Chapter 5
“Electrons in Atoms”

Section 5.1
Models of the Atom

OBJECTIVES:
• Identify the inadequacies in the Rutherford atomic model.

Section 5.1
Models of the Atom

OBJECTIVES:
• Describe the energies and positions of electrons according to the quantum mechanical model.

Section 5.1
Models of the Atom

OBJECTIVES:
• Describe how the shapes of orbitals related to different sublevels differ.

Ernest Rutherford’s Model

• Discovered dense positive piece at the center of the atom- “nucleus”
• Electrons would surround and move around it, like planets around the sun
• Atom is mostly empty space
• It did not explain the chemical properties of the elements...
**Niels Bohr’s Model**
- Why don’t the electrons fall into the nucleus?
- Move like planets around the sun.
  - In specific circular paths, or orbits, at different levels.
  - An amount of fixed energy separates one level from another.

**Bohr’s model**
- **Energy level** of an electron
  - analogous to the rungs of a ladder
- The electron cannot exist between energy levels, just like you can’t stand between rungs on a ladder
- A **quantum** of energy is the amount of energy required to move an electron from one energy level to another

**The Bohr Model of the Atom**
I pictured the electrons orbiting the nucleus much like planets orbiting the sun. However, electrons are found in specific circular paths around the nucleus, and can jump from one level to another.

**The Quantum Mechanical Model**
- Energy is quantized - it comes in chunks.
- A **quantum** is the amount of energy needed to move from one energy level to another.
- Since the energy of an atom is never "in between" there must be a quantum leap in energy.
- In 1926, Erwin Schrödinger derived an **equation** that described the energy and position of the electrons in an atom.

**The Quantum Mechanical Model**
- Things that are very small behave differently from things big enough to see.
- The **quantum mechanical model** is a mathematical solution.
- It is not like anything you can see.

**Equation for the probability of a single electron being found along a single axis (x-axis)**

\[ -\frac{\hbar^2}{8\pi m} \frac{d^2 \psi}{dx^2} + V(x) = E \psi \]
**The Quantum Mechanical Model**

- Has energy levels for electrons.
- Orbits are not circular.
- It can only tell us the **probability** of finding an electron a certain distance from the nucleus.

**Atomic Orbitals**

- Principal Quantum Number (n) = the energy level of the electron: 1, 2, 3, etc.
- Within each energy level, the complex math of Schrödinger’s equation describes several shapes.
- These are called **atomic orbitals** - regions where there is a high probability of finding an electron.
- Sublevels- like theater seats arranged in sections: letters s, p, d, and f

**Principal Quantum Number**

Generally symbolized by “n”, it denotes the shell (energy level) in which the electron is located.

Maximum number of electrons that can fit in an energy level:

$$2n^2$$

### Summary

<table>
<thead>
<tr>
<th># of shapes</th>
<th>Max electrons</th>
<th>Starts at energy level</th>
</tr>
</thead>
<tbody>
<tr>
<td>s</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>p</td>
<td>3</td>
<td>6</td>
</tr>
<tr>
<td>d</td>
<td>5</td>
<td>10</td>
</tr>
<tr>
<td>f</td>
<td>7</td>
<td>14</td>
</tr>
</tbody>
</table>

### By Energy Level

- **First Energy Level**
  - Has only s orbital
  - only 2 electrons
  - $$1s^2$$

- **Second Energy Level**
  - Has s and p orbitals available
  - 2 in s, 6 in p
  - $$2s^22p^6$$
  - 8 total electrons
By Energy Level
- Third energy level
- Has s, p, and d orbitals
- 2 in s, 6 in p, and 10 in d
- $3s^23p^33d^6$
- 18 total electrons
- Fourth energy level
- Has s, p, d, and f orbitals
- 2 in s, 6 in p, 10 in d, and 14 in f
- $4s^24p^44d^64f^4$
- 32 total electrons

By Energy Level
- Any more than the fourth and not all the orbitals will fill up.
- You simply run out of electrons
- The orbitals do not fill up in a neat order.
- The energy levels overlap
- Lowest energy fill first.

Section 5.2 Electron Arrangement in Atoms
- OBJECTIVES:
  - Describe how to write the electron configuration for an atom.

Section 5.2 Electron Arrangement in Atoms
- OBJECTIVES:
  - Explain why the actual electron configurations for some elements differ from those predicted by the aufbau principle.

