

## Chapter 4

# The Origin and Nature of Light

Question: How do astronomers take the temperatures of stars?

A thermometer is out of the question!

Consider what happens to a piece of iron when it is heated by a torch. As the iron heats up its color changes from dull red to orange, then yellow then white. The electromagnetic radiation that the rod gives off can be used to measure its temperature. Like the heated rod, a star's temperature can also be measured by measuring the radiation given off by the star. Both the rod and the star are examples of what astronomers call blackbodies.

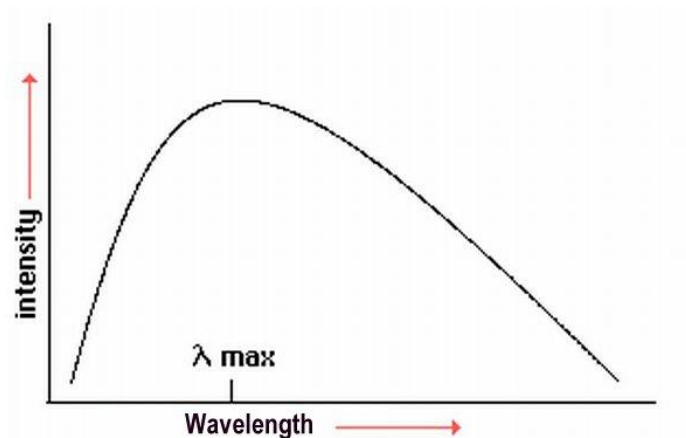
### Blackbody Radiation

**A. Blackbody:** An object that absorbs all of the electromagnetic radiation that strikes it. The absorbed radiation heats the object up and the radiation is then re-emitted at a different wavelength and intensity.

**B. Intensity:** describes the strength of radiation. It is amount of energy per unit area per second emitted by the blackbody.

**C. Wien's Law:** Distribution of radiation from a blackbody is called a blackbody curve. The wavelength (frequency) where the intensity is a maximum depends upon the temperature of the blackbody. Here the wavelength is measured in meters.

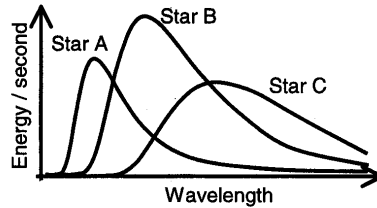
$$\lambda_{\max} = \frac{2.9 \times 10^{-3}}{\text{Temp(K)}}$$



### Concept Test

The following graph shows the blackbody curves for three different stars. Which of the stars is the hottest?

- a) Star A
- b) Star B
- c) Star C



**D. Stefan's Law:** Total amount of energy emitted (flux) by a blackbody is related to the fourth power of the temperature.

$$F = \sigma T^4$$

(F is the flux: total energy per square meter per second). Sigma is the Stefan-Boltzman constant =  $5.67 \times 10^{-8} \text{ J}/(\text{m}^2\text{K}^4 \text{ s})$

Blackbody curves are used to measure the temperatures of stars

### Spectra

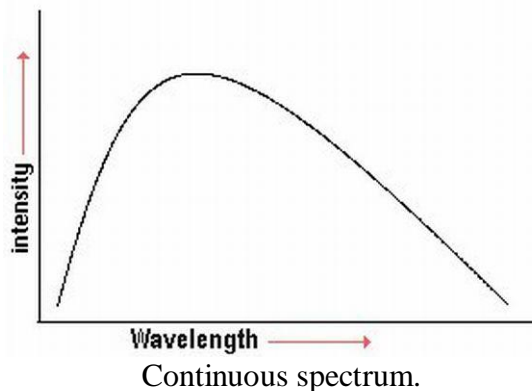
Spectra - how astronomers determine the chemical makeup of stars.

