10/27/17

Honors Student,

I have prepared this packet to assist you with the understanding of how atoms can be made to release light, and how line spectrum data was used by Neils Bohr to develop his model of the atom.

Support resources are available at a website I am developing for Honors Chemistry:
https://pthshonorschemistry.weebly.com/
On the home page, choosing "Semester 1 Units," then "Chapter Four Electron Arrangement" will take you to the specific information for this unit. (https://pthshonorschemistry.weebly.com/chpt-4---electron-arrangement.html)

In this packet you will first find pages from two Chemistry Textbooks that explain the Bohr Model of the Atom, and how Bohr created his model from line spectrum data. Please read through these pages, paying special attention to the graphics shown. Using what we have discussed in class, what you have read, and what you have possibly watched on the website, answer the questions I have on the following pages in the packet.

Do your best and do your own work (do not copy from a classmate). I will address any questions and clarify any confusion in class on Monday. We will test on Chapter Four the week of November 6th.

Chemistry - Matter & Change
Glencoe Science

Atomic Emission Spectra

Have you ever wondered how light is produced in the glowing tubes of neon signs? This process is another phenomenon that cannot be explained by the wave model of light. The light of the neon sign is produced by passing electricity through a tube filled with neon gas. Neon atoms in the tube absorb energy and become excited. These excited atoms return to their stable state by emitting light to release that energy. If the light emitted by the neon is passed through a glass prism, neon's atomic emission spectrum is produced. The **atomic emission spectrum** of an element is the set of frequencies of the electromagnetic waves emitted by atoms of the element. Neon's atomic emission spectrum consists of several individual lines of color corresponding to the frequencies of the radiation emitted by the atoms of neon. It is not a continuous range of colors, as in the visible spectrum of white light.

Reading Check **Explain** how an emission spectrum is produced.

Each element's atomic emission spectrum is unique and can be used to identify an element or determine whether that element is part of an unknown compound. For example, when a platinum wire is dipped into a strontium nitrate solution and then inserted into a burner flame, the strontium atoms emit a characteristic red color. You can perform a series of flame tests by doing the MiniLab.

Figure 5.8 shows an illustration of the characteristic purple-pink glow produced by excited hydrogen atoms and the visible portion of hydrogen's emission spectrum responsible for producing the glow. Note how the line nature of hydrogen's atomic emission spectrum differs from that of a continuous spectrum.

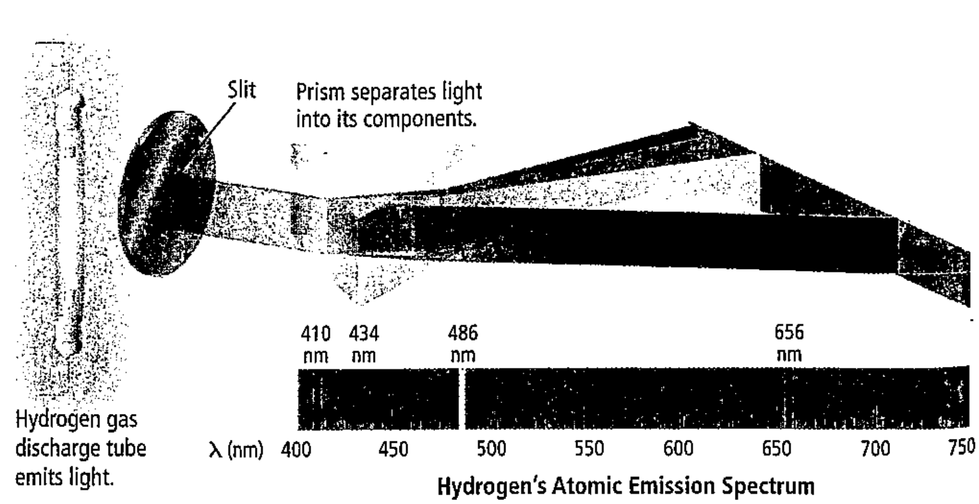


Figure 5.8 The purple light emitted by hydrogen can be separated into its different components using a prism. Hydrogen has an atomic emission spectrum that comprises four lines of different wavelengths.
Determine *Which line has the highest energy?*

