Advanced Placement Physics "B" Date: ________________

Mr. Mirro

Day #138 Period #205
Day #138 Period #206

Type: Double – Theory

Objective:

Teaching notes:

First students will read through and discuss class notes, then students will work in teams to jigsaw the Web Quest Activity below.

Web Quest Activity:

http://mhsweb.ci.manchester.ct.us/Library/webquests/atomicmodels.htm

Website(s) for enrichment:

http://www.classzone.com/books/earth_science/terc/content/investigations/es0501/es0501page01.cfm

http://www.thebigview.com/spacetime/quantumtheory.html

Emission Spectra: http://phys.educ.ksu.edu/vqm/html/emission.html

Bohr's Model: http://www.walter-fendt.de/ph14e/bohrh.htm

Lastly, have students make diagrams, drawings, white boards, overheads, and/or computer paint style images to display and discuss for last 20 min of period.

Note: This activity will take both periods.

Homework: Assignment (1- 105)

Read (30.1 through 30.3)

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Atomic Models I

Atomic Models: **Models of the Atom, Emission Line Spectra, Balmer Series, Lyman Series, Paschen Series, Wave Number, Rydberg Constant**

INTRODUCTION:

The impact of *modern physics* is most evident in the development of the *atomic model* of *matter*. Although the concept of the atom goes back to ancient Greece and Rome, the *nuclear atom* is a relatively new idea. It wasn't until the *20th century* that the *structure* of the atom was reasonably well *understood*.

- We use the term *atomic model* to indicate that we are trying to *describe* the *features* of the *atom*.
- The truth is, that we really do NOT know what an atom looks like because we have no instruments for its direct observation.

A model is like a road map; it gives us information to get from place to place, but it does not describe the actual landscape.

EARLY MODELS:

Early ideas about the existence of atoms are usually attributed to the philosophers of Ancient Greece. The hard core of this model contains two ideas:

i. matter is composed of very small indivisible corpuscles - probably first called 'atoms' by Leukippos.

ii. atoms are infinitely hard and differ in form, order, position, and size.

The influence of the Greek atomic model was persistent. During the seventeenth and eighteenth centuries, many scientists such as Frances Bacon, Rene Descartes, Pierre Gassendi, Robert Boyle, and Isaac Newton discussed the constitution of matter. With the exception of Descartes, who thought matter was continuous and infinitely divisible, all of them accepted one or both of the core ideas of the Ancient Greek model.

One major problem with the Ancient Greek model was that it provided no basis for distinguishing between types of atom, for example all the different elements.

DALTON'S MODEL:

The *first scientific model* of the atom was conceived by English chemist and physicist John Dalton at the beginning of the 19th century. Dalton's model allowed for the existence of the different elements but he carried forward the idea from the ancients that the atom was an indivisible entity.

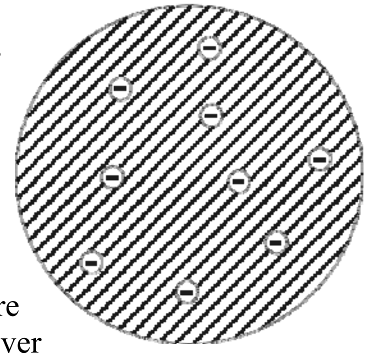
Dalton's model explained the *mathematics* of *chemical combinations*, but the *internal structure* of the atom and the *forces* between atoms remained a *mystery*. However, his model did make both the calculation of *masses* and the establishment of *combinatorial relations* between different types of atoms possible.

This meant that *chemistry* could now become a quantitative science which was a radical departure from the previous perspective. Unfortunately, from the second half of nineteenth century onwards, there was a great deal of experimental evidence amassed such as the *specific heat* which could not be explained by Dalton's model.

THOMSON'S 'EMBEDDED MASS' MODEL:

Around the turn of the century, the discoveries of Henri Becquerel, Frederic and Irene Cure led to the idea that an *atom* is *constructed* of *positively* and *negatively* charged *particles*.

However, in the early part of the 20^{th} century the English physicist John (Joseph) Thompson, pictured the atom not as a positively charged central nucleus – but as a *positive* charge spread throughout the atom forming a kind of "pudding" in which the *negative electrons* were suspended like "plums."

