Ch. 13 - Electrons in Atoms
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13.1 - The Development of the Atomic Models

The electron travels around the nucleus in the **Electron cloud**. The electron cloud is made up of several **atomic orbital’s** which are areas around the nucleus where the electron is most **likely** to be found.

Energy levels are like the **steps** on staircase, where each stair in one energy level. A person can climb up or down the staircase by going from step
to step. However, a person on a staircase can **NOT** stand **between** steps, nor can an electron exist **between** energy levels. Unlike the staircase, the energy levels of an atom are not equally spaced.

A quantum is a small **packet** of **energy**. A quantum is enough energy to make an electron jump from its current energy level up to the next one. The electrons in atoms are **quantized** meaning that an electron can **NOT** be moved only a **little** but must be moved to the next **energy level** or remain where it is. The term quantum leap refers to electrons **jumping** from one energy level to the next.
Continuous       Quantized

If your car was quantized with units of 10 mph, you could only travel at speeds ending with a zero (0 mph, 10, 20, 30, etc.)

13.1 - The Quantum Mechanical Model of the Atom

The modern understanding of the electrons in atom is called the quantum mechanical model and was developed in large part by Erwin Schrodinger.
Schrodinger’s model said you cannot know the exact location of the electron, as Bohr thought he could, but instead describes areas of probability where the electron is most likely to be found. So the electron cloud is very much like a blurry cloud of negative charge.

This is similar to a student during the school day. If I did not have a copy of your schedule I could not predict exactly where in the building you would be at any point in time, but I could predict that you are probably somewhere in the building.

The cloud is most dense when the likelihood of finding the electron is large and it is least dense when the likelihood of finding the electron is small. So it is difficult to say where
the cloud **ends** It is like walking in thick fog. You can not tell where the fog starts and ends.

**Atomic Orbitals (a.k.a. Sublevels)**

The energy levels we have been discussing are designated by the **principle quantum number, n.** $n$ can be 1, 2, 3, 4, 5, 6, or 7. Again the energy levels are like stairs in a staircase or levels of a 7 story building.

Sublevels are **parts** of energy levels. Each energy level has **one or more** sublevel. (Color chart)
The **first** energy level, \( n = 1 \), will only have only **1** sublevel, a **s** sublevel.

The **second** energy level, \( n = 2 \), will have **2** sublevels, a **s** and a **p** sublevels.

The **third** energy level, \( n = 3 \), will have **3** sublevels, a **s**, **p**, and **d** sublevels.
The fourth energy level, $n = 4$, will have $4$ sublevels, a $s$, $p$, $d$, and $f$ sublevels.

Since there are only four common sublevels the $n = 5$, $6$, and $7$ will also have $s$, $p$, $d$ and $f$ sublevels.

When referring to a sublevel we identify it with its energy level and sublevel. So the $p$ sublevel in the 3rd energy level is called $3p$
s sublevels are **spherical**

p sublevels are **dumbell**

d sublevels are **clover** and the **f** is beyond simple description.

s sublevels have **1** orbital and a maximum of **2** electrons
p sublevels have 3 orbitals and a maximum of 6 electrons

d sublevels have 5 orbitals and a maximum of 10 electrons

f sublevels have 7 orbitals and a maximum of 14 electrons

\[ n = 1 \quad 2 \quad 3 \quad 4 \]

How many orbital's do each of the energy levels contain? 1 4 9 16

Each orbital can hold 2 electrons, so calculate the

Maximum number of electrons each energy level can contain.______  _____  _____  _____
The general equation for the number of \textit{electrons} in an energy level is \( n \) and the general equation for the maximum number of electrons in an energy level is \( 2n^2 \)

All electrons have either \textit{clockwise} or \textit{counterclockwise} spins. Spinning electrons create magnetic fields. If two electrons have the \textit{same} spin near each other then they \textit{cancel} each other out.

13.2 Electron Configurations

We have two systems for describing the electrons around an atom. One is called \textit{orbital filling} diagrams and the other is called \textit{electron configurations}. Both are used to describe the arrangement of \textit{electrons} around a \textit{nucleus} of an atom. To use either of these methods we first need to learn three rules for how to place electrons.
Rule # 1 - **The Aufbau principle** - Electrons fill orbitals from *lowest* energy sublevel first. Electron orbitals fill just like a glass of milk -- from the bottom up.

Rule # 2 - **The Pauli exclusion principle** - No two electrons can have the same four quantum numbers. So, electrons in the same orbital must have opposite *spin*.

