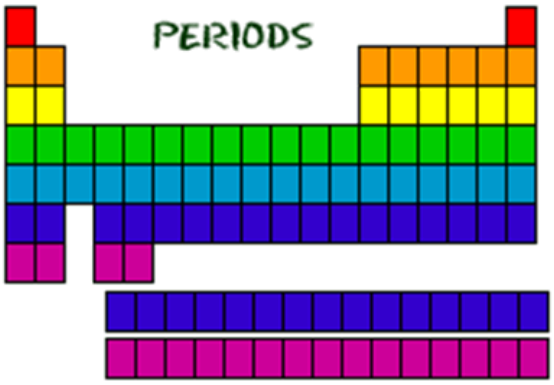
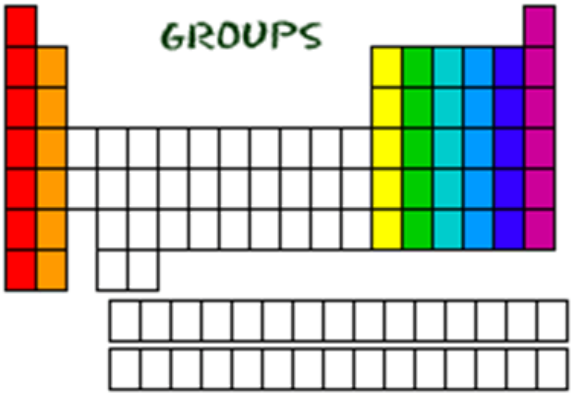
Name _____ Block _____

Notes: Unit 6
Electron Configuration and the Periodic Table

- In the 1790's Antoine Lavoisier compiled a list of the known elements at that time. There were only _____.
- By the 1870's _____ were known. And a system of organization was needed.
- John Newlands proposed an organization system based on increasing atomic mass in 1864.
- He noticed that both the chemical and physical properties repeated every 8 elements.
- He called this the _____.
- In 1869 both Lothar Meyer and Dmitri Mendeleev showed a connection between atomic mass and an element's properties.
- Mendeleev published first, and is given credit for this.
- He also noticed a periodic pattern when elements were ordered by increasing _____.
- By arranging elements in order of increasing atomic mass into columns, Mendeleev created the first Periodic Table.
- This table also predicted the existence and properties of undiscovered elements.
- After many new elements were discovered, it appeared that a number of elements were out of order based on their _____.
- In 1913 Henry Mosley discovered that each element contains a unique number of _____.
- By rearranging the elements based on _____, the problems with the Periodic Table were corrected.
- This new arrangement creates a periodic repetition of both physical and chemical properties known as the _____.

