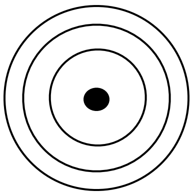
# NS D2 (pg 1 of 2) Electrons Inside the Atom – Quantum Theory

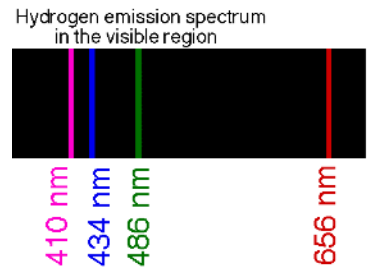
Name_______________________

## 1913 The planetary model of the atom – Niels Bohr



After the nucleus of the atom was discovered, scientists thought of the atom as a "microscopic solar system" in which electrons orbited the nucleus. But since electrons are negatively charged, and protons are positively charged, scientists wondered why the electrons didn't spiral into the nucleus. Bohr correctly assumed that the prevailing laws of physics were inadequate to describe the behavior of the incredibly tiny of atom and its parts. Bohr analyzed cathode ray tubes (CRT's) filled with different gases as a way to gain information about the inside of atoms. In particular he studied a CRT filled with hydrogen gas. Since hydrogen is the simplest atom, Bohr used the hydrogen atomic emission spectra to give support for the planetary model.

Bohr used the four bright lines (although this looked looked like 3 lines with our diffraction grating in class. The lines were red, green and two purple - no colors in-between) of the hydrogen atomic emission spectrum to postulate:




1. Only orbits of certain radii, corresponding to certain definite energies, are permitted for electrons in an atom.
2. An electron in a permitted orbit has a specific energy and is in an "allowed" energy state. An electron in an allowed energy state will not lose energy and therefore will not spiral into the nucleus.
3. Energy is only emitted or absorbed by an electron as it changes (quantum leaps) from one allowed energy state to another. This energy is emitted or absorbed as a very particular packet of energy, the color of which corresponds to its "leap" between energy states.

Much of our present understanding of the electronic structure of atoms has come from analysis of the light emitted (atomic emission spectra) or absorbed (atomic absorption spectra) by substances. To understand the basis for our current model of electronic structure, therefore, we must first learn more about light.

## What you need to know about light in chemistry class (and a bit more info in NoteSheet D1)

The chart hanging near the windows in class shows the various types of light, aka electromagnetic radiation (EMR) arranged order of increasing energy, a display called the electromagnetic spectrum. Notice that the 'light" spans an enormous energy range. Notice also that visible light is an extremely small portion of the electromagnetic spectrum. Red light is lower energy up to violet light, which is higher energy. (R O Y G B I V in order of low to high energy ) We can see visible light because it is in the wavelength range that is detectable by the molecules in the light receptors in our eyes. Some non-human animals can see non-visible light because the molecules in their eye receptors respond to higher or lower wavelengths of light. Because electromagnetic radiation carries (or radiates) energy through space, it is also known as radiant energy. There are many types of electromagnetic radiation in addition to visible light. These different forms—such as the radio waves that carry music to our radios, the infrared radiation (heat) from a glowing fireplace, the microwaves in your oven, and the X rays used by a dentist—may seem very different from one another, yet they share certain fundamental characteristics. All types of electromagnetic radiation move extremely fast at a speed of $3.00 \times 10^8$ m/s, the speed of light, symbolized by the letter c.

## Quantum Theory

In 1900, Max Plank proposed that light was quantized. This means that light can be thought of consisting of small fundamental packets, rather than a continuous stream. He called these discrete packets of energy, photons

## The Dual Nature of Radiant Light: Particle or Wave?

In the years around the development of Bohr's model, it became increasingly clear that depending on the experimental circumstances, radiation appears to have either particle-like properties or wave-like properties. Louis de Broglie proposed this same duality for matter. Subsequent experiments have demonstrated that all matter has wave properties, but for ordinary sized objects the wavelength is so small that it is impossible to observe under ordinary circumstances. For objects as small as an electron, its wave-like properties are important and measurable. Werner Heisenberg proposed the Uncertainty Principle (which would be better called the Indeterminability Principle) that concludes that the dual nature of matter places a fundamental limitation on how precisely we can describe both the location and momentum (mass x velocity) of any object.

