		<h2>Why is the atom structured the way it is?</h2>	Name _____ Date _____
<p>Essential Questions</p> <ol style="list-style-type: none"> How are the electrons organized? Why do the arrangement of the electrons determine the properties of the atom? 	<p>Vocabulary</p> <p><i>electron, orbital, energy level, ionization energy, electronegativity, effective nuclear charge, shielding,</i></p>	<p>Objectives</p> <ol style="list-style-type: none"> Describe the contributions scientists have made to the atom and atomic theory. Express the arrangement of all electrons in an atom using electron configuration and orbital notation Describe the mathematical relationships between energy, frequency, and wavelength of light Use the Periodic Table to identify the chemical family of an element and explain its properties 	

Lesson	Description	Homework
Winter Break (4.0)	History of Atomic Theory	Watch flip video over atomic theory and complete handout.
4.1	Properties of Periodic Table Ions, Isotopes and Weighted Atomic Mass	pg 7
4.2	Bohr's Model - Energy, Frequency, Wavelength Calculations	pg 14
4.3	Flame Test Lab Electron Configuration and Quantum Numbers	pg 29
4.4	Electron Configuration and Quantum Numbers cont'd	Watch flip over periodic trends
4.5	Major Quiz 4.0-4.4 Begin ADI Density Periodic Trends Lab	
4.6	ADI Density Periodic Trends Lab	
4.7	Periodic Trends: atomic radius, ionic radius, ionization energy, electronegativity; Ionization Energy magnets demo	pg 33
4.8	Exam 4	
4.9	Flex	

Unit 4 Vocabulary

- atomos
- Dalton's atomic theory
- law of definite proportions
- law of multiple proportions
- plum pudding model
- Bohr atomic model
- gold foil experiment
- quantum mechanical model
- periodic law
- periodic table
- group
- period
- block (s, p, d, f)
- chemical family
- metal
- nonmetal
- metalloid
- representative element
- alkali metal
- alkali earth metal
- transition metal
- inner transition metal
- halogen
- noble gas
- valence energy level
- valence electron
- ion / cation / anion
- isotope
- atomic mass unit
- relative abundance
- weighted atomic mass
- quantum / photon
- electromagnetic radiation (ER)
- ER spectrum
- visible light
- ultraviolet
- infrared
- microwave
- radio wave
- X-ray
- atomic emission spectrum
- wavelength (λ)
- frequency (ν)
- hertz (Hz)
- Planck's constant (h)
- speed of light (c)
- electron volt (eV)
- energy level (n)
- ground state
- excited state
- electron configuration
- sublevel / subshell
- atomic orbital
- Aufbau principle
- degenerate orbital
- Hund's rule
- Pauli exclusion principle
- isoelectronic species
- orbital notation
- n (principal quantum number)
- l (angular momentum number)
- m_l (magnetic quantum number)
- m_s (spin quantum number)
- periodic trend
- atomic radius
- ionic radius
- effective nuclear charge
- shielding effect
- electronegativity
- electron affinity
- (first, second, etc) ionization energy

Lesson 4.1: Properties of the Periodic Table; Ions, Isotopes and Weighted Atomic Mass**A. Organization of the Periodic Table****1. Groups or Families**

- a. Atoms of elements in the same group also have the same number of valence electrons and therefore - similar phys/chemical properties

2. Periods (series)

- a. Horizontal rows of elements - across a period, the atomic number increases by 1
- b. Atoms of elements in the same period have the same valence energy level (n)

3. Metals and Nonmetals

- a. A stair-step line on the table separates the metals from the nonmetals

- b. Metalloids (semimetals) straddle the line and have properties of both metals and nonmetals
- 4. Lanthanide and Actinide Series (Inner Transition Metals) (f-block)
 - a. Metals and manufactured metal elements
- 5. Group 1 – Alkali metals (the most reactive metal elements) (except hydrogen (H)) (s-block)
 - a. 1 valence electron; soft metals, very reactive, stored in oil; explosive in water
- 6. Group 2 – Alkaline earth metals (very reactive metal elements) (s-block)
 - a. 2 valence electrons; reactive, found in Earth's crust as ores
- 7. Group 17 – Halogens (the most reactive nonmetal elements)
 - a. 7 valence electrons; very reactive nonmetals; all diatomic
- 8. Group 18 – Noble gases (the least reactive elements – inert and very stable)
 - a. 8 valence electrons except for He (He has 2 valence electrons)

B. *Types of Elements*

Properties of Metals

1. Luster
2. Good conductors of heat and electricity
3. Malleable
4. Ductile
5. High tensile strength

Properties of Nonmetals

1. Many nonmetals are gases at room temperature
2. Solid nonmetals tend to be brittle and non-lustrous
3. Poor conductors of heat and electricity

Properties of Metalloids

1. Some properties of metals and some properties of nonmetals
2. Solids at room temperature
3. Semiconductors of electricity

Noble Gases

1. All are gaseous members of group 18
2. Generally unreactive and chemically stable

C. Mendeleev's Periodic Table (1869)

Organization

1. Vertical columns in atomic weight order
 - a. Mendeleev placed elements in rows with similar properties
2. Horizontal rows have similar chemical properties

B. Missing Elements

Gaps existed in Mendeleev's table

- a. Mendeleev predicted the properties of the "yet to be discovered" elements
 - (1) Scandium, germanium and gallium agreed with his predictions

Unanswered Questions

1. Why didn't some elements fit in order of increasing atomic mass? ex. Te and I
2. Why did elements exhibit periodic behavior?

D. Moseley and the Modern Periodic Table (1911)

1. The periodic table was found to be in atomic number order, not atomic mass order
2. The physical and chemical properties of the elements are periodic functions of their atomic numbers (= The Periodic Law - discovered by Moseley)
3. Elements with similar properties are found at regular intervals within the periodic table

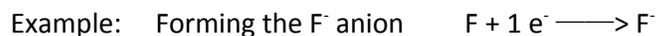
E. Ions, Isotopes, and Weighted Atomic Mass

Ions

1. **Ions** are charged atoms.
2. Charges arise due to loss or gain of valence electrons during chemical reactions, thus stabilizing the atom.
3. If the atom loses electrons, a positively charged **cation** is formed.



4. If the atom gains electrons, a negatively charged **anion** is formed.



5. In ordinary matter, cations and anions always occur together so that matter is neutral.

Isotopes

1. Atoms with same number of protons (#p⁺) but different numbers of neutrons (#n^o). Isotopes have the same atomic number (Z) but different mass numbers.
2. Recognizing Isotope Notation
 - i) X = element from periodic table
 - ii) A = mass number = #p + #n
 - iii) Z = atomic number = #p

NOTE: In a neutral atom, Z = #p = #e

Examples	Sodium-23	Na-23	(Z=11)	11 p, 12 n, 11 e
	Sodium 24	Na-24	(Z= 11)	11 p, 13 n, 11 e

Weighted (Average) Atomic Mass (AAM)

Steps for Calculating Average Atomic Mass (when given the masses and percent abundances):

1. Convert the percentages into decimals by dividing by 100 (move decimal over two spaces to the left)
2. Multiply the decimal of each isotope by its respective mass.
3. Add the numbers from step two together. (You can punch the numbers into your calculator as indicated in the equation below.

Formula:

$$\text{AAM} = \text{mass of isotope 1} * \text{decimal abundance} + \text{mass of isotope 2} * \text{decimal abundance} + \dots$$

Example: Calculate the average atomic mass of carbon given the following data.

Data:	98.89% Carbon 12	isotopic mass 12.000 amu
	1.11% Carbon 13	isotopic mass 13.0034 amu

$$\begin{aligned} \text{Solution: } & (12.000 \text{ amu} * 0.9889) + (13.0034 \text{ amu} * 0.0111) \\ & = 11.867 \text{ amu} + 0.144 \text{ amu} = 12.011 \text{ amu} \end{aligned}$$

Example: Calculate the average atomic mass of iron, using information in the table below.

Isotope name	Isotope abundance	Isotope mass (amu)
Iron-54	5.90%	53.94
Iron-56	91.72%	55.93
Iron-57	2.10%	56.94
Iron-58	0.280%	57.93

Steps for Calculating Percent Abundance from Average Atomic Mass (when given molar masses OR mass in amu for each isotope):

1. Because *decimal* abundances will always add up to 1, assign one isotope to have an abundance of "x" and the other isotope to have an abundance of "1-x."
2. Look up the average atomic mass of the atom on the periodic table if it's not given in the problem.
3. Set up your problem so it looks like the setup below:

$$(\text{Mass of Isotope A}) * x + (\text{Mass of Isotope B}) * (1-x) = \text{Average Atomic Mass}$$

1. Solve for "x." Multiply this number by 100 to turn it into a percentage. This is the percent abundance of isotope A.
2. The percent abundance of isotope B is 100% minus the percent abundance of isotope A.

