7.2 Atoms

The basic building blocks of matter are atoms.

7.2.1 Constituents of the atom

The basic constituents of atoms are listed in the following table.

Constituent	Symbol	Charge	Mass
electron	e^-	-1	9.1×10^{-31} kg
proton	p^+	$+1$	$1836 \times e^-$ mass
neutron	n	0	$\sim$ proton mass

7.2.2 Isotopes

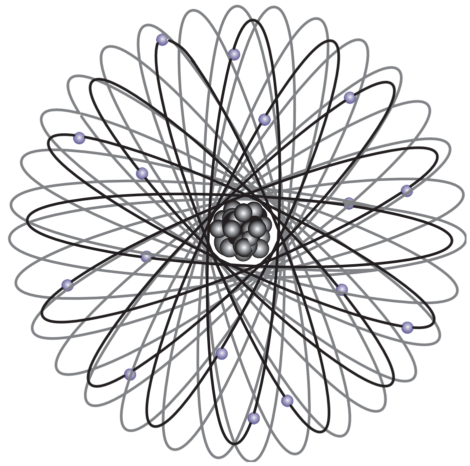
Isotopes are nuclei with the same number of protons but different numbers of neutrons (and therefore different atomic mass numbers). Examples:

$$^{235}_{92}\mathrm{U}_{143} \qquad ^{238}_{92}\mathrm{U}_{146} \qquad ^{4}_{2}\mathrm{He}_{2} \qquad ^{3}_{2}\mathrm{He}_{1}$$

The notation is $^{A}_{p}\mathrm{S}_{n}$, where S is the symbol for the element and $A = n + p$ (atomic mass number).

7.2.3 The Bohr atom

The simplest model of the atom is the "planetary" or *Bohr model*.



The typical picture of the Bohr atom is not to scale (that is, the relative sizes are not correct). The radius of the nucleus is about 100,000 times smaller than the radius of the entire atom and electrons are point particles without a physical size. Useful characteristic sizes to remember:

- Atoms: $1\ \text{Å} = 10^{-10}$ m
- Nuclei: $1\ \text{fm} = 10^{-15}$ m

where fm denotes a Fermi (traditional) or femtometer (standard metric). Typical atoms have radii of order $1-10$ Å and typical nuclei have radii of order $1-10$ fm.

7.2.4 Quantum mechanics

The correct theory of the atom is called *quantum mechanics;* the Bohr Model is an approximation to quantum mechanics that has the virtue of being much simpler.

Particle–wave duality At the core of quantum mechanics is the concept of *particle–wave duality:*

Particles Behaving As Waves

Just as we have seen that light behaves in some contexts as if it were a stream of particles and in others as a wave, quantum mechanics asserts that particles of matter can also behave as waves.

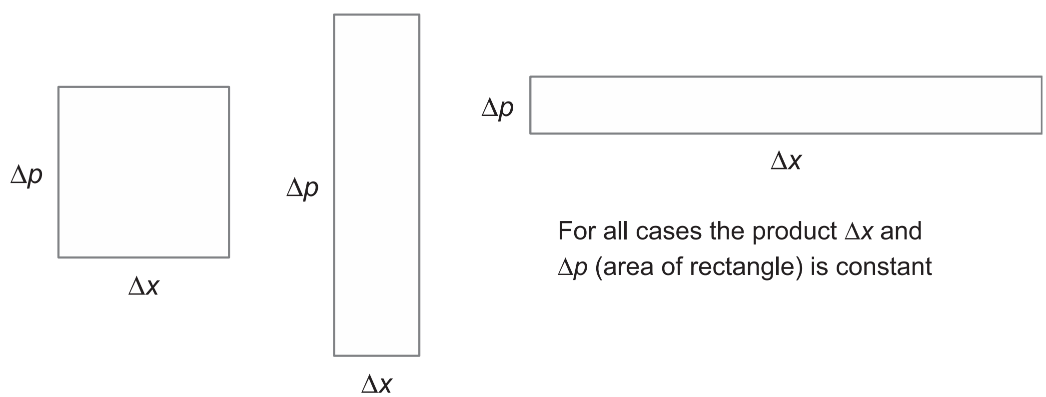
The quantum mechanical wavelength λ associated with a particle is called the *deBroglie wavelength* of the particle,

$$\lambda = \frac{h}{p} = 2\pi\frac{\hbar}{p} \qquad (\hbar \equiv h/2\pi). \tag{36}$$

Heisenberg uncertainty principle Because of the wave nature of matter, in quantum mechanics a particle obeys the *uncertainty principle.* Two common statements:

$$\Delta p \cdot \Delta x \gtrsim \hbar \qquad \Delta E \cdot \Delta t \gtrsim \hbar. \tag{37}$$

Position–momentum uncertainty in one dimension is illustrated in the following figure:



Example: Position uncertainty and deBroglie wavelengths

Consider a 1 kg rock moving at 2 m s^{-1}. Approximate Δp by (order of magnitude) $\Delta p \simeq p$. Then from the uncertainty principle

$$\Delta x \simeq \frac{\hbar}{\Delta p} \simeq \frac{\hbar}{p} = \frac{\hbar}{mv} = \frac{1.05 \times 10^{-34}\ \mathrm{kg\,m^2 s^{-1}}}{2\ \mathrm{kg\,m\,s^{-1}}} \simeq 5.3 \times 10^{-35}\ \mathrm{m}.$$