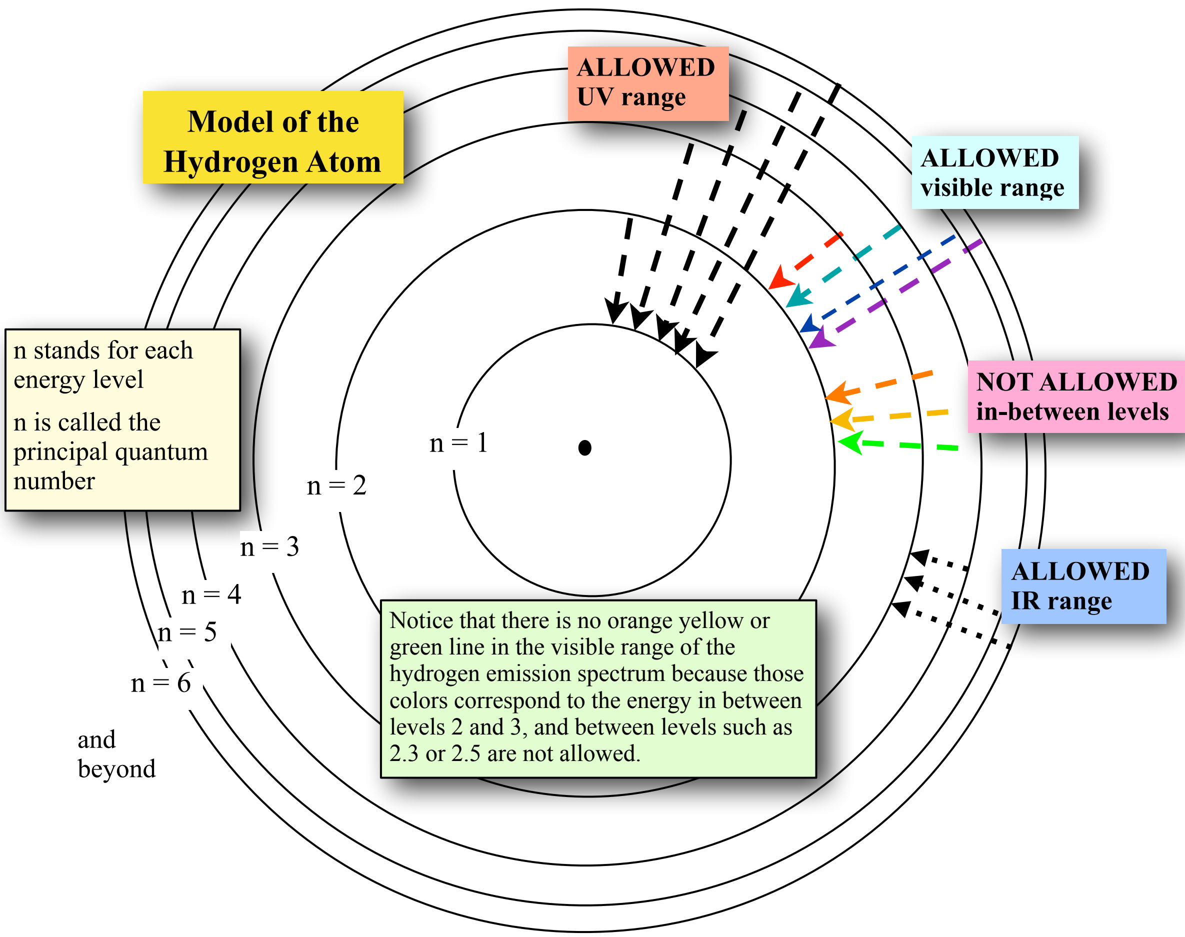
## The Bohr Model

In his third postulate, Bohr assumed that the electron could "jump up" from one lower-energy state to a higher-energy state (an excited state) by absorbing energy. This energy could be put into the atom as light, heat or electricity. Conversely, radiant energy is emitted when the electron drops back down to a lower energy state. The existence of spectral emission lines can be attributed to the quantized jumps of electrons between energy levels.

While the Bohr model offers an explanation for the line spectrum of the hydrogen atom, it can NOT explain the spectra of other atoms, except in a rather crude way. The Bohr model is only an important step along the way toward the development of a more comprehensive model. What is most significant about Bohr's model is that it introduces two important ideas that are also incorporated into our current wave mechanical model:

1. Electrons exist only in certain discrete energy levels, which are described by quantum numbers.
2. Energy is involved in moving an electron from one level to another. Energy must be put IN to move an electron to a higher energy level, and energy comes OUT when an electron moves to a lower energy level.

Thus when energy in the form of electricity or heat is put into an atom, the electrons absorb the energy that makes them "leap" further away from the nucleus (**excited state**) than they would ordinarily reside (**ground state**). When the energy source is no longer smacking that "*excited*" electron, the electron falls back down to where the electron would ordinarily be - in a lower energy orbital, called the "ground" state. When the electron returns to its ground state it must give off the energy that had been absorbed to get excited. This energy is emitted in the form of radiation, which is sometimes in the visible spectrum. Each of the bright lines of the hydrogen emission spectra corresponds to the fall from a particular energy level. Since there is no energy level 2.5 only 1, 2, or 3, we cannot see any yellow or orange light for the hydrogen spectrum.