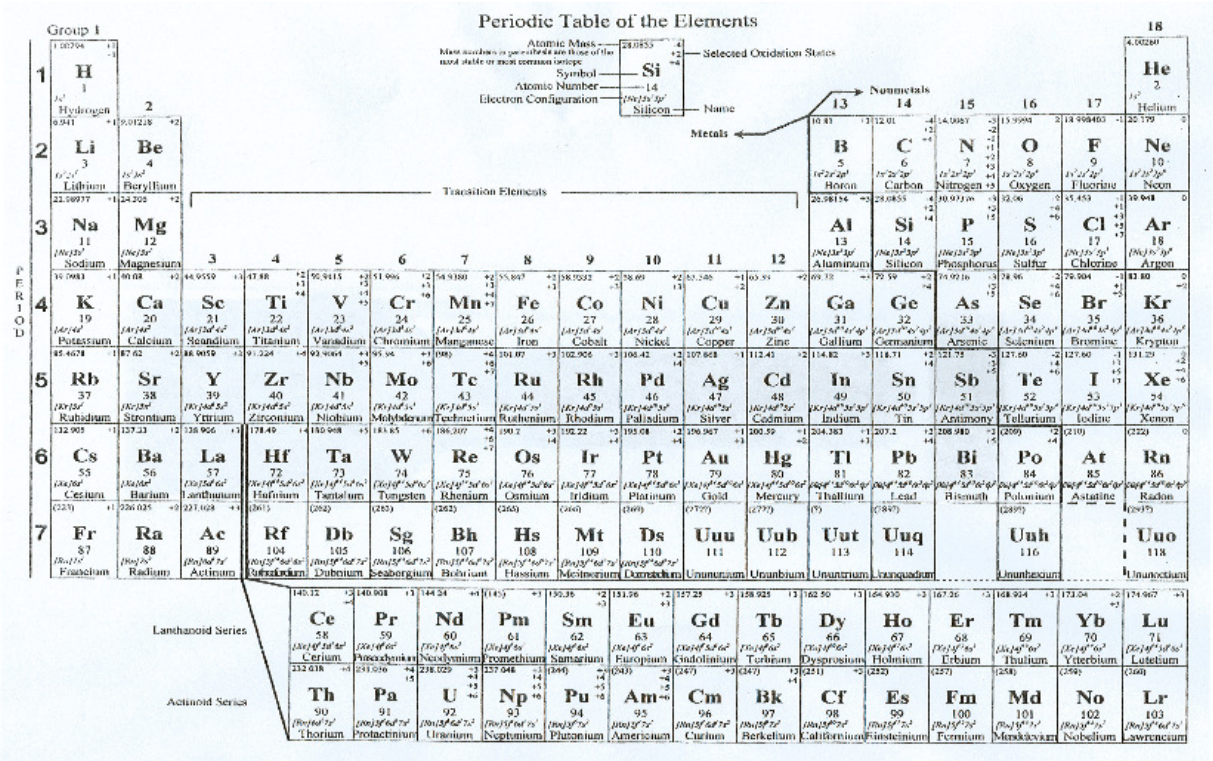




Periods are the _____
Valence electrons across a period are in the same energy level

Groups/Families are the _____
There are equal numbers of valence electrons in a group.

- Valence electrons are



- Hydrogen Group-

- Alkali Metals-

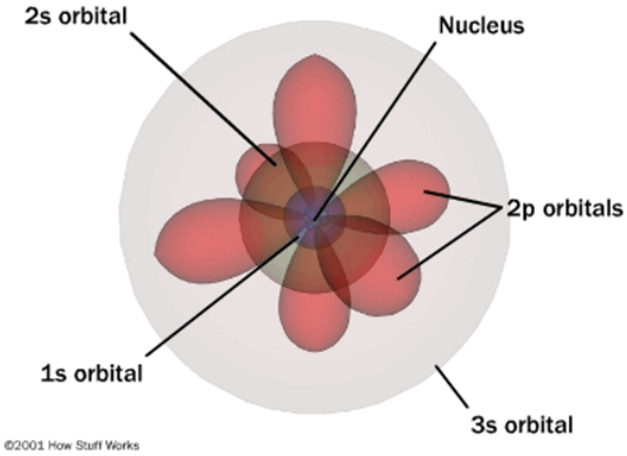
- Alkali Earth Metals-

- Noble Gases-

- Halogens-

Quantum Model Notes

- **Bohr** proved that the _____ an electron is from the nucleus means more energy it has and that there is no _____ energy
- **Heisenberg's Uncertainty Principle-** Can determine either the _____ _____ of an electron, **cannot** determine both.
- **Schrödinger's Equation** - Developed an equation that treated the _____ atom's electron as a wave.
    - Only limits the electron's energy values, does not attempt to describe the electron's path.
- Describe _____ of finding an electron in a given area of orbit.
- The **Quantum Model**- atomic orbitals are used to describe the possible position of an electron.



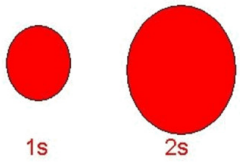
Orbitals

- The location of an electron in an atom is described with 4 terms.
    - **Energy Level-** Described by _____. The higher the level, the more energy an electron has to have in order to exist in that region.
    - **Sublevels-** energy levels are divided into sublevels. The # of sublevels contained within an energy level is equal to the integer of the _____.
    - **Orbitals-** Each sublevel is subdivided into orbitals. Each orbital can hold _____ electrons.
    - **Spin-** Electrons can be spinning clockwise (+) or counterclockwise (-) within the orbital.