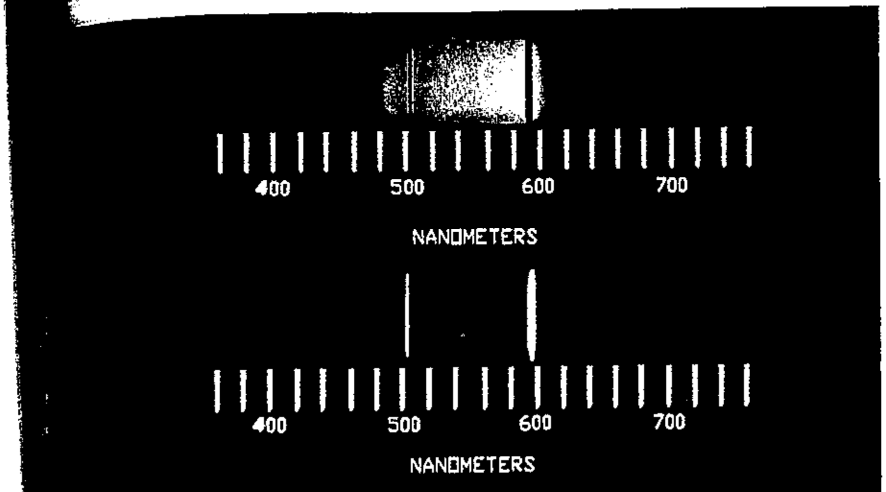


Figure 5.9 The first spectrum is an absorption spectrum. It is composed of black lines on a continuous spectrum. The black lines correspond to certain frequencies absorbed by a given element, helium in this case. They can be matched to the colored lines present in helium's emission spectrum, shown below the absorption spectrum.

onnection to Astronomy An atomic emission spectrum is character-ic of the element being examined and can be used to identify that ment. The fact that only certain colors appear in an element's atomic ission spectrum means that only specific frequencies of light are itted. Because those emitted frequencies are related to energy by the mula $E_{photon} = h\nu$, only photons with specific energies are emitted. is was not predicted by the laws of classical physics. Scientists had ected to observe the emission of a continuous series of colors as ited electrons lost energy. Elements absorb the same specific fre-encies of light as the frequencies they emit, thus creating an absorp-n spectrum. In an absorption spectrum, the absorbed frequencies ear as black lines, as shown in **Figure 5.9.** By comparing the black es to the emission spectrum of elements, scientists are able to deter-ne the composition of the outer layers of stars.

Section 5.2

Objectives

- **Compare** the Bohr and quantum mechanical models of the atom.
- **Explain** the impact of de Broglie's wave-particle duality and the Heisenberg uncertainty principle on the current view of electrons in atoms.
- **Identify** the relationships among a hydrogen atom's energy levels, sublevels, and atomic orbitals.

Review Vocabulary

atom: the smallest particle of an element that retains all the properties of that element; is composed of electrons, protons, and neutrons

New Vocabulary

ground state
quantum number
de Broglie equation
Heisenberg uncertainty principle
quantum mechanical model of the atom
atomic orbital
principal quantum number
principal energy level
energy sublevel

Quantum Theory and the Atom

MAIN Idea Wavelike properties of electrons help relate atomic emission spectra, energy states of atoms, and atomic orbitals.