**A. Spectral Lines:** produced by atoms absorbing or emitting at specific wavelengths.

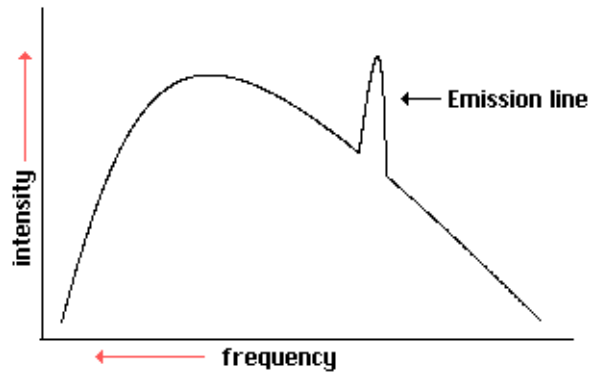
1. Atoms of each chemical element have unique spectral lines because each element has a unique set of electron orbits.

**B. Types of spectra**

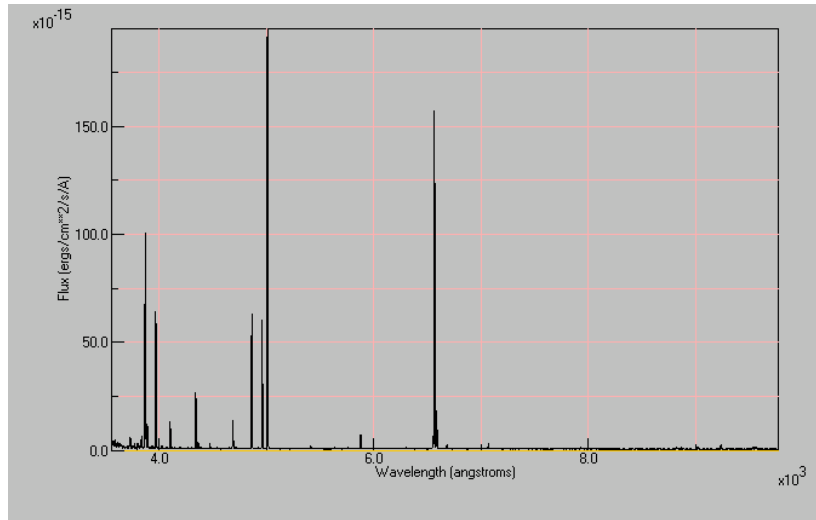
1. Continuous: contains all wavelengths.



2. Emission Lines: bright lines on a dark background or superimposed on a continuous spectrum.

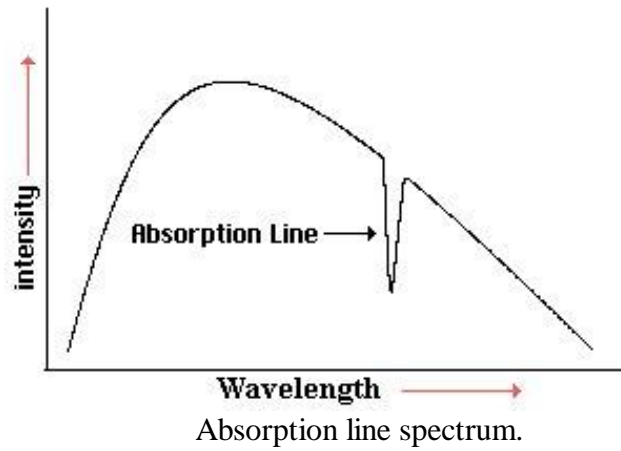


Emission line superimposed on a continuous spectrum.



Emission spectrum from a nebula.

