Section 4.1
Development of Atomic Theory

• Who came up with the first theory of atoms?
• What did Dalton add to the atomic theory?
• How did Thomson discover the electron?
• What is Rutherford’s atomic model?

“Elements of chemistry- Atoms, the building blocks of matter” Video
In this lesson we will take a look at the scientists who explored the wonders of the atom.

- **Aristotle (384-322 BC)**
  - Aristotle denounced the idea of the atom in favor of the classical elements: *fire, earth, air and water*.
  - Aristotle believed that all matter could be broken down to one of the classical elements.
  - The *atomos* idea was lost for approximately 2000 years.
The Greek philosopher Searched for a description of matter more than 2400 years ago.

He asked: Could matter be divided into smaller and smaller pieces forever, or was there a limit to the number of times a piece of matter could be divided?

Democritus (400 BC)

- Atomos Theory:
  - Atoms were small
  - Hard particles that were all made of the same material but were different shapes and sizes.
  - Came up with the idea that atoms exist.
    - He named the smallest piece of matter “atomos,” meaning “not to be cut.”

- No evidence to support idea
Dalton’s Model (1803)

- In the early 1800s, the English Chemist John Dalton
- Performed a number of experiments that eventually led to the acceptance of the idea of atoms.

Dalton’s Atomic Theory

1. All matter is composed of extremely small particles called atoms.
2. All atoms of a given element are identical in size, mass, and other properties.
3. Atoms cannot be divided, created, or destroyed.
4. Atoms of different elements can combine in simple, whole-number ratios to form compounds.
5. In chemical reactions, atoms are combined, separated, or rearranged.
Thomson's Model (1897)

- Provided the first hint that an atom is made of even smaller particles.
  - They are divisible
- Studied the passage of an electric current through a gas.
- As the current passed through the gas, it gave off rays of negatively charged particles.

Thomson’s Cathode-Ray Tube Experiment

By applying a magnet to the cathode ray tube and seeing the ray deflect, he determined the ray was negatively charged.
- He had discovered electrons.
Thomson’s Plum Pudding Model

- Atoms were made from a positively charged substance with negatively charged electrons scattered about, like raisins in a pudding.
- Thomson called the negatively charged “corpuscles,” today known as electrons.

Since the gas was known to be neutral, he reasoned that there must be positively charged particles in the atom.
Rutherford’s model

- Tested Thomson’s model.
- Most of the mass of an atom is concentrated at the atom’s center (Nucleus).

The gold foil experiment
- Aim positively charged alpha particles at thin gold foil.
- Prediction: most particles would travel straight through because Thomson’s model showed the positive and negative particles evenly distributed.

Rutherford’s Gold Foil Experiment

- Rutherford’s experiment
  Involved firing a stream of tiny positively charged particles at a thin sheet of gold foil (2000 atoms thick)
• Experiment did not match prediction.
• Most particles did pass straight through, but some were greatly deflected.
• Rutherford discovered the nucleus!

“It was almost as incredible as if you fired a 15 inch shell at a piece of tissue paper and it came back and hit you”.

“What does the inside of an atom look like?”

Small, positive particle (now we call them “protons”)

Tiny pieces of gold (atoms)
“What does the inside of an atom look like?”

Watch what happens when we shoot the particles at the gold atoms.

Observation:
Most of the “+” particles passed through. But a few also bounced back.

What was inside these “atoms?”
“What does the inside of an atom look like?”

\[ \text{Conclusion: An atom contains mostly empty space with a small, heavy nucleus with “+” charges.} \]
Rutherford's model (1911)

- Concluded that an atom had a small, dense, positively charged center that repelled his positively charged “bullets.”

What did he discover?

Nucleus

Rutherford's model (1911)

- The nucleus is tiny compared to the atom as a whole.
- All of an atom’s positively charged particles were contained in the nucleus.
- Negatively charged particles were scattered outside the nucleus around the atom’s edge.

If the nucleus of an atom were the size of a marble, its electron cloud would be the size of a football stadium!
In 1913, the Danish scientist Niels Bohr proposed an improvement. He placed electron in a specific energy level.

Electrons move in definite orbits around the nucleus.
- Like planets circle the sun.

These orbits, or energy levels, are located at certain distances from the nucleus.
Today's atomic model is based on the principles of wave mechanics. Electrons do not move about an atom in a definite path, like the planets around the sun.

It is impossible to determine the exact location of an electron. The probable location of an electron is based on how much energy the electron has.