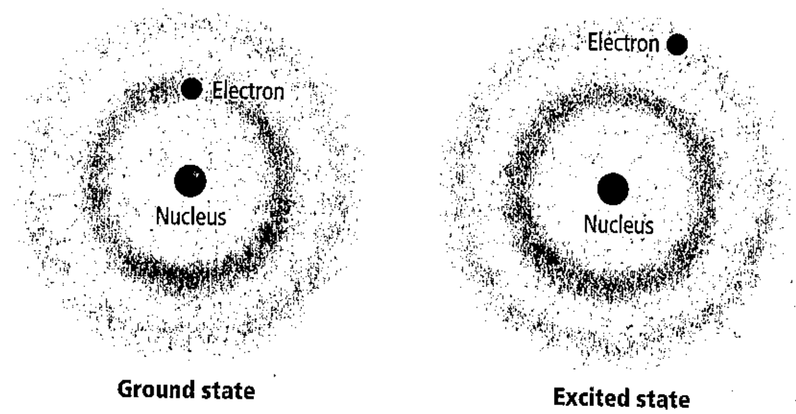
Real-World Reading Link Imagine climbing a ladder and trying to stand between the rungs. Unless you could stand on air, it would not work. When atoms are in various energy states, electrons behave in much the same way as a person climbing up the rungs of a ladder.

Bohr's Model of the Atom

The dual wave-particle model of light accounted for several previously unexplainable phenomena, but scientists still did not understand the relationships among atomic structure, electrons, and atomic emission spectra. Recall that hydrogen's atomic emission spectrum is discontinuous; that is, it is made up of only certain frequencies of light. Why are the atomic emission spectra of elements discontinuous rather than continuous? Niels Bohr, a Danish physicist working in Rutherford's laboratory in 1913, proposed a quantum model for the hydrogen atom that seemed to answer this question. Bohr's model also correctly predicted the frequencies of the lines in hydrogen's atomic emission spectrum.

Energy states of hydrogen Building on Planck's and Einstein's concepts of quantized energy, Bohr proposed that the hydrogen atom has only certain allowable energy states. The lowest allowable energy state of an atom is called its **ground state.** When an atom gains energy, it is said to be in an excited state.

Bohr also related the hydrogen atom's energy states to the electron within the atom. He suggested that the electron in a hydrogen atom moves around the nucleus in only certain allowed circular orbits. The smaller the electron's orbit, the lower the atom's energy state, or energy level. Conversely, the larger the electron's orbit, the higher the atom's energy state, or energy level. Thus, a hydrogen atom can have many different excited states, although it contains only one electron. Bohr's idea is illustrated in **Figure 5.10.**



■ **Figure 5.10** The figure shows an atom that has one electron. Note that the illustration is not to scale. In its ground state, the electron is associated with the lowest energy level. When the atom is in an excited state, the electron is associated with a higher energy level.

Table 5.1 Bohr's Description of the Hydrogen Atom

Bohr's Atomic Orbit	Quantum Number	Orbit Radius (nm)	Corresponding Atomic Energy Level	Relative Energy
First	$n = 1$	0.0529	1	E_1
Second	$n = 2$	0.212	2	$E_2 = 4E_1$
Third	$n = 3$	0.476	3	$E_3 = 9E_1$
Fourth	$n = 4$	0.846	4	$E_4 = 16E_1$
Fifth	$n = 5$	1.32	5	$E_5 = 25E_1$
Sixth	$n = 6$	1.90	6	$E_6 = 36E_1$
Seventh	$n = 7$	2.59	7	$E_7 = 49E_1$

In order to complete his calculations, Bohr assigned a number, n, called a **quantum number,** to each orbit. He also calculated the radius of each orbit. For the first orbit, the one closest to the nucleus, $n = 1$ and the orbit radius is 0.0529 nm; for the second orbit, $n = 2$ and the orbit radius is 0.212 nm; and so on. Additional information about Bohr's description of hydrogen's allowed orbits and energy levels is given in **Table 5.1.**

The hydrogen line spectrum Bohr suggested that the hydrogen atom is in the ground state, also called the first energy level, when its single electron is in the $n = 1$ orbit. In the ground state, the atom does not radiate energy. When energy is added from an outside source, the electron moves to a higher-energy orbit, such as the $n = 2$ orbit shown in **Figure 5.11.** Such an electron transition raises the atom to an excited state. When the atom is in an excited state, the electron can drop from the higher-energy orbit to a lower-energy orbit. As a result of this transition, the atom emits a photon corresponding to the energy difference between the two levels.

$$\Delta E = E_{\text{higher-energy orbit}} - E_{\text{lower-energy orbit}} = E_{\text{photon}} = h\nu$$