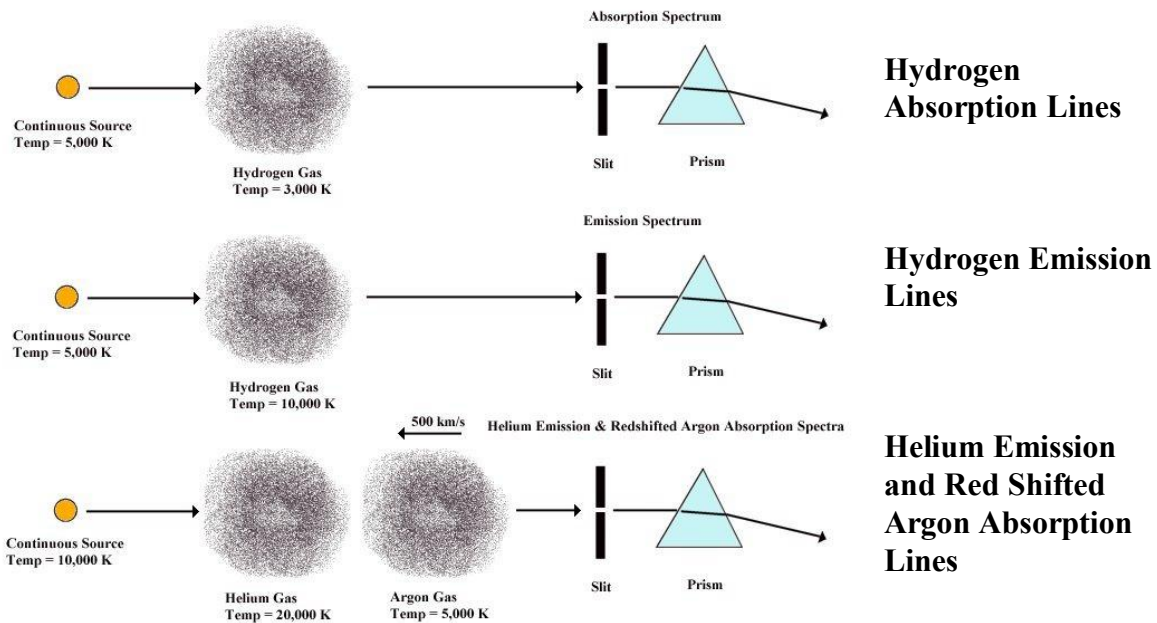
3. Absorption Lines: dark lines on a bright (continuous) background.



Example: Fraunhofer lines (absorption lines in the Sun's spectrum).

### C. Kirchhoff's Laws

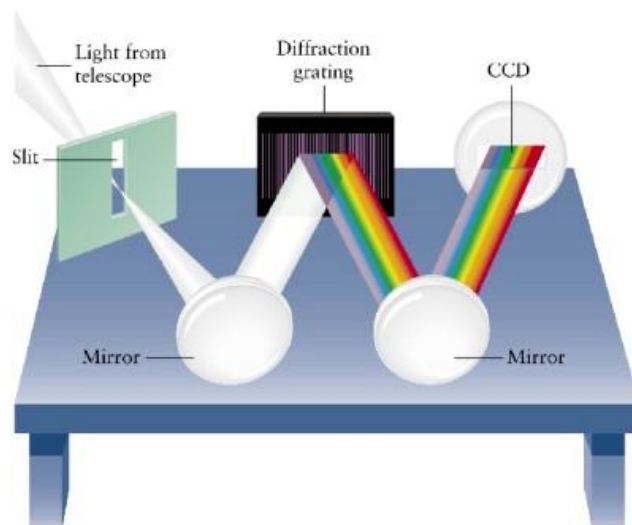
1. A luminous solid, liquid or dense gas emits light at all wavelengths producing a continuous spectrum.
2. A low density, hot gas emits light whose spectrum consists of a series of bright emission lines characteristic of the chemical composition.
3. A cool, rarefied gas absorbs certain wavelengths from a continuous spectrum, leaving dark absorption lines, superimposed on the continuous spectrum that are characteristic of the chemical composition of the gas.



**D. Spectrographs:** Instruments that break light into its component colors using prisms or gratings. Film or CCD cameras, are used to record the spectrum.

