Absorption line spectra of hydrogen

A given atom will absorb and emit the same frequencies of electromagnetic radiation (E-M). A hydrogen atom gas will produce an absorption line spectrum if it is between you (your telescope+spectrograph) and a continuous light source, and an emission line spectrum is seen from a different angle. If you observed the star (a white light source) directly, you would see a continuous spectrum, without breaks. If you look at the star through the gas (telescope to the right of the gas cloud, points to the star through the cloud), you will see a continuous spectrum with breaks where specific wavelengths of energy were absorbed by the gas cloud atoms and then re-emitted in a random direction, spreading them out of our telescope beam. We call this absorption spectrum (continuous + dips). If you look at the gas, but not the star (telescope below the gas cloud, points through the gas, but away from the star), you will only see a portion of the scattered light being resented by the gas. The star's continuous radiation will not fall on our telescope beam, because we are pointed away from the star. This is called emission spectrum (peaks only, not continuous). The absorbed and emitted E-M radiation frequencies correspond to the energy levels allowed in the atom. The energy levels allowed in an atom depend mainly on the configuration of the electric field. Hydrogen, with a proton in the nucleus, has a different field configuration than helium with two protons – that's why the two atoms have different energy levels and different characteristic lines of absorption and emission. [NMSU, N. Vogt] Continuous spectrum: A gas can be collisionically animated. Imagine a hot gas. The atoms are flying around, colliding with each other and sometimes the energy of the movement during the collision will collide with an electron at a higher energy level (or completely ionize the atom, releasing the electron completely). When this electron falls back to the lowest energy, a photon is emitted. This conversion of kinetic energy into radiant energy cools the gas. There is a connection between the emission lines of a gas and the continuous spectrum of a solid. As you cluster atoms together (as in a solid), the energy levels allowed in an atom begin to become distorted due to the influence of the electrical field of neighboring atoms. Slightly distort the energy level difference and you have a slightly different frequency emission/absorption line. A distribution of distortions leads to a distribution of lines than eventually blends into a continuum. Here we see how a spectrum fills as the density of our medium increases (from gas to solid) by clustered particles. [NMSU, N. Vogt] Absorption spectrum: What are star spectra (the observed light of stars)? Stars have absorption line spectra. We can think stars as a continuous hot source with a cool atmosphere of absorbent gas. The wavelengths that are absorbed depend on the chemical composition of the gas in the stellar stellar In the 1800s, sunlight was dispersed and looked more or less like a Planck spectrum (a blackbody curve) with some light missing, or absorption lines at certain wavelengths. [NMSU, N. Vogt] Stellar black body spectra have a characteristic shape, with a steep ascent, a peak in or near the visual passing band, and a slow decrease in infrared. Warmer stars have higher peak amplitudes, and peak at shorter wavelengths. Blackbody curves are shown for three stars in the figure below, with temperatures ranging from 4,000K (a red and cold star) to 7,500 K (a hot blue-blue star). Small arrows mark the maximum wavelength for each star. [NMSU, N. Vogt] Emission spectrum: Wavelengths with missing light in a stellar spectrum have proved very interesting and important. Its importance was perceived after the discovery and investigation of the chemists. If a gas is heated to the point where it glows, the resulting spectrum has light at discrete wavelengths that become able to match the wavelengths of light lost in stellar spectra. So by studying the spectra of various elements in a laboratory here on Earth, we can determine the composition of distant stars! [NMSU, N. Vogt] Development of the emission spectrum of the current atomic theory of hydrogen when an electric current is passed through a glass tube containing hydrogen gas at low pressure the tube emits blue light. When this light is passed through a prism (as shown in the figure below), four narrow bands of bright light are

