You have mastered this topic when you can:

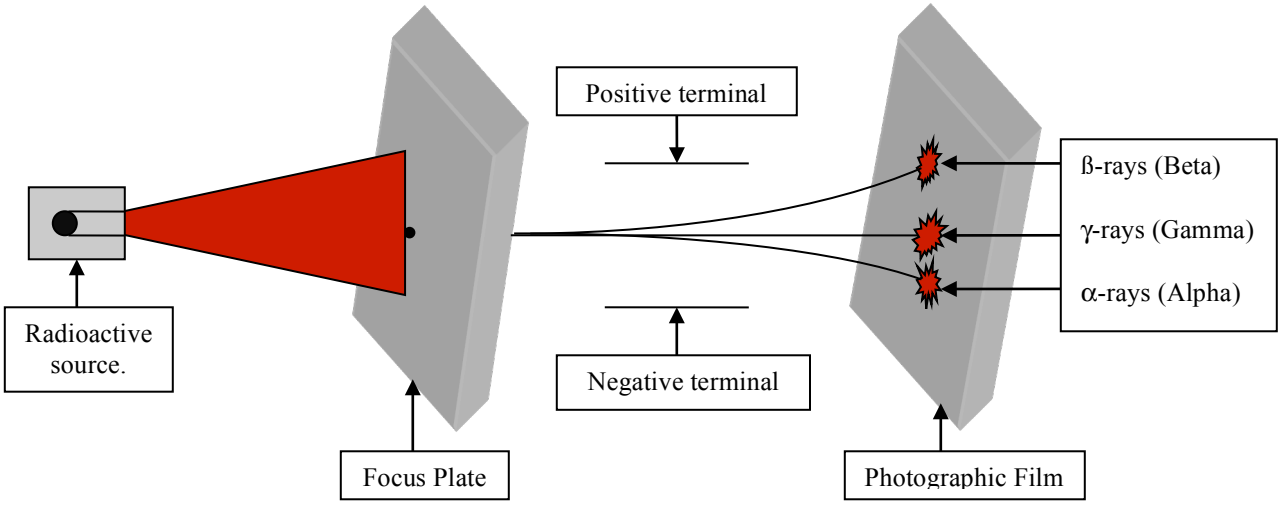
1) name and describe the atomic models developed by **RUTHERFORD** and **BOHR**.
2) describe **CHADWICK'S** contribution to the structure of the atom.
3) identify the **ATOMIC NUMBER** of an **ELEMENT** and an **ELEMENT** from an **ATOMIC NUMBER** using the **PERIODIC TABLE**.
4) define or describe **NUCLEUS**, **LINE SPECTRUM**, **SUBATOMIC PARTICLE**, **PROTON**, **ELECTRON**, **TRANSITION**, **GROUND STATE** and **PRINCIPLE QUANTUM NUMBER**.

EARLY ATOMIC THEORIES CONT' [VIII.1.A.a), pgs. 139 – 142]

I) **ERNEST RUTHERFORD**

A) The models developed by both ***Nagaoka*** and ***Thomson*** explained the existing evidence equally well, thus both models were equally valid. ***Ernest Rutherford*** favoured the model theorized by ***Thomson***. ***Rutherford's*** studies of radioactive decay in the early 1900's resulted scientists to favour one over the other. His research led him to modify ***Thomson's Muffin model*** creating his own model.

B) Radioactive decay results when an unstable atom breaks apart emitting radiation in the form of particles and or energy. ***Rutherford*** conducted many experiments investigating radiation using the apparatus described below. In it, the radiation released by a radioactive source was directed toward a focus plate. The focus plate concentrated the radiation into a narrow beam, which was passed between positive and negative terminals. When the focused beam of radiation passed between the positive and negative terminals it was split into three distinct beams.



