>Source of Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Uranium</td>
<td>The planet Uranus</td>
</tr>
<tr>
<td>Titanium</td>
<td>Titans (mythology)</td>
</tr>
<tr>
<td>Chlorine</td>
<td><em>Chloros</em>, “greenish-yellow” (Greek)</td>
</tr>
<tr>
<td>Iodine</td>
<td><em>Ioeides</em>, “violet” (Greek)</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Magnesia, a mineral</td>
</tr>
<tr>
<td>Californium</td>
<td>California</td>
</tr>
<tr>
<td>Curium</td>
<td>Marie and Pierre Curie</td>
</tr>
</tbody>
</table>

The **symbol** for most elements is the one- or two-letter abbreviation of the name of the element. Only the first letter of an element's symbol is capitalized. If the symbol has a second letter, it is written as lowercase.

Co (cobalt)  
CO (carbon and oxygen)

Although most of the symbols use letters from current names, some of the symbols of the elements are based on their Latin names.

K potassium (kalium)  
Na sodium (natrium)  
Fe iron (ferrum)

Some elements have formulas that are not single atoms. **Seven** of these elements have **diatomic** (2-atoms) molecules.

<table>
<thead>
<tr>
<th>Element</th>
<th>Formula</th>
<th>Element</th>
<th>Formulas</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>H₂</td>
<td>Chlorine</td>
<td>Cl₂</td>
</tr>
<tr>
<td>Oxygen</td>
<td>O₂</td>
<td>Fluorine</td>
<td>F₂</td>
</tr>
<tr>
<td>Nitrogen</td>
<td>N₂</td>
<td>Bromine</td>
<td>Br₂</td>
</tr>
<tr>
<td></td>
<td></td>
<td>Iodine</td>
<td>I₂</td>
</tr>
</tbody>
</table>
**ELEMENTS & THE PERIODIC TABLE**

- Arrangement of elements based on their atomic masses was first proposed by the Russian chemist, **Dmitri Mendeleev** in 1869.
- The elements are generally classified as **metals, nonmetals** and **metalloids**.

<table>
<thead>
<tr>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
<th>8A</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>Be</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>He</td>
</tr>
<tr>
<td>Li</td>
<td>B</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Ne</td>
</tr>
<tr>
<td>Na</td>
<td>Mg</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Ar</td>
</tr>
<tr>
<td>K</td>
<td>Ca</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Kr</td>
</tr>
<tr>
<td>Rb</td>
<td>Sr</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Xe</td>
</tr>
<tr>
<td>Cs</td>
<td>Ba</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>Rn</td>
</tr>
<tr>
<td>Fr</td>
<td>Ra</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td>**</td>
</tr>
</tbody>
</table>

**Metals**
- Mostly solid
- Have shiny appearance
- Good conductors of heat and electricity
- Are malleable and ductile
- Lose electrons in a chemical reaction

**Nonmetals**
- Can be solid, liquid or gas
- Have dull appearance
- Poor conductors of heat and electricity
- Are brittle (if solid)
- Gain or share electrons in a chemical reaction

**Metalloids** are elements that possess some properties of metals and some of non-metals. The most important metalloids are silicon (Si) and germanium (Ge) which are used extensively in computer chips.
The periodic table is composed of **periods** (rows) and **groups or families** (columns).

Elements in the same family have similar properties, and are commonly referred to by their traditional names.

- Elements in groups **1-2** and **13-18** are referred to as **main-group or representative elements**.
- Group 1 elements are called **alkali metal**. These are **soft** metals that are **very reactive**, and often react explosively with other elements.
- Group 8 elements are called **noble gases**. These are **un-reactive** gases that are commonly used in light bulbs.
- Group 7 elements are called **halogens**. These are the most **reactive nonmetals**, and occur in nature only as compounds.
- Group 2 elements are called **alkaline-earth metals**. These metals are less reactive than alkali metals.
- The group of metals in-between the main group elements are called the **transition metals**.
When reacting, atoms often lose or gain electrons to form charged particles, called ions.

