The electrons in an atom exist in various energy levels.

When an electron moves from a lower energy level to a higher energy level, energy is absorbed by the atom. When an electron moves from a higher to a lower energy level, energy is released (often as light).

Neils Bohr was able to determine the energy levels of hydrogen by the visible light energy that is released when the electron drops from $3 \rightarrow 2$ (red light), $4 \rightarrow 2$ (blue-green), $5 \rightarrow 2$ (blue-violet) and $6 \rightarrow 2$ (violet).

Transitions to level $n = 1$ are too high energy to see (UV).

A shorthand notation is the electron configuration:

$$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6,$$ etc.

Three rules define how the orbitals fill:

**The Pauli Exclusion Principle**
Each orbital can be occupied by no more than two electrons.

**The Aufbau Principle**
The electrons occupy the lowest energy orbitals available. The “Ground State” for an atom is when every electron is in its lowest energy orbital.

**Hund’s Rule**
When more than one orbital exists of the same energy (p, d, and f orbitals), place one electron in each orbital before doubling them up.

The *valence electrons* are the outermost electrons… those farthest from the nucleus. They have the largest principal quantum number, $n$.

These electrons occupy the $s$ and $p$ orbitals in the highest energy level. These orbitals are called the *valence orbitals*.

The columns of the periodic table are labeled, I, II, III, IV, V, VI, VII and VIII (ignoring the transition and rare earth elements). This label tell you the number of valence electrons of every element in that column (except He.)

The valence electrons are important in how atoms bond.