Examples:

Gallium consists of two naturally occurring isotopes, Ga-69 and Ga-71, with respective masses of 68.926 and 70.925 amu. Calculate the abundance of each isotope.

Copper is made up of two isotopes, Cu-63 (62.9296 amu) and Cu-65 (64.9278 amu). Given copper's atomic weight of 63.546, what is the percent abundance of each isotope?

4.1 Practice and Homework

1. Fill in the table below with the correct numbers.

Symbol	name	Atomic number	Mass number	charge	# of nucleons	# of protons	# of neutrons	# of electrons
${}_{11}^{23}\text{Na}$								
${}_{19}^{39}\text{K}$								
${}_{19}^{41}\text{K}$								
${}_{19}^{41}\text{K}^{1+}$								
		12	25	0				
				1-	35			18
						7	7	10

2. Naturally occurring magnesium consists of three stable isotopes:

isotope	amu	Abundance
Mg-24	23.985	78.99%
Mg-25	24.986	10.00%
Mg-26	25.983	11.01%

What is the average atomic weight of magnesium? Ans: _____

3. Naturally occurring silicon consists of three stable isotopes:

isotope	amu	Abundance
Si-28	27.977	92.21%
Si-29	28.976	4.70%
Si-30	?	3.09%

What is the average atomic mass of silicon? Ans: _____

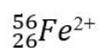
4. Compare the following four symbols:



(A)



(B)



(C)



(D)

- Which pairs of symbols (A, B, C, D) represent isotopes of each other?
- Explain how these isotopes (from #4a) are similar to each other. Explain how these isotopes are different from each other.
- Compare the four elements above in terms of their number of electrons, protons, and neutrons.

5. Oxygen comes in two stable isotopes, O-16 and O-18. The atomic mass of oxygen-16 is 15.989 amu.

The atomic mass of oxygen-18 is 17.999 amu. If the average atomic mass of oxygen is 15.999 amu, determine the percent abundance of each isotope.

Lesson 4.2: Bohr's Model - Energy, Frequency, Wavelength Calculations

A. Why Bohr's Model?

The "gold foil" experiment by Ernest Rutherford and colleagues established the presence of a small, incredibly dense, positively charged nucleus, and (by inference, since matter is neutral) moving, negatively charged electrons around it. However, this "Rutherford atomic model" has two problems:

- According to classical physics, any charged particle moving in a curved path emits radiation, since it is constantly accelerating. Thus, electrons moving around the nucleus should eventually lose energy (after all... radiation is energy), be attracted to the nucleus and eventually collide with it - which does not happen - or else matter would be self-destructing constantly.
- The radiation produced by the electrons, as they move faster and faster inward, should change in energy very quickly so that you obtain a continuum of energy values. However, the radiation emitted by electrons is *discrete* - that is, only specific values for these energies are ever observed.

To try to account for these observations, Bohr proposed an atomic model that expanded on Rutherford's. But first, let's lay down a few definitions.

B. The Nature of Electromagnetic Radiation

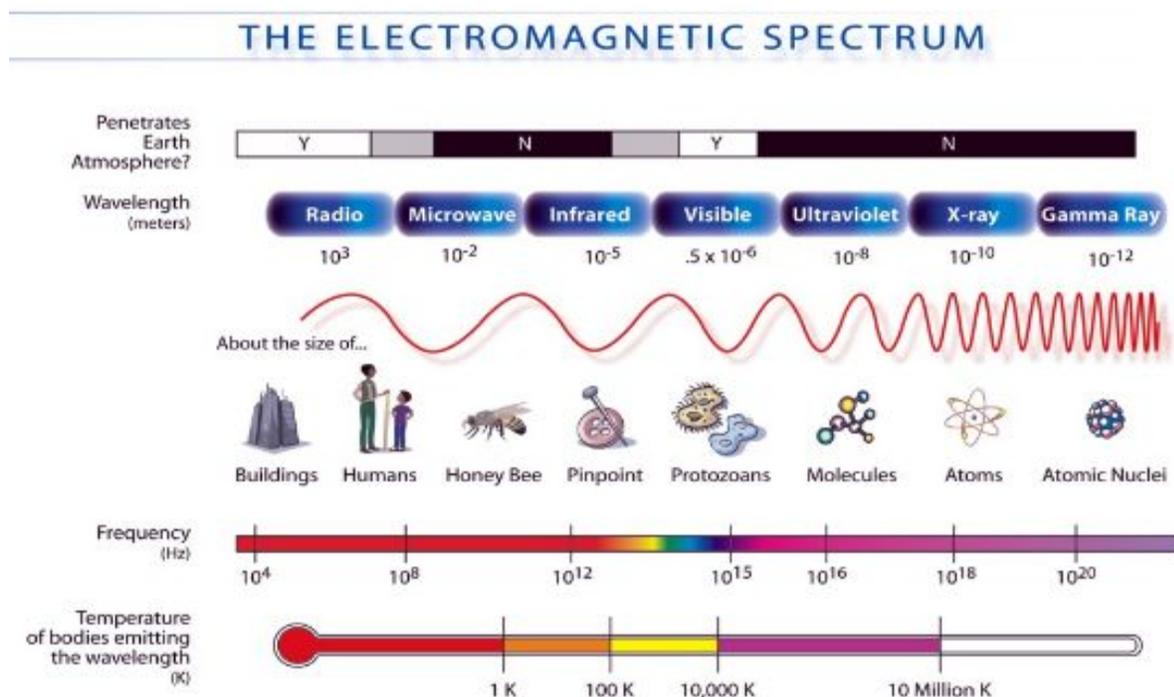
Electromagnetic radiation encompasses various types of energy such as visible light, radio waves, microwaves, X-rays and gamma rays. All electromagnetic radiation has a few things in common:

- They are all released into space by stars;
- They are all generated by a changing electric or magnetic field (hence the name)
- They all travel at the same speed, which is the speed of light, c ($= 3.00 \times 10^8$ m/s)
- They all travel as waves, and therefore have all the properties of waves -
 - a) They each have a constant wavelength - that is, the length between two crests (tops) of the wave is a constant. The unit of wavelength is generally m (meters) although since all these wavelengths are very small it is more convenient to use nm (nanometers), especially for visible light. 1 nanometer = 10^{-9} meters, or 10^9 nanometers = 1 meter.
 - b) They each have a constant frequency - that is, the number of waves that pass a certain point in space in one second is a constant. The unit of frequency is /s, or s^{-1} , or Hz (hertz).
1 Hz = $1 s^{-1} = 1 /s$.

As the next graphic illustrates¹, there are certain relationships that exist between the frequency, and the wavelength of various types of electromagnetic (EM) radiation. As the wavelength

¹ NASA. NASA, n.d. Web. 13 Dec. 2016.

increases, or gets longer, the frequency decreases, and vice versa. A high frequency is correlated to a high energy (you can think of this as damaging power), thus, a long wavelength means lower energy, and a short wavelength means a high frequency *and* a high energy.



There are a few relevant mathematical formulas that illustrate these relationships between energy, frequency, and wavelength. Use the legend below to get familiar with their symbols:

E = energy (in Joules)

h = Planck's constant = $6.626 \times 10^{-34} \text{ J}\cdot\text{s}$

ν = frequency (in Hz, or s^{-1} , or /s) ** pronounced "nu"

c = speed of light = $3.00 \times 10^8 \text{ m/s}$

λ = wavelength (in m) ** if wavelength is given in units other than m, you need to convert
** pronounced "lambda"

Formulas:

$$E = h \nu \quad (\text{Planck's equation})$$

$$c = \lambda \nu$$

Also:

$$1 \text{ eV (electron volt)} = 1.602 \times 10^{-19} \text{ J}$$

Examples:

- Derive a formula that relates energy to wavelength. (i.e. $E = \dots$)

b) What is the energy (in eV) of radiation that has a frequency of 3.99×10^{13} Hz? What type of EM radiation is this energy?

c) The wavelength of UV radiation is 250. nm. What is its frequency and energy in electron volts?

