Chemistry 10: Electron "Orbitals"

March 30-April 3

Time Allotment: 40 minutes per day

Student Name: ________________________________

Teacher Name: ________________________________

Academic Honesty

I certify that I completed this assignment independently in accordance with the GHNO Academy Honor Code.

Student signature:

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I certify that my student completed this assignment independently in accordance with the GHNO Academy Honor Code.

Parent signature:

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Packet Overview

Date	Objective(s)	Page Number
Monday, March 30	*Describe line spectra, and review the basic light equations.*	3
Tuesday, March 31	*Define the Bohr model of the atom.*	6
Wednesday, April 1	*Review last week's material.*	10
Thursday, April 2	*Review this week's material*	13
Friday, April 3	*Test on the wavelength equations, orbitals, and the electromagnetic spectrum.*	17

Additional Notes:

Hi students,

I still miss you all.

This week is going to be tough. The electron orbitals unit overlaps with the quantum mechanics unit, and the readings are more theoretical and less practical in nature. For that reason, there will be **less calculation and practice problems** and more reading and annotation to do for the next two days. On Wednesday, you'll start reviewing for the test Friday and I'd encourage you to spend Thursday brushing up on what you find difficult on Wednesday. **Send emails to me Wednesday (and Thursday).** Since there is no class review session, you don't have the opportunity to ask me one-on-one questions in person. Email me. Don't be confused for long periods. Good luck :)

Monday, March 30

Chemistry Unit: Electrons and orbitals
Lesson 1: Line Spectra

Unit Overview

Pure substances, when releasing energy, do **not** emit all the colors of the rainbow, as most (white) lights do. This collection of colors, which is unique for every substance, is called its **line spectrum**. A spectrum of colors that contains **all** the colors of the rainbow is called a **continuous spectrum**.

Each type of light (color) has its own unique energy level, which can be calculated using the equation $E = h\nu$.

An old equation (learned 2-3 weeks ago) allows one to find the wavelength, frequency, or speed of light (given two of the variables): $\nu = \frac{c}{\lambda}$

E = energy
h = Planck's constant
ν = frequency
λ = wavelength
c = speed of light

A Question for Lesson 1: Why do glowing, pure substances **not** emit all the colors of the rainbow?

Objective: Be able to do this by the end of this lesson.
Use the equations $E = h\nu$ and $\nu = \frac{c}{\lambda}$ to find the energy of light colors, given the wavelength (λ) of the light it emits.

Introduction to Lesson 1

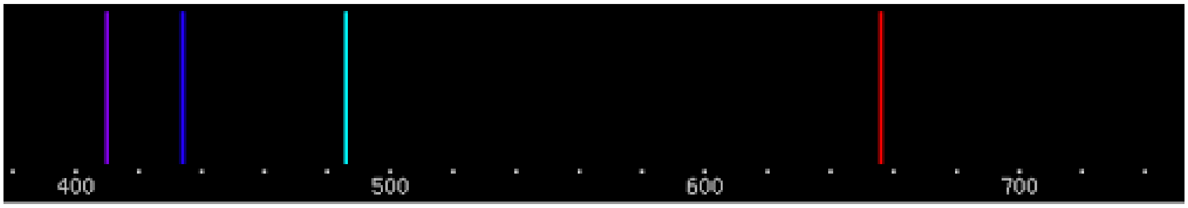
Line Spectra

What happens when you shine light into a prism?

If you recall the glasses we used before (spring) breaking, you'll remember that visible light (the fluorescents in the ceiling, the sun) contains **all the colors of the rainbow**. This is true for many light sources, especially the ones you'd see walking around campus or downtown.

However, under some circumstances, light is emitted that does **not** contain all the colors of the rainbow[1]. When you take a sample of Hydrogen gas (H_2) and pump it full of energy (by burning it or running an electrical charge through it), it emits (to the naked eye) red light. If you were to wear the prism glasses from two weeks ago, however, you'd see:

[1] Chemists refer to this as a "continuous spectrum"—in other words, the rainbow is not "broken up" into just some of the colors of the rainbow.



[If this packet is black and white, you should see violet, blue-violet, blue-green, and bright red lines]

Now, in past lessons we've discussed that the energy possessed by light[2] can be expressed by the equation:

$$E = h\nu$$

In other words, Energy is equal to Planck's constant (h) multiplied by the light's frequency.

