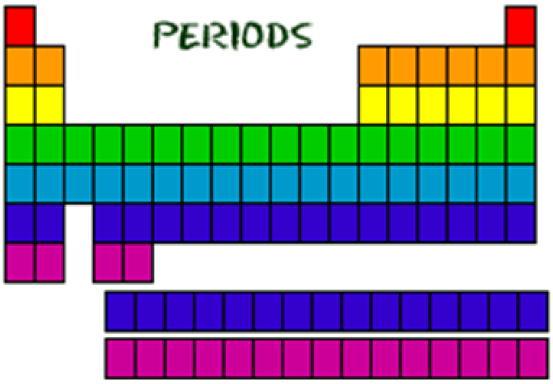
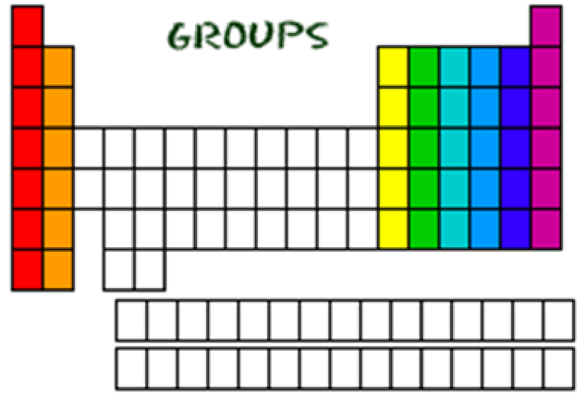
Name ________KEY______ Block _______

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 _____ 23 elements___________.
- By the 1870's _______ 70 elements________ 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 ______ Law of Octaves_____.
- 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 ____ atomic mass_____________.
- 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 ____properties_____.
- In 1913 Henry Mosley discovered that each element contains a unique number of _____protons_______.
- By rearranging the elements based on ____atomic number______, the problems with the Periodic Table were corrected.
- This new arrangement creates a periodic repetition of both physical and chemical properties known as the ____periodic law____.