C. *The Quantum Nature of Light*

Electrons in atoms emit and absorb energy in the EM spectrum, so that energy has a wavelength and a frequency, and has wave properties. This is the classical understanding of light. In reality, EM radiation behaves both as a wave and as particles.

An example of this particle nature is the fact that electrons in atoms only emit and absorb energy with very specific frequency values somewhere in the EM spectrum, so since frequencies can only be certain numbers, energy can only be certain numbers. We say this energy is “quantitized”. A “quantum” of energy is a “packet” with a discrete energy and a discrete frequency; a quantum of light is a “photon.” Max Planck introduced the basic “quantitization” idea in 1900, but it was Einstein who applied this idea to electromagnetic radiation and named the photon, in 1905.

To understand quantization, try thinking of it like buying clothes. One sweater may cost \$30 at your favorite store. This is probably the only price for that style of sweater at that store. You can't buy half a sweater - they only come in units of “sweaters”, so your total bill is dependent on the number you bought. This means that the total price for multiple sweaters can only be discrete values - \$30, \$60, \$90, \$120, \$150, etc. Similarly, energy is quantitized - since electrons can only have discrete specific frequencies, the possible energies can only be specific values.

Caution: the formula $E = h \nu$ applies only to the energy of one light photon (or one quantum of energy), so to calculate the energy of a mole of photons (or quanta) would eventually involve Avogadro's number. Calculate the energy of one photon first. However, the wavelength and the frequency of a mole of photons is the same as for one photon (of the same type of radiation).

Examples:

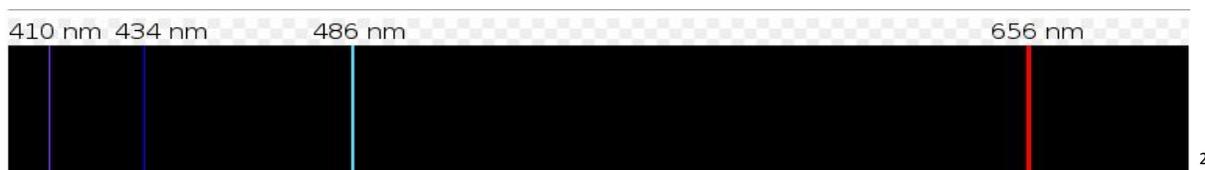
- a) One photon of red light has a wavelength of about 680 nm. What is the frequency of this light?
- b) What is the energy, in Joules and in electron volts, of one mole of photons with a wavelength of 770 nm?

D. Bohr's Atomic Model

When high voltage electricity runs through a sample of pure, gaseous element, it glows brightly. The color of that light is characteristic of that element. When you view that light in a prism or a spectrocope, you obtain a discontinuous spectrum consisting of various colored lines against a dark background. This is known as an atomic emission spectrum. Every colored line correlates with a specific wavelength, and any spectral lines that are invisible to you lie in the infrared or ultraviolet range.

Different elements give off different atomic emission spectra, and they are unique enough that you can use them to identify an unknown. This is how cesium (Cs) was discovered - Kirchhoff and Bunsen used a spectrocope, their own invention, to examine the spectrum of spa mineral water and found unique, sky-blue lines ("caesium" is Latin for "sky-blue"). If you analyze a gaseous mixture of two different elements in a spectrocope, it would produce a spectrum that is the combination of those two individual spectra, since a mixture retains the properties of each component.

When electricity runs through a gaseous hydrogen sample, three different visible light colors are seen in its emission spectrum: red, blue green, and violet. No other colors are seen. If electrons were constantly moving into different random locations inside the atom, one would expect to see all different, random colored lines instead as the electrons emit energy - so why is the atomic emission spectrum of an element so predictable? Bohr was attempting to figure this out.

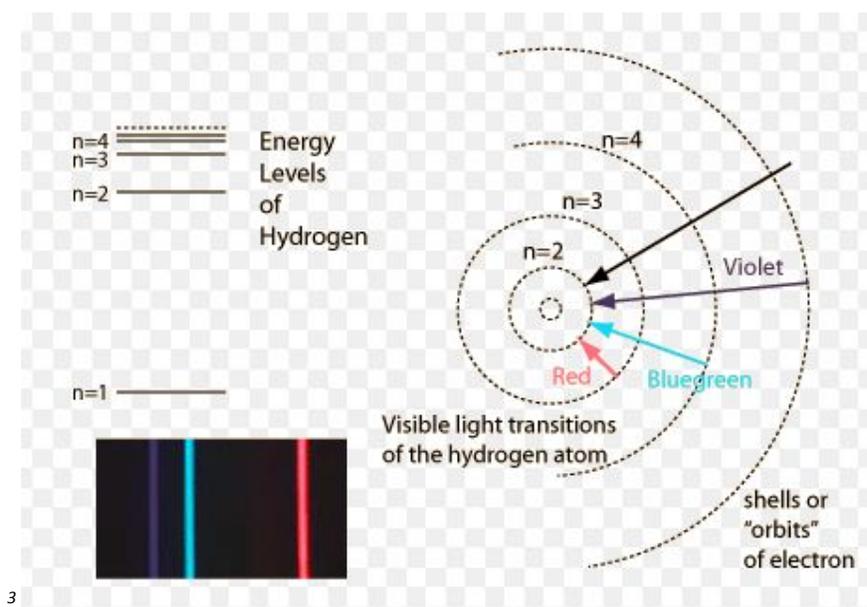


Bohr's atomic model solved this problem neatly. If the atomic emission spectrum of an element is predictable, and if electrons must have a specific energy to stay at a specific distance away from the nucleus, perhaps electrons can move between these distances and absorb or emit energy in the process. The model states:

- Electrons orbit the nucleus at discrete distances ("stationary states") away from the nucleus.

² "File:Emission Spectrum-H Labeled.svg." *File:Emission Spectrum-H Labeled.svg - Wikimedia Commons*. N.p., n.d. Web. 14 Dec. 2016.

- Each orbit corresponds to a discrete energy - an "energy level". (symbol n) Energy levels can be "empty" (i.e. no electron has that energy) or "full" (i.e. some maximum number of electrons have that energy)
- The energy of an electron depends on the size of the orbit, and a smaller orbit means a lower energy. A hydrogen atom with its (sole) electron in $n = 1$ is perfectly stable.
- Electrons usually exist in "the ground state," or in an unexcited state. While it is in the ground state, the electron is not emitting radiation.
- Electrons in the ground state will absorb energy to "jump up" to higher energy levels, thus becoming "excited" when they do.
- Excited electrons will emit energy to "fall down to" lower energy levels, or back to the ground state.
- Since the energy levels are discrete, the energy that electrons absorb and release is discrete. (This idea was taken from Planck.)
- When excited electrons fall back down to lower energy levels, they emit light of discrete energies and frequencies.
- If the light is visible, you see a photon corresponding to that energy and frequency, following the equation $E = h\nu$. Each color thus corresponds to a line in that element's atomic emission spectrum.

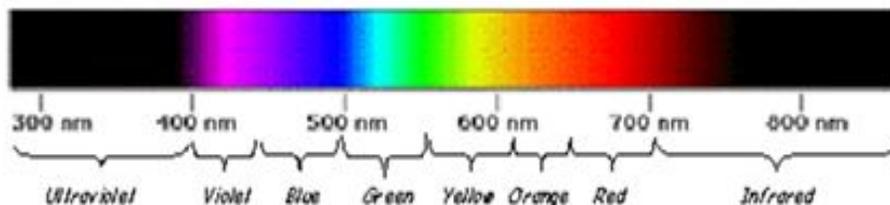


³ "Modern Atomic Theory." *JAHSCHEM - Modern Atomic Theory*. N.p., n.d. Web. 14 Dec. 2016.

4.2 Practice and Homework

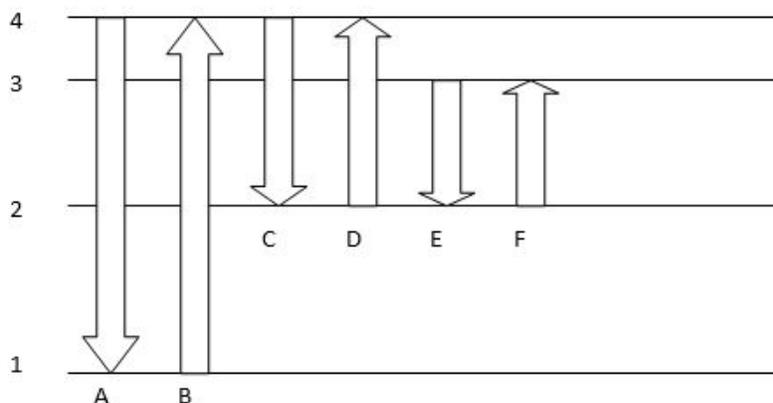
1. Calculate each quantity and fill in the chart with your answers.