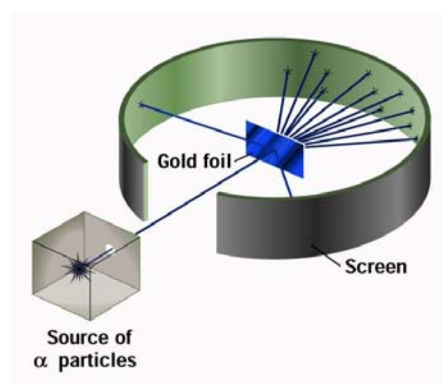


While exploring the nature of cathode rays, Thomson was able to show that these rays consisted of negatively charged particles of tiny mass – which he called 'electrons' – for which he calculated the charge-to-mass ratio. The discovery of electrons required a new model of the atom. Although electrons had a negative chare, it was widely accepted that atoms overall have zero charge. This implied that each atom must contain an equal number of positive and negative charges. Therefore Thomson proposed that 'the atom' constituted of electrons embedded in a sphere of uniform positive charge. While building on the notion that the atoms of different elements were unique, Thomson's model superseded Dalton's model in that it sparked a discussion over the internal structure of the atom.

Although the "plum-pudding" model was widely accepted, it was eventually *discredited* a few years later in 1911 when a New Zealand physicist named Ernest Rutherford published experimental results that the model could NOT explain.

RUTHERFORD'S 'NUCLEAR' MODEL:

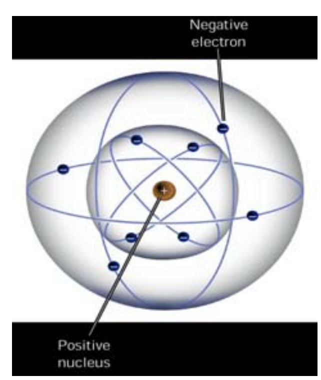
Rutherford and his two assistants, Hans Geiger and Ernest Marsden, bombarded thin metal foils, such as gold, with massive positively charged particles known as alpha particles [$_2He^4$].



When Rutherford and his team counted scintillations of scattered alpha particles on a zinc sulfide screen, they observed the following:

- Most of the particles passed through the foils without being deflected.
- A very small number of the particles were deflected through large angles.
- An even smaller number of particles were deflected through 180 degrees.

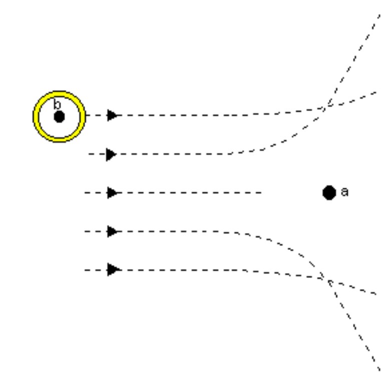
If atoms were as predicted by Thomson's model to have a uniform positive spread with negatively charged electrons sprinkled around, all the alpha particles would easily pass through the foil with only an occasional slight deflection in their paths.



On the basis of these experiments, Rutherford drew the following conclusions:

- Most of the atom is empty space.
- Most of the mass of the atom is concentrated in a dense, positively charged nucleus.
- The negative electrons orbit the nucleus in a planetary fashion.

According to Coulomb's law, the interaction of alpha particles with a very small massive nucleus would result in the scattering of the particles by the central repulsive force.



The scattering angle depends upon the mass of the gold atoms in the foil, the mass of the alpha particle and force upon impact

The positively charged particle follows a hyperbolic path and produces the observed angular pattern of scattering.

However, the ***limitation*** of ***Rutherford's Model*** was that it DID NOT account for:

(1) the ***lack*** of ***emission*** of ***radiation*** as ***electrons*** moved about the ***nucleus***.

(2) the ***unique spectrum*** of each ***element*** and

(3) the ***stability*** of the ***atom***.

The electrons moving around the nucleus, in Rutherford's Model, are accelerated (centripetally) and should radiate energy of changing frequency. Since atoms emit radiation of specific frequencies and do not collapse spontaneously, the Rutherford Model required modification.

MAX PLANCK

Around 1900, ***Max Planck*** from the University of Kiel concerned himself with observations of the radiation of heated materials. He attempted to draw conclusions from the radiation to the radiating atom. On basis of empirical data, he developed a new formula which later showed remarkable agreement with accurate measurements of the spectrum of heat radiation.