1) ***Rutherford's*** experiment revealed that radiation consists of three distinct rays.
   a) Since α-rays were deflected toward the negative electrode he concluded they are positive.
   b) Since ß-rays were deflected toward the positive electrode he concluded they are negative.
   c) Since γ-rays did not bend toward either electrode he concluded they are neutral.
   d) Further research revealed that α-rays and ß-rays are ***particles*** while γ-rays are energy waves. α-particles are small, high-energy positively charged particles. ß-particles are ***electrons***, which are negatively charged ***subatomic particles***.

B) ***Rutherford's*** more famous experiment conducted in 1911 shot α-particles at a piece of gold foil several atoms thick. Based on ***Thomson's muffin model***, he predicted all the α-particles would pass through the gold foil experiencing minor deflections of course. This hypothesis seemed reasonable because according to ***Thomson's muffin model***, the accepted model of the day, atoms are spheres of mostly empty space in which negatively charged ***electrons*** are suspended in a ***positive spongy matrix***. The results of ***Rutherford's*** experiments, however, were very different. Most of the α-particles passed through the gold foil unaffected while a significant number of them deflected off course at large angles with a stunning few being deflected back toward the source of the α-particles. How you would explain these results?

1) ***Rutherford*** explained the results by concluding the following: **FIRST:** Because most of the α-particles passed through the foil unaffected, the ***atom consists of mostly empty space***. **SECOND:** Since α-particles are positive and a significant number were deflected off course at large angles with a small but significant few deflecting back toward their source, the ***atom's positive charge must be isolated in an extremely small core***. These conclusions suggested an initial model where the ***neutrally charged atom*** consists of essentially ***empty space*** with its positive charge concentrated in an ***extremely small core*** he called the **NUCLEUS** around which ***electrons*** travel in circular ring shaped ***orbits*** like planets ***orbiting*** the sun. In 1914 further research led ***Rutherford*** to conclude that the ***nucleus*** consisted of positively charged ***subatomic particles*** he called **PROTONS,** symbolized as p^+. Research revealed that a ***proton's*** mass is 1.673 x 10^{-24} g and its charge is +1.602 x 10^{-19} C (C = Coulomb which is the standard unit of electric charge). Review ***Nagaoka's*** and ***Thomson's*** models described above then, based on ***Rutherford's*** results, predict which model was more accurate.

C) **RUTHERFORD'S NUCLEAR MODEL (1914).** ***Rutherford's model*** of the atom is called the **NUCLEAR MODEL** and it is summarized here.

1.

2.

3.

4.

V) **HENREY MOSELEY** discovered an important property of the ***nucleus*** in 1913. His experiments revealed that the atoms of each element have a unique amount of ***positive nuclear charge***. Since no two elements have the same ***positive nuclear charge***, ***Moseley*** knew that the ***positive nuclear charge*** could be used to identify elements. As a result, he created the term **ATOMIC NUMBER = *Z***. The ***atomic number*** of an element is equal to the ***positive nuclear charge*** of the atom. In 1914 ***Rutherford*** concluded that the ***atom's positive nuclear charge*** was housed in ***subatomic particles*** he called ***protons*** **(p^+)**. He theorized that each ***proton, (p^+)*** contains a single positive charge. Since atoms have a unique amount of ***positive nuclear charge***, they must also have a unique number of ***protons***. This means the ***atomic number of an element is equal to the number of protons*** **(p^+)** found in the ***nucleus*** of its atoms.

A) The ***atomic number***, which equals the ***number of protons***, is found at the top left corner of each box on your ***periodic table***. It is used to identify an element because it is exclusive to the element: **i.e.** The ***atomic number*** of oxygen is Z = 8, therefore all oxygen atoms have exactly 8 ***protons***, not 7 nor 9.