Orbital Diagrams

Energy Level

- Indicates relative sizes and energies of atomic orbitals. Whole numbers, ranging from _____.
- The energy level is represented by the letter _____.

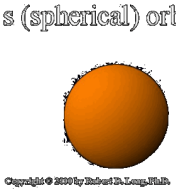


Sublevels

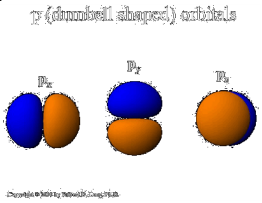
- Number of sublevels present in each energy level is equal to the n.
- Sublevels are represented by the letter _____.
- In order of increasing energy:

Orbitals

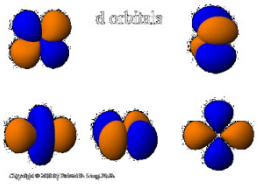
- Represented by m_l
- S Sublevel- Only _____ orbital in this sublevel level.



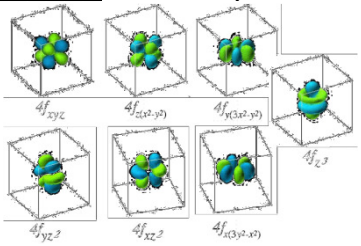
- P Sublevel- _____ orbitals present in this sublevel.
  - Each orbital can only have 2 electrons.



- D Sublevel- _____ orbitals present in this sublevel.



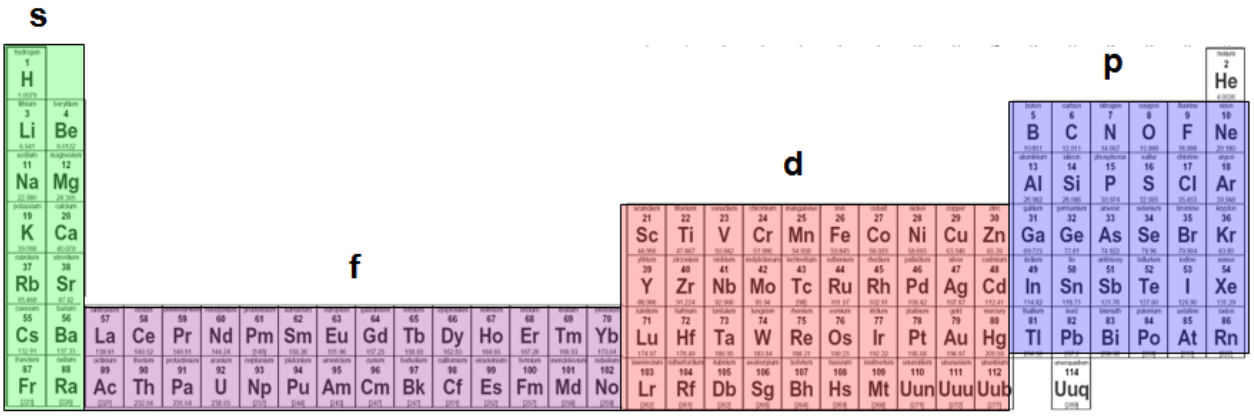
- F Sublevel- _____ orbitals present in this sublevel.



Energy Level	Sublevels Present	# of Orbitals	Total # of Orbitals in Energy Level	Total # of Electrons in Energy Level
1				
2				
3				
4				

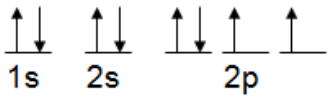
Orbital Diagrams

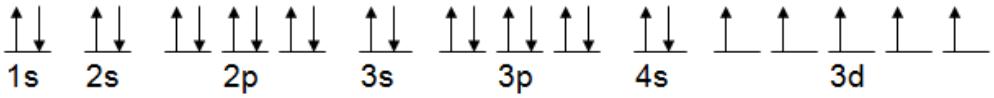
- An orbital diagram shows the arrangement of electrons in an atom.
- The electrons are arranged in energy levels, then sublevels, then orbitals. Each orbital can only contain 2 electrons.
- Three rules must be followed when making an orbital diagram.
    - Aufbau Principle- An electron will occupy the _____ energy orbital that can receive it.
        - To determine which orbital will have the lowest energy, look to the periodic table.
    - Hund's Rule- Orbitals of equal energy must each contain _____ _____ before electrons begin pairing.
    - Pauli Exclusion Principle- If two electrons are to occupy the same orbital, they must be spinning _____.