The result of this formula was so that energy is always emitted or absorbed in discrete units, which he called quanta. Planck developed his quantum theory further and derived a universal constant, which came to be known as ***Planck's constant***. The resulting law states that the energy of each quantum is equal to the frequency of the radiation multiplied by the universal constant: $E = h\, f$, where h is 6.63×10^{-34} J s. The discovery of quanta revolutionized physics, because it contradicted conventional ideas about the nature of radiation and energy.

THE BOHR MODEL

In atomic physics ***Niels Bohr*** created a model that depicted the atom as a small, positively charged nucleus surrounded by electrons that travel in circular orbits around the nucleus — similar in structure to the solar system, but with ***electrotatic*** forces providing ***attraction***, rather than gravity. This was an improvement on the earlier plum-pudding model and Rutherford model. Since the Bohr model is a ***quantum physics*** based modification of the Rutherford model, many sources combine the two, referring to the Rutherford-Bohr model.

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Atomic Models II

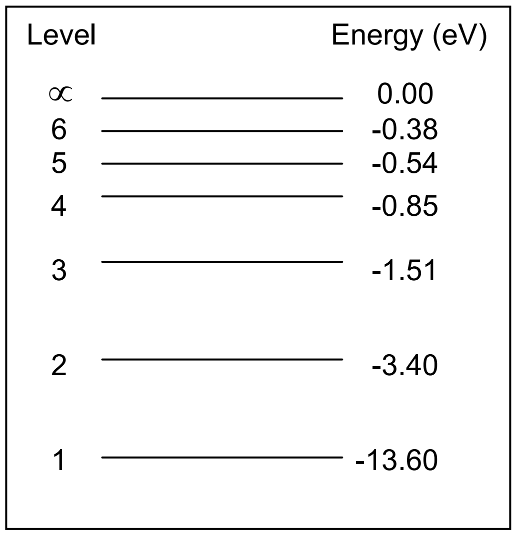
BOHR’S QUANTUM MODEL

In the years following Rutherford’s observations and the nuclear (planetary) model of the atom, Niels Bohr added an important piece to the atomic puzzle, which overcame the some of the difficulties that the planetary model exhibited.

Bohr’s model for the hydrogen atom began to explain, or at least shed more “light” [*pun intended*] on the limitations of the Rutherford Model.

The energy levels for the hydrogen atom can be obtained by the formula $E_n = \frac{1}{n^2}(-13.6 \text{ eV})$

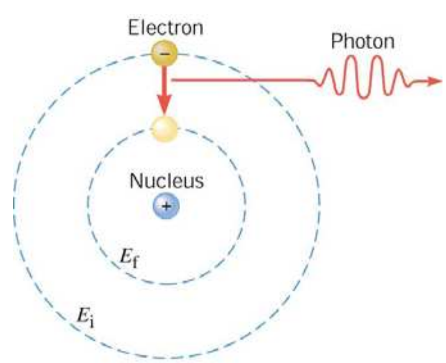
N	Energy (eV)
∝	0.00
6	-0.38
5	-0.54
4	-0.85
3	-1.51
2	-3.40
1	-13.60



For Hydrogen

Where the energy *absorbed by* or *emitted from* the electron, as it is either *excited* or *relaxes*, can be determined by the *difference in potential energy* between the respective *energy levels*.

To incorporate Einstein’s *photon* concept, Bohr theorized that a “**photon**” is emitted only when the *electron* changes orbits from a *larger* one with a *higher energy* (E_i) to a *smaller* one with a *lower energy* (E_f) as shown on the right.



Therefore,

$$\boldsymbol{E_{pho} = \Delta E \Rightarrow E_i - E_f}$$
$$\boldsymbol{\Rightarrow |E_1 - E_2| = hf}$$

BUT…how do electrons get into the higher-energy orbits in the first place ?

They get there by picking up “**energy**” when atoms collide, which happens more often when a *gas* is *heated*, or by acquiring energy when a *high voltage* is *applied* to a *gas*.

The Bohr model is a primitive model of the hydrogen atom. As a theory, it can be derived as a first-order approximation of the hydrogen atom using the broader and much more accurate quantum mechanics, and thus may be considered to be an obsolete scientific theory.

Introduced in 1913, the model give some insight to Rydberg's formula for the spectral emission lines of atomic hydrogen; while the Rydberg formula had been known experimentally, it did not gain a theoretical underpinning until the Bohr model was introduced. Not only did the Bohr model explain the reason for the structure of the Rydberg formula, but it provided a justification for its empirical results in terms of fundamental physical constants.

