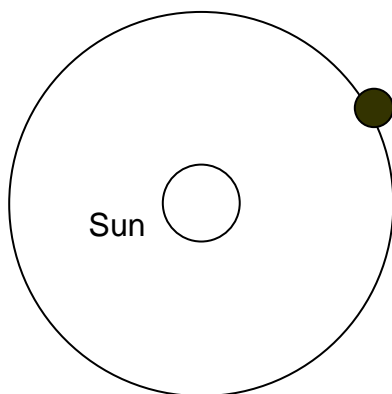


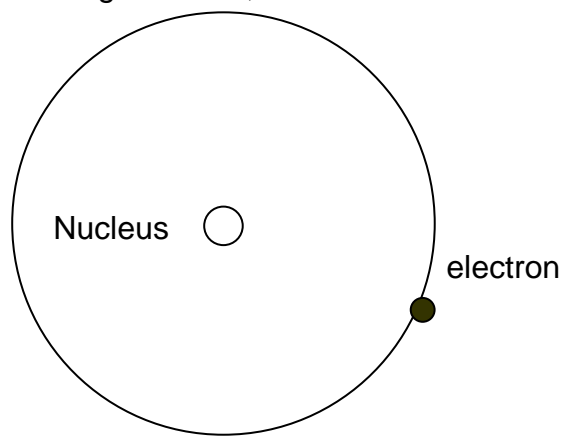
Physics 30 Lesson 31 The Bohr Model of the Atom

I. Planetary models of the atom

After Rutherford's gold foil scattering experiment, all models of the atom featured a nuclear model with electrons moving around a tiny, massive nucleus. A simple way to visualise the nuclear model was as planets orbiting a central Sun. As the Sun of our solar system attracted the planets, the positive nucleus of the atom would attract the negative electrons. While the Sun and planets involve gravitation, the nucleus and electrons involve electrostatic forces.



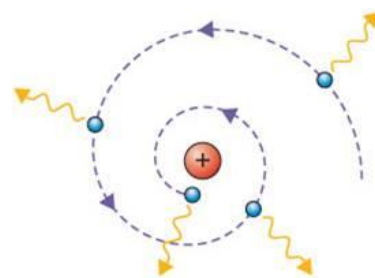
Gravitational Attraction



Planetary Model of Atom

The nuclear atom was a nice combination of gravitational ideas and sub atomic particles, but it had several major flaws:

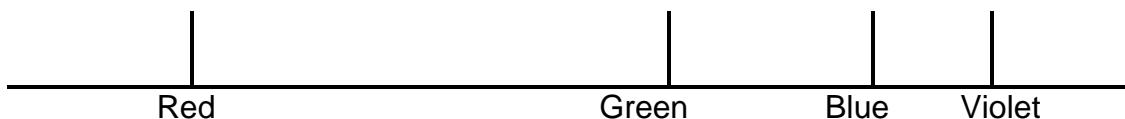
1. Did all of the electrons travel in the same orbit? Why did they not bump into one another? What was the electron structure?
2. From the known bonding characteristics of different chemical compounds, how are the electrons involved in the bonding process?
3. Why do the positive protons stay together in the nucleus? Their strong mutual repulsion should tear the nucleus apart.
4. The final, and most important flaw, concerned the nature of accelerating charges. James Maxwell had shown that accelerating electric charges radiate EM radiation (Lesson 24). If the electrons were in circular orbits, they would continually experience a centripetal **acceleration** and **should continually radiate energy in the form of electromagnetic waves**. Further, since their kinetic energy is being converted into radiant energy, the electrons should spiral into the nucleus. From observation, we know that atoms have stable structures for long periods of time and they do not radiate energy on their own.



A good model of the atom must include Rutherford's findings – i.e. that the majority of an atom's volume is empty space containing the electrons of the atom and there is a small, massive, positively charged nucleus. In addition, any model of the atom would have to account for the absorption and emission spectra of elements and molecules. Refer to Pearson pages 773 to 778.

II. Regularities in the hydrogen spectrum

Over many years, different people worked on measuring and understanding the properties of the lightest and simplest element – hydrogen. The visible bright line spectrum formed by excited hydrogen gas consists of four lines – a red line ($\lambda = 656.21 \text{ nm}$), a green line ($\lambda = 486.07 \text{ nm}$), a blue line ($\lambda = 434.01 \text{ nm}$) and a violet line ($\lambda = 410.12 \text{ nm}$).