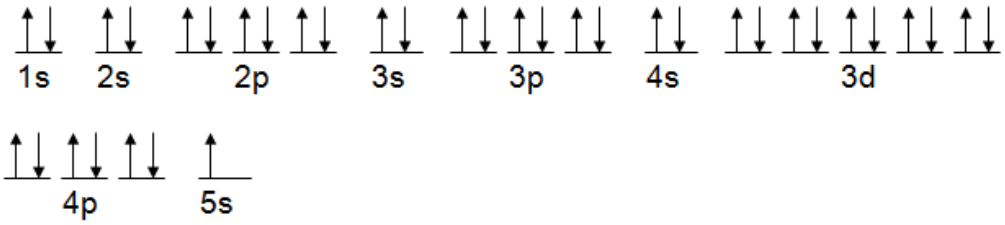


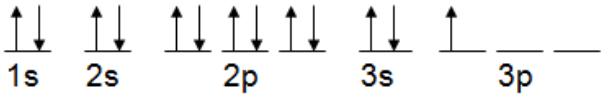
- Energy Levels (n) determined by the ROWS
- Sub Levels (s,p,d,f)- determined by the sections
- Orbitals - determined by the # of columns per sublevel

Orbital Diagrams

$_{8}O$ = 

$_{25}\mathbf{Mn}$ = 

$_{37}Rb$ = 

$_{13}Al$ = 

- S
- As
- Pb
- N
- Sc

Orbital Diagrams WS

Give the orbital diagram for the following elements:

1. Mg

2. Cu

3. Sb

4. N

5. Na

6. Al

7. W

8. Ag

9. B

Electron Configurations and Oxidation States

- Electron configurations are shorthand for orbital diagrams. The electrons are not shown in specific orbitals nor are they shown with their specific spins.
- Draw the orbital diagram of oxygen:

- The electron configuration should be:

- Manganese

- Rubidium

- Aluminum

- The Noble Gas shortcut can be used to represent the electron configuration for atoms with many electrons. Noble gases have a full s and p and therefore can be used to represent the inner shell electrons of larger atoms.
- For example: Write the electron configuration for Lead.

$$Pb = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^2$$

- Write the electron configuration for Xenon.

$$Xe = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$$

- Substitution can be used:

$$Pb = [Xe] 6s^2 4f^{14} 5d^{10} 6p^2$$

- Manganese

- Rubidium

- Aluminum

- Valence electrons, or outer shell electrons, can be designated by the s and p sublevels in the highest energy levels
- Write the noble gas shortcut for Bromine

$$Br = [Ar]4s^2 3d^{10} 4p^5$$

- Write only the s and p to represent the valence level.

$$Br = 4s^2 4p^5$$

- This is the Valence Configuration. Bromine has 7 valence electrons.
- Silicon

- Uranium

- Lead

Octet Rule and Oxidation States

- The octet rule states the electrons need _____ valence electrons in order to achieve maximum stability. In order to do this, elements will gain, lose or share electrons.
- *Write the Valence configuration for oxygen*

O = $2s^2 2p^4$- **6 valence electrons**

- *Oxygen will gain 2 electrons to achieve maximum stability*

O = $2s^2 2p^6$- **8 valence electrons**

  - Now, oxygen has 2 more electrons than protons and the resulting charge of the atom will be -2
  - The symbol of the _____ formed is now O^{-2}.
- Elements want to be like the Noble Gas family, so they will gain or lose electrons to get the same configuration as a noble gas.
- When an element gains or losses an electron, it is called an _____.
- An ion with a positive charge is a _____.
- An ion with a negative charge is an _____.