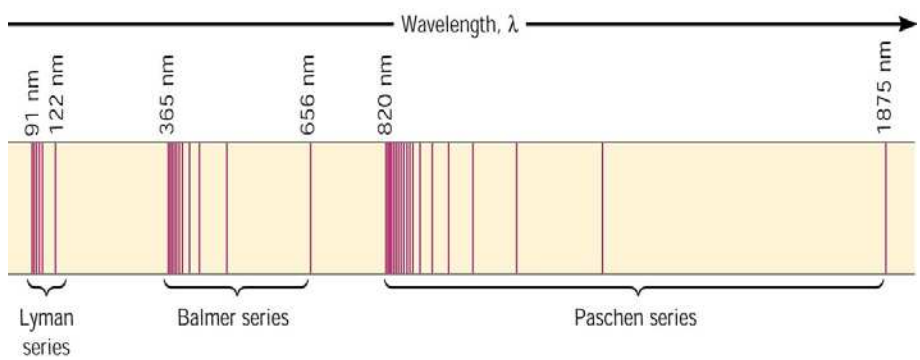
EMISSION LINE SPECTRA, BALMER SERIES, LYMAN SERIES, PASCHEN SERIES, WAVE NUMBER AND RYDBERG CONSTANT

For over 50 years, it had been known by observation that light from a glowing gas passed through a diffraction grating could be dispersed into its component colors. These sharp lines, known as the atomic spectra, are unique for that gas and act as their atomic or spectral fingerprint.

The visible wavelengths that appeared in the emission spectrum of hydrogen had been summarized by the *Balmer formula*,

$$\frac{1}{\lambda_n} = R\left(\frac{1}{(2)^2} - \frac{1}{n^2}\right)$$

Note: for Balmer $n_f = 2$



shown in the diagrams (above right) and (bottom right) as the *Balmer Series*.

Where $R = 1.1 \times 10^7\ m^{-1}$ is the *Rydberg constant* and λ_n represents the wave number (n).

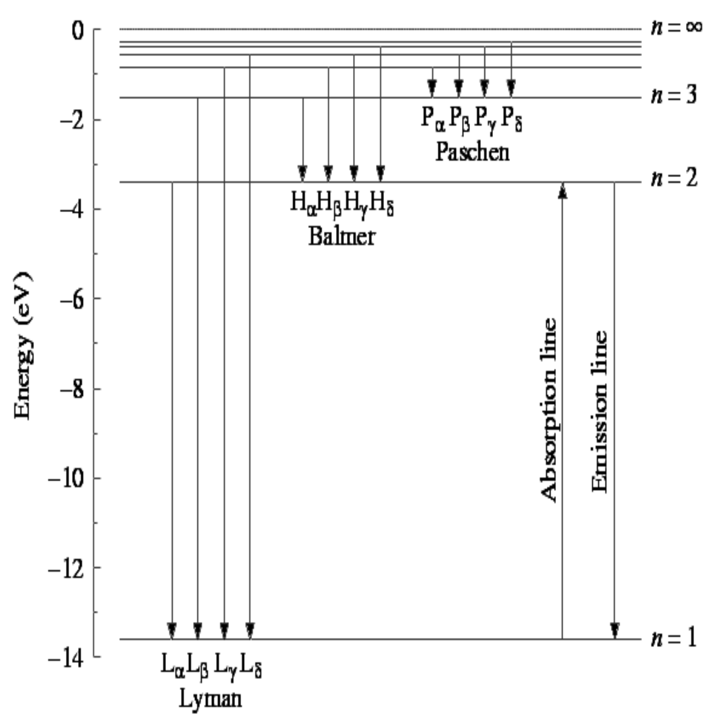
BUT…why do atoms emit (or absorb) radiation only in discrete wavelengths ?

Bohr's model of the atom explains the spectroscopists' observations.

Using the simplest model of the atom, *hydrogen* which has one electron, Bohr postulated that the *electron orbits* the nuclei at certain *discrete radii*.

When the electron is in one of these "special orbits", it does NOT lose energy as the *classical theory* would *predict*.

However, if the electron *absorbs* a certain amount of energy, it is *excited* to a higher orbit – one with greater radius.