B) **Required Practice 1:** Answer these questions on your own paper. **{Answers are on page 7 these of notes.}**

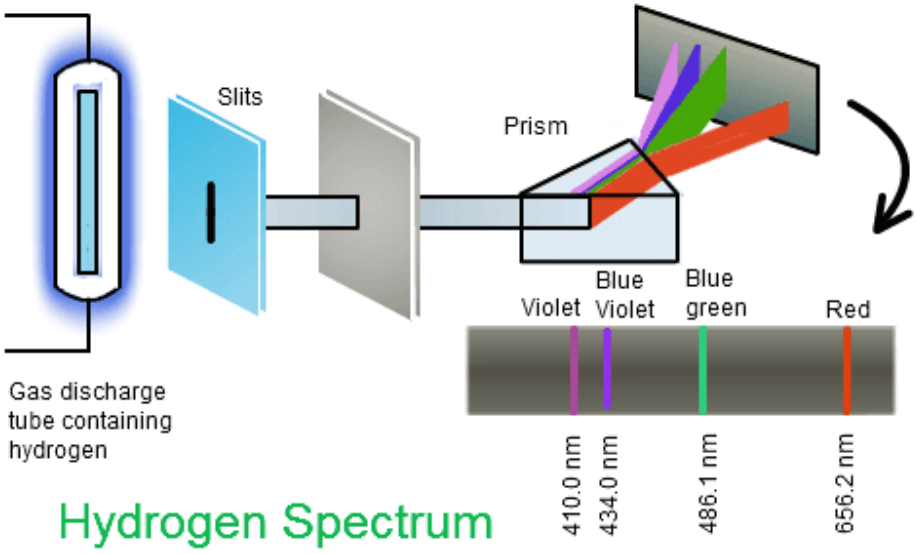
1. Complete this table.

	Z	Element
a	3	
b		Sulphur
c	21	
d		Kr
e	74	

2. Name the element that has:
   a. Z = 4 b. Z =13 c. Z = 24 d. Z = 40 e. Z = 84
3. Name, describe and draw the Rutherford's atomic model.
4. Draw Rutherford models for elements having Z = 1 to 5.
5. How are the models of Thomson, Nagaoka and Rutherford similar and how are they different?

NEILS BOHR [pg. 144]

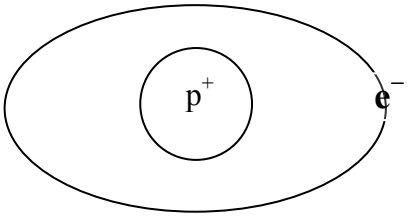
I) **NEILS BOHR** knew that when elements are energized by heat or electricity they release flashes of light energy. When the flash of light released from an energized atom is passed through a prism it is broken apart producing a pattern of different coloured lines known as an **EMMISION LINE SPECTRUM**. The diagram below illustrates how an energized gaseous elements releases light (blue glow). This light is focused by traveling through two focus plates, each contains a vertical slit. The focused beam of light is passed through a prism, which splits it into four visible colours.



A) ***Bohr's*** (~1913) experimental analysis of hydrogen explained the **EMMISION LINE SPECTRA** of elements and modified ***Rutherford's Nuclear Model***. ***Rutherford's Nuclear Model*** stated that each atom has positively charged ***protons*** **(p^+)** residing in a ***tiny central nucleus*** with negatively charged ***electrons*** **(e^-)** traveling around the ***nucleus*** in ***randomly oriented circular orbits*** like bees circling their hive (see the diagram at the bottom of page 3 of these notes). ***Bohr*** realized there were problems with this model. He knew that oppositely charged particles attract each other so he wondered why the negatively charged ***electrons*** **(e^-)** didn't fall into the positively charged ***nucleus***. The answer to ***Bohr's*** question came as a result of his study of the relationship between elements and light. ***Bohr*** knew that energized atoms release light, a form of energy, and that ***electrons*** carry energy. As a result he reasoned that ***electrons*** are involved producing the flash of light that creates an element's ***emission line spectrum***. The element with the simplest atomic structure is hydrogen. Its ***emission ine spectrum*** is given above. ***Bohr's*** research led him to develop a theory that explained hydrogen's ***line spectrum***. In doing so, he modified ***Rutherford's Nuclear Model*** and sparked the creation of one of the most profound theories in chemistry and physics – **QUANTUM MECHANICS**.