Element	Total # of electrons	Valence Configuration	Gain or Lose e^-	How Many?	Ion Symbol	New Valence Configuration	Total # of e^-
Cs	55	$6s^1$	L	1	Cs^+	$5s^25p^6$	54
Cl							
Rb							
Ca							
Na							

Element or Ion	Symbol	# protons	# electrons	# neutrons	Valence configuration
Calcium (+2)	Ca^{2+}	20	18	20	$3s^23p^6$
Aluminum (+3)					
Barium (0)					
Sulfur (-2)					
Potassium (+)					

Give the <u>noble gas shortcut</u> configuration for the following elements:

1. Pb

2. Eu

3. Sn

4. As

Give ONLY the <u>outer shell configuration</u> for the following elements:

1. Ba

2. Po

3. S

4. F

Quantum Number Notes

- The quantum mechanical model uses three quantum numbers, n, *l* and m_l to describe an orbital in an atom. A fourth quantum number, m_s, describes an individual electron in an orbital.
- ***n* =** _____
    - This describes the energy level and can be described as an integer from 1-7. The larger the number, the larger the orbital. As the numbers increase, the electron will have greater energy and will be less tightly bound by the nucleus.
- $\mathscr{l}$ = _____
    - This describes the shape of the orbital level and can be described as an integer from 0 to n-1.
    - 0 is used to describe s orbitals; 1 is used to describe p orbitals.
    - 2 is used to describe d orbitals; 3 is used to describe f orbitals.
- m_l= _____
    - This describes the orientation of the orbital in space and can be described as an integer from ***-l* to *l***
    - s sublevels have one orbital, therefore possible values of m_l include 0 only.
    - p sublevels have three orbitals, therefore possible values of m_l include -1, 0, 1.
    - d sublevels have five orbitals, therefore possible values of m_l include -2, -1, 0, 1, 2.
    - What about f?

- m_s= _____
    - This describes the spin of the electron in the orbital.
    - The possible values for m_s are +½ and ½.
    - The positive spin reflects the first electron in a specific orbital and negative spin reflects the second electron in an orbital.
- *Examples:*
    - Give the four quantum numbers for the 8th electron in Argon:

    - What is the maximum number of electrons that can have the following quantum numbers: $n = 2$ and $m_s = -½$

    - Which of the following quantum numbers would **NOT** be allowed in an atom
        $n = 2$, $l = 2$ and $m_l = -1$
        $n = 4$, $l = 2$ and $m_l = -1$
        $n = 3$, $l = 1$ and $m_l = 0$
        $n = 5$, $l = 0$ and $m_l = 1$

Quantum Numbers Worksheet

Rules for assigning quantum numbers:

***n*:** can be 1, 2, 3, 4, … Any positive whole number

ℓ*:** can be 0, 1, 2, … (n***-1) Any positive whole number, up to (***n***-1)

m_l**:** can be (-*ℓ*), (-*ℓ*+1), (-*ℓ*+2), … 0, 1, … (*ℓ*-1), ***l*** Any integer, from – *ℓ* to *ℓ*.

m_s**:** can be +½ or - ½

1) If ***n***=2, what possible values does *ℓ* have?

2) When *ℓ* is 3, how many possible values of m_l are there?

3) What are the quantum numbers for the 17th electron of Argon?

4) What are the quantum numbers for the 20th electron of Chromium?

5) What are the quantum numbers for the 47th electron of Iodine?

6) Give the quantum numbers for ALL of the electrons in Nitrogen.

Determine if the following sets of quantum numbers would be allowed in an atom. If not, state why and if so, identify the corresponding atom.

7) $n = 2, \ell = 1, m_l = 0, m_s = +\frac{1}{2}$

8) $n = 4, \ell = 0, m_l = 2, m_s = -\frac{1}{2}$

9) $n = 1, \ell = 1, m_l = 0, m_s = +\frac{1}{2}$

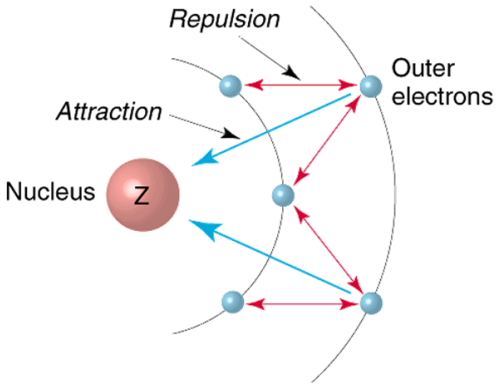
10) $n = 6, \ell = 2, m_l = -2, m_s = +\frac{1}{2}$