observed against a black background. These narrow bands have the characteristic wavelengths and colors shown in the table below. Wavelength Color 656.2 red 486.1 blue-green 434.0 blue-violet 410.1 violet Plus four series of lines were discovered in the hydrogen emission spectrum, looking for the infrared spectrum at longer wavelengths and the ultraviolet spectrum at shorter wavelengths. Each of these lines fits into the same general equation, where n1 and n2 are integers and HR is 1.09678 x 10-2 nm-1. Explanation of the Emission Spectrum Max Planck presented a theoretical explanation of the radiation spectrum emitted by an object that glows when heated. He argued that the walls of a shiny solid could be imagined to contain a series of resonators that oscillated at different frequencies. These resonators gain energy in the form of heat from the walls of the object and lose energy in the form of electromagnetic radiation. The energy of these resonators at any given time is proportional to the frequency with which they oscillate. To fit the observed spectrum, Planck had to assume that the energy of these oscillators could assume only a limited number of values. In other words, the energy spectrum for these oscillators was not Continuous. Because the number of energy values of these oscillators is limited, they are theoretically countable. The energy of the oscillators in this system is therefore said to be quantified. Planck Planck the notion of quantization to explain how light was emitted. Albert Einstein extended Planck's work to the light that had been emitted. At a time when everyone agreed that light was a wave (and therefore continuous), Einstein suggested that it behaved as if it were a stream of small beams, or packets, of energy. In other words, light has also been quantized. Einstein's model was based on two assumptions. First, he assumed that light was composed of photons, which are small and discreet beams of energy. Second, he assumed that a photon's energy is proportional to its frequency. E = hv In this equation, h is a constant known as planck constant, which is equal to 6,626 x 10-34 J-s. Example: Let's calculate the energy of a single red light photon with a wavelength of 700.0 nm and the energy of a mole from these photons. The red light with a wavelength of 700.0 nm has a frequency of 4,283 x 1014 s-1. Replacing this frequency in the Planck-Einstein equation gives the following result. A single photon of red light carries an insignificant amount of energy. But a mole from these photons carries about 171,000 joules of energy, or 171 kJ/mol. The absorption of a mole from red light photons would provide enough energy to raise the temperature of one liter of water by more than 40oC. The fact that hydrogen atoms emit or absorb radiation at a limited number of frequencies implies that these atoms can only absorb radiation with certain energies. This suggests that there are only a limited number of energy levels within the hydrogen atom. These energy levels are countable. The energy levels of the hydrogen atom are quantized. The Bohr Model of the Niels Bohr Atom proposed a model for the hydrogen atom that explained the hydrogen atom's spectrum. The Bohr model was based on the following assumptions. The electron in a hydrogen atom travels around the nucleus in a circular orbit. The energy of the electron in an orbit is proportional to its distance from the nucleus. The more the electron is from the nucleus, the more energy it has. Only a limited number of orbits with certain energies are allowed. In other words, the orbits are quantized. The only orbits allowed are those for which the angular momentum of the electron is an integral multiple of the Planck constant divided by 2p. Light is absorbed when an electron jumps into a higher energy orbit and emitted when an electron falls into a lower energy orbit. The energy of the light emitted or absorbed is exactly equal to the difference between the energies of the orbits. Some of the key elements of this hypothesis are illustrated in the figure below. Three points deserve special attention. First, Bohr acknowledged that his first assumption violates the principles of classical mechanics. But he knew it was impossible to explain the spectrum of the hydrogen atom within the confines of classical physics. He was therefore to assume that one or more of the principles of the classic classic may not be valid on the atomic scale. Second, he assumed that there are only a limited number of orbits in which the electron can reside. He based this assumption on the fact that there are only a limited number of lines in the hydrogen atom spectrum and his belief that these lines were the result of light being emitted or absorbed as an electron moved from one orbit to another in the atom. Finally, Bohr restricted the number of orbits in the hydrogen atom by limiting the allowed values of the electron's angular momentum. Any object that moves along a straight line has an impulse equal to the product of its mass (m) times the speed (v) with which it moves. An object moving in a circular orbit has an angular moment equal to its mass (m) times the velocity (v) times the radius of the orbit (r). Bohr assumed that the angular momentum of the electron can assume only certain values, equal to Planck's entire times divided by 2p. Bohr then used classical physics to show that the energy of an electron in either of these orbits is inversely proportional to the square of the integer n. The difference between the energies of any two orbits is therefore given by the following equation. In this equation, n1 and n2 are integers and HR is the proportionality constant known as the Rydberg constant. Planck's equation states that the energy of a photon is proportional to its frequency. E = hv Replacing the relationship between frequency, wavelength and speed of light in this equation suggests that the energy of a photon is inversely proportional to its wavelength. The inverse of the