Wavelength (m)	Frequency (Hz)	Energy (J)	Electromagnetic Radiation Range
0.001			
	7.0×10^{13}		
5.0×10^{-7}			
		2.0×10^{-15}	
	1.2×10^{22}		



Examine the continuous spectrum above to help you answer question #2.

2. Which color of *visible* light (i.e. ROYGBIV) has...
- the shortest wavelength?
 - the longest wavelength?
 - the lowest energy?
 - the greatest energy?
3. The “Balmer series” describes spectral line emissions in a hydrogen atom which correspond to “falls” by excited electrons from higher energy levels to $n = 2$. For each of these four “falls”, calculate the energy (in electron volts) and the frequency of the released radiation. The wavelength of each colored line is in the spectrum graphic on pg. 12.

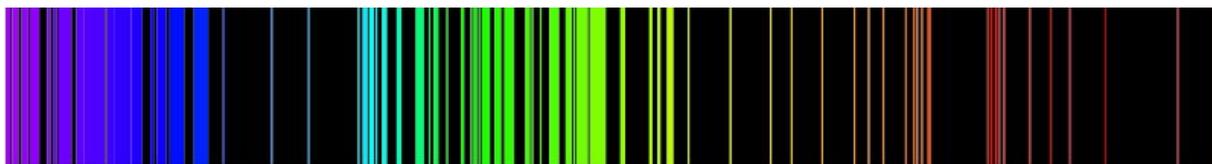


4. The figure above represents the four lowest energy levels of an atom. ($n = 1$ to 4). The six lettered arrows represent changes in the energy level of an electron.
- Why do these energy levels mean that the atom will show an emission spectrum of discrete lines rather than a continuous spectrum of emitted light?
 - Which three of the lettered energy changes involve absorption of energy by the atom?
 - Which three of the lettered energy changes involve emission of light energy by the atom?

Lesson 4.3 and 4.4: Electron Configuration and Quantum Numbers

A. Problems with Bohr's Atomic Model

A scientific model is accepted to be true if it is able to explain phenomena. While Bohr's model does an excellent job of explaining behavior in a one-electron system, such as H, Li^+ , Be^+ , etc., it cannot successfully predict multi-electron system spectra, such as iron⁴:



The Bohr model cannot predict the intensities of the lines, nor their finer structure - for instance, how can you explain why some lines are thicker or thinner, or why some lines appear to be the same color but have a different placement in the spectrum? Furthermore, Bohr's "stationary states" were purely circular (i.e. in one plane), but atoms are three dimensional.

⁴ *Wikipedia*. Wikimedia Foundation, n.d. Web. 14 Dec. 2016.

More relevant for today, the Bohr atomic model also violates the Heisenberg uncertainty principle. This principle states that the more precisely you know the location of a particle (such as an electron in an energy level), the less precisely you know its momentum. Momentum is the quantity of motion that a particle has. In Bohr's model, the position of an electron in an atom is certain, since it's a certain distance from the nucleus, and the energy of that electron is certain, because that distance is related to its energy. This is a contradiction because the velocity of an electron is related to energy: kinetic energy (KE) = $\frac{1}{2}mv^2$. In the Bohr model, you know both position and velocity; according to the uncertainty principle, you can know one, but not the other.

So, we use another atomic model to better explain experimental observations of the atom. We don't completely reject Bohr's statements; that is to say, we cite parts of it in this newer model, and refine the parts we don't.

B. The Quantum Mechanical Model

In 1926, Erwin Schrodinger published mathematical equations describing the probability of finding an electron in a certain position. In other words, you cannot be completely certain where the electrons are. The equations don't define the trajectory of the electron, as the Bohr model does, so concerns about violating the Heisenberg uncertainty principle are moot. These equations are the basis for the quantum mechanical model of the atom. The equations are based on an idea put forth by Louis deBroglie in 1924 stating that electrons in atoms have both wave and particle properties; so, moving matter has a wavelength, which is significant for electrons because of their very small mass. Schrodinger's equations treat electrons like they are waves.

The quantum mechanical model retains Bohr's idea of energy levels, but states that electrons move within an "electron cloud" around the nucleus. Electrons could be anywhere within this electron cloud, but, depending on their energy, they are more likely to be in that corresponding energy level. The energy levels are areas of higher density within the cloud.

Quantum Numbers

The probable location of an electron is based on four factors (variables in Schrodinger's equations), each represented by a different letter: its principal energy level (n), its sublevel (l), its atomic orbital (m_l), and its spin (m_s). These four values are collectively known as quantum numbers and each has a name: principal energy level, angular momentum number, magnetic quantum number and spin quantum number. There are very specific rules about what these numbers can be. Each electron in an electron cloud has a unique set of "allowable" quantum numbers that fit these rules. Sets of quantum numbers that don't follow these rules are "not allowable", meaning that electron with that set of numbers doesn't exist.

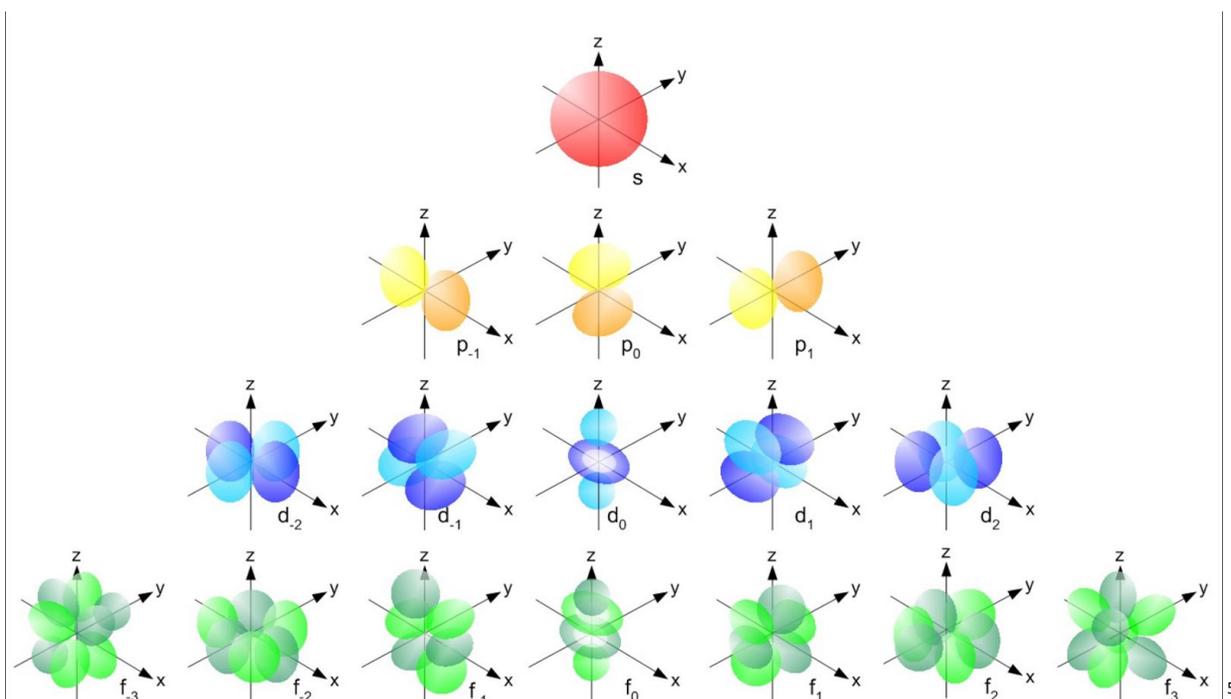
- A principal energy level (n) is one of seven levels inside the electron cloud. The first energy level, $n = 1$, is the smallest, the least energetic, and the closest to the nucleus. Each level corresponds to a period in the periodic table.

Allowable values of n must be positive integers greater than 1, i.e. $n = 1, n = 2, n = 3$, etc.

- A sublevel (l) is a volume of space within an energy level. An energy level can contain several sublevels.