Planck's constant is $6.62607004 \times 10^{-34}$ J/m
The light's frequency should be expressed in $\frac{1}{\text{seconds}}$ (Hz).

Take a look at those emission lines. The light lines that appear (and you'd see all four of these in the prism glasses, but no other colors) have wavelengths of 410 nm, 434 nm, 486 nm, and 656.2 nm.

Remember that to calculate the frequency of light based on its wavelength you can use the equation

$$\nu = \frac{c}{\lambda}$$

λ is wavelength. c is the speed of light (3.00×10^8 m/s)

Using the formula above, calculate the energy of each light color above. You'll have to calculate the frequency of the light given the wavelength first.

Red light: λ = 656.2 nm. ν = ____________________ (write units always!!)

E = ____________________ (write units always!!)

Blue-green light: λ = 486 nm ν = ____________________

E = ____________________

Blue-violet light: λ = 434 nm ν = ____________________

[2] (and this includes x-rays, UV rays, gamma rays, radio waves, microwaves, and infrared waves)

E = ________________

Violet light: λ = 410 nm ν = ________________

E = ________________

Now answer these challenging questions.

Which light type is the most energetic? (which has the most energy)

Why does the hydrogen only emit light of certain colors (4, specifically)? Why doesn't it produce light of **all** colors when you pump it up with energy?

Read and annotate the following selection.

When scientists first detected the line spectrum of hydrogen in the mid-1800s, they were fascinated by its simplicity. In 1885 a Swiss schoolteacher named Johann Balmer, and later a man named Rydberg, showed that the wavelengths of these four visible lines of hydrogen fit an intriguingly simple formula:[3]

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$$

Easy there. Lambda (λ) means what it usually means (wavelength) and "n_1" and "n_2" stand for any two integers[4]. R_H is a number that just fits. Rydberg tried different numbers and found that 10,973,731.6 m^{-1} always works. Since he (and we) didn't understand exactly why it works, he simply wrote it down and included it in the equation. It has to be in there for the equation to work.

So, this equation was designed to describe the four lights that appear when hydrogen is burned. Here's the interesting thing. Plug in 2 and 3 for n_1 and n_2, then use basic algebra to solve for λ. You'll get 656.2 x 10^{-9}, or **656.2 nm.**

Plug in 2 and 4 for n_1 and n_2, then use basic algebra to solve for λ. You'll get **486 nm.**

---

[3] He did this by trial and error, by looking at the data.

[4] i.e. n_1 could be 1, while n_2 could be 2; or n_1 could be 6, while n_2 could be 4. n_1 and n_2 *couldn't* be 6.2, because 6.2 isn't an integer (whole number).

Plug in 2 and 5 for n_1 and n_2, then use basic algebra to solve for λ. You'll get **434 nm.**

And finally, plug in 2 and **6** for n_1 and n_2, then use basic algebra to solve for λ. **You'll get 410 nm.**

This really (relatively) simple equation ties the wavelength of light being emitted to two integer (non-zero) numbers. For Hydrogen, the four colors emitted by hydrogen atoms being pumped with energy correspond to the four pairs: **2 and 3**, **2 and 4**, **2 and 5**, and **2 and 6.**

What does this mean? We'll have to wait for tomorrow, but do the closing problems below.

Memorize these definitions below.

Reread your notes and email your teacher quick if you have any questions.

continuous spectrum: the rainbow of colors, containing light of all visible wavelengths.
line spectrum: a spectrum containing radiation of only specific wavelengths (e.g. the rainbow of colors, but with some colors missing). (plural "line spectra").

Closing:

1. Some stars in the sky appear red. Another star appears blue-white. **Which star is hotter**?
2. Suppose visible light was used to eject electrons from a certain metal surface. Do you think this would work if UV light was used?

Tuesday, March 31

Chemistry Unit: Electrons and orbitals
Lesson 2: Hydrogen electron orbit radii and energy "levels"

Unit Overview
Bohr postulated three points about electron orbits, which corresponded directly to electron "levels" defined in the Rydberg equations as "n" and predicting exactly the energy quantity and wavelength of light released or gained when an electron moved from one "n"-level to another. The "n"-levels, while somewhat vague, have some basis in reality in predicting how far an atom is from the nucleus as well as the energy released or gained (as radiation and/or light) when an electron moves between orbit levels.

Objective: Be able to do this by the end of this lesson.
Describe the importance of the principal quantum number and its effect on wavelength and energy level.
Understand the Bohr model of the atom.