In 1885, a mathematician named Johann Jakob Balmer found a simple empirical formula that would give the wavelengths of these four lines (an empirical formula describes the phenomenon, but it does not explain the phenomenon). His equation is:

$$\frac{1}{\lambda} = R_H \left(\frac{1}{2^2} - \frac{1}{n^2} \right)$$

where n = any integer greater than 2 and less than 7 (i.e. 3, 4, 5, 6)
 $R_H = 1.0972 \times 10^7 / \text{m}$ (later called the Rydberg constant)

When, for example, we calculate the wavelength for $n = 3$:

$$\frac{1}{\lambda} = R_H \left(\frac{1}{2^2} - \frac{1}{n^2} \right)$$

$$\frac{1}{\lambda} = 1.0972 \times 10^7 / \text{m} \left(\frac{1}{2^2} - \frac{1}{3^2} \right) = 1.524 \times 10^6 / \text{m}$$

$$\lambda = \frac{1}{1.524 \times 10^6 / \text{m}} = \mathbf{656.2 \text{ nm (red spectral line)}}$$

When $n = 4, 5$ and 6 are substituted into the equation, we obtain the wavelengths of the green, blue and violet lines respectively. Balmer's values were accurate to within 0.02%. However, being a mathematician, Balmer wanted to know why values above $n = 6$ could not be used. When values above $n = 6$ are substituted, the calculated wavelengths produced are not in the visible part of the spectrum – they fall into the **ultraviolet** section of the spectrum. It was several years before scientists, using ultraviolet sensitive film, were able to detect that the lines in the ultraviolet section actually existed.

Once the UV lines were verified, the Balmer equation was modified to handle other possibilities. The new equation was called the **Rydberg equation**:

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

where n_f = final integer (1, 2, 3 ...)
 n_i = initial integer (1, 2, 3 ...)

Note: This equation works for hydrogen only. For other atoms a different R constant is required.

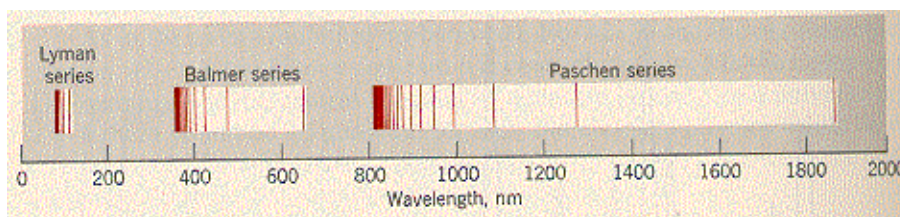
When, for example, we calculate the wavelength when $n_i = 3$ and $n_f = 1$ for the Rydberg equation:

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

$$\frac{1}{\lambda} = 1.0972 \times 10^7 / \text{m} \left(\frac{1}{1^2} - \frac{1}{3^2} \right) = 9.753 \times 10^6 / \text{m}$$

$$\lambda = \frac{1}{9.753 \times 10^6 / \text{m}} = \mathbf{102.5 \text{ nm (light is ultra violet)}}$$

The work by Balmer and Rydberg encouraged other scientists to search for spectral lines outside the normal visible light range. In 1908, for example, F. Paschen found a whole series of spectral lines generated in the **infrared** section of the spectrum for excited hydrogen gas. The lines corresponded to the values that you obtain when $n_f = 3$ and n_i is give values of 4, 5, 6, 7 etc. In 1914, Lyman found spectral lines in the ultra violet part of the spectrum. Lyman's lines corresponded to $n_f = 1$ and n_i is give values of 2, 3, 4, 5, etc. Further research by other scientists led to the detection of other spectral line series. In each case the detection of the series came after the mathematical calculation of the values using the Rydberg equation. (See Pearson page 776.)



The fact that a precise, and fairly simple, mathematical relationship described the wavelengths of spectral lines generated by hydrogen gas indicated that the spectral lines were governed by some physical relationship or law. Niels Bohr would use the production of spectral lines for hydrogen to enhance and refine his model of the atom.