The corresponding deBroglie wavelength is

$$\lambda = 2\pi \frac{\hbar}{p} = 2\pi \Delta x \simeq 3.3 \times 10^{-34} \text{ m}.$$

These are completely negligible lengths compared with the size of the rock, so quantum mechanics has no detectable influence on the motion of the rock. But now consider an electron moving at 1/100 the speed of light. The corresponding momentum is

$$p = mv = (9.1 \times 10^{-31} \text{ kg})(3 \times 10^{6} \text{ m s}^{-1}) = 2.7 \times 10^{-24} \text{ kg m s}^{-1},$$

and the position uncertainty may be estimated as

$$\Delta x \simeq \frac{\hbar}{p} = \frac{1.05 \times 10^{-34} \text{ kg m}^2\text{s}^{-1}}{2.7 \times 10^{-24} \text{ kg m s}^{-1}} \simeq 3.9 \times 10^{-11} \text{ m}.$$

The corresponding deBroglie wavelength is

$$\lambda = 2\pi \Delta x \simeq 2.5 \times 10^{-10} \text{ m}.$$

These lengths are now *comparable to the size of an atom,* so an electron behaves with large wavelike character, and the uncertainty in its position when bound inside an atom is comparable to the size of the atom itself.

7.2.5 Quantized energy levels

The basic feature of quantum mechanics that is incorporated in the Bohr Model, which is completely different from the analogous planetary model, is that the energy of the particles in the Bohr atom is restricted to certain discrete values. One says that the energy is *quantized.* This means that only certain orbits with certain radii and energies are allowed; orbits in between simply don't exist.

Origin of Energy Quantization in the Wave Nature of Matter

The quantization of energy levels in an atom (or any bound microscopic system) ultimately follows from the wave nature of matter implied by particle–wave duality in quantum mechanics:

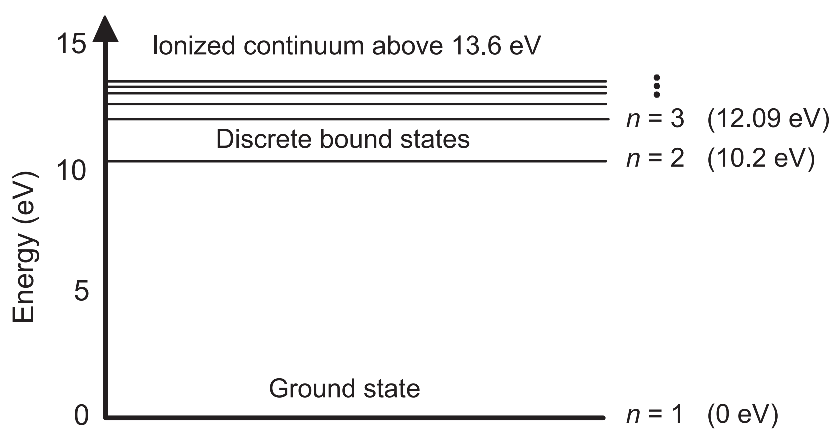
- Only deBroglie wavelengths such that an integer or half-integer number of λs fit within the characteristic volume are allowed.
- Since $p = h/\lambda$ and the kinetic energy is

  $$E = \tfrac{1}{2}mv^2 = \frac{p^2}{2m} = \frac{h^2}{2m\lambda^2},$$

  if only certain λ are allowed, then only the corresponding discrete energies are allowed.

Thus, energy in bound systems is *quantized.*

Discrete levels of the hydrogen atom are illustrated in the following figure.



These levels are labeled by an integer n that is called the *principle quantum number.* The lowest energy state is generally termed the *ground state.* The states with successively more energy than the ground state are called the *first excited state,* the *second excited state,* and so on.

Electron-Volts

In atomic and nuclear physics it is common to use an energy unit called the *electron–Volt* (eV). It is defined to be the energy that an electron would acquire when accelerated across an electrical potential of one Volt. The conversion factor to SI units is

$$1 \text{ eV} = 1.602 \times 10^{-19} \text{ J}.$$

It is also common to define the kilo electron-Volt (keV) and the Mega electron–Volt (MeV) by

$$1 \text{ keV} = 10^3 \text{ eV} \qquad 1 \text{ MeV} = 10^6 \text{ eV}.$$

It is often useful to express Planck's constant using eV for the energy unit:

$$h = 6.625 \times 10^{-34} \text{ J s} \times \frac{1 \text{ eV}}{1.602 \times 10^{-19} \text{ J}} = 4.135 \times 10^{-15} \text{ eV s}.$$

As we explain below, in the hydrogen atom there is a continuum of levels above an excitation energy of 13.6 eV.

The general formula for the energy of an atomic level in hydrogen relative to its ground state is

$$E = 13.6 \left(1 - \frac{1}{n^2}\right) \text{ eV}, \tag{38}$$

if we take the ground state to have $E = 0$. (Note that it is also common to take the ionization energy as the zero, in which case $E = -13.6 \text{ eV}/n^2$ and the ground state is at -13.6 eV.)