Introduction to Lesson 2

What do these numbers mean? If two specific integers can be used in a simple equation to **predict** the fact that Hydrogen will produce red light of wavelength 656.2 nm, and the same can be done for the three other wavelengths hydrogen produces when energized, what does that tell us about the atom?

Bohr thought it meant three things:

1. Only orbits **of certain radii**, corresponding to certain definite energies, are permitted for the electron (in a hydrogen atom).
   a. An electron in a permitted orbit, will not lose energy and therefor be pulled into the nucleus.
   b. As the electron changes from one permitted orbit to another, it will absorb or emit energy.

Mark closely what Bohr is saying. Summarize the first point on the next page.

____
____
____

Check your answer below.

Before you move on: **Speculate as to how Bohr's model explains the fact that hydrogen gas emits a line spectrum rather than a continuous spectrum.**

____
____
____
____

Just checking:

1. Would you expect an electron orbiting closer to the nucleus to have less energy or more energy than an electron orbiting further from the nucleus?

Summary of the first point of Bohr's model:

1) In the hydrogen atom, an electron can be orbiting at specific distances, but can't orbit between those distances. For example, it's permitted to orbit 10, 8, or 3 angstroms away from the nucleus, but for some reason it's impossible for the electron to orbit 7 angstroms away from the nucleus. If it did, it would fall into the 8-angstrom orbit or fall back to the 3-angstrom orbit.
2) An electron in the 10-angstrom orbit will have a specific amount of energy, while if it fell to the 8- or the 3-angstrom orbit it would have two different specific amounts of energy.

Continue to read and annotate:

Bohr took the three equations:

$E = h\nu$, $\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$, and $\nu = \frac{c}{\lambda}$

and rearranged them to form the equation:

$$E = (-h\, c\, R_H)\left(\frac{1}{n^2}\right)$$

to calculate the energy possessed by an electron at different radiuses. Go on, try it yourself. It's very fun, and is totally within your ability[5].

What does this long equation mean? Notice those three constants: h (Planck's constant), *c* (the speed of light), and R_H (Rydberg's constant). n, of course, as usual indicates some integer number (take your pick).

Now, things get a bit more solid. "n" refers to some number. It's the same "n" from the Rydberg equation. Try solving this equation, using 1 through 7 for "n". **Don't calculate it** (unless you want to)—**just set it up.**

If n = 1: $E = (-h\, c\, R_H)\left(\frac{1}{n^2}\right) = -(h\, c\, R_H)\frac{1}{\_\_}$ (solve for n^2 here)

If n = 2: $E = (-h\, c\, R_H)\left(\frac{1}{n^2}\right) = -(h\, c\, R_H)\frac{1}{\_\_}$

If n = 3: $E = (-h\, c\, R_H)\left(\frac{1}{n^2}\right) = -(h\, c\, R_H)\frac{1}{\_\_}$

[5] A general and badly worded but still-helpful hint: use the third equation to solve for lambda, then use the first equation to solve for nu (*ν*), then plug the results into the second equation to relate E to the right half.

If n = 4: $E = (-h\,c\,R_H)\left(\frac{1}{n^2}\right) = -(h\,c\,R_H)\frac{1}{\_\_}$

If n = 5: $E = (-h\,c\,R_H)\left(\frac{1}{n^2}\right) = -(h\,c\,R_H)\frac{1}{\_\_}$

If n = 6: $E = (-h\,c\,R_H)\left(\frac{1}{n^2}\right) = -(h\,c\,R_H)\frac{1}{\_\_}$

If n = 7: $E = (-h\,c\,R_H)\left(\frac{1}{n^2}\right) = -(h\,c\,R_H)\frac{1}{\_\_}$

Now, do some brief analysis. Which value for n—1, 2, 3, 4, 5, 6, or 7—has the **lowest** energy? The highest? Remember the negative sign—big negative numbers are **smaller** than medium negative numbers[6].

Bohr said that these values for n—and you can keep calculating them for n past 6—correspond to distance from the nucleus. In other words, as an electron is further from the nucleus, it must have **more** energy (less negative E). However, if n continues to get larger (you can try this by plugging in a big number—try 2238093—for n), the energy of the electron becomes infinitesimally small. If, in theory, the electron was infinitely far from the nucleus (try plugging in ∞ for n), its energy would be zero.

n signifies some kind of energy **level** for the electron—it corresponds to the electron's distance from the nucleus, and seems to indicate that electrons can only have distinct quantities of energy corresponding to whole numbers (e.g. the electron can't gain 0.1 n worth of energy, but only 1 whole n of energy at a time). We call it the **principal quantum number.**

Let's take our heads out of this problem and turn to another.