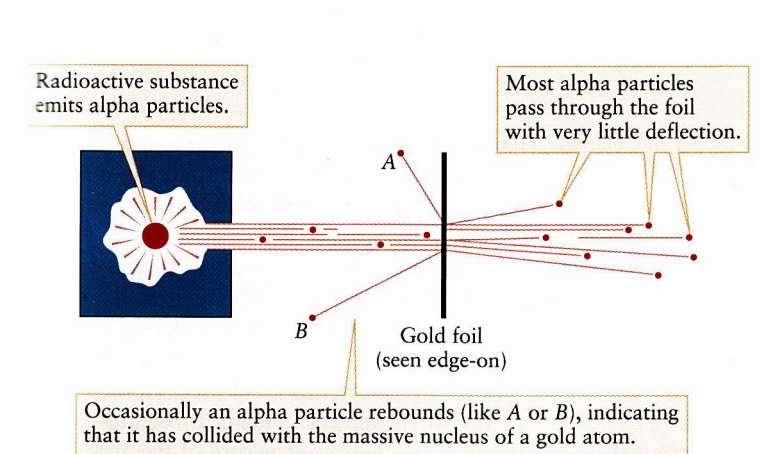
1. Prism Spectrograph: uses a prism to break light into its component colors. Works on the principle of refraction. Different colors are refracted by different amounts.

2. Grating Spectrograph: Uses a diffraction grating to break light into its component colors. Works on the principle of diffraction where different colors of light are diffracted by a different amount.



## E. Atoms

1. Each chemical element is made of a specific type of atom. Each atom consists of a nucleus composed of positively charged protons and electrically neutral neutrons. The nucleus is surrounded by negatively charged electrons in a number equal to the number of protons.
2. Rutherford proposed that most of the atom's mass was concentrated in the nucleus by observing the interaction of alpha particles (helium nuclei) with gold foil.

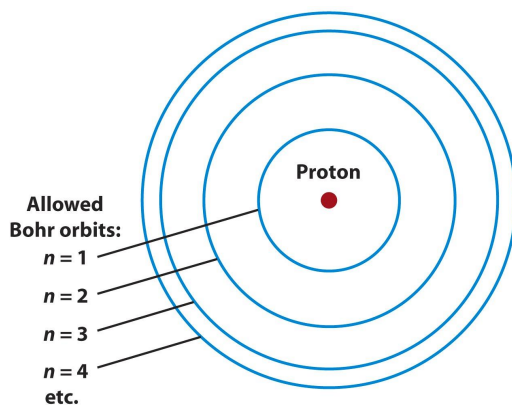


**Figure 5-17**

**Rutherford's Experiment** Alpha particles from a radioactive source are directed at a thin metal foil. This experiment provided the first evidence that the nuclei of atoms are relatively massive and compact.

## F. Bohr Model of the Hydrogen Atom

1. Proposed by Niels Bohr in order to explain the appearance of the spectrum of hydrogen.

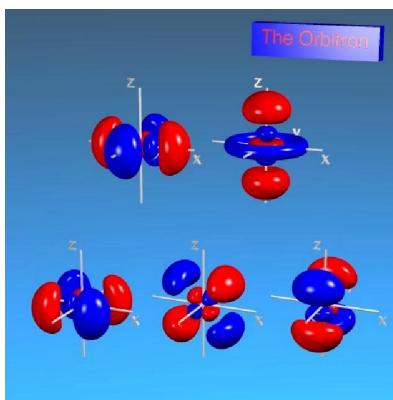


2. A single electron in a circular orbit around a proton (nucleus).

- a. An atom in its lowest energy state is in the ground state (preferred state).
- b. When the electron gains enough energy to break free of the nucleus, the atom is said to be ionized. Any atom that has lost one or more electrons is called an ion.
- c. Between the ground state and ionization the electron exists in discrete energy states called orbitals.
- d. Bohr model is not a correct picture of the hydrogen atom but it did predict the correct energies of the hydrogen atom.