Note, you are not required to know the Balmer or Rydberg equations, they were used above for instructive purposes only. You are required to know that hydrogen has a spectrum that can be described by a simple mathematical relationship.

III. Bohr's postulates

Niels Bohr received his PhD in 1911 in the field of Physics from the University of Copenhagen. In 1912, he spent a year working under Ernest Rutherford at the University of Manchester in England. In the same year, Bohr would propose a model of the hydrogen atom that incorporated:

- ⇒ J. J. Thomson's cathode ray tube experiments
- ⇒ Rutherford's nuclear model
- ⇒ Planck's quantum theory
- ⇒ Einstein's photon theory
- ⇒ the emission and absorption spectra for hydrogen
- ⇒ Rydberg's equation for predicting the absorption wavelengths of hydrogen
- ⇒ the Franck-Hertz experimental results

In order to account for the existence of stable electron orbits and separate emission spectra, Bohr made three major assumptions called the **postulates**. (A postulate is an assumed idea upon which a theory is based.) The postulates are:

1. Electrons move in circular orbits, but they do not radiate energy. These orbits are called **stationary states**.
2. **Electrons can “jump” from one stationary state to another, but they cannot exist in between them.** When an electron absorbs energy it jumps up in energy level. When an electron releases energy it jumps down in energy level. The energy emitted or absorbed has a frequency determined by the relation
$$h f = | E_f - E_i |$$
3. Of all possible orbits around the nucleus, only a few are allowed. Each orbit has a characteristic **energy** and **radius** given by the following equations. (See Pearson pages 774 and 775.)

$$E_n = \frac{E_1}{n^2}$$

Where $n = 1, 2, 3, 4, \dots$

n is the **principal quantum number**.

$$r_n = n^2 r_1$$

For the ground state $n = 1$.

For hydrogen:

$$E_1 = -13.6 \text{ eV or } -2.18 \times 10^{-18} \text{ J}$$

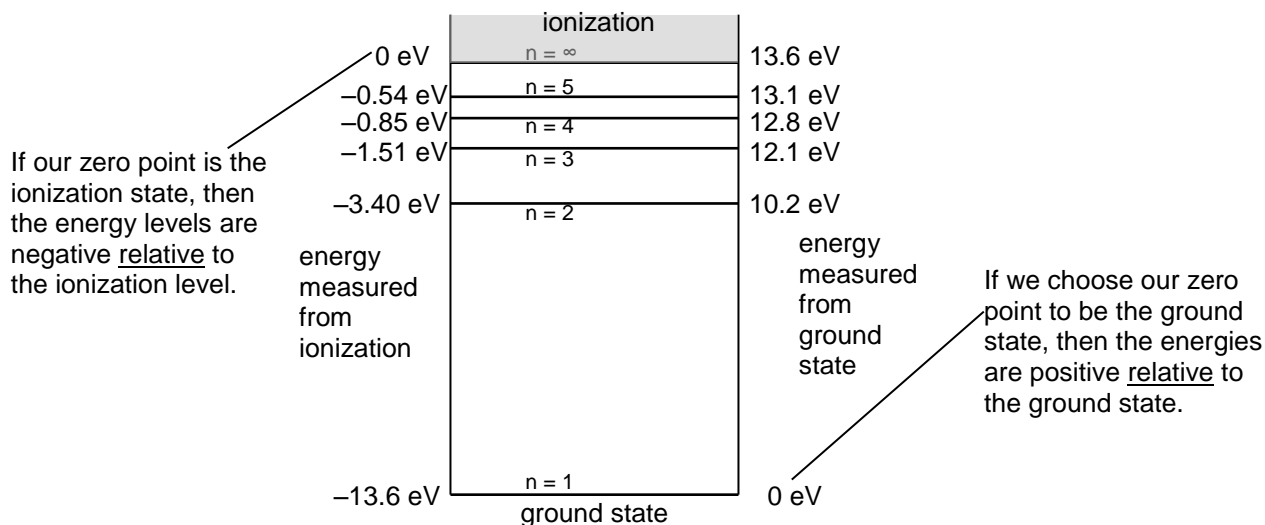
$$r_1 = 5.29 \times 10^{-11} \text{ m}$$

Bohr's postulates were intended to work for all atoms, not just hydrogen, but it did not work out that way. For multi-electron atoms, the interactions between electrons require a far more sophisticated and involved idea – see Lesson 34.