1. Does the energy of the electron go up or down when it moves from the n=3 level to the n=7 level? Therefore, does it emit (release) or absorb (gain) energy? Consider the electron to be the system, does it release energy into its surroundings, or does it absorb energy from its surroundings?

    a. So if an electron moves further away from the nucleus (e.g. from the n=3 to the n=7 level), it is an **(endothermic/exothermic)** process (circle one).

[6] E.g. -2300 is **less than** -49.

I hope this was fun. Take some time now and write: What have you learned about the structure of atoms? About the behavior of electrons around the nucleus?

(Write four sentences. Try to be concise, but all-inclusive)

____
____
____
____

What do the line spectrum data emitted by Hydrogen atoms have to do with Bohr's model of the atom and the electron? Why does Hydrogen only emit some of the colors of the rainbow, instead of all colors like most glowing substances?

____
____
____
____
____

Wednesday, April 1

Chemistry Unit: Electrons and orbitals
Lesson 3: Review.

Unit Overview:
You will be reviewing last week's and this week's (this unit's) material.

Lesson 3 Socratic Question: Keep these questions in mind as you study this lesson!
(At every point) What does this tell me about the atom?

Objective: Be able to do this by the end of this lesson.
Put asterisks/mark each section that you require further practice in.
Form questions and/or send an email to Mr. Luke asking for further practice in sections you found difficult.

Unit: Electron "Orbitals" Study Guide:
You may use 1) a periodic table and 2) a calculator to answer these questions. Try not to use the periodic table with ionization energies, your notes, or any other resource when completing these questions.

Fill in the blanks for each element.

Mg
# of valence electrons: ________
# of electrons it wants to gain/lose (circle one): ________
Charge once it has gained/lost electrons: ________

N
\# of valence electrons: _______
\# of electrons it wants to gain/lose (circle one): _______
Charge once it has gained/lost electrons: _______

I
\# of valence electrons: _______
\# of electrons it wants to gain/lose (circle one): _______
Charge once it has gained/lost electrons: _______

Br
\# of valence electrons: _______
\# of electrons it wants to gain/lose (circle one): _______
Charge once it has gained/lost electrons: _______

Answer the questions.

1. How many distinct energy levels ("layers") do electrons fall into? _____

2. What are the four known orbit shapes that electrons can take within these layers? (simply name them) ___________________________

3. Using the terms "atomic radius" and "atomic mass", explain why ionization energy decreases as you move down a column in the periodic table.

____

____

For each element, write its electron configuration, showing the principal quantum number, the orbital shape, and the number of electrons each one contains. E.g.: $1s^1$ shows that the atom has an s-shaped orbital with a principal quantum number (n) of 1, and that level is holding only ^{1} electron.

Lithium: ___________________________
electron configuration

Carbon: ___________________________
electron configuration

Fluorine: ___________________________
electron configuration

Neon: ___________________________
electron configuration

Magnesium: ___________________________
electron configuration

Argon: ___________________________
electron configuration

Write the electron configuration for the ion. Remember you have to subtract or add from the particular number of electrons that atom has normally.

Al^{3-} : ______________________________
electron configuration

Li^{+}: ______________________________
electron configuration

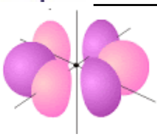
Br^{-}: ______________________________
electron configuration

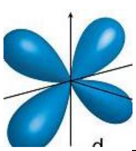
Kr: ______________________________
electron configuration

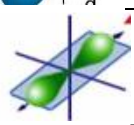
Orbital Shapes

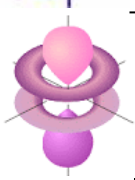
Identify each shape by name (s, p, d, or f). Lines and shaded planes are only there to give you a sense of the 3D shape.