Example: Energy of the $n = 4$ level in hydrogen

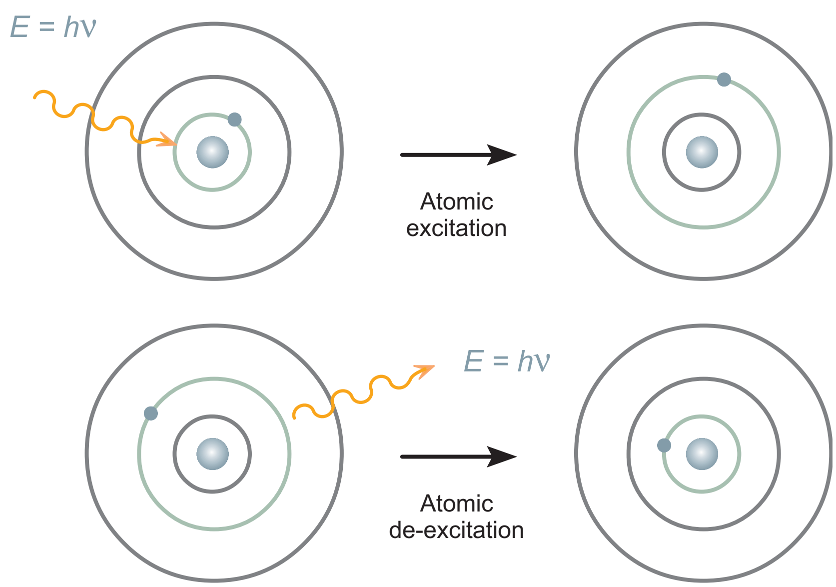
For the $n = 4$ (3rd excited state) in a hydrogen atom the energy is

$$E_4 = 13.6 \left(1 - \frac{1}{n^2}\right) \text{ eV} = 13.6 \left(1 - \frac{1}{4^2}\right) \text{ eV} = 12.75 \text{ eV}.$$

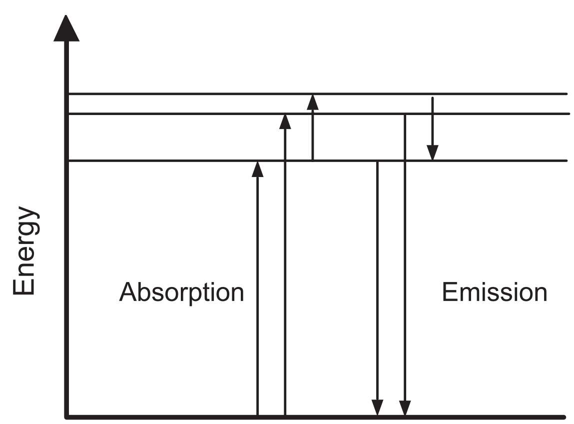
relative to the ground state.

7.2.6 Excitation and de-excitation of atomic levels

Because the energy levels are quantized, atomic systems can make discrete jumps between levels by absorbing and emitting quantized bundles of energy. For example, the following figure illustrates absorption and emission of photons, leading to excitation and de-excitation of an atom.



The corresponding picture in terms of our energy-level diagram is



The energy associated with these photons is exactly equal to the difference in energy ΔE between the two levels, and the frequency ν of the emitted or absorbed photon is given by $\Delta E = h\nu$:

$$\nu = \frac{\Delta E}{h} \qquad \Delta E \equiv E_{\text{final}} - E_{\text{initial}}.$$

Equivalently, the wavelength of the emitted or absorbed photon is

$$\lambda = \frac{hc}{\Delta E}.$$

Example: An emission transition in hydrogen

In a hydrogen atom the $n = 2$ level is at 10.2 eV and the $n = 3$ level is at 12.09 eV relative to the ground state. What is the wavelength of the photon emitted in a transition from the $n = 3$ level to the $n = 2$ level?

The transition energy is

$$\Delta E = 12.09 \text{ eV} - 10.2 \text{ eV} = 1.89 \text{ eV}.$$

For atomic problems it is useful to calculate once and for all that

$$hc = (4.1357 \times 10^{-15} \text{ eV s})(3 \times 10^{8} \text{ m s}^{-1}) \times \frac{1 \text{ nm}}{10^{-9} \text{ m}} = 1.2398 \times 10^{3} \text{ eV nm}.$$

Then we have

$$\lambda = \frac{hc}{\Delta E} = \frac{1.2398 \times 10^{3} \text{ eV nm}}{1.89 \text{ eV}} = 656 \text{ nm}.$$

This is in the red part of the visible spectrum. (Later we will see that this corresponds to the first transition in the Balmer emission series for hydrogen, denoted H_α.)

Ionization Continuum

Beyond an energy called the *ionization potential,* the single electron of the hydrogen atom is no longer bound to the atom and the energy levels form a continuum. For hydrogen, this continuum starts at 13.6 eV above the ground state, as illustrated in figure before the preceding one.

7.2.7 Ionization and the formation of plasmas

Atoms and molecules are electrically neutral: the number of negatively charged electrons is exactly equal to the number of positively charged protons. But in astrophysics the atoms are often subjected to high temperatures or collisions and may gain or lose their electrons.

The gain or loss of electrons by an atom is called *ionization.* The loss of electrons, which is the more common process in astrophysical environments, converts an atom into a positively charged ion, while the gain of electrons converts an atom into a negatively charged ion.

