Arrangement of Electrons

Chapter 4
Properties of Light

- Light’s interaction with matter helps to understand how electrons behave in atoms
- Light travels through space & is a form of electromagnetic radiation
- Light has properties of waves & particles
- Light travels in waves
- Other examples of electromagnetic radiation are x-rays, gamma rays, & radio waves
- Electromagnetic waves consist of electric & magnetic fields oscillating at right angles to each other & to the direction of motion the wave
- Require no medium to travel through
Waves have 4 characteristics:
- Wavelength
- Frequency
- Speed
- Amplitude
Wavelength ($\lambda$): distance between 2 consecutive peaks. Measure in meters.

Frequency ($\nu$): how fast wave oscillates up & down. It is measured by the number of times a light wave completes a cycle up & down motion in one second. Measured in hertz $1$ hertz = $1$ cycle/second = $1$ s$^{-1}$
Amplitude: the height of the wave. Measured from the middle of the wave.
WAVES

- Speed (c): all electromagnetic radiation travels at the speed of light.
- Speed of light = \(3.00 \times 10^8\) m/s
- All electromagnetic radiation travels at this same rate when measured in a vacuum
- Since light moves at a constant speed, there is a relationship between wavelength & frequency
Waves and Frequency

- Waves & frequency are inversely related
- Different frequencies of light are different colors of light.
- The longer the wavelength, the shorter the frequency and low energy.

\[ \lambda \nu = c \]  

- \( \lambda \) = wavelength  
- \( \nu \) = frequency  
- \( c \) = speed of light
EXAMPLES

1. Yellow light has a wavelength of 589 nm. What is the frequency?

2. The US Navy system uses radio waves with a frequency of 76 s\(^{-1}\). What is the wavelength in meters?
Long Wavelengths = Low Frequency = Low ENERGY

Short Wavelengths = High Frequency = High ENERGY

\[ \nu_1 = 4 \text{ cycles/second} = 4 \text{ hertz} \]

\[ \nu_2 = 8 \text{ cycles/second} = 8 \text{ hertz} \]

\[ \nu_3 = 16 \text{ cycles/second} = 16 \text{ hertz} \]
Electromagnetic Spectrum

- Composed of all forms of electromagnetic radiation
- The visible spectrum is a small portion of the electromagnetic spectrum
- The visible spectrum is continuous since one color fades into the other
- Each color has a different wavelength & frequency
Electromagnetic Spectrum

- Red has the longest wavelength & shortest frequency
- Violet has the shortest wavelength & longest frequency
- The rest of the spectrum is invisible such as microwaves, radio waves, & x-rays
*Figure 5.10 Electromagnetic Spectrum*

- Frequency Increases with longer wavelength.
- Low energy (700 nm) to high energy (380 nm) in the visible light range.

- Wavelengths range from radio waves to gamma rays.
- Electromagnetic spectrum with frequency and wavelength labels.
- When light strikes the surface of a metal it causes valence electrons to leave the metal
- This is known as the photoelectric effect
- It is the transfer of energy from the light to the electron
- In order to eject electrons the frequency of light has to increase
Photoelectric Effect

- As the electrons were ejected their kinetic energy increased with the frequency.
- Increasing the intensity increased the number of electrons ejected, but didn’t change the kinetic energy.
Wave-Particle Duality

- Max Planck & Albert Einstein tried to explain why energy was proportional to frequency & not amplitude.
- They both developed theories giving the following principles:
  - Light is made up of particles called photons.
  - The energy of a photon depends only on its frequency.
Wave-Particle Duality

- Matter only absorbs electromagnetic radiation in whole numbers of photons.
- In metals, electrons will only eject from the metal if the photon has the minimum energy required to eject it.
- The energy corresponds to the minimum frequency.
Based on Planck and Einstein’s work, a relationship between quantum energy & frequency developed:

\[ E = hv \]

- \( E \) = energy (J)
- \( v \) = frequency (s\(^{-1}\))
- \( h \) = Planck’s constant \( 6.626 \times 10^{-34} \text{ J}\cdot\text{s} \)
The blue in fireworks is often from Copper (I) chloride heated to 1200°C. Blue light has a wavelength of 450 nm. What is the quantum of energy that may be emitted by light at 450 nm?
Example