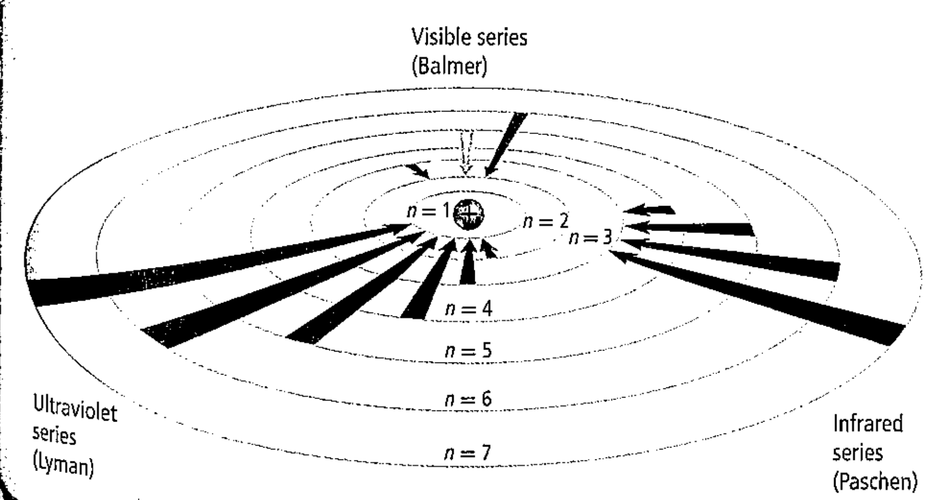


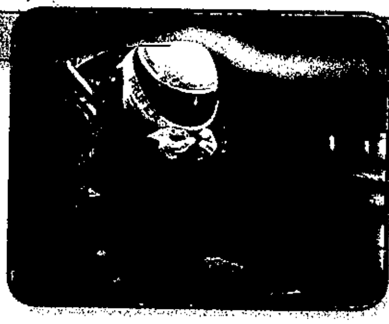
Figure 5.11 When an electron drops from a higher-energy orbit to a lower-energy orbit, a photon is emitted. The ultraviolet (Lyman), visible (Balmer), and infrared (Paschen) series correspond to electrons dropping to $n = 1$, $n = 2$, and $n = 3$, respectively.

Concepts In Motion

Interactive Figure To see an animation of the Balmer Series, visit .

Chemistry
Pearson

5.1 Revising the Atomic Model



CHEMISTRY & YOU

Q: *Why do scientists use mathematical models to describe the position of electrons in atoms?* Wind tunnels and models are often used to simulate the forces from the moving air on a design. Shown here is a life-sized model of a speed skier. It is a physical model. However, not all models are physical. In fact, the current model of the atom is a mathematical model.

Key Questions

What did Bohr propose in his model of the atom?

What does the quantum mechanical model determine about the electrons in an atom?

How do sublevels of principal energy levels differ?

Vocabulary

- energy level
- quantum
- quantum mechanical model
- atomic orbital

Energy Levels in Atoms

What did Bohr propose in his model of the atom?

Thus far, the atomic model presented in this textbook has considered atoms as consisting of protons and neutrons making up a nucleus surrounded by electrons. After discovering the atomic nucleus, Rutherford used existing ideas about the atom and proposed an atomic model in which the electrons move around the nucleus like the planets move around the sun.

Limitations of Rutherford's Atomic Model Rutherford's atomic model explained only a few simple properties of atoms. It could not explain the chemical properties of elements. For example, Rutherford's model could not explain why metals or compounds of metals give off characteristic colors when heated in a flame. It also could not explain why an object such as the iron scroll shown in Figure 5.1 first glows dull red, then yellow, and then white when heated to higher and higher temperatures. Explaining what leads to the chemical properties of elements required a model that better described the behavior of electrons in atoms.

Figure 5.1 Glowing Metal
Rutherford's model failed to explain why objects change color when heated. As the temperature of this iron scroll is increased, it first appears black, then red, then yellow, and then white. The observed behavior could be explained only if the atoms in the iron gave off light in specific amounts of energy. A better atomic model was needed to explain this observation.