Astrophysics notation for degree of (positive) ionization is summarized in the following table.

Suffix	Ionization	Examples	Chemist's notation
I	Not ionized	H I, He I	H, He
II	Singly ionized	H II, He II	H^+, He^+
III	Doubly ionized	He III, O III	He^{++}, O^{++}

Plasmas

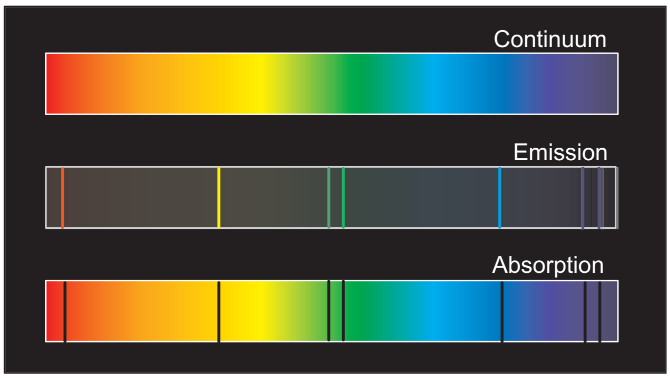
If most of the atoms or molecules in a region are ionized, the resulting state of matter corresponds to a gas that is electrically neutral on a global scale, but composed microscopically of a soup of positively charged ions and the negatively charged electrons stripped from the atoms to form the ions. Such a state of matter is called a *plasma.* Most matter in stars and in many clouds of gas is in the form of a plasma.

7.3 Spectra

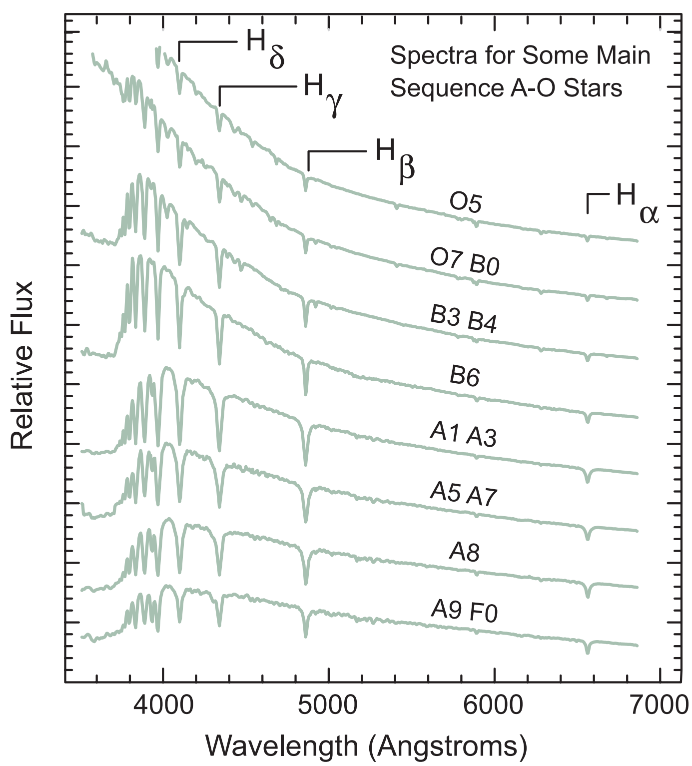
Isolated atoms, ions, or molecules can absorb and emit packets of electromagnetic radiation having discrete energies dictated by the detailed atomic structure of the atoms. When the corresponding light is passed through a prism or spectrograph it is separated spatially according to wavelength by either dispersion or diffraction. This is called a *spectrum,* and it is perhaps the most powerful tool that an astronomer has to unlock the secrets of the Universe.

7.3.1 Examples of spectra

The corresponding spectrum may exhibit a continuum, or may have superposed on the continuum bright lines (emission spectrum) or dark lines (absorption spectrum). The following figure shows the characteristic spectra that would be recorded on photographic plates.



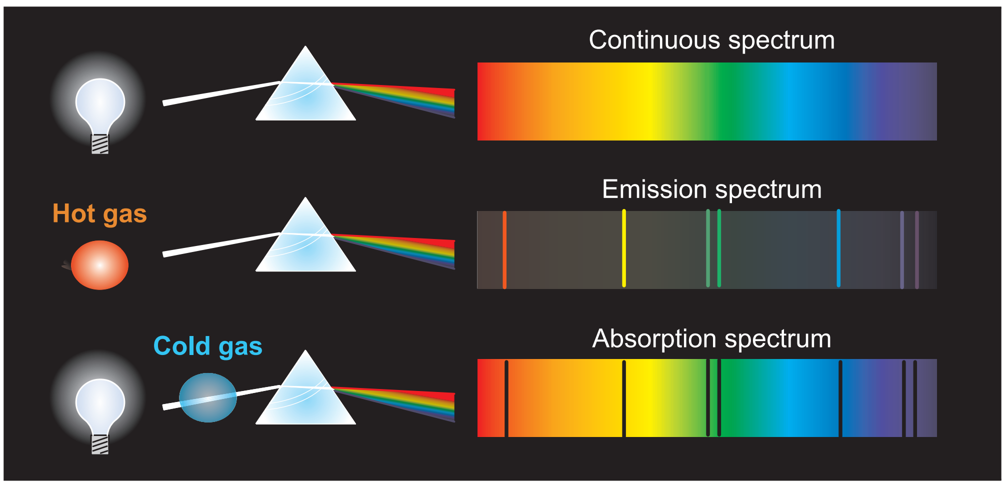
Some actual spectra recorded electronically are shown in the following figure



The spectral designations on the spectra will be discussed later. The notations H_α and so on denote hydrogen spectral lines in the Balmer series.