### G. Modern Picture of the Hydrogen Atom - Quantum Mechanics

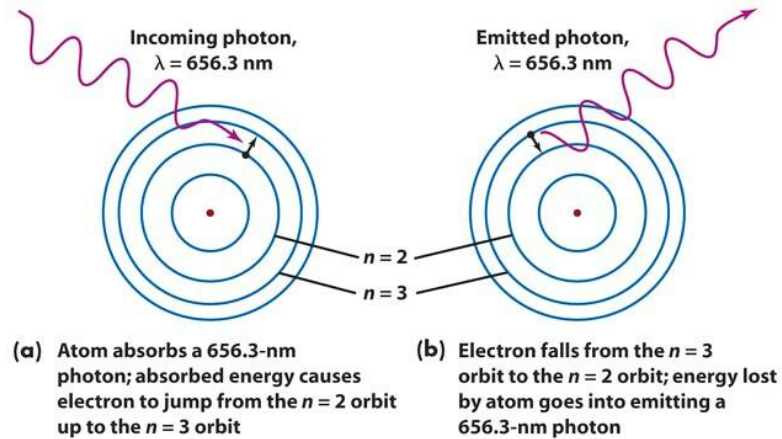
1. Instead of circular orbits the electrons are smeared out in an electron cloud. But each orbital still has a precise energy.
2. The average distance of the cloud from the nucleus is taken as the radius of the electron's orbit.



Ray traced representation of the Hydrogen 4D orbitals

### I. Absorption and Emission of a Photon

1. Absorption: **Occurs when a photon having an energy equal to the energy difference between two of the orbitals interacts with the atom and is absorbed by the electron.** The electron will temporarily reside in a higher energy level. Electrons can also gain energy through collisions (process called collisional excitation) with other atoms or particles
2. Emission: After about  $10^{-8}$  seconds the electron falls back to the ground state by emitting a photon having an energy equal to the difference in energy between the higher energy level state and the level that it falls to (electrons can fall to levels other than the ground state).