The Bohr Model In 1913, Niels Bohr (1885–1962), a young Danish physicist and a student of Rutherford, developed a new atomic model. He changed Rutherford's model to incorporate newer discoveries about how the energy of an atom changes when the atom absorbs or emits light. He considered the simplest atom, hydrogen, which has one electron. **Bohr proposed that an electron is found only in specific circular paths, or orbits, around the nucleus.**

Each possible electron orbit in Bohr's model has a fixed energy. The fixed energies an electron can have are called **energy levels.** The fixed energy levels of electrons are somewhat like the rungs of the ladder in Figure 5.2a. The lowest rung of the ladder corresponds to the lowest energy level. A person can climb up or down the ladder by stepping from rung to rung. Similarly, an electron can move from one energy level to another. A person on the ladder cannot stand between the rungs. Similarly, the electrons in an atom cannot exist between energy levels. To move from one rung to another, a person climbing the ladder must move just the right distance. To move from one energy level to another, an electron must gain or lose just the right amount of energy. A **quantum** of energy is the amount of energy required to move an electron from one energy level to another energy level. The energy of an electron is therefore said to be quantized.

READING SUPPORT

Build Vocabulary: *Latin Word Origins* *Quantum* comes from the Latin word *quantus*, meaning "how much." ***What other commonly used English word comes from this root?***

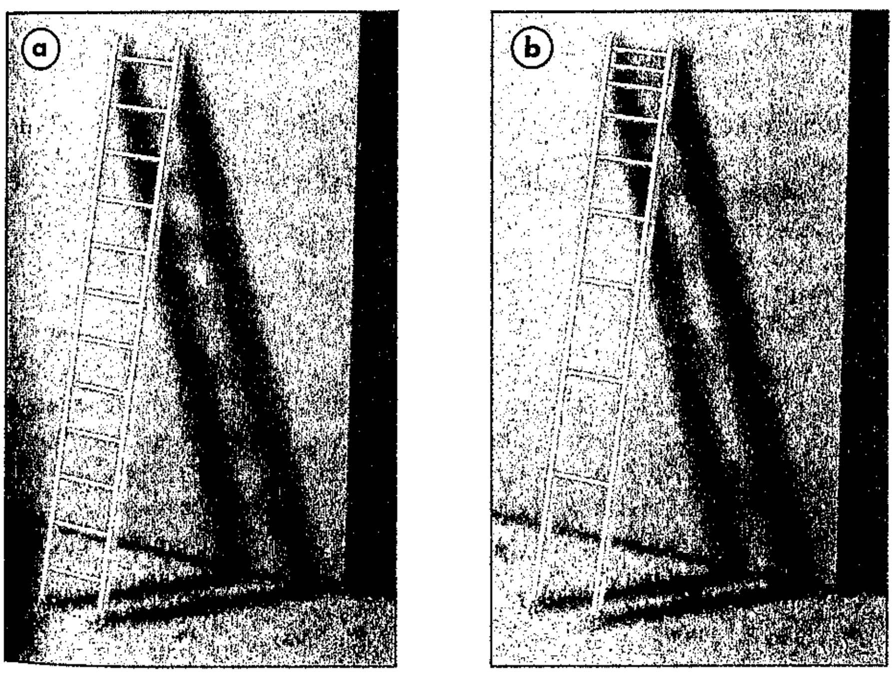


Figure 5.2 Energy Levels
The rungs of a ladder are somewhat like the energy levels in Bohr's model of the atom. **a.** In an ordinary ladder, the rungs are equally spaced. **b.** The energy levels in atoms are unequally spaced, like the rungs in this unusual ladder. The higher energy levels are closer together.
Compare ***For the ladder in* b*, compare the amount of energy it would take to move from the first rung to the second rung with the amount of energy it would take to move from the second rung to the third rung.***

The amount of energy an electron gains or loses in an atom is not always the same. Like the rungs of the strange ladder in Figure 5.2b, the energy levels in an atom are not equally spaced. The higher energy levels are closer together. It takes less energy to climb from one rung to another near the top of the ladder in Figure 5.2b, where the rungs are closer. Similarly, the higher the energy level occupied by an electron, the less energy it takes the electron to move from that energy level to the next higher energy level.