7.3.2 Origin of different types of spectra

The origin of the different types of spectra is illustrated in the following figure.

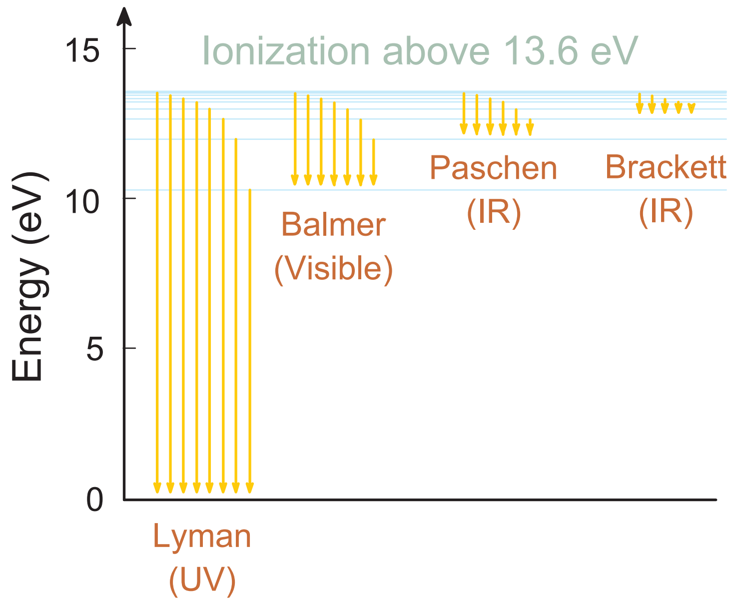


We may summarize the content of this figure as follows:

- Emission spectra are produced by thin gases in which the atoms do not experience many collisions. The emission lines, or peaks in intensity for an electronically recorded spectrum, correspond to photons of discrete energies that are emitted when excited atomic states make transitions back to lower-lying states.
- A continuum spectrum results when the gas pressures are higher. Generally, solids, liquids, or dense gases emit light at all wavelengths when heated.
- An absorption spectrum occurs when light passes through a cold, dilute gas and atoms in the gas absorb at characteristic frequencies; since the re-emitted light is unlikely to be emitted in the same direction as the absorbed photon, this produces dark lines (absence of light) in the line spectrum, or intensity dips in the electronically recorded spectrum.

7.3.3 The hydrogen spectrum

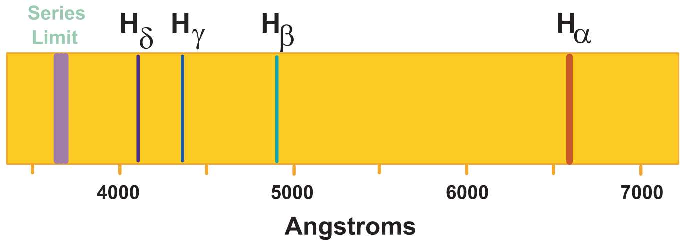
The spectrum of hydrogen is particularly important in astronomy because most of the visible Universe is made of hydrogen. Emission or absorption processes in hydrogen produce series, which are sequences of lines corresponding to atomic transitions, each ending or beginning with the same atomic state in hydrogen. The following figure illustrates some hydrogen spectral series.



- The Balmer Series involves transitions starting (for absorption) or ending (for emission) with the first excited state of hydrogen.
- The Lyman Series involves transitions that start or end with the ground state of hydrogen.
- The Paschen and Brackett series involve the second and third excited states, respectively.

The series illustrated are emission series (reverse the direction of the arrows to get absorption series). Because of the details of hydrogen's atomic structure, the Balmer Series is in the visible spectrum and the Lyman Series is in the UV. The other two series come in the IR.

The lines of the Balmer series and their labeling are illustrated in the following figure:



Example: Lyman absorption series

What is the wavelength corresponding to the lowest energy transition in the Lyman absorption series?

In a hydrogen atom the $n = 2$ level is at

$$E_2 = 13.6\left(1 - \frac{1}{n^2}\right)\ \text{eV} = 13.6\left(1 - \frac{1}{4}\right)\ \text{eV} = 10.2\ \text{eV}$$

relative to the ground state, so the transition energy for $2 \to 1$ is $\Delta E = 10.2$ eV and the corresponding wavelength is

$$\lambda = \frac{hc}{\Delta E} = \frac{1.2398 \times 10^3\ \text{eV nm}}{10.2\ \text{eV}} = 121.5\ \text{nm},$$

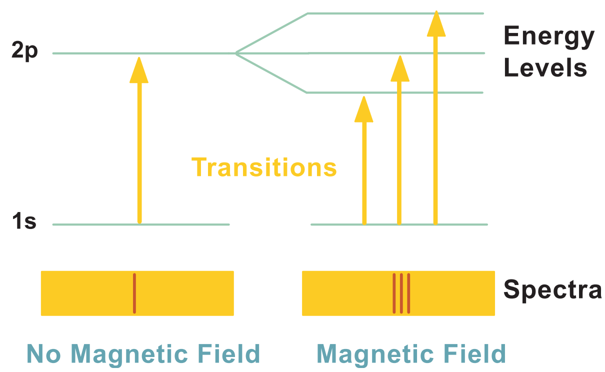
which is in the UV part of the spectrum.