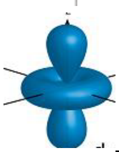
 _____________

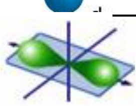
 _____________

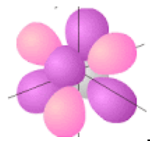
 _____________

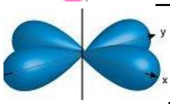
 _____________

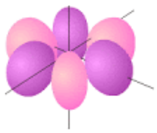
 _____________

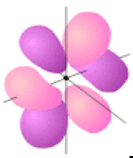
 _____________

 _____________

 ____________

 ____________

 ____________

 ____________

Write the name of the element that corresponds to the electron configuration:

$1s^22s^2$ Element: ______________

$[Ne]3s^23p^1$ Element: ______________

$[Ar]3s^23p^2$ Element: ______________

$1s^22s^22p^2$ Element: ______________
$[Kr]5s^1$ Element: ______________

$[Xe]6s^24f^{14}5d^6$ Element: ______________

$[Ar]4s^2$ Element: ______________

Write the eléctron configuration for the ions below:

Fe^{2+} ______________

Ag^+ ______________

Li^+ ______________

N^{3-} ______________

H^- ______________

<u>Thursday, April 1</u>

Chemistry Unit: Electrons and orbitals
Lesson 4: Review (continued)

Unit Overview:
You will be reviewing last week's and this week's (this unit's) material.

Lesson 3 Socratic Question: Keep these questions in mind as you study this lesson!
(At every point) What does this tell me about the atom?

Objective: Be able to do this by the end of this lesson.
Put asterisks/mark each section that you require further practice in.
Form questions and/or send an email to Mr. Luke asking for further practice in sections you found difficult.

Unit: Electron "Orbitals" Study Guide (part II)
You may use 1) a periodic table and 2) a calculator to answer these questions. Try not to use the periodic table with ionization energies, your notes, or any other resource when completing these questions.

Use the equation below

$$\nu = \frac{c}{\lambda}$$

to solve for the frequency or wavelength of different types of radiation:

Example #1: What is the frequency of electromagnetic radiation having a wavelength of 210.0 nm?

Example #2: What is the frequency of violet light having a wavelength of 4000 Å?

(1 Å = 1.0×10^{-10} meters)

Example #3: What is the frequency of EMR having a wavelength of 555 nm? (EMR is an abbreviation for electromagnetic radiation.)

Example #4: What is the wavelength (in nm) of EMR with a frequency of $4.95 \times 10^{14}\ s^{-1}$?

Example #5: What is the wavelength (in both cm and Å) of light with a frequency of 6.75×10^{14} Hz?

Use the equation below

$$E = h\nu$$

to solve for the energy contained of each type of radiation below.

Example #1: What is the energy of electromagnetic radiation having a frequency of 1427583133.33333 MHz? (remember your SI prefixes—look up what M (Mega) stands for if you must).

Example #2: What is the energy of violet light having a frequency of 749481145 MHz?

Example #3: What is the energy of EMR having a frequency of 540166590.990991

MHz?

Example #4: What is the energy of EMR with a frequency of 4.95×10^{14} s^{-1}?

Example #5: What is the energy of light with a frequency of 6.75×10^{14} Hz?

Summarize the three important points Bohr outlined in his model of the atom.

Re-write or copy onto flashcards all definitions (in bold). Definitions from last packet are found below.

Ionization energy: the minimum motive force required to remove a specific electron from an atom (thereby making it an ion).
Periodic trends: patterns you see in the periodic table when looking at the elements' properties or behaviors.
Orbitals: 3D patterns or shapes the electron's path takes around the nucleus. They have many different known shapes (see diagrams in the packet) and have names like "The 2s

orbital", which tell you what energy level they're on (e.g. level 2) as well as their possible shape (the s orbital is a sphere shape).

Calculate answers to the questions below, using the Rydberg equation and the given values for n.

In a hydrogen atom, how much energy does an electron emit when it falls from the n=7 to the n=3 orbital?

In a hydrogen atom, how much energy does an electron need in order to rise from the n=3 to the n=7 orbital?

Is the n=7 orbital further from the nucleus or closer to the nucleus than the n=3 orbital?

Answer the questions below.

What orbital shapes (s, p, d, etc) does the **n=3** level contain? How many of each orbital shape does it hold?

How many electrons can each orbital shape contain?

s
p
d
f

How many **total electrons** can the n=4 level hold? Helpful hint: pick the noble gas that has all the 4-level orbital shapes in it and write its electron configuration.

What is the significance of the variable "n" in the Rydberg equation? What information does it tell us about the electron in a hydrogen atom?

What is the significance of the fact that "n" is an integer, and never a mixed number like 6.2, 4.63, etc?