Progress of the Atom

<table>
<thead>
<tr>
<th></th>
<th>Indivisible</th>
<th>Electron</th>
<th>Nucleus</th>
<th>Orbit</th>
<th>Electron Cloud</th>
</tr>
</thead>
<tbody>
<tr>
<td>Democritus</td>
<td>X</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Dalton</td>
<td>X</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Thomson</td>
<td></td>
<td>X</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Rutherford</td>
<td>X</td>
<td>X</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Bohr</td>
<td></td>
<td>X</td>
<td>X</td>
<td></td>
<td>X</td>
</tr>
<tr>
<td>Wave</td>
<td>X</td>
<td>X</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Timeline of Progress of the Atom

Greek Model
400 BC
Democritus

Aristotle's 4 elements

Dalton Model
1803

Thomson Model
1897

Rutherford Model
1911

Bohr Model
1922

Wave Model
Modern

Section 4.2: Structure of Atoms

• What is the difference between protons, neutrons, and electrons?
• What do atoms of an element have in common with other atoms of the same element?
• Why do isotopes of the same element have different atomic masses?
• What unit is used to express atomic mass?
What is in an Atom?

- Atoms are the building blocks of everything.
- Atoms can be divided, but not easily on Earth.

118 different types of Atoms

If an apple is magnified to the size of the Earth, the atoms in the apple are approximately the size of the original apple.

- R. Feynmann
### Parts of the Atom

**What is in an Atom?**

<table>
<thead>
<tr>
<th>Particle</th>
<th>Charge</th>
<th>Mass (kg)</th>
<th>Location</th>
</tr>
</thead>
<tbody>
<tr>
<td>Proton</td>
<td>+1</td>
<td>$1.67 \times 10^{-27}$</td>
<td>In the nucleus</td>
</tr>
<tr>
<td>Neutron</td>
<td>0</td>
<td>$1.67 \times 10^{-27}$</td>
<td>In the nucleus</td>
</tr>
<tr>
<td>Electron</td>
<td>-1</td>
<td>$9.11 \times 10^{-31}$</td>
<td>Outside the nucleus</td>
</tr>
</tbody>
</table>

- Contains: Proton, Electron, and Neutron
- Proton and neutron are 2000 times heavier than electron
- Mostly empty space
- No overall charge
**Protons**
- An element is defined by the number of protons
- NO TWO ELEMENTS WILL HAVE THE SAME NUMBER OF PROTONS.
- Positive (+) charge
- Located in the nucleus

**Electrons**
- Electron have a negative (-) charge
- Located outside the nucleus
- The electric force between protons in the nucleus and electrons outside the nucleus holds the atom together.
Neutrons

- Neutrons have no charge (0) and are neutral.
- Neutrons contribute to the overall mass of the atom.

Nucleus

- Contains the protons and neutrons
- Where most of the mass of the atom is concentrated
- Has an overall positive (+) charge
Atomic Number and Mass Number

- Atomic number (Z) = the number of protons
  - This is also the number of electrons if the atom is neutral.
  - No two elements share the same atomic number or number of protons

Carbon: Z = 6
Hydrogen: Z = 1
Uranium: Z = 92
• Mass number (A) = the total number of subatomic particles in the nucleus, that is the sum of the protons and the neutrons.
  – Mass number is denoted by \( A \)
  – Example: A fluorine atom has 9 protons and 10 neutrons, so \( A = 19 \) for fluorine.

Atomic Number and Mass Number

Carbon: \( A = \boxed{12} \)
Chlorine: \( A = \boxed{35} \)
Aluminum: \( A = \boxed{27} \)

Finding # of Neutrons

We can find the number of neutrons by subtracting the Atomic Number from the Mass Number.

\[
\text{Fluorine (F) has } 10 \text{ neutrons. } 19 - 9 = 10 \\
\text{Aluminum (Al) has } 14 \text{ neutrons. } 27 - 13 = 14
\]
Identify the following two atoms:

- **Hydrogen**
- **Helium**

### Atoms Components:
- Proton (+)
- Neutron
- Electron (-)

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Try This:

**Sodium atom**

- Nucleus
- Electrons in various orbitals
Atomic mass

- **Average atomic mass (amu)** is the weighted average of the masses of all the commonly occurring isotopes of an element.
- A **unified atomic mass unit (u)** equals 1/12 the mass of a carbon-12 atom.

<table>
<thead>
<tr>
<th>Element</th>
<th>Atomic Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Ni</td>
<td>58.6934</td>
</tr>
<tr>
<td>Fe</td>
<td>55.845</td>
</tr>
<tr>
<td>K</td>
<td>39.0983</td>
</tr>
</tbody>
</table>

Isotopes

- An isotope is an atom that has the same number of protons but a different number of neutrons (relative to other atoms of the element).
- They vary in mass and mass number (A).
- \( A - Z = n \)
Isotopes
Same element, but
different mass due to
different number of neutrons.

Carbon-12
6 PROTONS
6 NEUTRONS

Carbon-14
6 PROTONS
8 NEUTRONS

Isotopes
CHANGE IN WEIGHT, BUT
NOT CHARGE

TWO ISOTOPES OF HYDROGEN.
ON THE LEFT, ONE NEUTRON
ON THE RIGHT, THERE ARE TWO
Molar mass: the mass, in grams, of one mole of a substance.