Example: Web links to examples of wavelength filtering and false-color images

- http://antwrp.gsfc.nasa.gov/apod/ap061020.html (Gum Nebula in false color)
- http://antwrp.gsfc.nasa.gov/apod/ap051223.html (H-alpha in Rosette Nebula)
- http://antwrp.gsfc.nasa.gov/apod/ap060324.html (Rosette Nebula in false color)
- http://hubblesite.org/gallery/behind˙the˙pictures/meaning˙of˙color/eagle.shtml (Example of Hubble Space Telescope false coloring)
- http://hubblesite.org/gallery/behind˙the˙pictures/meaning˙of˙color/hubble.shtml (Interactive: NGC 1512 through 7 filters)
- http://hubblesite.org/gallery/behind˙the˙pictures/meaning˙of˙color/filters.shtml (Summary: NGC 1512 through 7 filters)

7.3.4 Zeeman effect

In the presence of magnetic fields, spectral lines are split. This is called the *Zeeman effect*, and is illustrated in the following figure.



The pattern of spectral line splitting is a a signature that a magnetic field is present, and the magnitude of the splitting is a measure of the strength of the field.

7.3.5 The Doppler effect

The Doppler effect for sound results when relative motion of the source causes the wavelength of the sound waves to be decreased ahead of the source and stretched out behind. Light is also a wave, and relative motion of the source leads to a corresponding Doppler effect for light.

In this case it is not the pitch but the color (wavelength) that is shifted by the motion of the source. The following figure illustrates.

Observer A and observer B see the same wavelength

Observer A sees red shift; observer B sees blue shift

The wavelength is changed to

- larger values if the motion of the source is away from the observer (*redshift*)
- smaller values if the motion is toward the observer (*blueshift*).

For velocities well below the speed of light, the shift in wavelength for the Doppler effect is given by

$$z \equiv \frac{\Delta\lambda}{\lambda_0} \equiv \frac{\lambda - \lambda_0}{\lambda_0} = \frac{v}{c}, \tag{39}$$

where

$\lambda =$ observed wavelength $\qquad \lambda_0 =$ wavelength at rest $\qquad z \equiv$ redshift parameter.

(For velocities approaching the speed of light, a more complicated formula that uses relativity to relate the velocity to the shift in wavelength must be used.)

7.3.6 Radial velocities

The Doppler formula provides a method to measure radial velocities (the component v_r of velocity along the line of sight), since

$$v_r = \frac{\Delta\lambda}{\lambda_0} c = \frac{\lambda - \lambda_0}{\lambda_0} c. \tag{40}$$

By determining the shift of known spectral lines from their wavelength when emitted at rest, v_r can be determined.

From Eq. (40), the radial velocity will be a

- Positive number if the $\lambda > \lambda_0$ (the Doppler shifted line corresponds to longer wavelength than the unshifted line). This is called a *redshift*.
- Negative number if the $\lambda < \lambda_0$ (the Doppler shifted line corresponds to shorter wavelength than the unshifted line). This is called a *blueshift*.

To summarize:

$$v_r > 0 \leftrightarrow \text{ Increase in } \lambda \leftrightarrow \text{ Motion away from observer } \leftrightarrow \text{ Redshift}$$

$$v_r < 0 \leftrightarrow \text{ Decrease in } \lambda \leftrightarrow \text{ Motion toward observer } \leftrightarrow \text{ Blueshift}$$

Example: Radial velocity from Doppler shift of spectral lines for stars

The hydrogen Balmer series H_α line is observed in the laboratory to occur at 656.285 nm. For the star Arcturus, the H_α line is observed at 656.274 nm. Assuming this to be a Doppler shift caused by radial motion of Arcturus, what is the radial velocity of Arcturus relative to us? From the Doppler formula

$$v_r = \frac{\lambda - \lambda_0}{\lambda_0} c = \left(\frac{656.274 \text{ nm} - 656.285 \text{ nm}}{656.285 \text{ nm}} \right) \times 3 \times 10^5 \text{ km s}^{-1}$$
$$= -5.03 \text{ km s}^{-1} \qquad \text{(blueshift).}$$

(The accepted value is $v_r = -5.2$ km s^{-1} for Arcturus, from SIMBAD.[1])

For the star Canopus, the H_α line is observed at 656.33 nm. Then the radial velocity of Canopus relative to us is

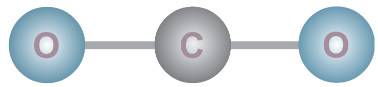
$$v_r = \frac{\lambda - \lambda_0}{\lambda_0} c = \left(\frac{656.33 \text{ nm} - 656.285 \text{ nm}}{656.285 \text{ nm}} \right) \times 3 \times 10^5 \text{ km s}^{-1}$$
$$= +20.6 \text{ km s}^{-1} \qquad \text{(redshift).}$$

(The accepted value is $v_r = +20.5$ km s^{-1} for Canopus, from SIMBAD.)