Some final notes.

This is part of a developing unit on quantum mechanics. The line spectra of hydrogen, the equations for the energy, wavelength, frequency, and speed of light, and the new equations we've discussed this week form the basis of what is known as "quantum theory". At its heart lies the theory that energy is quantized, energy is discrete—it exists in "atoms" of energy, if you will, certain indivisible units of energy that cannot be split. Yet how much energy is that? We don't know—it seems to vary, depending on the wavelength/frequency of the energy in question.

It really seems that "n" is a funny variable. It corresponds directly to distance from the nucleus (n=3 means that the electron is further from the nucleus than with n=2). It corresponds directly with the amount of energy an electron has (n=3 means that electron has more energy than with n=2). However, the variable has no units—its units are **not** distance! Nor are they Joules. We've simply found that energy of an electron and distance from the electron are linked to whole-number—integer—increases and decreases. That seems to be its significance.

<u>Friday, April 3</u>

Chemistry Unit: Electrons and orbitals: Test

Before turning to the next page, clear your work space. Take out a calculator and a pencil.

Do not use the periodic table with ionization energies, **your notes**, or **any other resource** when completing these questions. Remember, this is a test, and should be treated as such. Good luck!

When you are ready, tear out the next two pages and begin. A fresh periodic table is provided for you, as well as an answer sheet.

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Periodic Table of the Elements

Atomic Number
Symbol
Name
Atomic Mass

1 IA 1A	2 IIA 2A	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8	9 VIII 8	10	11 IB 1B	12 IIB 2B	13 IIIA 3A	14 IVA 4A	15 VA 5A	16 VIA 6A	17 VIIA 7A	18 VIIIA 8A
1 **H** Hydrogen 1.008																	2 **He** Helium 4.003
3 **Li** Lithium 6.941	4 **Be** Beryllium 9.012											5 **B** Boron 10.811	6 **C** Carbon 12.011	7 **N** Nitrogen 14.007	8 **O** Oxygen 15.999	9 **F** Fluorine 18.998	10 **Ne** Neon 20.180
11 **Na** Sodium 22.990	12 **Mg** Magnesium 24.305											13 **Al** Aluminum 26.982	14 **Si** Silicon 28.086	15 **P** Phosphorus 30.974	16 **S** Sulfur 32.066	17 **Cl** Chlorine 35.453	18 **Ar** Argon 39.948
19 **K** Potassium 39.098	20 **Ca** Calcium 40.078	21 **Sc** Scandium 44.956	22 **Ti** Titanium 47.867	23 **V** Vanadium 50.942	24 **Cr** Chromium 51.996	25 **Mn** Manganese 54.938	26 **Fe** Iron 55.845	27 **Co** Cobalt 58.933	28 **Ni** Nickel 58.693	29 **Cu** Copper 63.546	30 **Zn** Zinc 65.38	31 **Ga** Gallium 69.723	32 **Ge** Germanium 72.631	33 **As** Arsenic 74.922	34 **Se** Selenium 78.972	35 **Br** Bromine 79.904	36 **Kr** Krypton 83.798
37 **Rb** Rubidium 85.468	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.906	40 **Zr** Zirconium 91.224	41 **Nb** Niobium 92.906	42 **Mo** Molybdenum 95.95	43 **Tc** Technetium 98.907	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.906	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.868	48 **Cd** Cadmium 112.411	49 **In** Indium 114.818	50 **Sn** Tin 118.711	51 **Sb** Antimony 121.760	52 **Te** Tellurium 127.6	53 **I** Iodine 126.904	54 **Xe** Xenon 131.294
55 **Cs** Cesium 132.905	56 **Ba** Barium 137.328	57-71	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.948	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.207	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.217	78 **Pt** Platinum 195.085	79 **Au** Gold 196.967	80 **Hg** Mercury 200.592	81 **Tl** Thallium 204.383	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.980	84 **Po** Polonium [208.982]	85 **At** Astatine 209.987	86 **Rn** Radon 222.018
87 **Fr** Francium 223.020	88 **Ra** Radium 226.025	89-103	104 **Rf** Rutherfordium [261]	105 **Db** Dubnium [262]	106 **Sg** Seaborgium [266]	107 **Bh** Bohrium [264]	108 **Hs** Hassium [269]	109 **Mt** Meitnerium [278]	110 **Ds** Darmstadtium [281]	111 **Rg** Roentgenium [280]	112 **Cn** Copernicium [285]	113 **Nh** Nihonium [286]	114 **Fl** Flerovium [289]	115 **Mc** Moscovium [289]	116 **Lv** Livermorium [293]	117 **Ts** Tennessine [294]	118 **Og** Oganesson [294]