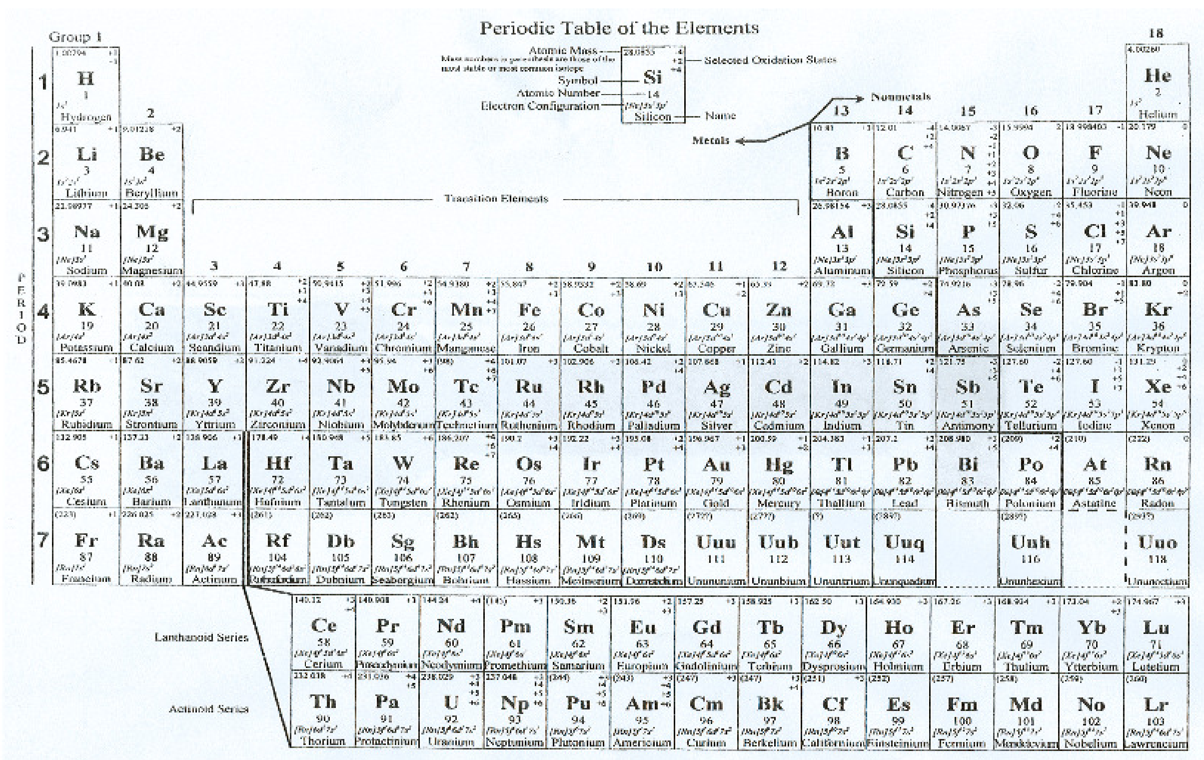
Periods are the __ rows___

Valence electrons across a period are in the same energy level

Groups/Families are the __ columns_

There are equal numbers of valence electrons in a group.

- Valence electrons are the electrons in the highest energy level of the atom (the electrons on the outside)



- Hydrogen Group-H

  Can act both as a metal and nonmetal

- Alkali Metals- Li, Na, K, Rb, Cs, Fr (group 1)
  Highly reactive; not found uncombined in nature; form stable compounds

- Alkali Earth Metals (**Alkalines**)- Be, Mg, Ca, Sr, Ba, Ra (group 2)

  Are less reactive than group 1; form basic solutions when reacted with water; usually found combined with other nonmetals in the Earth's crust

- Noble Gases- He, Ne, Ar, Kr, Xe, Rn (group 18)

  This family is considered inert, because they do not easily react; they all have a full valence shell, making them stable

- Halogens- F, Cl, Br, I, At

  The elements in this family form salts when they combine with other elements; at room temperature they exist as solids, liquids, and gases; these are the most reactive non-metals

Added:

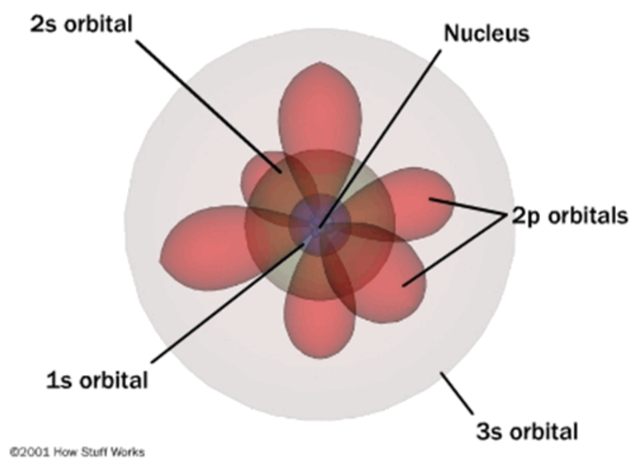
- **Metals-** solids at room temperature (except Mercury); malleable (able to be bent); ductile (able to be pulled into a fine wire); shiny (luster); good conductor of heat and electricity
- **Nonmetals-** many are gases, Bromine is a liquid at room temperature; not malleable, not ductile; not shiny; poor conductor
- **Semimetals (metalloids)** - properties similar to both metals and nonmetals. Si- shiny, high melting pt., poor conductor of electricity (compared to most metals), but can conduct electricity at temperatures where most metals would have melted.
- **Transition Metals (Groups 3-12):**
  These have various colors (most of the elements that we think of as metals). These elements are very hard, with high melting points and boiling points.

Quantum Model Notes

- **Bohr** proved that the ___further away____ an electron is from the nucleus means more energy it has and that there is no _in between______ energy

Heisenberg's Uncertainty Principle- Can determine either the __ velocity OR the

- position___________________________ of an electron, **cannot** determine both.
- **Schrödinger's Equation** - Developed an equation that treated the __ hydrogen__ atom's electron as a wave.
    - Only limits the electron's energy values, does not attempt to describe the electron's path.
- Describe __ probability____ 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.