7.3.7 Molecules

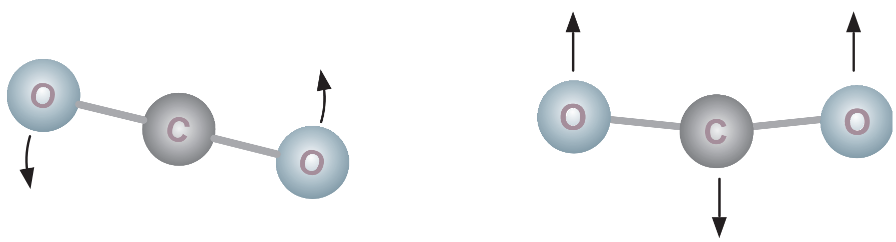
Molecules are often represented by simple ball and stick models in which the balls stand for atoms and the sticks stand for the chemical bonds that bind the atoms together to form the molecules. Carbon dioxide (CO_2) molecule:

[1]SIMBAD is an astronomy database at *http://simbad.harvard.edu/sim-fid.pl*. If you go to SIMBAD and type "Arcturus" into the top left field and click Submit, it will return data on the star Arcturus.



Carbon dioxide molecule

Atoms can absorb energy by exciting electrons. Molecules can do the same, but they also have additional ways to absorb energy. A molecule can rotate, and it can vibrate the lengths of its chemical bonds or their orientation angles. This is illustrated in the following figure.



Carbon dioxide rotational mode

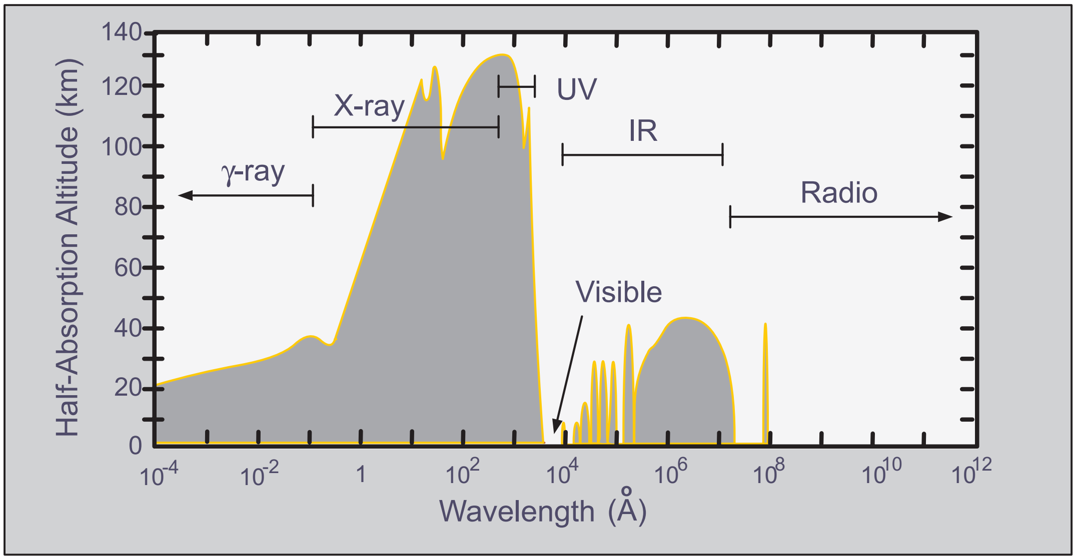
Carbon dioxide vibrational mode

Molecules can interact with light and give rise to characteristic spectra. Because of the energies associated with the rotations and vibrations, the spectra for molecules typically involve infrared (IR) or microwave wavelengths.

> Because molecules are usually fragile, molecular spectra are important mostly in objects that are relatively cool such as planetary atmospheres, the surfaces of very cool stars, and various interstellar regions (regions between the stars).

7.3.8 Atmospheric windows

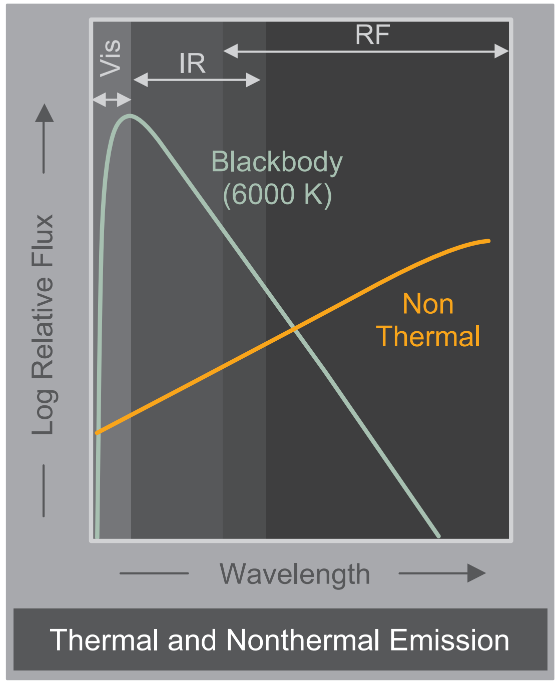
One important practical consequence of the interaction of electromagnetic radiation with matter and of the detailed composition of our atmosphere is that only light in certain wavelength regions can penetrate the atmosphere well. These regions are called atmospheric windows. The following figure shows the amount of absorption at different wavelengths in the atmosphere.