For example, when reacting, lithium atom (3 protons and 3 electrons) loses one electron to form a Li⁺ ion:

\[ \text{Li} \rightarrow \text{Li}^+ + e^- \]

The charge of an ion is determined by the imbalance of the number of protons and electrons in the atom. Therefore, for lithium ion, the charge is determined as follows:

\[ \text{Ion charge} = \text{number of protons} - \text{number of electrons} \]
\[ \text{Ion charge} = 3 - 2 = 1^+ \]

Similarly, when reacting, fluorine atom (9 protons and 9 electrons) gains an electron to form a F⁻ ion:

\[ \text{F} + e^- \rightarrow \text{F}^- \]

\[ \text{Ion charge} = \text{number of protons} - \text{number of electrons} \]
\[ \text{Ion charge} = 9 - 10 = 1^- \]

Positive ions are called cations. Cations are formed when a neutral atom loses electrons.

Negative ions are called anions. Anions are formed when a neutral atom gains electrons.

In nature, cations and anions occur together so that matter remains charge-neutral. For example, in table salt, sodium cations (Na⁺) occur together with chloride anions (Cl⁻).

For most main-group elements, the periodic table can be used to predict the number of electrons lost or gained to form ions.
**ISOTOPES & ATOMIC MASS**

- Atoms of the *same element* that possess *different* number of *neutrons* are called *isotopes*.

- *Isotopes* of an element have the *same atomic number* \( (Z) \), but *different mass number* \( (A) \).

\[
\begin{array}{ccc}
\text{Ne-20} & \text{Ne-21} & \text{Ne-22} \\
(90.48\%) & (0.27\%) & (9.25\%)
\end{array}
\]

- For a given element, the relative amount of each isotope in a naturally occurring sample of that element is always the same.

- The *mass* of an atom is measured *relative* to the mass of a chosen *standard* \( (\text{carbon-12 atom}) \), and is expressed in *atomic mass units* \( (\text{amu}) \).

- The *average atomic mass* of an element—called *atomic mass*—is the *mass* of that element’s natural occurring *isotopes weighted* according to their *abundance*. Therefore the atomic mass of an element is closest to the mass of its most abundant isotope.

**Examples:**

1. Determine the number of protons, neutrons and electrons in \( ^{35} \text{Cl} \).

   \[
   \begin{align*}
   \text{number of } p^+ &= \\
   \text{number of } e^- &= \\
   \text{number of } n^0 &= 
   \end{align*}
   \]

2. Which two of the following are isotopes of each other?

\[
\begin{array}{cccc}
410 & X & 410 & Y & 412 & Z & 412 & R \\
186 & 185 & 183 & 185
\end{array}
\]

3. Based on the information below, which is the most abundant isotope of boron (atomic mass = 10.8 u)?

<table>
<thead>
<tr>
<th>Isotope</th>
<th>(^{10}\text{B})</th>
<th>(^{11}\text{B})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass (amu)</td>
<td>10.0</td>
<td>11.0</td>
</tr>
</tbody>
</table>
CALCULATING ATOMIC MASSES FROM ISOTOPIC DATA

- The atomic mass of an element can be calculated from the mass and abundance of its isotopes as shown below:

\[
\text{Atomic mass of an element} = \left( \frac{\text{Abundance of isotope 1}}{\text{Mass of isotope 1}} \right) + \left( \frac{\text{Abundance of isotope 2}}{\text{Mass of isotope 2}} \right)
\]

**Examples:**

1. Calculate the average atomic mass of silver from the following isotopic data:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (amu)</th>
<th>Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(^{107}\text{Ag})</td>
<td>106.90509</td>
<td>51.84</td>
</tr>
<tr>
<td>(^{109}\text{Ag})</td>
<td>108.90476</td>
<td>48.16</td>
</tr>
</tbody>
</table>

\(^{107}\text{Ag}=106.90509 \times (0.5184)= 55.42 \text{ amu}\)

\(^{109}\text{Ag}=108.90476 \times (0.4816)= 52.45 \text{ amu}\)

Atomic mass of Ag = \text{55.42 amu} + \text{52.45 amu} = 107.87 amu

2. Calculate the average atomic mass of gallium from the following isotopic data:

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (amu)</th>
<th>Abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ga-69</td>
<td>68.9256</td>
<td>60.11</td>
</tr>
<tr>
<td>Ga-71</td>
<td>70.9247</td>
<td>39.89</td>
</tr>
</tbody>
</table>

3. A fictitious element is composed of isotopes A and B with masses of 61.9887 and 64.9486 amu, respectfully. The atomic mass of the element is 64.52 amu. What can you conclude about the natural abundances of the two isotopes? Briefly explain your answer.