wavelength of electromagnetic radiation is therefore directly proportional to the energy of this radiation. By properly defining the units of the constant, HR, Bohr was able to show that the wavelengths of light given or absorbed by a hydrogen atom should be given by the following equation. Thus, once he introduced his basic assumptions, Bohr was able to derive an equation that corresponded to the relationship obtained from the analysis of the hydrogen atom spectrum. According to the Bohr model, the wavelength of light emitted by a hydrogen atom when the electron falls from a high-energy orbit (n = 4) into a lower energy orbit (n = 2). Overwriting the appropriate HR, n1, and n2 values in the equation shown above gives the following result. Solving for the wavelength of this light gives a value of 486.3 nm, which agrees with the experimental value of 486.1 nm for the blue line in the visible spectrum of the hydrogen atom. Wave-particle duality The wave-particle duality theory developed by Louis-Victor de Broglie eventually explained why the Bohr model was successful with atoms or ions that contained an electron. It also provided a basis for understanding why this model failed for more complex systems. Light acts like a particle and a wave. In many ways light acts as a wave, with a feature, wavelength, wave, Range. Light carries energy as if it had discrete photons or energy packs. When an object behaves like a moving particle, it has an energy proportional to its mass (m) and the speed at which it moves through space(s). E = ms2 When behaving like a wave, however, it has an energy proportional to its frequency: By simultaneously assuming that an object can be a particle and a wave, de Broglie configured the following equation. By rearranging this equation, it derived a relationship between one of the wave properties of matter and one of its properties as a particle. As noted in the previous section, the product of an object's mass times the speed at which it moves is the impulse (p) of the particle. Thus, broglie's equation suggests that the wavelength (l) of any moving object is inversely proportional to its momentum. De Broglie concluded that most particles are too heavy to observe their wave properties. When the mass of an object is very small, however, wave properties can be detected experimentally. De Broglie predicted that the mass of an electron was small enough to display the properties of particles and waves. In 1927 this prediction was confirmed when electron diffraction was experimentally observed by C. J. Davisson. De Broglie applied his theory of wave-particle duality to the Bohr model to explain why only certain orbits are allowed for the electron. He argued that only certain orbits allow the electron to satisfy its particle and wave properties at the same time because only certain orbits have a circumference that is an integral multiple of the electron's wavelength, as shown in the figure below. The Bohr Model vs. Reality At first glance, the Bohr model looks like a two-dimensional model of the atom because it restricts the movement of the electron to a circular orbit in a two-dimensional plane. In reality, the Bohr model is a one-dimensional model, because a circle can be defined by specifying only one dimension: its radius, r. As a result, only one coordinate (n) is required to describe the orbits of the Bohr model. Unfortunately, electrons are not particles that can be restricted to a one-dimensional circular orbit. They act to some extent as waves and therefore exist in three-dimensional space. The Bohr model works for atoms of an electron or ions just because certain factors present in more complex atoms are not present in these atoms or ions. To build a model that describes the distribution of electrons into atoms that contain more than one electron we have to allow electrons to occupy three-dimensional space. We therefore need a model that uses three coordinates to describe the distribution of electrons in these atoms. Wave and Orbital functions We still talk about the Bohr model of the atom, even though the only thing this model can do is explain the atom's spectrum of because it was the last model of the atom for which a simple physical physical image be built. It is easy to imagine an atom consisting of solid electrons rotating around the nucleus in circular orbits. Erwin Schrödinger combined the equations for wave behavior with the Broglie equation to generate a mathematical model for the distribution of electrons in an atom. The advantage of this model is that it consists of mathematical equations known as wave functions that meet the requirements placed on the behavior of electrons. The downside is that it is difficult to imagine a physical model of electrons as waves. Schrödinger's model assumes that the electron is a wave and tries to describe regions in space, or orbital, where electrons are more likely to be found. Instead of trying to tell us where the electron is at any time, the Schrödinger model describes the probability that an electron can be found in a certain region of space at a given time. This model no longer tells us where the electron is; he just tells us where he might be. Be.

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