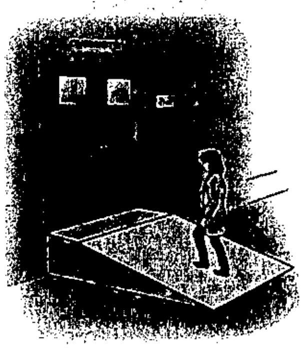
The Bohr model provided results in agreement with experiments using the hydrogen atom. However, the Bohr model failed to explain the energies absorbed and emitted by atoms with more than one electron.

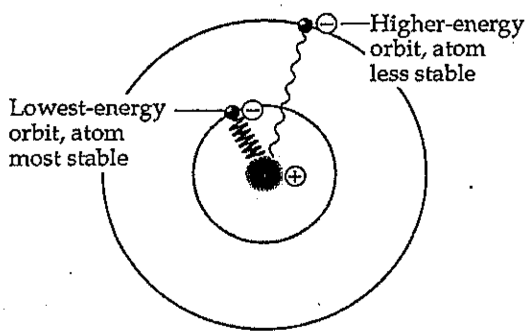
BOHR MODEL OF THE ATOM

The **ramp** is an example of a continuous situation ~~of a continuous situation~~ in which any energy state is possible up the ramp.

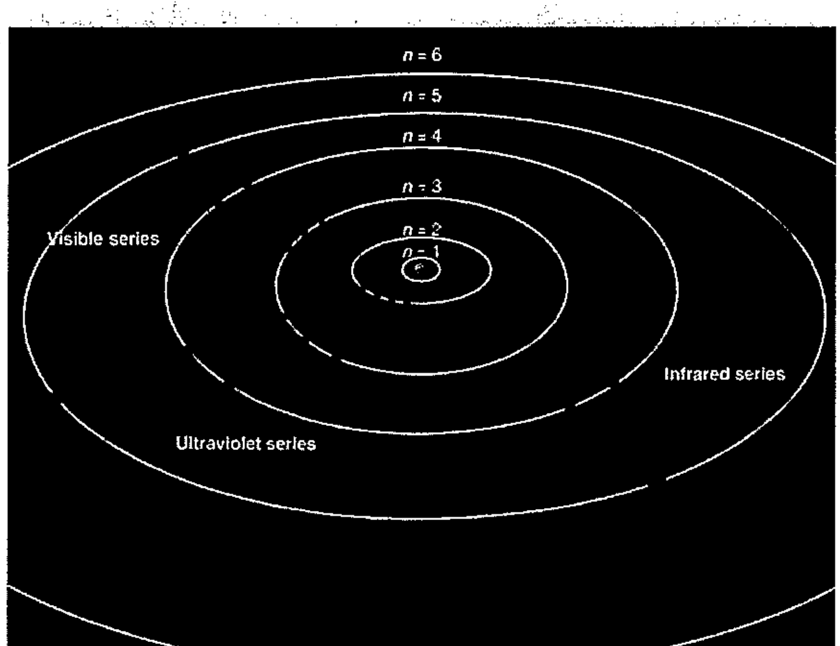
Like a set of stairs, the energy states of an electron is quantized

– i.e. electrons are only found on a specific step





Bohr Model of Hydrogen



The Atomic Model Through Time

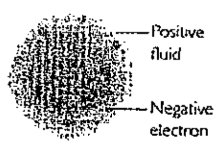
EXPERIMENT 1
In 1803, John Dalton studied how elements combine chemically to form compounds. He observed that elements combine in whole-number ratios to form compounds and that matter is not created or destroyed in chemical reactions. Dalton reasoned that elements are made of tiny, indivisible spherical particles called atoms.

SOLID SPHERE MODEL

The atom is a solid sphere that cannot be divided up into smaller particles or pieces.