- Avogadro’s number: $6.022 \times 10^{23}$
- This is defined as the number of particles in 12 g of Carbon-12.
- Avogadro’s number is not useful for larger objects...

The Mole!

$6.0221367 \times 10^{23}$ particles!
Mass for 1 mole of atoms

The average atomic mass = grams for 1 mole

Average atomic mass is found on the periodic table

<table>
<thead>
<tr>
<th>Element</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 mole of carbon atoms</td>
<td>12.01 g</td>
</tr>
<tr>
<td>1 mole of oxygen atoms</td>
<td>16.00 g</td>
</tr>
<tr>
<td>1 mole of hydrogen atoms</td>
<td>1.01 g</td>
</tr>
</tbody>
</table>

Unit for molar mass: g/mole or g/mol
Example: Molar Mass

Example:
Find the molar mass for CaBr₂

Ca 1 × 40.08 g/mole = 40.08 g/mole
Br 2 × 79.91 g/mole = 159.82 g/mole

1 mole of CaBr₂ molecules would have a mass of 199.90 g

Example: Molar Mass

Example:
Find the molar mass for Sr(NO₃)₂

Sr 1 × 87.62 g/mole = 87.62 g/mole
N 2 × 14.01 g/mole = 28.02 g/mole
O 6 × 16.00 g/mole = 96.00 g/mole

1 mole of Sr(NO₃)₂ molecules would have a mass of 211.64 g
Converting Between Moles & Grams

• You can convert between moles and grams.

Use \( \frac{g}{mol} \) conversion factor.

Section 4.3: Modern Atomic Theory

• What is the modern model of the atom?
• How are the energy levels of an atom filled?
• What makes an electron jump to a new energy level?
Joke for the Day

- Why shouldn’t you trust an atom?
  - They make up everything.

4.3 Modern Atomic Theory

- Electrons can be found only in certain energy levels, not between levels.
- Electron location (not precise) is limited to energy levels.
- Electron act like waves.
- The whole shaded region is called an electron cloud.
Energy Level

• Describe the path the electron takes around the nucleus
• Farther from nucleus is more energy
• Gain energy they move away (absorbing photon)
• Lose energy they move toward (release photon)
• Only certain energies are allowed in each atom

Energy Levels

- Like an elevator
- it can only be on certain floors
- Never in between
- Energy levels get closer together the higher you go
- Each has room for a certain number of electrons
Electron structure

Consider an atom of Potassium

Electron structure = 2, 8, 8, 1

• The inner shell has __ electrons
• The next shell has __ electrons
• The next shell has __ electrons
• The next shell has the remaining __ electron

Orbitals

• Regions where you have a chance of finding the electron
• There are different types of orbitals – s, p, d, f
• Each has its own shape or shapes
• Each shape has room for two electrons
• Each can be found in the energy levels
S orbital

- Shaped like a sphere
- One possible orientation in space
- Maximum of 2 electrons

P orbital

- 3 dumbbell-shaped regions
- Each p orbital can hold a maximum of 2 electrons, so all three together can hold a total of 6 electrons.
d orbital

- Five different shapes
- More complex
- Each can hold 2 electrons
- Total of 10 electrons

f orbital

- Seven different shapes
- Much more complex
- Each can hold 2 electrons
- Total of 14 electrons
Orbitals

- Orbitals determine the number of electrons that each level can hold.
- Fill up orbitals in the order 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p - until you run out of electrons

<table>
<thead>
<tr>
<th>Energy level</th>
<th>Number of orbitals by type</th>
<th>Total number of orbitals</th>
<th>Number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>1 s 1 p 1 d 1 f</td>
<td>1 = 1</td>
<td>2</td>
</tr>
<tr>
<td>2</td>
<td>1 s 3 p 1 d 1 f</td>
<td>1 + 3 = 4</td>
<td>8</td>
</tr>
<tr>
<td>3</td>
<td>1 s 1 p 3 d 5 f</td>
<td>1 + 3 + 5 = 9</td>
<td>18</td>
</tr>
<tr>
<td>4</td>
<td>1 s 1 p 3 d 5 7 f</td>
<td>1 + 3 + 5 + 7 = 16</td>
<td>32</td>
</tr>
</tbody>
</table>

- 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p
Valence electrons

- The electrons in the outermost energy level
- Responsible for most of the chemical properties
- When two atoms interact, the outside electrons are the ones affected
- Can be determined by counting to the group the element is in, skipping the transition elements.
How many Valence Electrons?

8

- Identify the valence electrons.
Atoms are stable if the “valence” shell is full with 8 electrons in it.

Electron Transitions

• Electrons jump between energy levels when an atom gains or loses energy.
• Ground state- lowest state of energy of an electron.
• Excited state- electron gains energy by absorbing a particle of light, called a photon. Electron can jump “up” an energy level.
How Atoms Emit Light

1. A collision with a moving particle excites the atom.
2. This causes an electron to jump to a higher energy level.
3. The electron falls back to its original energy level, releasing the extra energy in the form of a light photon.