II) **EXPLAINING HYDROGEN'S EMMISION LINE SPECTRUM**

A) Recall that ***Rutherford's Nuclear Model*** of hydrogen has one ***proton*** in the ***nucleus*** and one ***electron*** traveling it in a ***randomly oriented circular orbit*** as illustrated below.



B) When an hydrogen atom is energized by heat or electricity, it emits a flash of light. To explain this flash of light, ***Bohr*** reasoned that hydrogen's single ***electrons*** **(e^-)** absorbs the heat or electrical energy and becomes ***excited***. Since ***electrons***, like all matter, have a natural tendency to exist with the lowest amount of energy possible, the ***excited high-energy electron*** will release the energy it gained as a flash of light.

1) Passing the flash of light through a prism produces the four-coloured ***line spectrum*** that is unique to hydrogen [See the ***emission line spectrum*** the previous page.] ***Bohr's*** reasoning explained the source of the light flashes from energized hydrogen but not the different colours of its ***line spectrum***. His curiosity continued: How does hydrogen's single ***excited high-energy electron*** create four different coloured lines?

a) ***Bohr*** knew that ***electrons*** **(e^-)** are energy carriers and that light travels in waves. Waves have peaks and valleys. The distance from one peak to the next is called a wavelength. High-energy light waves have a short wavelength while low-energy light waves have a long wavelength. Light consists of many colours, some of which are seen in a rainbow: red, orange, yellow, green, blue, indigo and violet. Each colour of light has a unique fixed wavelength, which means it carries a unique fixed amount of energy. Hydrogen's ***emission line spectrum*** consists of four different coloured lines: violet, blue-violet, blue-green and red. When arranged from shortest to longest wavelength, the four colours are ordered: violet, blue-violet, blue-green and red. Since light waves with short wavelengths are high-energy waves while light waves with long wavelengths are low-energy waves, violet light carries more energy than blue-violet light, which carries more energy than blue-green light, which carries more energy than red light.

C) ***Bohr*** explained hydrogen's ***emission line spectrum*** by modifying ***Rutherford's Nuclear Model*** as outlined here. The different colours of light found in hydrogen's ***emission line spectrum*** are caused by its single ***electron*** first absorbing then emitting ***different quantities*** of energy. ***Bohr theorized*** that an ***electron*** **(e^-)** is an energy carrier capable of carrying different ***fixed quantities of energy***. Since an ***electron*** carries different ***fixed quantities of energy***, its energy is said to be **QUANTIZED**. The ***quantity*** of energy carried by an ***electron*** determines its distance from the nucleus. Since an ***electron*** can carry ***different fixed quantities of energy***, it can reside at ***different fixed distances*** from the ***nucleus***, each distance being determined by the ***quantity of energy*** the electron carries. ***Bohr*** called those distances **ENERGY LEVELS**, although chemists and physicists continue to call them orbits. Each hydrogen atom has an infinite number of ***energy levels*** (***orbits***). The ***quantity of energy*** each ***orbit*** is capable of carrying increases as its distance from the ***nucleus*** increases, this means ***low-energy orbits*** are close to the nucleus while ***high-energy orbits*** are far from the nucleus. Since each ***orbit*** holds a unique ***fixed quantity of energy***, each ***orbit*** is given a unique designation called the **PRINCIPLE QUANTUM NUMBER**, symbolized **n**. The first ***orbit*** is designated n = 1, the second is n = 2, the third is n = 3, etc. The larger the ***principle quantum number***, **n**, the greater the energy carried by the ***orbit***. The n = 1 is the lowest energy ***orbit***; n = 2 is a higher energy ***orbit*** and n = 3 is an even higher energy ***orbit***. ***When an electron, e^-, is in the lowest-energy orbit possible, it is said to be in its GROUND STATE.*** An ***electron*** cannot exist outside an ***orbit***, thus when it gains or emits energy it '*appears*' in a different ***orbit*** **– IT NEVER EXISTS IN THE SPACE BETWEEN ORBITS**. ***The process of 'changing' orbits is called*** **TRANSITION**.