Electron Configurations...
- ...are the way electrons are arranged in various orbitals around the nuclei of atoms. Three rules tell us how:
  - Aufbau principle - electrons enter the lowest energy first.
    - This causes difficulties because of the overlap of orbitals of different energies – follow the diagram!
  - Pauli Exclusion Principle - at most 2 electrons per orbital - different spins
Pauli Exclusion Principle
No two electrons in an atom can have the same four quantum numbers.

To show the different direction of spin, a pair in the same orbital is written as:

Quantum Numbers
Each electron in an atom has a unique set of 4 quantum numbers which describe it.
1) Principal quantum number
2) Angular momentum quantum number
3) Magnetic quantum number
4) Spin quantum number

Electron Configurations
• **Hund’s Rule** - When electrons occupy orbitals of equal energy, they don’t pair up until they have to.

  • Let’s write the electron configuration for Phosphorus
    • We need to account for all 15 electrons in phosphorus

  ![Electron Configuration Diagram]

  - The next electrons go into the 2s orbital
  - only 11 more...

  - The first two electrons go into the 1s orbital
  - Notice the opposite direction of the spins
  - only 13 more to go...

  ![Electron Configuration Diagram]

  - The next electrons go into the 2p orbital
  - only 5 more...
Orbitals fill in an order:
- Lowest energy to higher energy.
- Adding electrons can change the energy of the orbital. Full orbitals are the absolute best situation.
- However, half filled orbitals have a lower energy, and are next best
  • Makes them more stable.
  • Changes the filling order

Write the electron configurations for these elements:
- Titanium - 22 electrons
  • $1s^22s^22p^63s^23p^64s^23d^2$
- Vanadium - 23 electrons
  • $1s^22s^22p^63s^23p^64s^23d^3$
- Chromium - 24 electrons
  • $1s^22s^22p^63s^23p^64s^23d^1$ (expected)
  • But this is not what happens!!

Chromium is actually:
- $1s^22s^22p^63s^23p^64s^23d^6$
- Why?
  • This gives us two half filled orbitals (the others are all still full)
  • Half full is slightly lower in energy.
  • The same principal applies to copper.

Copper's electron configuration:
- Copper has 29 electrons so we expect: $1s^22s^22p^63s^23p^64s^23d^6$
- But the actual configuration is:
  • $1s^22s^22p^63s^23p^64s^23d^6$
  • This change gives one more filled orbital and one that is half filled.
- Remember these exceptions: $d^4$, $d^6$
Section 5.3
Physics and the Quantum Mechanical Model

**OBJECTIVES:**
- **Describe** the relationship between the wavelength and frequency of light.
- **Identify** the source of atomic emission spectra.
- **Explain** how the frequencies of emitted light are related to changes in electron energies.
- **Distinguish** between quantum mechanics and classical mechanics.

**Light**
- The study of light led to the development of the quantum mechanical model.
- Light is a kind of electromagnetic radiation.
- Electromagnetic radiation includes many types: gamma rays, x-rays, radio waves...
- **Speed of light** = $2.998 \times 10^8$ m/s, and is abbreviated “c”.
- *All electromagnetic radiation* travels at this same rate when measured in a vacuum.
**Figure 5.10 Electromagnetic Spectrum - Page 139**

**Parts of a wave**

- Crest
- Wavelength
- Origin
- Amplitude
- Trough

**Electromagnetic radiation propagates through space as a wave moving at the speed of light.**

**Equation:**

\[
\frac{c}{\nu} = \text{const.}
\]

- \(c\) = speed of light, a constant (2.998 \times 10^8 m/s)
- \(\lambda\) (lambda) = wavelength, in meters
- \(\nu\) (nu) = frequency, in units of hertz (Hz or sec\(^{-1}\))

**Wavelength and Frequency**

- Are inversely related
  - As one goes up the other goes down.
- Different frequencies of light are *different colors* of light.
- There is a wide variety of frequencies
- The whole range is called a spectrum

**Sample Problem 5.1 - Page 140**

**SAMPLE PROBLEM 5.1**

*Calculating the Wavelength of Light*

Calculate the wavelength of the yellow light emitted by the sodium lamp shown above if the frequency of the radiation is 5.10 \times 10^{14} Hz or 5.10 x 10^{14} Hz.