Allowable values of l depend on n . l for an electron can be a positive integer ranging from 0 to $n - 1$. For example, if $n = 1$, $l = 0$. If $n = 2$, l can be 0 or 1. Each of these l values corresponds to a particular shape and orientation in space: s , p , d , or f , and to the s , p , d , and f blocks of the periodic table. $l = 0$ corresponds to s , $l = 1$ corresponds to p , $l = 2$ corresponds to d , and $l = 3$ corresponds to f .

What this means is that in $n = 1$, the only sublevel present is the s sublevel. In $n = 2$, there are two sublevels: the s and the p . $n = 3$ has three sublevels: s , p , and d . $n = 4$ has four sublevels: s , p , d , and f . We often combine energy level and sublevel: for example, an electron in "3s" is in $n = 3$ and in the s sublevel.



- An atomic orbital (m_l) is that specific volume of space within a sublevel where the electron (probably) lies. The spin quantum number (m_s) simply defines the direction of the electron as it spins on its own axis.

Allowable values of m_l depend on l , and its integers range from $-l$ to $+l$. For example, if $l = 1$ (so, the p sublevel), then m_l can be -1 , 0 , or 1 . Therefore, there are three atomic orbitals within a p sublevel. If $l = 2$ (so, the d sublevel), then m_l can be -2 , -1 , 0 , 1 , or 2 , meaning there are five orbitals within a d sublevel. Thus, an s sublevel has one orbital; a p sublevel has three; a d sublevel has five; and a f sublevel has seven.

An orbital can only hold up to two electrons, and because electrons repel each other, they

⁵ "What Is Electron Configuration?" *The Periodic Table- E Configuration*. N.p., n.d. Web. 18 Dec. 2016.

move with opposite “spin” away from each other. So, two electrons in the same orbital will have the same n value, the same l value, and the same m_l value, but they won't have the same m_s value. This is called the Pauli exclusion principle: no two electrons in the same atom can have the same four quantum numbers. There are only two allowable m_s values: $+\frac{1}{2}$ and $-\frac{1}{2}$.

The rules for quantum numbers are summarized in the following chart⁶:

n	l	m_l	Number of Orbitals	Name	Number of electrons
1	0	0	1	1s	2
2	0	0	1	2s	2
	1	-1, 0, +1	3	2p	6
3	0	0	1	3s	2
	1	-1, 0, +1	3	3p	6
	2	-2, -1, 0, +1, +2	5	3d	10
4	0	0	1	4s	2
	1	-1, 0, +1	3	4p	6
	2	-2, -1, 0, +1, +2	5	4d	10
	3	-3, -2, -1, 0, +1, +2, +3	7	4f	14

Examples:

1. Are the following sets of quantum numbers allowable? If they aren't, state the error(s).

a) $n = 1, l = 0, m_l = 0, m_s = +\frac{1}{2}$

b) $n = 2, l = 3, m_l = -2, m_s = -\frac{1}{2}$

c) $n = 2, l = -1, m_l = 2, m_s = +\frac{1}{2}$

d) $n = 3, l = -2, m_l = -3, m_s = -\frac{1}{2}$

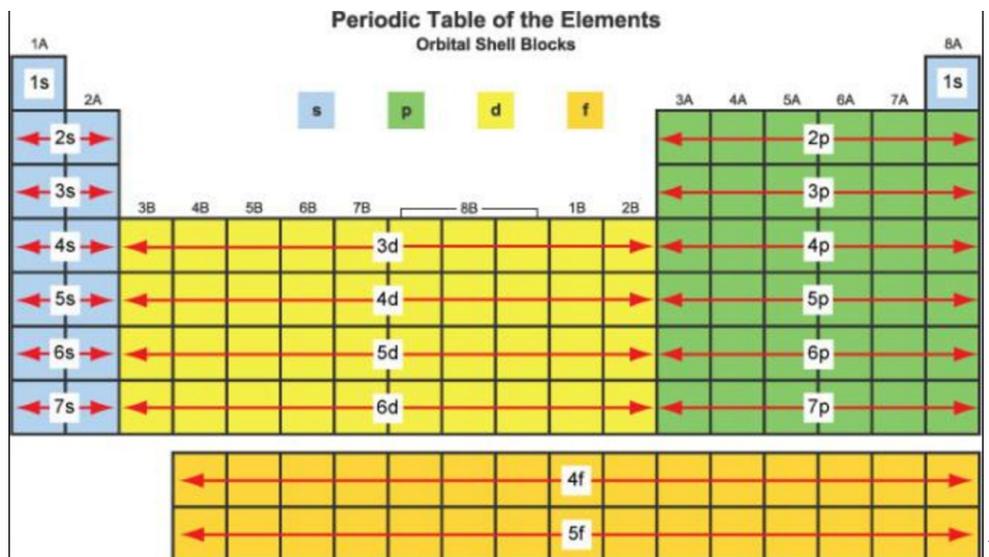
e) $n = 4, l = 1, m_l = -1, m_s = +1$

Electron Configurations

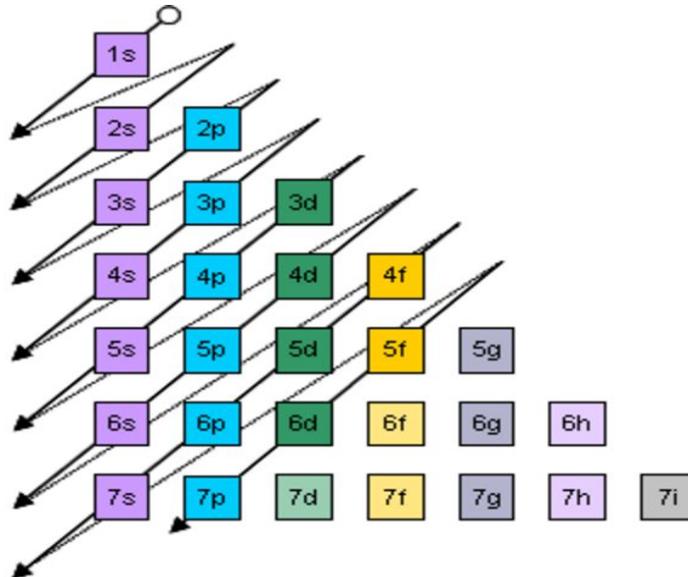
If a set of quantum numbers defines the probable location of one electron in an atom, an electron configuration shows the probable locations of *all* the electrons. Writing an electron configuration generally follows the principle that electrons will fill orbitals in order of increasing

⁶ "Quantum Numbers, Atomic Orbitals, and Electron Configurations." *Quantum Numbers, Atomic Orbitals, and Electron Configurations*. N.p., n.d. Web. 18 Dec. 2016.

energy: the “Aufbau principle”. The periodic table follows the Aufbau principle very closely - if you know where the s, p, d, and f blocks are, and know that the Aufbau principle makes you read the periodic table from left to right, then writing electron configurations is easier than you think.



Alternatively, many students like to use the following “road map”⁸ to write configurations using the Aufbau principle. Read the map starting at the circle, and follow the arrow.



Remember that an atomic orbital can only hold two electrons, and you're all set. You can also use an electron configuration to identify an atom, or identify the valence energy level and

⁷ Helmenstine, Anne Marie. "What Are Element Blocks?" *About.com Education*. N.p., 29 Nov. 2014. Web. 18 Dec. 2016.

⁸ Bigler, Jeff. "Electron Configurations." *Electron Configurations*. N.p., n.d. Web. 18 Dec. 2016.

number of valence electrons in that atom. Also, we operate under the assumption that all the lowest possible energy levels are filled up to the proper number of electrons, hence the common term “ground state electron configuration”. An electron configuration where an electron or two seem to lie in higher than normal orbitals is an “excited state electron configuration”. Ions will have electrons added to or removed from the highest orbital.

Examples:

1. Write the ground state electron configuration for an atom of each of the following elements:

- a) Hydrogen
- b) Nitrogen
- c) Titanium
- d) Platinum
- e) Bohrium

2. Which element is denoted by each ground state electron configuration? What is its valence energy level and how many valence electrons are there?

- a) $1s^2 2s^2 2p^6 3s^2$
- b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$
- c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$

3. Isoelectronic species have similar electron configurations. Which of these are isoelectronic?

- a) Li^+ , H^- , He
- b) Ca^{2+} , Ne , S^{2-}

4. Which of these are excited state electron configurations? In all cases, identify the atom.

- a) $1s^2 3p^1$
- b) $1s^2 2s^2 2p^3 3s^1$
- c) $1s^2 2s^2 2p^6 3s^2 3p^1$

For elements nearer the bottom of the periodic table, it is very tedious to write out the full, unabbreviated electron configuration. There is a shortcut: find the symbol of the noble gas coming immediately before your element, and write it in square parentheses, i.e. [Ne] for Al and [Kr] for Mo. Start at that noble gas and follow the Aufbau principle as normal. This type of configuration is called a noble gas core electron configuration, or sometimes just called the “shortcut” or “abbreviated” configuration.”