Rule # 3 - **Hund's Rule** - When electrons occupy *orbitals* of the same *energy level* one electron will enter *each* orbital before the second enters any. This is similar to being on a bus full of people you don't know. Passengers will put one person in each empty seat, before two people will sit next to each other.
Configurations with unpaired electrons are attracted to magnetic fields are called **Paramagnetism** (para). Ex. Li, N

Configurations with only paired electrons are weakly repelled by magnetic field called **diamagnetism** (di). Ex. He, Be
Electron Configurations and the Periodic Table

We can relate the quantum numbers to the periodic table.

The energy level, \( n \), is the **period** (horizontal row) on the periodic table.

The sublevel, \( l \), is the **block** of the periodic table. Groups 1A and 2A make up the **s** block.

Groups 3A - 8A make up the **p** block. All "B" groups make up the **d** block and the two rows at the bottom make up the **f** block.

If you read the periodic table from left to right, just like a book, it corrects all order changes due to sublevels overlapping.
Valence electrons are only the electrons in the highest energy level.

13.3 Light and Atomic Spectra

Visible light is one form of electromagnetic energy. Other forms include radio, infrared, ultraviolet, and X-ray waves. Electromagnetic energy travels in the form of waves. Look at p. 373 of your text to see the entire electromagnetic spectrum. Notice that visible light is a small part of the electromagnetic spectrum. The rainbow’s order is ROY G. BV.

Take out a separate sheet of paper
All waves can be described by **FOUR** characteristics: -

**Amplitude, Wavelength, Frequency** and **speed**.

![Electromagnetic Spectrum Diagram]

- **Amplitude**: This is a wave's height. Amplitude is the distance from the **origin** to the **crest**.
- **Wavelength**: Represents the distance between two consecutive crests or troughs.
- **Frequency**: The number of waves that pass a given point in a given time.
- **Speed**: The velocity at which the wave travels through the medium.
Next is wavelength. **Wavelength** is the distance between **crest**. A diagram showing wavelength is below. Wavelength is represented by “λ” and is read "lambda".

The third is frequency. The **frequency** of a wave is the **number** of times that the wave completes a cycle each **second**. Frequency is symbolized by **ν**, which is called “**nu**”

- Per second = 1/s = s⁻¹ = 1 Hertz = 1 Hz.
The last is **speed**. Electromagnetic radiation moves through space at

\[ 3.00 \times 10^8 \text{ m/s} \]

\[ c = \nu \lambda \]

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

1) Calculate the frequency of an x-ray having a wavelength of \( 2.5 \times 10^{-9} \) m.

2) Calculate the wavelength of a wave with the frequency of \( 5.0 \times 10^{14} \) Hz.
13.3 The Quantum Concept

Plank discovered the relationship between the frequency of light and the amount of energy it contains.

\[ E = h \nu \]

where:
- \( E \) = energy
- \( h \) = Plank's constant = \( 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \)
- \( \nu \) = frequency

3) How much energy is there in an infrared quantum with a frequency of \( 4.5 \times 10^{13} \text{ s}^{-1} \)?
4) How much energy is there in a X-ray quantum with a wavelength of \(5.42 \times 10^{-9}\) m?

The Explanation of Atomic Spectra

We normally assume that the atom is at its **ground state**, which means that all electrons are in their **lowest energy** levels possible.
When an electron of an atom receives energy, electrons in atoms enter an **excited state** where an electron **leaves** its **low** level orbital and moves to one that is **higher** in energy.

Then that electron is attracted back to its nucleus and it **falls** back down to a **lower** energy level. When this happens the **energy** is **given off** by the electron in the form of **electromagnetic** energy. Some of this electromagnetic energy is at the right wavelength to be visible light.

Each element gives off different colors of light when the electron falls back down. This light is called a **spectrum**. By studying known element's...
spectrum, if we have an unknown element's spectrum, we can identify it. This is how astronomers can identify what distant stars are made of, by looking at their spectrum.

The Photoelectric Effect

Albert Einstein confirmed this idea when he observed the photoelectric effect. If you were to shine a red light on a sheet of metal nothing would happen regardless of the intensity (amplitude) of the red light. However, if you were to shine a violet light on a piece of metal regardless of the intensity (amplitude) electrons are ejected from the surface of the metal. He suggested that quanta of energy behave like tiny
particles, which he called photons. The idea of light ejecting electrons on a metal surface is the fundamentals of the solar cell.