# Emission & Absorption

**Unit:** Quantum and Particle Physics

**NGSS Standards/MA Curriculum Frameworks (2016):** N/A

**AP® Physics 2 Learning Objectives/Essential Knowledge (2024):** 15.3.A, 15.3.A.1, 15.3.A.2, 15.3.A.2.i, 15.3.A.2.ii, 15.3.A.2.iii, 15.3.A.3, 15.3.A.4, 15.3.A.4.i, 15.3.A.4.ii, 15.3.A.4.iii

**Mastery Objective(s):** (Students will be able to…)

- Describe the emission or absorption of photons by atoms.
- Relate atomic emission and absorption spectra.
- Identify an element based on its emission or absorption spectra.
- Calculate the frequency/wavelength of light emitted using the Rydberg equation.

**Success Criteria:**

- Calculations are correct

**Language Objectives:**

- Be able to explain and draw & label representations of an atom.

**Tier 2 Vocabulary:** atom, charge, nucleus

## Notes:

energy state: the position of an electron that corresponds to the interaction energy between that electron and the nucleus of the atom. In general, higher energy states correspond with greater distances from the nucleus.

absorption: when photons are absorbed (taken in) by an atom, pushing the atom to a higher energy state.

emission: when photons are emitted (given off) by an atom, causing the atom to spontaneously move to a lower energy state.

Absorption and emission are the processes by which an atom can take in or give off energy. The specific amount of energy that is absorbed or emitted corresponds to the difference in energy between the energy states of electrons in the atom before *vs.* after the absorption or emission.

Because the photons produced by these transitions have a specific amount of energy, those photons therefore have a specific frequency and wavelength:

$$E = hf$$
$$c = \lambda f$$

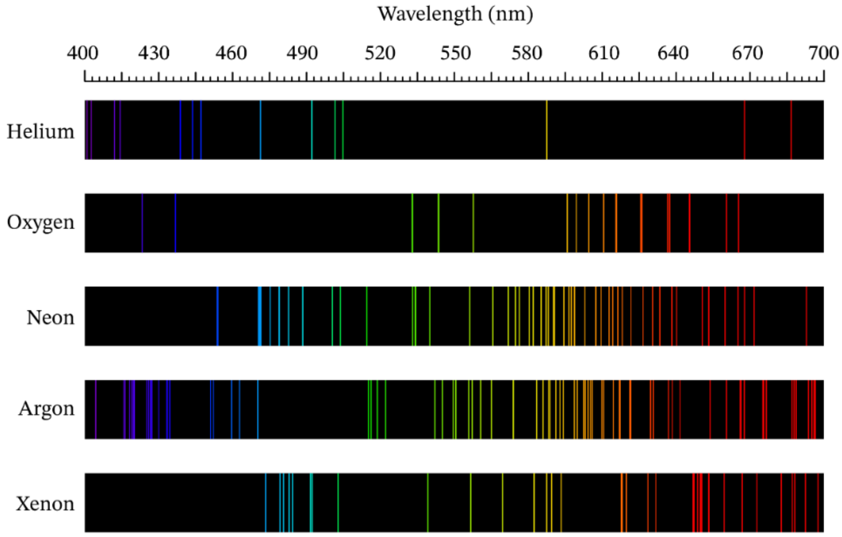
## Absorption and Emission Spectra

The set of electrons in an atom and their possible energy states define the specific frequencies and wavelengths that can be produced. In this way, the spectrum of wavelengths produced when energy is absorbed or emitted is a way of determining the elements in a source of light.

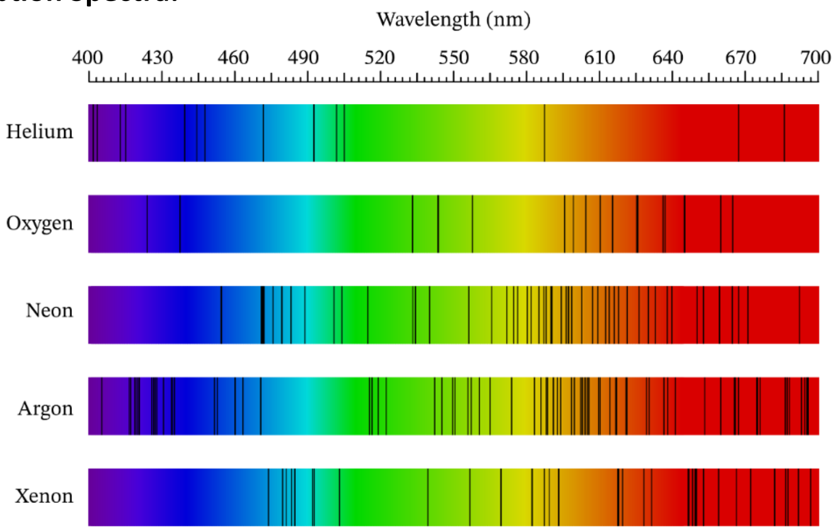
spectroscopy: the study and interpretation of wavelengths of light

Here are emission and absorption spectra for selected elements. Notice that for a given element, the wavelengths emitted and absorbed are the same.

**Emission Spectra:**




**Absorption Spectra:**



## Development of Spectroscopy

Spectroscopy was discovered in the 17th century, when prisms were first used to separate light by wavelength. Isaac Newton performed the first experiments with optics and is generally considered to be the founder of modern spectroscopy.

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In 1885, Johann Balmer studied the emission lines produced by the hydrogen atom. He devised an empirical equation to relate these emission lines.

Rydberg Formula (1888): Swedish physicist Johannes Rydberg developed a generalized formula that could describe the wave numbers of all of the spectral lines in hydrogen (and similar elements).

There are several series of spectral lines for hydrogen, each of which converge at different wavelengths. Rydberg described the Balmer series in terms of a pair of integers ($n_1$ and $n_2$, where $n_1 < n_2$), and devised a single formula with a single constant (now called the Rydberg constant) that relates them.

$$\frac{1}{\lambda_{vac}} = R_H \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)$$

The value of Rydberg's constant, $R_H$, is

$\dfrac{m_e e^4}{8\varepsilon_o^2 h^3 c} = 10\,973\,731.6 \text{ m}^{-1} \approx 1.1 \times 10^7 \text{ m}^{-1}$, where $m_e$ is the rest mass of the electron, $e$ is the elementary charge, $\varepsilon_o$ is the permittivity of free space, $h$ is Planck's constant, and $c$ is the speed of light in a vacuum.

Rydberg's equation was found to be consistent with other series discovered later, including the Lyman series (in the ultraviolet region; first discovered in 1906) and the Paschen series (in the infrared region; first discovered in 1908).

Those series and their converging wavelengths are:

| Series | Wavelength | $n_1$ | $n_2$ |
|---|---|---|---|
| Lyman | 91 nm | 1 | $2 \rightarrow \infty$ |
| Balmer | 365 nm | 2 | $3 \rightarrow \infty$ |
| Paschen | 820 nm | 3 | $4 \rightarrow \infty$ |

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The following diagram shows Lyman, Balmer and Paschen series transitions for the hydrogen atom, from higher energy levels ($n$ = 2 through $n$ = 6) back to lower ones ($n$ = 1 through $n$ = 3):