**Analyze** List the knowns and the unknowns.

- Knowns
  - Frequency \(\nu = 5.10 \times 10^{14}\) Hz or \(5.10 \times 10^{14}\) Hz
  - Speed of light \(c = 2.998 \times 10^8\) m/s
- Unknowns
  - Wavelength \(\lambda\)

**Equation**

\[
\frac{c}{\nu} = \lambda
\]

**Solve**

\[
\lambda = \frac{c}{\nu} = \frac{2.998 \times 10^8}{5.10 \times 10^{14}} = 5.85 \times 10^{-7}\text{ m}
\]

**Conclusion**

The wavelength of the yellow light is 5.85 \times 10^{-7} m.
Atomic Spectra
- White light is made up of all the colors of the visible spectrum.
- Passing it through a prism separates it.

If the light is not white
- By heating a gas with electricity we can get it to give off colors.
- Passing this light through a prism does something different.

Atomic Spectrum
- Each element gives off its own characteristic colors.
- Can be used to identify the atom.
- This is how we know what stars are made of.

Light is a Particle?
- Energy is quantized.
- Light is a form of energy.
- Therefore, light must be quantized.
- These smallest pieces of light are called photons.
- Photoelectric effect? Albert Einstein
- Energy & frequency: directly related.

- These are called the atomic emission spectrum
- Unique to each element, like fingerprints!
- Very useful for identifying elements
The energy ($E$) of electromagnetic radiation is directly proportional to the frequency ($\nu$) of the radiation.

Equation: $E = h \nu$

- $E$ = Energy, in units of Joules (kg⋅m$^2$/s$^2$)  
  (Joule is the metric unit of energy)
- $h$ = Planck's constant ($6.626 \times 10^{-34}$ J⋅s)
- $\nu$ = frequency, in units of hertz (Hz, sec$^{-1}$)

Examples
- What is the wavelength of blue light with a frequency of $8.3 \times 10^6$ Hz?
- What is the frequency of red light with a wavelength of $4.2 \times 10^5$ m?
- What is the energy of a photon of each of the above?

Explanation of atomic spectra
- When we write electron configurations, we are writing the lowest energy.
- The energy level, and where the electron starts from, is called its ground state - the lowest energy level.

Changing the energy
- Let's look at a hydrogen atom, with only one electron, and in the first energy level.

Changing the energy
- Heat, electricity, or light can move the electron up to different energy levels. The electron is now said to be "excited"
Changing the energy

- As the electron falls back to the ground state, it gives the energy back as light.

Changing the energy

- They may fall down in specific steps.
- Each step has a different energy.

Click to add an outline.

- Ultraviolet
- Visible
- Infrared

- The further they fall, more energy is released and the higher the frequency.
- This is a simplified explanation!
- The orbitals also have different energies inside energy levels.
- All the electrons can move around.

What is light?

- Light is a particle - it comes in chunks.
- Light is a wave - we can measure its wavelength and it behaves as a wave.
- If we combine $E=mc^2$, $c=\lambda \cdot v$, $E = 1/2 \cdot mv^2$ and $E = h\cdot v$ then we can get:
  
  $\lambda = h/mv$ (from Louis de Broglie)

- called de Broglie's equation
- Calculates the wavelength of a particle.

Wave-Particle Duality

J.J. Thomson won the Nobel prize for describing the electron as a particle.

His son, George Thomson won the Nobel prize for describing the wave-like nature of the electron.

The electron is a particle!

The electron is an energy wave!
Confused? You've Got Company!

"No familiar conceptions can be woven around the electron; something unknown is doing we don't know what."

Physicist Sir Arthur Eddington
The Nature of the Physical World
1934

The physics of the very small

- Quantum mechanics explains how very small particles behave
- Quantum mechanics is an explanation for subatomic particles and atoms as waves
- Classical mechanics describes the motions of bodies much larger than atoms

Heisenberg Uncertainty Principle

- It is impossible to know exactly the location and velocity of a particle.
- The better we know one, the less we know the other.
- Measuring changes the properties.
- True in quantum mechanics, but not classical mechanics

Heisenberg Uncertainty Principle

"One cannot simultaneously determine both the position and momentum of an electron."

You can find out where the electron is, but not where it is going.
OR...
You can find out where the electron is going, but not where it is!

Werner Heisenberg

It is more obvious with the very small objects

- To measure where a electron is, we use light.
- But the light energy moves the electron
- And hitting the electron changes the frequency of the light.

Before

Photon

Moving Electron

After

Photon wavelength changes

Electron velocity changes

Fig. 5.16, p. 145
End of Chapter 5