11) $n = 3, \ell = 3, m_l = -3, m_s = -\frac{1}{2}$

Periodic Trends- Notes

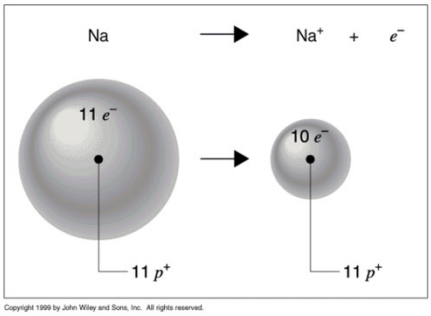
- Effective Nuclear Charge: The *s*, *p*, *d*, and *f* orbitals within a given shell have slightly different energies.
    - The difference in energies between subshells result in electron-electron repulsion which _____outer electrons from the nucleus.
    - The _____ nuclear charge felt by an electron is called the *effective nuclear charge* (Z_{eff}).

- Z_{eff} is lower than actual nuclear charge.
- *Zeff* increases toward nucleus
  ns > np > nd > nf
- This explains certain periodic changes observed.

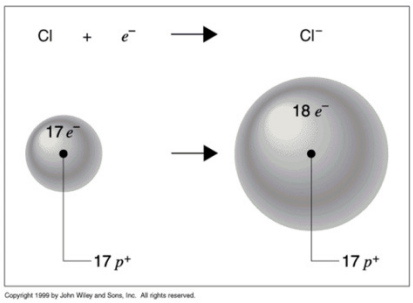


- Shielding: As you go down the periodic table, the number of shells increases which results in greater electron-electron repulsion.
    - The more shells there are, the further from the nucleus the valence electrons are.
    - Therefore, more shielding means the electrons are _____ attracted to the nucleus of the atom.
- Atomic Radius is defined as half the distance between adjacent nuclei of the same element.
    - As you move DOWN a group an entire energy level is added with each new row, therefore the atomic radius _____.
    - As you move LEFT-TO-RIGHT across a period, a proton is added, so the nucleus more strongly attracts the electrons of a atom, and atomic radius _____.

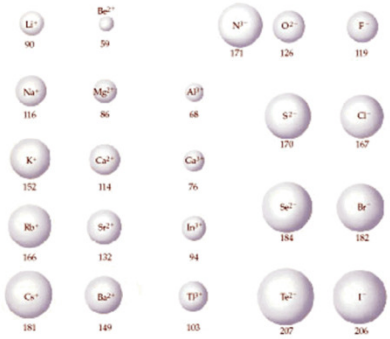
- Ionic Radius is defined as half the distance between adjacent nuclei of the same ion.
    - For _____ an electron was lost and therefore the ionic radius is smaller than the atomic radius.



    - For _____ an electron was gained and therefore the ionic radius is larger than the atomic radius.



    - As you move down a group an entire energy level is added, therefore the ionic radius increases.
    - As you move left-to-right across a period, a proton is added, so the nucleus more strongly attracts the electrons of a atom, and ionic radius _____.
        - However! This occurs in 2 sections. The cations form the first group, and the anions form the second group.

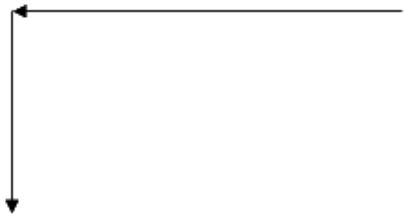


- Isoelectronic Ions: Ions of different elements that contain the same number of electrons.

- Ionization energy is defined as the energy required to _____ the first electron from an atom.
    - As you move down a group atomic size increases, allowing electrons to be further from the nucleus, therefore the ionization energy _____.
    - As you move left-to-right across a period, the nuclear charge increases, making it harder to remove an electron, thus the ionization energy _____.