EXPERIMENT 2
In 1897, J. J. Thomson, a British scientist, zapped atoms with electricity. He observed that negatively charged particles were removed. Thomson reasoned that atoms contain negatively charged particles, which he called electrons.

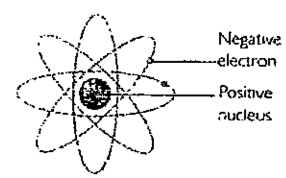
PLUM PUDDING MODEL



The atom can be divided into a fluid (the "pudding") and electrons. Most of the atom is made of fluid. The fluid spreads out in the atom and is positively charged. The electrons are very tiny and negatively charged.

EXPERIMENT 3
In 1911, Ernest Rutherford, a New Zealand-born scientist, shot tiny positively charged particles, called alpha particles, at thin gold foil. He observed that most of the alpha particles went through the foil, but a few bounced back. Rutherford reasoned that there must be something small, massive, and positively charged in an atom, which he called the nucleus.

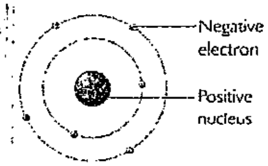
NUCLEAR MODEL



The atom can be divided into a nucleus and electrons. The nucleus occupies a small amount of space in the center of the atom. The nucleus is dense and positively charged. The electrons circle around the nucleus. The electrons are tiny and negatively charged. Most of the atom is empty space.

EXPERIMENT 4
In 1913, Neils Bohr, a Danish scientist, developed a model of the atom that explained the light given off when elements are exposed to flame or electric fields. He observed that only certain colors of light are given off. For example, hydrogen atoms give off red, blue-green, and blue light. Bohr reasoned that the electrons orbit around the nucleus at different distances like planets orbiting the Sun. The electrons in these orbits have different energies. When an electron falls from an outer to an inner orbit, the color of the light given off depends on the energies of the two orbits.

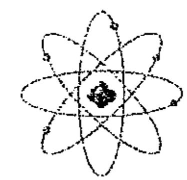
SOLAR SYSTEM MODEL



The atom can be divided into a nucleus and electrons. The nucleus is in the center of the atom. The nucleus is massive and positively charged. The electrons circle around the nucleus in specified orbits. The electrons are tiny and negatively charged. Different electrons are in orbits at different distances from the nucleus.

EXPERIMENT 5
In 1918, Rutherford made a further contribution. He found he could use alpha particles as bullets to knock off small positively charged particles, which he called protons. He reasoned that the nucleus must be a collection of protons.

PROTON MODEL

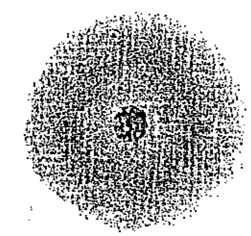


The nucleus contains protons. The protons are tiny and positively charged. The electrons circle around the nucleus. The electrons are tiny and negatively charged. Most of the atom is empty space.

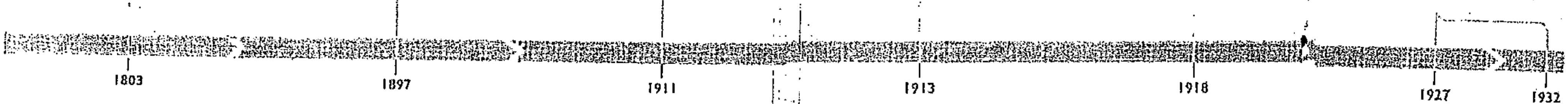
EXPERIMENT 6
In 1927, Werner Heisenberg, a German scientist, proposed a cloud model of the atom. Heisenberg suggested that the location of an electron could not be specified precisely. Instead, it is only possible to talk about the probability of where an electron might be. This led to a cloud model of the atom; the electron cloud indicates where you will most likely find a single electron.

EXPERIMENT 7
In 1932, a British physicist, James Chadwi[ck] found that the nucleus also included uncharged, or neutral, particles, which he called neutrons. He reasoned that the neutrons were important in holding the positively charged protons together.