Lanthanide Series	57 **La** Lanthanum 138.905	58 **Ce** Cerium 140.116	59 **Pr** Praseodymium 140.908	60 **Nd** Neodymium 144.242	61 **Pm** Promethium 144.913	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.964	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.925	66 **Dy** Dysprosium 162.500	67 **Ho** Holmium 164.930	68 **Er** Erbium 167.259	69 **Tm** Thulium 168.934	70 **Yb** Ytterbium 173.055	71 **Lu** Lutetium 174.967
Actinide Series	89 **Ac** Actinium 227.028	90 **Th** Thorium 232.038	91 **Pa** Protactinium 231.036	92 **U** Uranium 238.029	93 **Np** Neptunium 237.048	94 **Pu** Plutonium 244.064	95 **Am** Americium 243.061	96 **Cm** Curium 247.070	97 **Bk** Berkelium 247.070	98 **Cf** Californium 251.080	99 **Es** Einsteinium [254]	100 **Fm** Fermium 257.095	101 **Md** Mendelevium 258.1	102 **No** Nobelium 259.101	103 **Lr** Lawrencium [262]

Alkali Metal | Alkaline Earth | Transition Metal | Basic Metal | Semimetal | Nonmetal | Halogen | Noble Gas | Lanthanide | Actinide

Name: ___________________

Answer Sheet

For #1-10: if you do not put units on your answer, you may lose full credit for that question.

1 _______________
2 _______________
3 _______________
4 _______________
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Challenge: _______
11 _____
12 _____
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40. What do the line spectrum data emitted by a Helium atom have to do with Bohr's model of the atom and the electron? Why is Helium only emitting some of the colors of the rainbow, instead of all colors like most glowing substances? Use the terms electrons, energy levels, and orbitals in your answer.

41. What is the significance of the variable "n" in the Rydberg equation? What information does it tell us about the electron in a hydrogen atom?

Test:
Line Spectra, Bohr's Model, and Electron Orbitals

What's *nu*?
Answer the questions below. Use this page as scratch paper and show your work.

1. The yellow light given off by a sodium vapor lamp used for public lighting has a wavelength of 589 nm. What is the frequency of this radiation?

2. A certain microwave has a wavelength of 0.032 meters. Calculate the frequency of this microwave.

3. A radio station broadcasts at a frequency of 590 KHz. What is the wavelength of the radio waves?

4. Microwave ovens emit microwave energy with a wavelength of 12.9 cm. What is the energy of exactly one photon of this microwave radiation?

5. Calculate the energy of one photon of yellow light that has a wavelength of 589nm.

Energy and Frequency
Answer the questions below. Use this page as scratch paper and show your work

6. A hypothetical wave has 6.6 J of energy. What is its frequency?

7. UVA penetrates more deeply into the skin, but UVB more readily causes sunburn. These UVB photon travel with a frequency of about 1.0 x 1015. How much more energy do they carry?

8. The wavelength of a doctor's x-ray is only 0.01 nm. Calculate its frequency.

9. A certain electromagnetic wave has a wavelength of 625 nm. What is the energy of the wave?

10. The blue color of the sky results from the scattering of sunlight by air molecules. The blue light has a frequency of about 7.5 x 1014 Hz.

Challenge (Bonus)
Come back to this problem when you finish the test.

A student removes the spinning plate from his microwave oven. He places a chocolate bar inside on a paper plate and zaps it for 10 seconds. Removing the candy, he sees two melted spots approximately 6cm apart. The microwave says on the back that it operates at 24.5 GHz. Considering that the speed of light in air is very close to the speed of light in a vacuum, he calculates the wavelength of a microwave.

Orbital Shapes
Answer the questions below.