- Electronegativity is defined as the relative ability of an atom to attract electrons in a _____.
    - As you move down a group atomic size increases, causing available electrons to be further from the nucleus, therefore the electronegativity _____.
    - As you move left-to-right across a period, the nuclear charge increases, making it easier to gain an electron, thus the electronegativity _____.
- Reactivity is defined as the ability for an atom to react/combine with other atoms.
    - With reactivity we must look at the metals and non-metals as two separate groups.
- Metal Reactivity- metals want to lose electrons and become cations
    - As you move down a group atomic size increases, causing valence electrons to be further from the nucleus, therefore these electron are more easily lost and reactivity _____.
    - As you move left-to-right across a period, the nuclear charge increases, making it harder to lose electrons, thus the reactivity _____.
- Non-metal Reactivity- non-metals want to gain electrons and become anions
    - As you move down a group atomic size increases, making it more difficult to attract electrons, therefore reactivity _____.
    - As you move left-to-right across a period, the nuclear charge increases, making it easier to attract electrons, thus the reactivity _____.

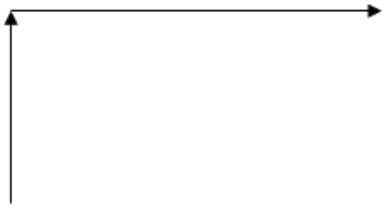
Atomic Radius



Ionic Radius

Ionization Energy

Electronegativity



Reactivity

Periodic Trends

1. Explain why a magnesium atom is smaller than both sodium AND calcium.

2. Would you expect a Cl^- ion to be larger or smaller than a Mg^{2+} ion? Explain.

3. Which effect on atomic size is more significant, the nuclear charge or the energy level that electrons are filling? Explain.

4. Explain why the sulfide ions (S^{2-}) is larger than a chloride ion (Cl^-).

5. Compare the ionization energy of sodium to that of potassium and EXPLAIN.

6. Explain the difference in ionization energy between lithium and beryllium.

7. Order the following ions from largest to smallest: Ca^{2+}, S^{2-}, K^{+}, Cl^{-}. Explain your order.

1. Rank the following atoms/ions in each group in order of decreasing radii and **explain** your ranking for each.

   a. I, I^{-}

   b. K, K^{+}

   c. Al, Al^{3+}

2. Which element would have the greatest electron affinity: B or O? Explain. **Hint: a positive electron affinity means that the element wants to form a negative charge.**

Review

Give the <u>Orbital Diagram</u> for the following elements:

1. Chromium

2. Nitrogen

3. Strontium

4. Aluminum

Give the <u>COMPLETE electron configuration</u> for the following elements:

5. Argon

6. Phosphorous

7. Iron

8. Uranium

Complete the table.

Element	Total # of electrons	Valence Configuration	Gain or Lose e^-	How Many?	Ion Symbol	New Valence Configuration	Total # of e^-
Phosphorous							
Chlorine							
Cesium							
Lithium							

Give the 4 quantum numbers for the last electron of the following elements:

9. Phosphorous

10. Manganese

11. Silver

12. Promethium

13. Iodine

Determine if the following sets of quantum numbers would be allowed in an atom. If not, explain why and if so, identify the corresponding atom.

14. n = 2, l = 1, m_l = 0, $m_s = +\frac{1}{2}$

15. n = 4, l = 0, m_l = 2, $m_s = -\frac{1}{2}$

16. n = 1, l = 1, m_l = 0, $m_s = +\frac{1}{2}$

Give the element with the LARGER radius, ionization energy, electronegativity and reactivity.

ELEMENTS	ATOMIC RADIUS	IONIZATION ENERGY	ELECTRONEGATIVITY	REACTIVITY
Sodium and Aluminum				
Chlorine and Iodine				
Oxygen and Fluorine				
Magnesium and Calcium				

Circle the element / ion with the larger radius.

17. Mg or Mg^{2+}

18. S or S^{2-}

19. N^{3-} or F^{-}

20. Sr^{2+} or Br^{-}

21. Cl^{-} or Mg^{2+}

22. B or F

For each of the following families, give their relative reactivity, the number of valence electrons, and at least one additional piece of information (such as how they are found in nature or what other group the generally react with).

23. Alkaline Earth Metals

24. Alkali Metals

25. Halogens

26. Noble Gases