## Schrödinger and Dirac

de Broglie and Heisenberg set the stage for a new theory of atomic structure called Quantum Mechanics. In this model, the "location" of electrons is defined as a probability density, or an area where the electron is "likely" to be found. We call these locations, electron orbitals (instead of orbits). In 1926, Erwin Schrödinger and Paul Dirac proposed a mathematical "wave equation" which when solved with advanced calculus gives us the information that we know today about the "locations" of electrons inside of the atom. We categorize those "locations" with electron configurations.

## **NS D2** (bonus…) Why don’t electrons crash into the nucleus???

*From what I learned in chemistry, the protons in the nucleus pull the electrons in and push on each other through electromagnetic forces, but are held in place by the strong nuclear forces in its gluons. Not much was said, however, about what keeps the electrons orbiting. I've always just assumed it was other electrons that prevented an electron from becoming part of the nucleus. In the form of Hydrogen that only has one electron, what keeps that electron from being pulled completely into the nucleus?*

If the electron obeyed classical mechanics and the electron was only subject to electrostatic (+/−) attraction to the nucleus, the electron would never fall into the nucleus despite the fact that the electron would be constantly attracted to the nucleus. This is exactly analogous to why the Earth doesn't fall into the Sun: the Earth has too much angular momentum, so by the time the Sun has made Earth "fall" significantly, Earth is already on another part of Earth’s orbit. Thus the Earth (like the prospective electron) keeps "falling" in circles around the Sun.

Electrons, alas, do not feel only electrostatic forces but must comply with the full electromagnetic theory of Maxwell, which dictates that accelerating charges (like circling electrons) must radiate their energy as electromagnetic waves. This energy is taken out of the orbital motion, which would steadily collapse into the nucleus. And within a fraction of a second, too.

This was a puzzle for a very long time and this puzzle was the glaring flaw in the planetary model of the atom when the model was first proposed by Rutherford and Moseley. You can only cure the puzzle by making the electron a quantum-mechanical beast, a weird hybrid between a particle and a wave.

Essentially, the electron fails to fall into the nucleus because an electron’s position, like any wave, cannot be tightly confined without giving the electron a very small wavelength (and that would confer the electron a large momentum, allowing the electron to break out of the nucleus). Electron waves, like all waves, like to spread out, and they can also interfere with themselves to make complicated interference patterns around the nucleus. A huge triumph of theoretical physics occurred when Schrödinger and Dirac proposed an equation describing the way in which these electron waves can add with themselves constructively to make standing wave patterns whose energies exactly matched those of the planetary Bohr model and therefore experimental facts. These standing waves are the only stable states of the electron waves in the atom, which is why the electrons don’t collapse intoto the nucleus.

*You see, while you are learning, lots...in some ways we are only scratching the surface.*

## **NS D2** (bonus…) Why don’t electrons crash into the nucleus???

*From what I learned in chemistry, the protons in the nucleus pull the electrons in and push on each other through electromagnetic forces, but are held in place by the strong nuclear forces in its gluons. Not much was said, however, about what keeps the electrons orbiting. I've always just assumed it was other electrons that prevented an electron from becoming part of the nucleus. In the form of Hydrogen that only has one electron, what keeps that electron from being pulled completely into the nucleus?*

If the electron obeyed classical mechanics and it was only subject to electrostatic attraction to the nucleus, it would never fall into the nucleus despite the fact that it would be constantly attracted to it. This is exactly analogous to why the Earth doesn't fall into the Sun: it has too much angular momentum, so by the time the Sun has made it "fall" significantly, it is already on another part of its orbit. Thus the Earth (like the prospective electron) keeps "falling" in circles around the Sun.

Electrons, alas, do not feel only electrostatic forces but must comply with the full electromagnetic theory of Maxwell, which dictates that accelerating charges (like circling electrons) must radiate their energy as electromagnetic waves. This energy is taken out of the orbital motion, which would steadily collapse into the nucleus. And within a fraction of a second, too.

This was a puzzle for a very long time and it was the glaring flaw in the planetary model of the atom when it was first proposed by Rutherford and Moseley. You can only cure it by making the electron a quantum-mechanical beast, a weird hybrid between a particle and a wave.

Essentially, the electron fails to fall into the nucleus because its position, like any wave, cannot be tightly confined without giving it a very small wavelength (and that would confer it a large momentum, allowing it to break out of the nucleus). Electron waves, like all waves, like to spread out, and they can also interfere with themselves to make complicated interference patterns around the nucleus. It was then a huge triumph of theoretical physics when Schrödinger proposed an equation describing the way in which these electron waves can add with themselves constructively to make standing wave patterns whose energies exactly matched those of the planetary Bohr model and therefore experimental facts. These standing waves are the only stable states of the electron waves in the atom, which is why it doesn't collapse.

*You see, while you are learning, lots...in some ways we are only scratching the surface.*