It may be of interest to note that Bohr was not very good at writing down his ideas. His PhD was delayed by many years because he had developed such a fear of writing. Finally, he dictated his ideas to his wife who then wrote the necessary thesis. Throughout his professional life, Bohr found writing to be very painful. When you read Albert Einstein's papers there is an eloquent flow of words and ideas. When you try to read Bohr's papers, you are fortunate if you get beyond the first paragraph. When asked why he had such difficulties, he said, "Grammar teachers."



Below is an energy level diagram for hydrogen. Note that the sign of the energy depends on where we place our zero value.



Note, this is a very important diagram to understand.

In order to jump from one energy level to another, the atom had to **absorb** or **emit** the difference in the energy levels.

$$\Delta E = |E_f - E_i|$$

$$hf = |E_f - E_i|$$

$$\frac{hc}{\lambda} = |E_f - E_i|$$

Example 1

An electron drops from the fourth energy level of hydrogen to the second energy level.

a) What is the energy released?

$$\Delta E = |E_f - E_i|$$

$$\Delta E = |-3.4\text{eV} - (-0.85\text{eV})|$$

$$\Delta E = 2.55\text{eV}$$

b) What is the frequency and wavelength of the emitted photon?

$$\Delta E = hf$$

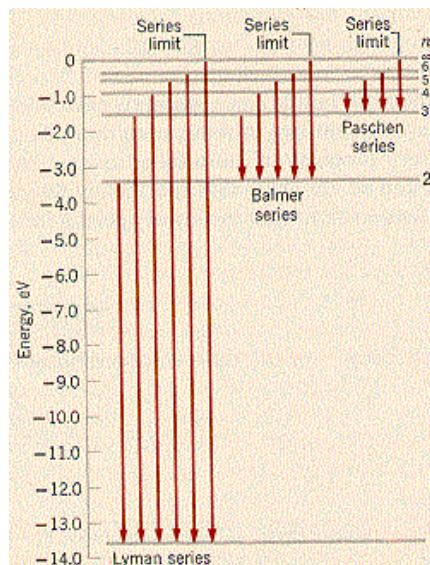
$$f = \frac{\Delta E}{h} = \frac{2.55\text{eV}}{4.14 \times 10^{-15}\text{eV}\cdot\text{s}} = 6.16 \times 10^{14}\text{ Hz}$$

$$\lambda = \frac{c}{f} = \frac{3.00 \times 10^8\text{m/s}}{6.17 \times 10^{14}\text{Hz}} = 487\text{ nm}$$

which corresponds to the green line in the emission spectrum of the Balmer series

The Bohr model of the atom explains why absorption and emission line spectra occur for hydrogen and other elements. The atom can only absorb frequencies/energies of light that correspond to differences between the atom's energy levels, resulting in dark line absorption spectra. Similarly, atoms can only emit photons that have energies that correspond to energy transitions from higher to lower energy states.

We will occasionally work with other fictitious "elements" which we will treat in a Bohr-like fashion.



IV. Strengths of the Bohr model of the atom

1. Bohr's model of the atom could explain the size of the hydrogen atom. The radius calculated corresponded to known values of the hydrogen atom.
2. Bohr's model gave an ionization value ($-E_1$) that corresponded to the known ionization value for the hydrogen atom.
3. Bohr's model could account for the formation of the spectral lines in the hydrogen spectra. While Balmer, Lyman, Paschen, and others had used a mathematical relationship to calculate the known wavelengths, the relationship had no scientific basis. Bohr's atom gave the same results and a strong scientific basis for the observed emission and absorption spectral lines.
4. Bohr's model could be expanded with some modifications to account for a) larger charges in the nucleus and b) possible shielding by inner electrons on outer electrons in larger atoms in the periodic table. When Bohr applied his theories to the periodic table as a whole, he was able to explain why elements were grouped vertically according to chemical and physical properties. He found all the elements in a vertical column had the same number of electrons in their outer most energy level. Bohr also found that the energy level quantum number corresponded to the horizontal row (period) of the periodic table. Bohr's atom provided greater understanding for the workings of the periodic table designed by Mendeleev and Moseley.
5. Bohr's atom used the nucleus idea of Rutherford but now gave an explanation for the structure of the electron cloud.