7.4 Non-Thermal Emission

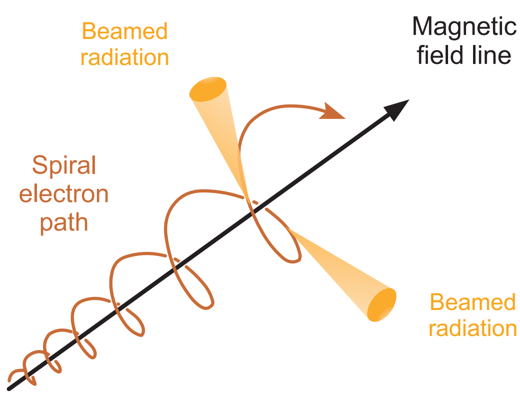
The Planck law describes what we term thermal emission. It characterizes the emission of radiation from a hot gas that is in approximate thermal equilibrium, and the resulting spectrum is a blackbody spectrum.

We also may observe emission of radiation that is nonthermal in character and does not follow the blackbody (Planck) law. It is characterized by a spectrum that increases in intensity at very long wavelengths. The following figure illustrates the difference between a blackbody spectrum and nonthermal emission.



Nonthermal emission is less common than thermal emission in astronomy. Nonthermal spectra indicate violent processes and large accelerations of charged particles. These in turn are associated with some of the most interesting objects in astronomy: pulsars, quasars, and active galactic nuclei, for example.

The most common form of nonthermal radiation in astronomy is associated with *synchrotron radiation.* Its origin is illustrated in the following figure.



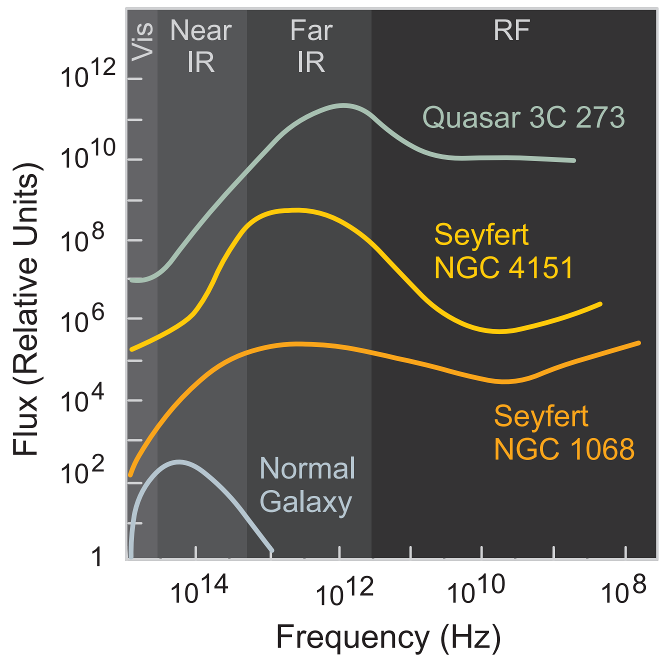
There are 3 essential points in understanding synchrotron radiation:

- First, high-velocity electrons (and other charged particles) follow curved paths in strong magnetic fields; thus, they are accelerated.
- Second, by fundamental laws of electricity and magnetism, an accelerated charged particle must radiate energy in the form of electromagnetic waves.
- Because of the nature of the emission process, synchrotron light is often *strongly polarized* (constraint on the direction in which the light wave is vibrating). Most sources of light are not strongly polarized, so this is a distinctive property.

The wavelength of the emitted radiation is related directly to how fast the particle spirals in the magnetic field. As the particle emits radiation, it slows and emits longer wavelength synchrotron radiation. This accounts for the broad distribution in wavelength of synchrotron radiation as compared with thermal radiation.

> Increasing the strength of the magnetic field increases the energy of the radiation, because the tightness of the spirals is governed by the field strength. Thus, high-frequency synchrotron radiation indicates strong magnetic fields.

Normal galaxies radiate primarily a thermal spectrum. However, there are many strong nonthermal sources associated with a class of galaxies called *active galaxies.* The following figure illustrates the largely blackbody spectrum for a normal spiral galaxy, and that of three active galaxies that are decidely not normal galaxies: a quasar and two Seyfert galaxies (we shall discuss these kinds of galaxies when we consider active galaxies).



Energy Sources for Nonthermal Synchrotron Emission

Electrons radiate continuously as they spiral in magnetic fields, so the energy driving the huge sustained synchrotron emission from these nonthermal sources must be replenished constantly. Strong, polarized, nonthermal emission is an indirect sign not only of strong magnetic fields, but also of a very large energy source. Quasars and Seyfert galaxies contain *active galactic nuclei* or AGNs. The huge nonthermal emission from AGNs implies a strong and very compact energy source, generally believed to be a massive rotating black hole.

Show APOD, March 9, 2002, for quasar examples.