ELECTRON CLOUD MODEL



An electron cloud surrounds the nucleus. The cloud is made up of fast-moving electrons:
The nucleus is made up of protons: and neutrons:



* Beginning of Assignment *

In your own words, explain

HOW DO ATOMS PRODUCE LIGHT?

In your response, include the following terms:

ground state, excited state, relax, energy, electrons, visible light, nucleus.

Name ______________________ Class __________ Date __________

Refinements of the Atomic Model

Section Review **4.1**

DIRECTIONS: Write on the line at the right of each statement the letter preceding the word or expression that best completes the statement.

Please answer (all questions that are not crossed out)

1. Electromagnetic radiation has some properties of particles when it (a) travels through space; (b) is transferred to matter; (c) interacts with photons; (d) interacts with other radiations. ______ 1
2. The wave model of light was not able to explain (a) light's frequency; (b) the continuous spectrum; (c) interference; (d) the photoelectric effect. ______ 2
3. In wave motion, the product of frequency and wavelength is equal to (a) the number of waves passing a given point in a second; (b) the speed of the wave; (c) the distance between successive wave crests; (d) the time for one full wave to pass a given point. ______ 3
4. The common characteristic shown by X rays, visible light, infrared radiation, and radio waves is that they all have the same (a) energy; (b) wavelength; (c) speed; (d) frequency. ______ 4
5. Red light has a longer wavelength than blue light. Compared to the blue line on the hydrogen spectrum, the red line would represent (a) higher energy and lower frequency; (b) higher energy and higher frequency; (c) lower energy and higher frequency; (d) lower energy and lower frequency. ______ 5
6. A line spectrum is produced when an electron moves from one energy level (a) to a higher energy level; (b) to a lower energy level; (c) into the nucleus; (d) to another position in that same sublevel. ______ 6
7. ~~The drop of an electron from a high energy level to the ground state in a hydrogen atom would be most closely associated with (a) long wavelength radiation; (b) low frequency radiation; (c) infrared radiation; (d) high frequency radiation.~~ ______
8. The change of an atom from excited state to ground state always involves (a) absorption of energy; (b) emission of electromagnetic radiation; (c) release of visible light; (d) an increase in electron energy. ______ 8
9. An orbital may be defined as (a) the most stable state of an atom; (b) the circular path followed by an electron around the nucleus; (c) the positively charged central part of an atom; (d) a highly probable location of an electron within the atom. ______ 9
10. The quantum model of the atom locates the electron (a) at a specific distance from the nucleus; (b) in a definite path around the nucleus; (c) within a region of high probability; (d) at any distance from the nucleus. ______ 10
11. The size and shape of an electron cloud is most closely related to the electron's (a) charge; (b) mass; (c) spin; (d) energy. ______ 11

DIRECTIONS: Complete the following statements, forming accurate sentences.

12. ~~A quantum of electromagnetic energy is called a(n)~~ ______________________ 12
13. ~~The spectral lines of hydrogen that occur in the ultraviolet region of the electromagnetic spectrum are called the~~ ______________________ 13
14. ~~An optical instrument that separates light entering it into component wavelengths is a(n)~~ ______________________ 14
15. The lowest total energy of the electron in a hydrogen atom occurs when the electron is in the state called the ______________________ 15

Honors Chem – Chpt 4

1. How did Bohr expand on Rutherford's model of the atom?

2. Compare the energy of an electron in the ground state and an electron in the excited state.

3. Explain how the gaseous neon atoms in a neon sign emit light.

4. List the six colors of the visible spectrum in order of ***increasing*** energy.

5. Determine the type of radiation (gamma rays, infrared waves, or radio waves) that has the:

   a. longest wavelength ______________________
   b. highest frequency ______________________
   c. greatest energy ______________________

6. What is the wavelength of electromagnetic radiation having a frequency of 5.00×10^{12} s^{-1}? What kind of electromagnetic radiation is this?

7. The laser in a CD player uses light with a wavelength of 7.70×10^{-7} m (780 nm). What is the frequency of this light?