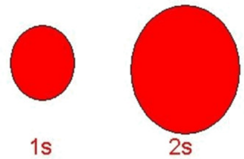
Orbitals

- The location of an electron in an atom is described with 4 terms.
    - **Energy Level-** Described by __ integers__. 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 _____ energy level_______.
    - **Orbitals-** Each sublevel is subdivided into orbitals. Each orbital can hold __2____ 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 ________1 to 7________.
- The energy level is represented by the letter __*n*__.



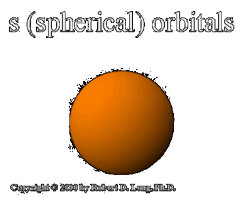
Sublevels

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

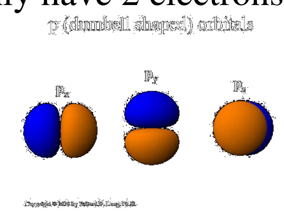
  s<p<d<f

Orbitals

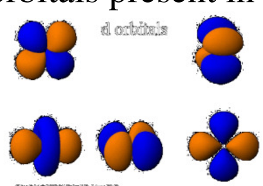
- Represented by m_l
- S Sublevel- Only ___**1**____ orbital in this sublevel level.



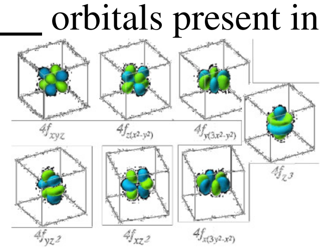
- P Sublevel- ______**3**_____ orbitals present in this sublevel.
  - Each orbital can only have 2 electrons.



- D Sublevel- ______**5**_____ orbitals present in this sublevel.



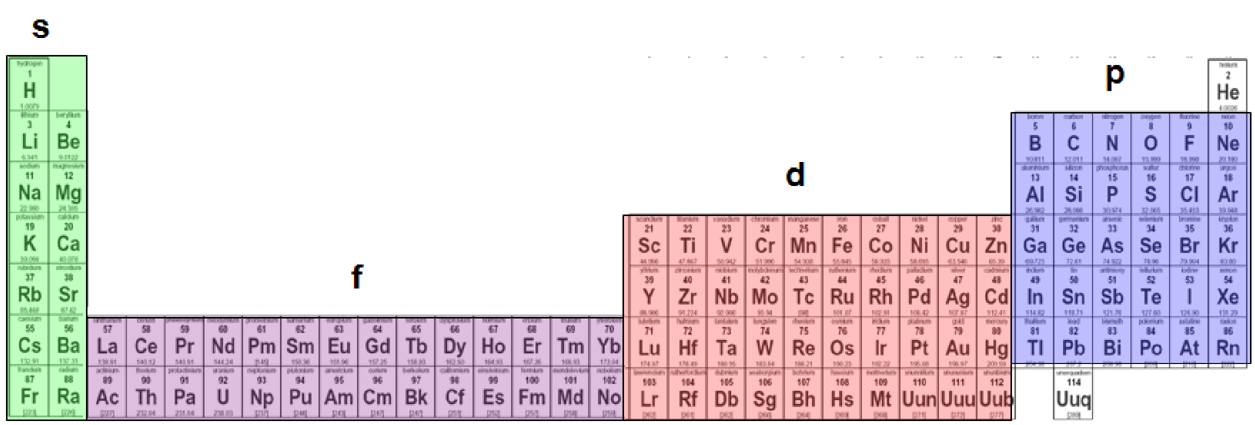
- F Sublevel- _______**7**______ orbitals present in this sublevel.