1) **IN SUMMARY:** When an ***electron*** **(e^-)** in its ***ground state*** is energized, it becomes excited and must ***transition*** to a ***higher-energy orbit***. The more energy the ***electron*** gains the farther from the ***nucleus*** it will ***transition***. Since ***electrons*** have a natural tendency to exist carrying the lowest amount of energy possible, they will ***transition*** back to the ***lowest-energy orbit*** available to them as soon as possible. In order to ***transition*** back to the ***lowest-energy orbit*** available, the excited ***high-energy electron*** must release the energy it absorbed as a flash of light. ***Transition*** *between different* ***high-energy orbits*** *and* ***low-energy orbits*** *releases different amounts of energy, which appears as different colours of light*. Since each colour

represents a different wavelength of light and thus a different amount of energy, ***the colour of each line of the line spectrum indicates the energy difference between the two orbits involved in the electron's transition***. This information was used to confirm that each individual ***orbit*** is capable of carrying a different ***quantized*** (fixed) amount of energy (the difference in energy between the ***orbits*** involved in a ***transition***) and the amount of energy each orbit is capable of carrying.

2) ***Bohr*** used this ***theory*** to explain hydrogen's ***emission line spectrum***. Hydrogen's single ***electron*** **(e^-)** exists in its ***ground state***, the first ***orbit***, n = 1. When an hydrogen atom is energized by subjecting it to heat or electrical energy, its single ***electron*** **(e^-)** absorbs some of the energy and becomes ***excited*** causing it to ***transition*** (*'changes'*) from its ***ground state***, the n = 1 ***orbit***, to either the n = 3, n = 4, n = 5 or n = 6 ***orbit***. The greater the quantity of energy it absorbs the farther from the nucleus is its ***transition***. Since every ***electron*** has a natural tendency to exist in the lowest possible state of energy, Hydrogen's excited ***high-energy electron*** will emit the energy it absorbed by ***transitioning*** ('*changes'*) back to its ***ground state***, the n = 1 ***orbit***. ***Transition*** of hydrogen's ***electron*** from the four different ***high-energy orbits*** it can occupy to the lowest-energy ***orbit*** available as it moves to its ***ground state orbit*** causes it to release different amounts of energy in the form of flashes of different colours of light.

a) Study Figure 5 on page 39 of your text. The colours produced are explained here.
Transition from the n = 6 ***orbit*** to the n = 2 ***orbit*** causes the ***electron*** to emit a flash of violet light.
Transition from the n = 5 ***orbit*** to the n = 2 ***orbit*** causes the ***electron*** to emit a flash of indigo light.
Transition from the n = 4 ***orbit*** to the n = 2 ***orbit*** causes the ***electron*** to emit a flash of blue-green light.
Transition from the n = 3 ***orbit*** to the n = 2 ***orbit*** causes the ***electron*** to emit a flash of red light.
Transition from the n = 2 ***orbit*** to the n = 1 ***orbit*** causes the ***electron*** to emit a flash of ultraviolet radiation. We cannot see ultraviolet radiation thus it does not appear in hydrogen's ***emission line spectrum***.

III) **THE BOHR'S ATOMIC THEORY**

A) ***Bohr*** used the knowledge he acquired from his analysis of hydrogen's line spectrum to modify ***Rutherford's Nuclear Model***. ***Bohr's*** atomic model, often called the **SOLAR SYSTEM/PLANETARY MODEL**, is summarized below.

B) **JAMES CHADWICK** knew that when an atom is bombarded with α-particles, neutral emissions are released from its ***nucleus***. In 1932, ***Chadwick*** discovered that the neutral emissions consisted of ***subatomic particles*** he named **NEUTRONS = *n***. Further experimentation revealed that the charge of one ***neutron*** is 0 C (Coulombs), and its mass is 1.675×10^{-24} g. ***Chadwick's*** discovery of the ***neutron*** allowed ***BOHR'S ATOMIC THEORY*** to more accurately describe the structure of an atom and explain its properties.