V. Problems with the Bohr model of the atom

1. Bohr's model only works well for hydrogen. It must be tremendously modified to accommodate other elements with larger nuclei and more electrons.
2. Bohr's model could not explain why the number of spectral lines increased for some elements when they were placed in electric and magnetic fields.
3. Bohr's model was not able to explain the relative intensity of some of the spectral lines (i.e. why some lines were brighter than others).

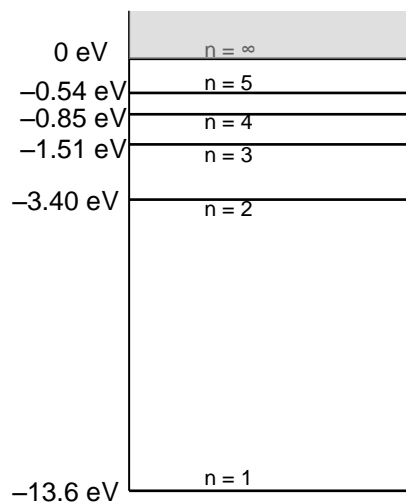
Nevertheless Bohr's model of the atom was a spectacular success. His model is still used as the starting point for teaching the atom in the field of chemistry. His model will explain many of the known properties of the group 1,2, 13-18 elements. The Bohr model breaks down for transition elements and members of the Lanthanide and Actinide series.

VI. Hand-in assignment

1. What is the major problem with Rutherford's idea of the planetary atom?
2. Photon emission:
 - A. What happens to electrons when an electric current is passed through a gas?
 - B. What is meant by an electron being in an excited state?
 - C. What happens when an electron jumps down to a lower energy level?
 - D. How does photon emission explain the emission spectrum of a gas?
3. Describe how the emission and absorption spectra of hydrogen are explained by the Bohr model of the atom.
4. What are the strengths of the Bohr model? What are the weaknesses of the Bohr model?

Use the energy level diagram for hydrogen to answer questions 5 to 11.

5. How much energy must a hydrogen electron in energy level $n = 1$ absorb in order to jump to energy level $n = 2$ and $n = 3$? (10.2 eV, 12.1 eV)
6. How does the total energy of photons emitted by an electron as it jumps back down to its ground state compare with the energy absorbed by the electron when it jumped to the higher level?
7. Why is a blue photon emitted when a hydrogen electron jumps from energy level $n = 5$ to $n = 2$, but a red photon is emitted when an electron jumps from level $n = 3$ to $n = 2$?



8. What is the ionization energy for the electron in hydrogen from its ground state?
9. What is the wavelength of a photon emitted when an electron in a hydrogen atom falls:
 - A. From $n = 4$ to $n = 2$. (487 nm)
 - B. From $n = 5$ to $n = 1$. (95.1 nm)
10. Electrons are accelerated through hydrogen gas at room temperature in a Franck-Hertz experiment by a potential difference of 12.3 V. What wavelengths of light can be expected to be emitted by the hydrogen? (122 nm, 103 nm, 657 nm)
11. What energy is needed to ionize hydrogen from the $n=2$ state? How likely is this to occur? Explain. (3.40 eV)

12. Use the energy level diagram below to answer the following questions.

$n = \infty$	_____	0.000 eV
$n = 9$	_____	-0.450 eV
$n = 8$	_____	-0.570 eV
$n = 7$	_____	-0.744 eV
$n = 6$	_____	-1.013 eV
$n = 5$	_____	-1.458 eV
$n = 4$	_____	-2.278 eV
$n = 3$	_____	-4.050 eV
$n = 2$	_____	-9.113 eV
$n = 1$	_____	-36.450 eV

- A. What is the wavelength of the photon emitted when an electron falls from the sixth to second energy level? (153 nm)
- B. What is the frequency of the photon emitted when an electron falls from $n = 7$ to $n = 3$? (7.987×10^{14} Hz)
- C. What is the energy absorbed by an electron to jump from $n = 1$ to $n = 8$? (35.880 eV)
- D. What is the ionization energy of this element?