**Given** \( \lambda \)

\[
\lambda = 450 \text{ nm} \\
= 4.5 \times 10^{-7} \text{ m} \quad (1 \text{ nm} = 10^{-9} \text{ m})
\]

\[
\nu = \frac{c}{\lambda} \\
= \frac{2.9979 \times 10^8 \text{ m/s}}{4.5 \times 10^{-7} \text{ m}} \\
= 6.6 \times 10^{14} \text{ s}^{-1}
\]
\[ \Delta E = h \nu \]
\[ = (6.626 \times 10^{-34} \text{J} \cdot \text{s}) (6.6 \times 10^{14} \text{s}^{-1}) \]
\[ = 4.417 \times 10^{-19} \text{ J} \]
HYDROGEN LINE SPECTRUM

- When an electron (e-) absorbs (gains) energy (in whole photons or “quanta”) it “jumps” to a higher energy level
  - This is called the EXCITED STATE

- When an e- emits (loses) energy it falls to a lower energy level and the energy emission is given off as photons (light)
  - This is called the GROUND STATE
- Every time hydrogen falls to ground state or gets excited it emits a photon.
- This photon’s energy is the difference between the initial & ground states.
- An element’s line spectrum indicates its electrons exist only in specific energy states.
- An electron’s return to ground state produces a colored light.
- Hydrogen emits four specific colors of the visible spectrum.
- Hydrogen emits four visible wavelengths of light
- This indicates only certain energies are allowed for the electron in the hydrogen atom.

- Hydrogen produces a line spectrum meaning it contains only certain colors.

- A continuous spectrum produces continuous range of colors. Usually emitted by solids.
- Electrons orbit the nucleus in fixed energy ranges called orbits (energy levels)

- An electron can move from one energy level to another by gaining or losing discrete amounts of energy

- Electrons cannot be found between energy levels (think of energy levels like rungs on a ladder . . . )
- Scientists use the Bohr model to explain how color is produced during a flame test.

- The lowest energy level is closest to the nucleus, the highest is farthest away.

- The electron energy levels are quantized.

Hydrogen Spectrum

Increasing energy of orbits

A photon is emitted with energy $E = hf$.

- $n = 3$
- $n = 2$
- $n = 1$
There is NO net change in energy
- Energy absorbed = energy released = energy of light produced

-Sometimes (like the flame test) this light is in the small section of wavelengths called the visible spectrum and we can see it
An atomic orbital is the region of space in which there is a high probability of finding an electron.

Quantum Numbers

- Each electron must have a different location (atomic orbital) in the atom. The electrons are described by four quantum numbers.
  - Principal
  - Angular momentum (shape)
  - Magnetic (orientation of orbital)
  - Spin
Pauli’s Exclusion Theory: States no two electrons in an atom can have the same set of four quantum numbers.
Principal Quantum Number

- Energy Level where the electron is located
- Numbered 1 to 7 out from the Nucleus
  - Each level holds a maximum of $2n^2$ number of electrons
  - Principal level 1 holds 2 electrons
  - Principal level 2 holds 8 electrons
  - Principal level 3 holds 18 electrons
  - Principal level 4 holds 32 electrons
- These energy levels correspond to the periods on the periodic table
Electrons also occupy energy sublevels within each level. These sublevels are given the designations $s$, $p$, $d$, and $f$.

These designations are in reference to the sharp, principal, diffuse, and fine lines in emission spectra. Shapes of the atomic orbital include: spherical ($s$), dumbbell ($p$), double-dumbbell ($d$), quadruple-dumbbell ($f$)
- The number of sublevels in each level is the same as the number of the main level (up to four sublevels).

<table>
<thead>
<tr>
<th>Energy Level</th>
<th>Can Have Shapes</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>s</td>
</tr>
<tr>
<td>2</td>
<td>s, p</td>
</tr>
<tr>
<td>3</td>
<td>s, p, d</td>
</tr>
<tr>
<td>4-7</td>
<td>s, p, d, f</td>
</tr>
</tbody>
</table>
- The maximum number of electrons in each of the energy sublevels depends on the sublevel:
  - The s sublevel holds a maximum of 2 electrons.
  - The p sublevel holds a maximum of 6 electrons.
  - The d sublevel holds a maximum of 10 electrons.
  - The f sublevel holds a maximum of 14 electrons.

- The maximum electrons per level is obtained by adding the maximum number of electrons in each sublevel.
- Gives the order in which atomic orbitals are filled
- Electrons occupy the orbitals of lowest energy first
- The Periodic Table is a guide for the Aufbau Principle, going from left to right as you move down the periodic table
- Each element represents one electron, each period (row) represents one energy level.
Electron Configurations

- The arrangement of electrons in an atom showing the location of electrons by sublevel.

- Electron’s arrangements assume the lowest possible energies known as the element’s ground state.

- All configurations are unique based on the number of electrons involved.
- The sublevel is written followed by a superscript with the number of electrons in the sublevel.