Examples:

1. Write the noble gas core ground state electron configuration for:

- a) Manganese
- b) Palladium
- c) Tantalum
- d) Meitnerium
- e) Radon (no, it isn't [Rn])

2. Identify the element.

- a) [Kr] $5s^2 4d^{10} 5p^6$
- b) [Ne] $3s^2 3p^1$
- c) [Ar] $4s^1$
- d) [Xe] $6s^2 4f^{14} 5d^7$

Electron Configuration: Exceptions to Aufbau's Principle

In several cases, the correct electron configuration looks like it represents an excited state configuration when it actually doesn't. The noble gas core electron configuration for lanthanum is [Xe] $6s^2 5d^1$. For any element in the “lanthanide series” beginning with Ce (so, elements 58 - 71), electrons fill the 4f block in order after that *and include that 5d¹*. For example, the noble gas core electron configuration for cerium is [Xe] $6s^2 4f^1 5d^1$, and for dysprosium, [Xe] $6s^2 4f^9 5d^1$.

The “actinide” series (elements 90-103) works in exactly the same way, except that they include 6d¹: actinium is [Rn] $7s^2 6d^1$, curium is [Rn] $7s^2 5f^7 6d^1$, and fermium is [Rn] $7s^2 5f^{10} 6d^1$.

The electron configurations of chromium and copper do not exactly follow the Aufbau principle either, largely because there is no other way to try to explain their observed charges. Chromium has the electron configuration [Ar] $3d^5 4s^1$, and copper is [Ar] $3d^{10} 4s^1$. This is because a half filled or a completely filled inner energy level is more stable; also, chromium and copper do have observable charges of +1, which is understandable if there is a 4s electron that is removable. This exception also applies to any elements in the chromium or copper group.

Examples:

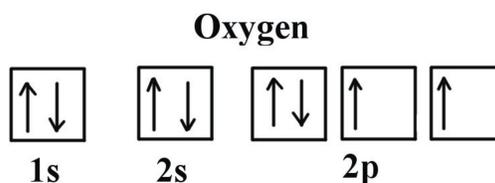
1. Write the noble gas core electron configuration for each of the following elements:

- Molybdenum
- Gadolinium
- Silver
- Americium
- Gold
- Seaborgium

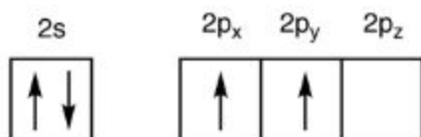
Orbital Notation

Another way of representing an electron configuration is by drawing the orbital notation. The orbital notation uses arrows pointing up or down to stand for electrons; the up or down directions denote spin ($+\frac{1}{2}$ for up, $-\frac{1}{2}$ for down), and either a short horizontal line, a circle, or a box to represent an atomic orbital. Only two arrows, or electrons, can be in a given orbital and must have opposite spin (remember, Pauli exclusion principle). Moreover, if electrons are entering, one at a time, into orbitals of equivalent energy (i.e. into the three orbitals of the 2p sublevel), then they enter individually into each of these orbitals until they are forced to pair up. This is called "Hund's rule." Orbitals of the same energy level and sublevel are called "degenerate orbitals."

For example, the orbital notation for oxygen ($1s^2 2s^2 2p^4$) is⁹:



In oxygen, there are three boxes for the 2p sublevel, since each box stands for each allowable value of m_l (-1, 0, or 1), whereas there is only one box for each of 1s and 2s because you can have only one orbital in each of those sublevels. You can draw an orbital notation using the noble gas core as well.



As with electron configurations, it's possible to write orbital notation showing only the valence orbitals and electrons. The image on the left is the valence orbital notation for carbon.¹⁰

⁹ "Hund's Rules." *Chemistry LibreTexts*. N.p., 21 July 2016. Web. 18 Dec. 2016.

¹⁰ "Hybridization." *OChemPal*. N.p., n.d. Web. 19 Dec. 2016.

Examples:

1. Draw the unabbreviated orbital notation for each of the following atoms.

a) Nitrogen

b) Silicon

c) Germanium

d) Technetium

e) Tellurium

2. Draw the noble gas core orbital notation for each of the following atoms.

a) Vanadium

b) Tungsten

c) Terbium

d) Plutonium

4.3 Practice and Homework

1. Are each of the following sets of quantum numbers allowable or not allowable? For the ones that are not allowable, state the mistake(s).

- a) $n = 4, l = 2, m_l = 0, m_s = -\frac{1}{2}$
- b) $n = 2, l = -1, m_l = 1, m_s = +\frac{1}{2}$
- c) $n = 3, l = 0, m_l = -2, m_s = -\frac{1}{2}$
- d) $n = 5, l = -3, m_l = 4, m_s = +\frac{1}{2}$
- e) $n = 4, l = 2, m_l = -1, m_s = -1$

2. Write the unabbreviated electron configurations for the following elements:

a) Sodium _____

b) Iron _____

c) Bromine _____

3. Write the noble gas core electron configurations for the following elements:

a) Cobalt _____

b) Silver _____

c) Europium _____

4. Determine what elements are denoted by the following electron configurations, and indicate the valence energy level and number of valence electrons.

a) $1s^2 2s^2 2p^6 3s^2 3p^4$

b) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

c) $[\text{Kr}] 5s^2 4d^{10} 5p^3$

5. Draw noble gas core orbital notations for each of the following atoms:

a) Silicon

b) Cobalt

c) Iridium

d) Erbium

Lesson 4.7: Periodic Trends

Because the electrons are laid out in such a regular fashion, you would expect the properties of the representative elements (s & p block) to follow regular trends. Before we talk about some of these trends, we need to review and discuss a few terms.

- Valence electrons: electrons in the highest principal energy level
- Core electrons: all other electrons

How far (in which energy level) the valence electrons are from the nucleus is going to help predict the properties of the elements. All the trends depend on the attraction of the valence electrons to the nucleus. This attraction can be explained in two ways: **effective nuclear charge (Z_{eff})** and **shielding**.

Effective nuclear charge (Z_{eff}) is the net positive charge experienced by a particular electron in an atom from the nucleus. The effective nuclear charge experienced by an electron depends on the number of electrons that shield the electron of interest. Shielding is the phenomenon where core electrons act to offset the positive charge of the nucleus as seen by an electron further away. Core electrons screen a portion of the nuclear charge from valence electrons.

The effective nuclear charge can be approximated using the equation;

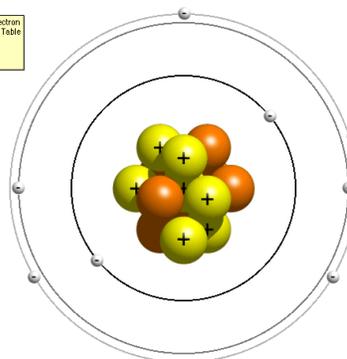
$$Z_{\text{eff}} = Z - S$$

where Z is the atomic number and S is the number of core electrons.

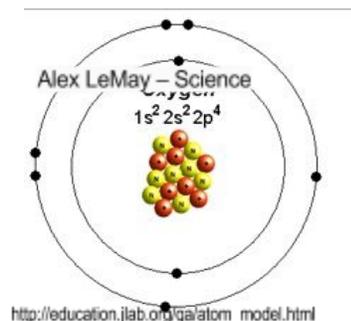
For example if we look at the 2nd period elements nitrogen and oxygen, nitrogen has 5 valence electrons, 2 core electrons, and 7 protons. Using the above equation the Z_{eff} for nitrogen is 5.

$$Z_{\text{eff}} = 7 - 2 = 5$$

Nitrogen's Electron Configuration Table
 $1s^2$
 $2s^2 2p^3$



This means that the net positive charge from the nucleus to the valence electrons is 5.