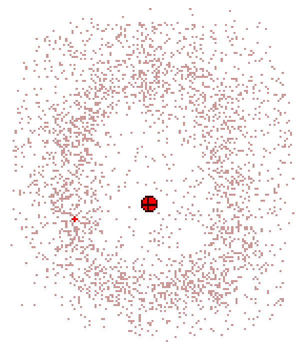


After spending a short time in this *excited state*, it returns to a lower orbit, *emitting* a photon (light) in the process. Since each allowed orbit, or energy level, has a *specific radius* and *energy*, the *photons emitted* in each jump also have *specific energies* and thus *wavelengths*.

$$\boxed{\lambda = \frac{c}{f} = \frac{c}{E_{pho}/h} = \frac{hc}{E_{pho}} = \frac{hc}{\Delta E}}$$

Because of its simplicity, and its correct results for selected systems, the Bohr model is still commonly taught to introduce quantum mechanics, before moving on to the more accurate but more complex valence shell atom.

THE CLOUD PROBABILITY MODEL



The cloud model represents a sort of history of where the electron has probably been and where it is likely to be going. The large dot in the middle represents the nucleus of an atom while the small dot, just left and down of the nucleus, represents an instance of the extra-nuclear electron.

Imagine, as the electron moves it leaves a trace of where it was.
This collection of traces quickly begins to resemble a cloud.

The probable locations of the electron were predicted by Erwin Schrödinger in 1925. Schrödinger's predictions happen to coincide with the locations specified in Bohr's model. However, Schrödinger built upon the thoughts of Bohr and yet took them in a new direction.

Erwin Schrödinger

Schrödinger developed a probability function for the hydrogen atom that basically describes a cloud-like region where the electron is "likely" to be found.

It can not say with any certainty, where the electron actually is at any point in time, yet can describe where it "ought" to be.

CLARITY THROUGH FUZZINESS, WAS ONE WAY HE DESCRIBE THE IDEA.

Schrödinger's theory of the quantum world is called *wave mechanics*. He worked out the exact solutions of the wave equation for the hydrogen atom, and the results perfectly agreed with the known energy levels of these atoms, seemingly without any of the complications and metaphysical speculations associated with the uncertainty principle. Moreover, the equation could also be applied to more complicated atoms, and even to particles not bound in atoms at all. It was soon found that in *every* case, Schrödinger's equation gave a correct description of a particle's behavior, provided it was not moving at a speed near that of light.

In spite of this success, the very meaning of the waves remained unclear. Schrödinger believed that the intensity of the wave at a point in space represented the 'amount' of the electron that was present at that point. In other words, the electron was spread out, rather than concentrated at a point. However, it was soon found that this interpretation was untenable, because observations revealed that particles never spread out. For example, it follows from the wave equation that when a wave, representing an electron, strikes a target, it spreads out in *all directions*. Experimentally, on the other hand, the electron scatters in some *specific* direction but *never breaks up*.

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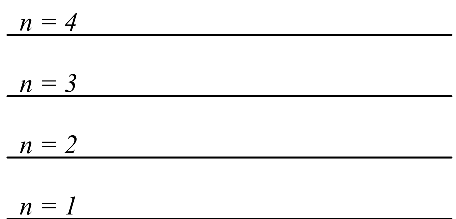
Atomic Models II

Ex 1: Consider the energy diagram for a simple hydrogen atom.

Level	Energy (eV)
∝	0.00
6	-0.38
5	-0.54
4	-0.85
3	-1.51
2	-3.40
1	-13.60

a. How much energy must a ground state electron absorb to be excited to the $n = 4$ state ?

b. How many possible frequencies of electromagnetic energy could be given off as an electron falls from the $n = 4$ energy level to the $n = 1$ energy level ?



a. Classify the possible electromagnetic radiation emitted as the electron returns from the $n = 4$ to $n = 1$ (ground state).

Ex 2: a. Find the wavelength of the first line ($n_i = 3$) in the Balmer series of the hydrogen spectrum. [Taffel31.2mod]

b. Determine the frequency of light for the first line and classify the electromagnetic radiation.

Ex 3: A photon of energy 14.0 eV enters a hydrogen atom in the ground state and ionizes it. [Taffel31.19.2]

a. With what kinetic energy will the electron be ejected from the atom ?

b. Compute the maximum speed that the electron will be ejected from the atom.