C) ***BOHR'S ATOMIC THEORY*** of the atom is described in these seven points. The diagram is at the top of the next page.

1.

2.

3.

4.

5.

Continued on the next page.

6.

7.

Principle Quantum # = n	**$2n^2$**	**Maximum # of electrons**
1	$n = \mathbf{1} \therefore 2n^2 = 2(\mathbf{1})^2 = 2$	2
2	$n = \mathbf{2} \therefore 2n^2 = 2(\mathbf{2})^2 = 2(\mathbf{4}) = 8$	8
3		
4		

D) **Required Practice 2:** Using correct terminology, answer these questions on your own paper. **{Answers are on page 7 of these notes.}**

1. Explain how chemists find emission line spectra of elements useful.
2. Define the terms ground state and transition.
3. Describe what happens to an electron when it is energized and when it releases energy.
4. What does the principle quantum number refer to?
5. Which orbit has the least energy: n = 2, n = 3, n = 4?
6. What do different colours of an element's line spectrum tell us?
7. How are the models of Rutherford and Bohr similar?
8. How are the models of Rutherford and Bohr different?

ANSWERS TO THE REQUIRED PRACTICE

Required Practice 1 from page 3

1. Answers in the table are in bold.

	Z	Element
a	3	**Lithium**
b	**16**	Sulphur
c	21	**Scandium**
d	**36**	Kr
e	74	**Tungston**

2a. Be; **2b.** Al; **2c.** Cr; **2d.** Zr; **2e.** Po. **3. Rutherford:** Nuclear Model: All atoms contain positively charged subatomic particles called protons which are housed in a central nucleus and negatively charged subatomic particles called electrons which are housed in randomly oriented orbits that surround the nucleus. The nucleus contains approximately 99.95 % of the atom's mass and approximately 0.05 % of its volume while the orbits contain approximately 99.95% of the atom's volume and approximately 0.05 % of its mass. See your teacher to check your drawings. **4.** See your teacher. **5. Similarities:** atoms are electrically neutral, contain positive and negative charge, and contain subatomic particles called electrons. The atoms are electrically neutral which means an atoms positive charge is equal to its number of electrons. **Differences:** Thomson's model has electrons suspended in a spongy matrix, while Nagaoka's and Rutherford's models have electrons orbiting a central positive core. Rutherford's model has the positive charge housed in subatomic particles called protons, which are located in an extremely tiny central nucleus.

Required Practice 2 from page 6

1. Emission line spectra can be used to identify the elements that of which a substance is composed. **2.** Ground state is the lowest energy level that an electron can occupy. Transition is the process of electrons changing energy levels. **3.** When an electron is energized it transitions to a higher-energy orbit; when an electron loses energy it transitions to a lower-energy orbit. **4.** Principle quantum numbers refer to the different energy levels an atom contains. **5.** The n = 2 orbit has the least energy. **6.** The colours of each line of an emission line spectrum indicate the different wavelengths of light that are released as electrons transition toward the ground state. The different wavelengths are an indication of the difference in energy between the two energy levels involved in the transition toward an electrons ground state. **7.** Similarities: Atoms are electrically neutral containing equal numbers of positive charge subatomic particles called protons and negative charged subatomic particles called electrons. Protons are found in a tiny central nucleus with electrons circling in orbits. The nucleus contains approximately 99.95 % of the atom's mass and approximately 0.05 % of its volume while the orbits contain approximately 99.95% of the atom's volume and approximately 0.05 % of its mass. **8.** Differences: Rutherford's model has orbits containing electrons arranged randomly around the nucleus like bees circling their hive. Bohr's model has an infinite number of circular orbits called energy levels that an electron can occupy. Low-energy orbits are close to the nucleus while high-energy orbits are far from the nucleus. Electrons can carry fixed quantities of energy, which determines the orbit they occupy. Energized electrons will transition to a higher energy orbit then transition release the energy as they transition back to a lower energy level.