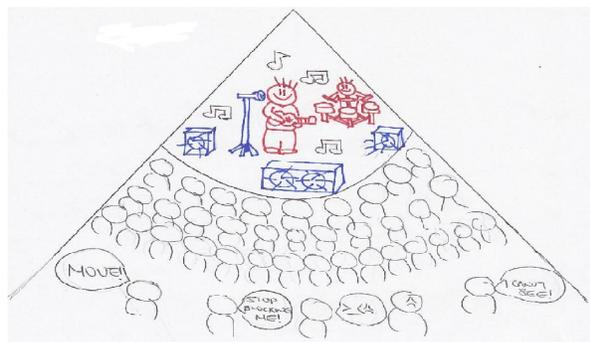


Now comparing this to oxygen, oxygen has one extra electron and one extra proton, but no new core electrons. The addition of that extra proton in the nucleus is going to increase the net positive charge and thus increase the attraction of the electrons to the nucleus, as can be seen by a Zeff of 6: $Z_{\text{eff}} = 8 - 2 = 6$.

The greater Zeff is the more attraction there is between the electrons and the nucleus implying there is a stronger pull on the electrons from the nucleus.

Xiue (Michelle) Xie, UC Davis

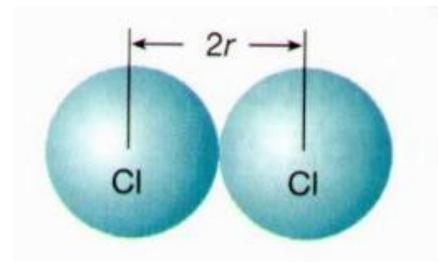
Two elements within the same group are going to have the same Zeff, however the attraction from the nucleus is not the same due to **shielding**. One way to imagine shielding is to think of being at a concert or theater performance. If you are in the front row, you have full access of sight and sound to the artists, however as the number of rows and spectators, energy levels, between you and the stage increases your view and quality of sound decreases. This is shielding. As the number of core electrons increase the net positive charge felt on the valence electrons from the nucleus is going to decrease. Shielding decreases moving down a group (increase in core electrons), but remains constant within a period (core electrons remain the same).



in short, Zeff is used to explain a trend seen within a period and shielding is used to explain a trend within a group.

Atomic Radius

Atomic radius is one-half the distance between the two nuclei in a molecule consisting of two identical atoms. Notice the trend is for atomic size to increase as you go down a column, and for atomic size to decrease as you move across a row.



As one moves across the 2nd period of the table, energy of the valence electrons remains approximately constant. The effective nuclear charge, however, increases. The electrons are being pulled toward the nucleus more strongly on the right side of the

table and thus the radii decrease as one moves from left to right.

$$Z_{\text{eff}} \text{ of Li} = 3 - 2 = +1 \quad Z_{\text{eff}} \text{ of F} = 9 - 2 = +7$$

The radius of the atom increases down a group due to an increase in the number of total energy levels. As energy levels are added, what were once valence electrons become core electrons and therefore the number of core electrons increase. This results in greater shielding and a decrease in the net positive pull from the nucleus. As shielding increases, pull decreases and the radius of the atom is going to decrease.

To sum it up, the radius of the atoms decrease from left to right across a period due to an increase in Z_{eff} and increase down a group due to an increase in shielding resulting in a reduction of the pull from the nucleus.

Ionic Radius

An ion is a charged atom that differs in the number of electrons, compared to the neutral atom. If the substance has fewer electrons than its neutral counterpart, it will have a net positive charge, if it has more electrons than protons it will have a net negative charge.

When an ion is positively charged, the net positive charge on the electrons is greater than its atom and this is going to increase the pull on the electrons from the nucleus. The increased attractions pull the electrons in closer and causes the radius to decrease in size. **Cations**, positive ions, are smaller than their atoms.

However, the more negative charge an ion has, the less tightly the electron is held. It is understood that like charges repel and as more electrons are added to the same energy level, without an increase in protons, the electrons are going to try to position themselves as far from each other as possible increasing the size of the electron cloud. Due to repulsions with other electrons, the electron cloud spreads out, causing the ion to be larger. **Anions**, negative ions, are larger than their atom. As can be seen from the image above anions are larger than cations. Metals tend to form cations and nonmetals tend to form anions.

Within a group the addition of principal energy levels is going to result in increased shielding from the nucleus and the ionic radius increases. If given a series of ions that all have the same electron configuration (isoelectronic) and you are asked to put them in order of size, place them in order of increasing atomic number, this will put them in order of largest to smallest.

		1A		2A		3A		4A		5A		6A		7A	
		Li 76	Be 31	B 20	C 15	N 146	O 140	F 133							
		1+ ●	2+ ●	3+ ●	4+ ●	3- ●	2- ●	1- ●							
		Na 102	Mg 72	Al 54	Si 41	P 212	S 184	Cl 181							
		1+ ●	2+ ●	3+ ●	4+ ●	3- ●	2- ●	1- ●							
		K 138	Ca 100	Ga 62	Ge 53	As 222	Se 198	Br 195							
		1+ ●	2+ ●	3+ ●	4+ ●	3- ●	2- ●	1- ●							
		Rb 152	Sr 118	In 81	Sn 71	Sb 62	Te 221	I 220							
		1+ ●	2+ ●	3+ ●	4+ ●	5+ ●	2- ●	1- ●							
		Cs 167	Ba 135	Tl 95	Pb 84	Bi 74									
		1+ ●	2+ ●	3+ ●	4+ ●	5+ ●									

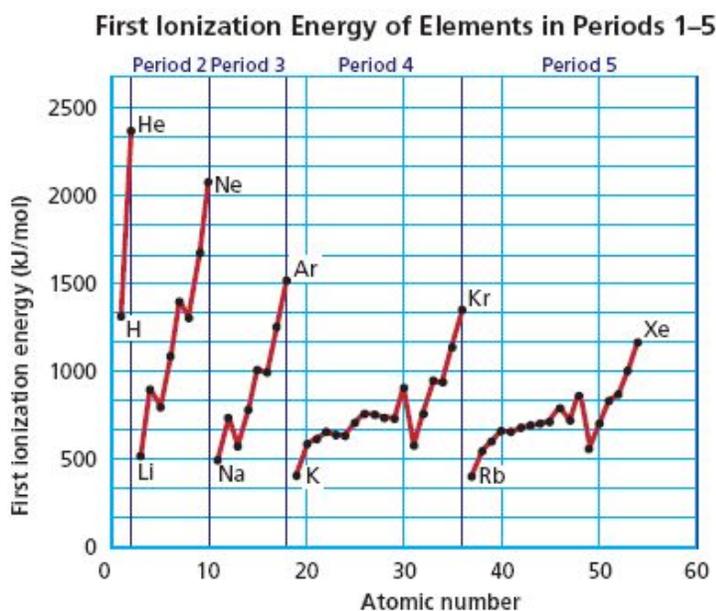
Chemical symbol — **K** 138 — Ionic radius
Charge — 1+ — Relative size

First Ionization energy

First Ionization energy is the energy required to remove a single valence electron from a neutral atom in the gaseous state. For the purposes of our class it will be the last electron placed in the atom. This roughly measures how strong the attraction are between the nucleus and the electrons. The equation belows shows the ionization of mystery element M.



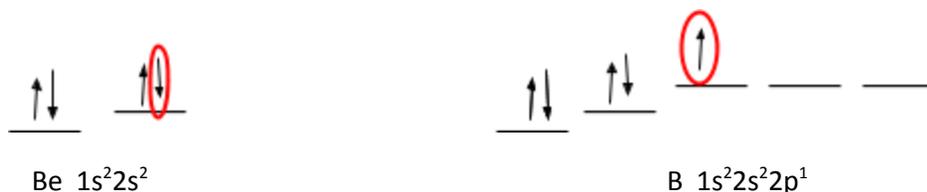
Because an electron is being removed the ionization results in the formation of a cation. Elements that have high ionization energies tend to form anions and elements with low ionization energies tend to form cations.



Referencing the image to the left, which represents periods 1-5, notice that the ionization energy increases across a period and decreases down a group.

The ionization energy decreases down a group because the outermost electron occupies a higher energy level. The higher the energy level the effects of shielding increases and less pull is felt from the nucleus, resulting in less energy required to remove the electron. Moving across a period the electron feels a greater attraction to the nucleus and the energy required to remove the electron increases, due to an increase in Z_{eff} .

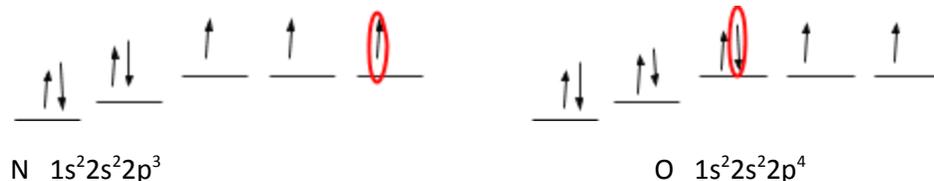
Notice there are two glitches found in every period greater than 1. B vs. Be, and O vs. N. To explain why these occur, we need to look at the electronic configurations of the elements.