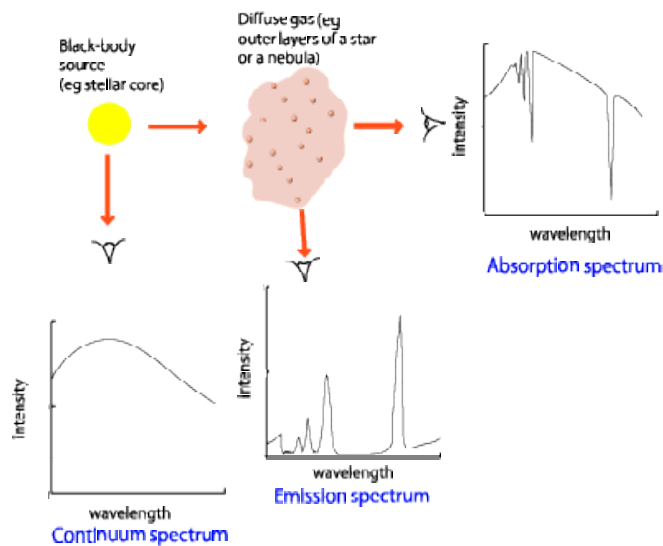


## J. Formation of Spectral Lines

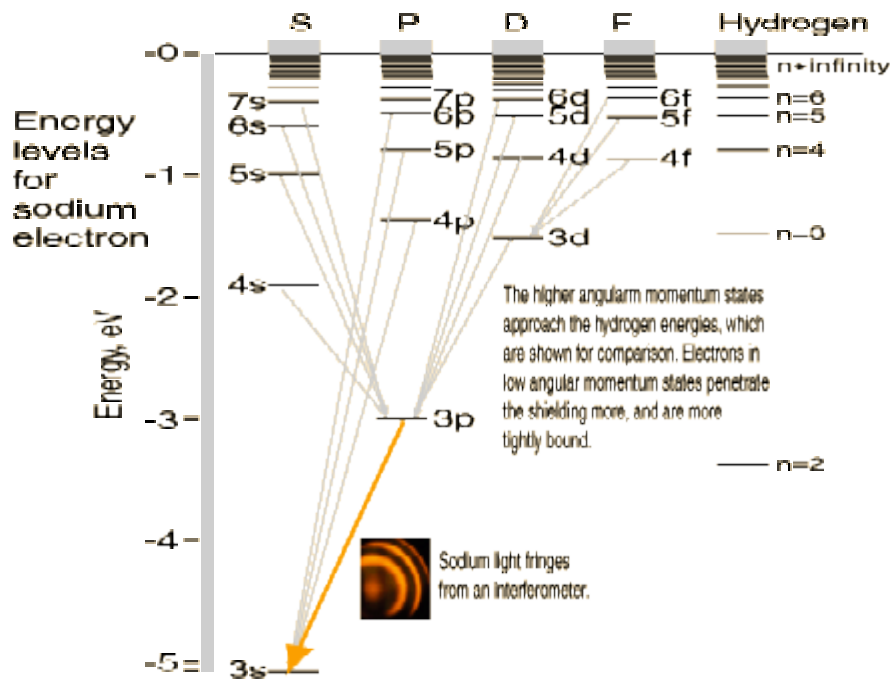
1. Cool gas illuminated from behind by a continuous source:

a. absorption lines are observed when the source and gas are in line with the observer.

b. emission lines are observed when the observer looks at the gas from the side.



2. Atoms of each chemical element have unique spectral lines because each element has a unique set of electron orbits.



### Concept Test

During Kirchhoff's early investigations into spectra he attempted to counter the absorption lines of calcium seen in the spectrum of the Sun with calcium salts burned in a Bunsen burner. He expected to see the dark absorption lines vanish but instead they got darker. Why?

- a) The Bunsen burner flame was hotter than the Sun.
- b) The Bunsen burner flame was the same temperature as the Sun.
- c) The Bunsen burner flame was cooler than the Sun.
- d) He was using the wrong chemical element.

## J. Atomic Spectra

### 1. Hydrogen

a. Balmer Lines: transitions to and from the  $n = 2$  level. Seen in the visible portion of the spectrum.

i.  $H\alpha$ : 656.3 nm

ii.  $H\beta$ : 486.1 nm

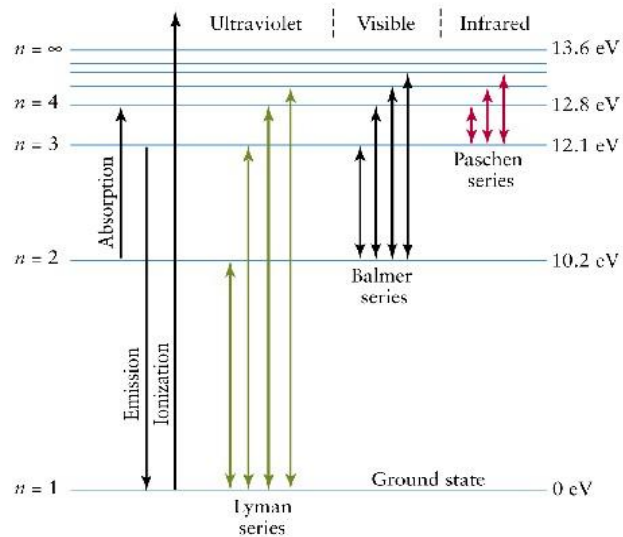
iii.  $H\gamma$ : 434.1 nm

iv. H $\delta$ : 409.6 nm

v. Balmer decrement ( $n = \infty$ ): 364.8 nm

b. Lyman Lines: transitions to and from the  $n = 1$  level. Seen in the ultraviolet.

c. Other series correspond to transitions to and from higher levels.



### Concept Test

In the Sun, the transition from level 4 to level 2 of hydrogen produces photons with a wavelength of 486.1nm. In a star twice as hot as the Sun, this transition would produce photons with

- a) half that wavelength.
- b) the same wavelength.
- c) twice that wavelength.
- d) four times that wavelength.

## K. Molecules

Spectral lines of molecules arise from changes in their rotational or vibrational states.