- If the $2p$ sublevel contains 2 electrons, it is written $2p^2$.

- The electron sublevels are arranged according to increasing energy.
Electron Configurations

- Assigning the address of all electrons in an atom

### Example

\[ 2p^6 \]

- **Principal Energy level**
- **Orbital Shape**
- **Number of electrons in that shape**
ELECTRON CONFIGURATIONS

- The periodic table can be used as a guide for electron configurations.

- The period number is the value of \( n \) (the principal energy level).

  • Groups 1A and 2A have the s-sublevel filling.
  • Groups 3A – 8A have the p-sublevel filling.
  • Groups 3B – 2B have the d-sublevel filling.
  • The lanthanides and actinides have the f-sublevel filling.
We can use the periodic table to predict which sublevel is being filled by a particular element.
Electron Configurations

- Use the Diagonal Rule to determine the order of filling the orbitals

- Using diagonal arrows, it allows you to determine the order in which sub orbitals fill with electrons

Start

1s

2s 2p

3s 3p 3d

4s 4p 4d 4f

5s 5p 5d 5f

6s 6p 6d

7s
Write configurations for O, Ni, Br, Sr

O = 1s² 2s² 2p⁴

Ni = 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d⁸

Br = 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁵

Sr = 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d¹⁰ 4p⁶ 5s²
Exceptions to the AUFBAU principle
- Exceptions are only in the d sublevel and only when you have a $Xs^2Yd^4$
- It is more stable to have a half filled d shell than a partially filled d shell
- The configuration would be $Xs^1Yd^5$
- The same goes if you have $Xs^2Yd^9$
- It would be more stable to have $Xs^1Yd^{10}$
Valence Electrons

- Electrons in an atom’s outermost principal energy level (furthest from nucleus).

- When an atom undergoes a chemical reaction, only the outermost electrons are involved.
Valence Electrons

- These electrons are generally further from the nucleus are of the highest energy and determine the chemical properties of an element--they are the “most important” electrons to chemists.

- Each element can have a maximum of 8 valence electrons.
Shorthand E- Configurations

-Since the valence electrons are the “important” electrons, we use a shorthand system to show an element's valence electrons.

-All noble gases (group 18) have 8 valence electrons (except for helium) and therefore have a very stable configuration (most atoms want 8 valence electrons).
Write configurations for K and Ar

K = 1s² 2s² 2p⁶ 3s² 3p⁶ 4s¹

Ar = 1s² 2s² 2p⁶ 3s² 3p⁶

Write configuration for K using shorthand

K = [Ar] 4s¹
Write the shorthand electron configuration of:

P

[Ne] 3s² 3p³

Br

[Ar]4s²3d¹⁰4p⁵

Ca

[Ar] 4s²

V

[Ar]4s²3d³
An orbital is the region of space where there is a high probability of finding an atom.

The higher the energy of an orbital, the larger its size.
- Each atomic orbital has a box (2 electrons per box)
- Hund’s Rule: Give each orbital (of equal energy) one electron before any get two electrons

**Nitrogen Orbital Diagram**

<table>
<thead>
<tr>
<th>1s</th>
<th>2s</th>
<th>2p__x</th>
<th>2p__y</th>
<th>2p__z</th>
</tr>
</thead>
<tbody>
<tr>
<td>\uparrow \downarrow</td>
<td>\uparrow \downarrow</td>
<td>\uparrow</td>
<td>\uparrow</td>
<td>\uparrow</td>
</tr>
</tbody>
</table>
HUND’S RULE

- Within a sublevel, place one electron per orbital before pairing them.
- “Empty Bus Seat Rule”
Write the orbital diagram and determine the number of unpaired electrons for iron.

Fe = 1s² 2s²  2p⁶  3s²
    ▲▼ ▲▼ ▲▼ ▲▼ ▲▼ ▲▼ ▲▼

3p⁶  4s²  3d⁶
    ▲▼ ▲▼ ▲▼ ▲▼ ▲▼ ▲▼ ▲▼ ▲▼ ▲▼

4 unpaired electrons
- Because valence electrons are so important in the formation of bonds chemists represent them visually using another shorthand method.
- An electron dot structure consists of an atom's symbol surrounded by dots that represent the atom's valence electrons.
- Example: Carbon $\text{[He]}2s^22p^2$ has 4 valence electrons.
- Place valence electrons one at a time on all four sides of the symbol, then (if needed) pair each electron up until all have been used.

- Exception: Helium has a full valence shell with 2 electrons
Draw the electron dot structures for:

Sr
\[ \text{Sr} \]

F
\[ \text{F} \]

Na
\[ \text{Na} \]

S
\[ \text{S} \]