In the above examples Be has a full 2s subshell and B has one electron in the 2p. The electrons in question are being removed from different subshells. Everything in nature tries to tend towards stability. Having a full subshell is preferable a partially filled subshell. Therefore it is going to require less energy to remove the electron from 2p because it is in a partially filled subshell and it is

higher in energy than the 2s. This decrease in ionization energy can always be seen between at the transition from the s to the p subshell.

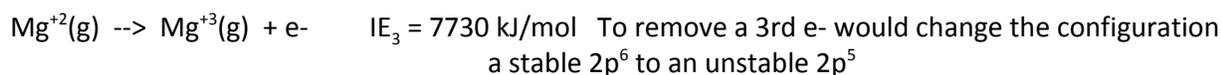
Now let's look at the discrepancy found between nitrogen and oxygen.



The ionization energy for O is lower than expected. A half filled sublevel (1 electron in each orbital) is preferable to a partially filled sublevel (any other configuration besides half filled and full). Notice that in removing the electron from oxygen, we are forming a half-filled orbital, which is especially stable. This makes the electron easier to remove in oxygen, thus giving oxygen a lower ionization energy than nitrogen. This discrepancy can be found whenever there is a change in configuration from half full to partially full. The higher the ionization energy is the harder is it to remove the electron and the more attraction there is from the nucleus to the electron.

Higher Ionization Energies (2nd, 3rd)

When considering taking off electrons past the first ionization energy, one needs to look at the placement of the valence electrons.



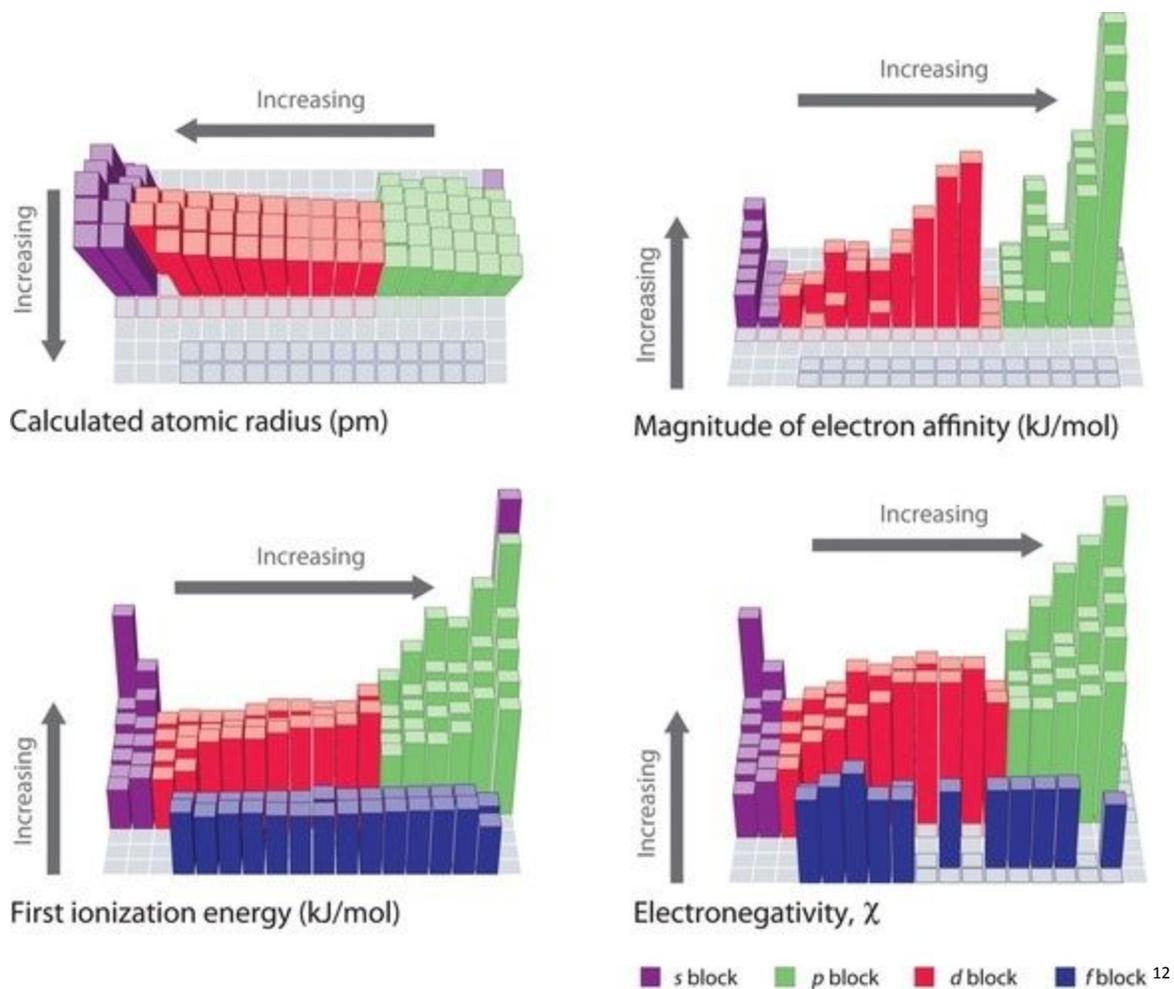
Notice three things from above:

1. All IE's are endothermic. An electron is always more stable around a nucleus, therefore it always takes energy to remove an electron.
2. Every subsequent removal of an electron takes more energy. This is because with each additional charge, the remaining electrons are held tighter due to an increase pull from the nucleus.
3. The value for IE_3 skyrockets. This is due to the fact that in the third IE, you are removing a core electron, and these are very stable and difficult to remove.

Electronegativity (EN)

Electronegativity was developed by Linus Pauling. It is the relative ability of an atom to attract electrons while simultaneously being in a chemical bond. It is assigned a numerical value of 3.98 or less. Fluorine's small size (due to an increase pull from the nucleus, or large Zeff) ensures that it will have the highest electronegativity. Conversely, francium's large size and high degree of shielding prevents it from attracting electrons and it will have the lowest electronegativity. Noble gases are typically assigned an electronegativity of 0.0 because of their full octets and stable configuration.

To sum up the periodic trends:



¹² https://qph.ec.quoracdn.net/main-qimg-3945d7d4517722d481258d0e23bd54d5-c?convert_to_webp=true

4.7 Periodic Trends Practice/Homework

1. Briefly explain each of the following statements.
 - a) The atomic radius of Mg is smaller than Ba.

 - b) The ionization energy of S is lower than P.

 - c) The energy required to remove the fourth electron in Al is significantly larger than the third electron.

 - d) The energy change associated with gaining an electron, the electron affinity, is positive for some atoms and negative for others.

2. Of cesium, Cs, hafnium, Hf and gold, Au which element has the smallest atomic radius? Explain your answer using the trends and the concepts of Zeff and shielding.

3. Without looking at the electron affinity table, arrange the following elements in order of decreasing electron affinities: C, O, Li, Na, Rb, and F.

4. Without looking at the ionization energy table, arrange the following elements in order of decreasing first ionization energies: Ba, Ga, Si, Cl, Tl, and S. Which of the elements listed above would you expect to have the highest second ionization energy? Explain your reasoning.

5. Which of the following cations is least likely to form? Sr^{2+} , Al^{3+} , or K^{2+}

6. Which of the following anions is least likely to form? F^- , N^{3-} , or Cl^{2-}

7. Which element is most electronegative among the following? C, N, O, Br, or C

8. The two ions K^+ and Ca^{2+} each have 18 electrons surrounding the nucleus. Which would you expect to have the smaller radius? Why?

9. The increase in the effective nuclear charge as one moves from left to right across a period can be used to explain the fact that
 - a. ionization energy and atomic radius both increase
 - b. ionization energy and atomic radius both decrease
 - c. ionization energy decreases and atomic radius increases
 - d. ionization energy increases and atomic radius decreases

10. Which one of the following atoms has the largest radius? F, Cl, O, or S

11. Metals tend to have:
 - a. low ionization energies because they gain electrons easily.
 - b. high ionization energies because they lose electrons easily.
 - c. high ionization energies because they gain electrons easily.
 - d. low ionization energies because they lose electrons easily.