11. The s orbital shape can hold
    a. 2 electrons
    b. 6 electrons
    c. 10 electrons
    d. 14 electrons

12. The d orbital shape can hold
    a. 2 electrons
    b. 6 electrons
    c. 10 electrons
    d. 14 electrons

13. The p orbital shape can hold
    a. 2 electrons
    b. 6 electrons
    c. 10 electrons
    d. 14 electrons

14. The f orbital shape can hold
    a. 2 electrons
    b. 6 electrons
    c. 10 electrons
    d. 14 electrons

15. Is the n=6 orbital further from the nucleus or closer to the nucleus than the n=5 orbital?
    a. Farther
    b. Closer

16. What orbital shapes (s, p, d, etc) does the **n=2** level contain? How many of each orbital shape does it hold?
    a. $1s^2 2s^2 2p^6$
    b. $1s^2 2s^2 2p^6 3s^2 3p^6$
    c. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$
    d. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10}\ 5s^2\ 5p^6$

17. What orbital shapes (s, p, d, etc) does the **n=4** level contain? How many of each orbital shape does it hold?
    a. $1s^2 2s^2 2p^6$
    b. $1s^2 2s^2 2p^6 3s^2 3p^6$
    c. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$
    d. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10}\ 5s^2\ 5p^6 4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$

18. What orbital shapes (s, p, d, etc) does the **n=5** level contain? How many of each orbital shape does it hold?
    a. $1s^2 2s^2 2p^6 3s^2 3p^6$
    b. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$
    c. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10}\ 5s^2\ 5p^6$
    d. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10}\ 5s^2\ 5p^6 4f^{14}\ 5d^{10}\ 6s^2\ 6p^6$

True or False: The Bohr Model

Write “T” or “F” on the answer sheet.

19. A hydrogen atom’s electron can only be found at certain specific distances from the nucleus.

20. A hydrogen atom’s electron can be found at any distance from the nucleus, so long as it is not outside of the pull of its proton.

21. An electron orbiting the hydrogen nucleus will naturally lose energy and be pulled into the nucleus, unless it is at a specific “allowed” distance from the nucleus.

22. An electron in a hydrogen atom, when it moves closer to the nucleus, must be losing energy.

23. An electron in a hydrogen atom, when it moves away from the nucleus, must be gaining energy.

24. When an electron in a hydrogen atom loses energy and moves closer to the nucleus, it *emits* energy.

25. One can move a hydrogen atom's electron further away from the nucleus by pumping energy into it (heat, electricity, etc).

Electron Configurations
Write the electron configuration for each atom below on the answer sheet.

26. N^{3-}

27. C

28. Ar

29. Cl

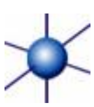
Identify the element, based on its electron configuration.

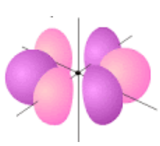
30. $[Ne]3s^2 3p^1$

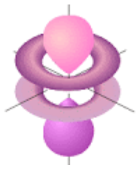
31. $1s^2 2s^2 2p^2$

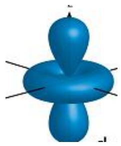
32. $[Ar]3s^2 3p^2$

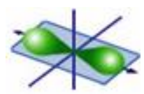
Identify the orbital shape.
Write "s", "p", "d", or "f" on the answer sheet.

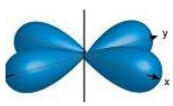
33. 

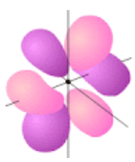
34. 

35. 

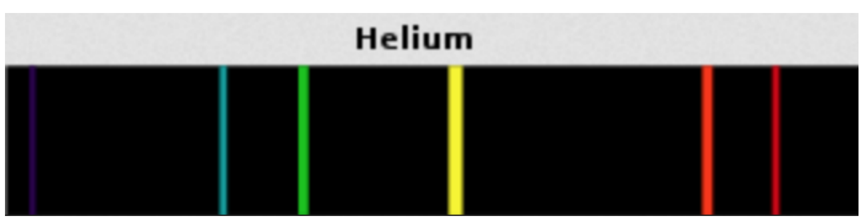
36. 

37. 

38. 

39. 

Written Answers
Answer the questions below, using full sentences.



The line spectrum for a helium atom, emitted upon heating.

40. What do the line spectrum data emitted by a Helium atom have to do with Bohr's model of the atom and the electron? Why is Helium only emitting some of the colors of the rainbow, instead of all colors like most glowing substances? Use the terms electrons, energy levels, and orbitals in your answer.

WRITE YOUR ANSWER ON THE ANSWER SHEET.

Turn to the next page.

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41. What is the significance of the variable “n” in the Rydberg equation? What information does it tell us about the electron in a hydrogen atom?

WRITE YOUR ANSWER ON THE ANSWER SHEET