Energy Level	Sublevels Present	# of Orbitals	Total # of Orbitals in Energy Level	Total # of Electrons in Energy Level
1	s	1	1	2
2	s, p	1, 3	4	8
3	s, p, d	1, 3, 5	9	18
4	s, p, d, f	1, 3, 5, 7	16	32

Orbital Diagrams

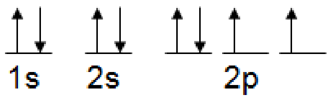
- An orbital diagram shows the arrangement of electrons in an atom.
- The electrons are arranged in energy levels, then sublevels, then orbitals. Each orbital can only contain 2 electrons.
- Three rules must be followed when making an orbital diagram.
    - Aufbau Principle- An electron will occupy the _____lowest____ 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 ____________ __one electron___ before electrons begin pairing.
    - Pauli Exclusion Principle- If two electrons are to occupy the same orbital, they must be spinning _____in opposite directions_________.



- 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

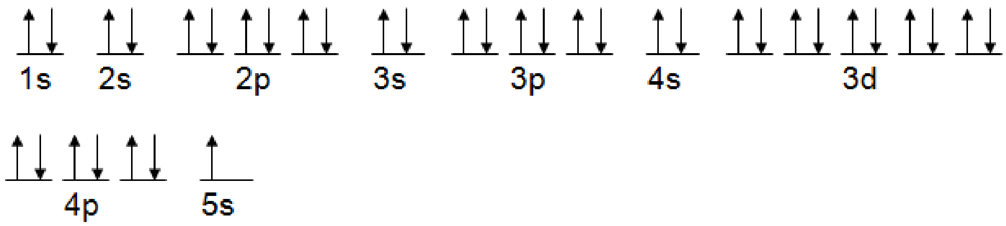
Orbital Diagrams

$_{8}O$ =

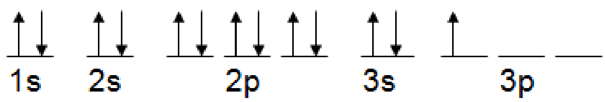


Mn =

$_{37}Rb$ =



$_{13}Al$ =



- S

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

- As

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

- Pb

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

- N

  $1s^2 2s^2 2p^3$

- Sc

  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$

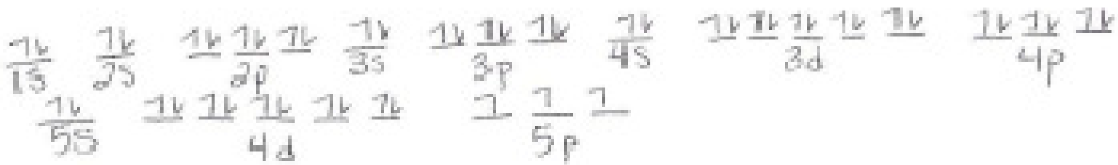
Orbital Diagrams WS

Give the orbital diagram for the following elements:

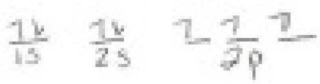
1. Mg

2. Cu

3. Sb

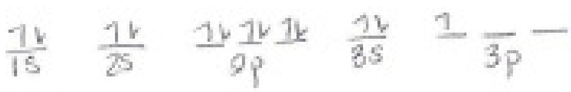


4. N

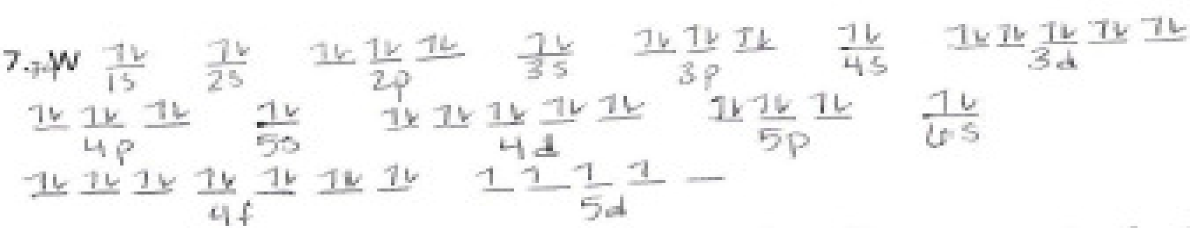


5. Na

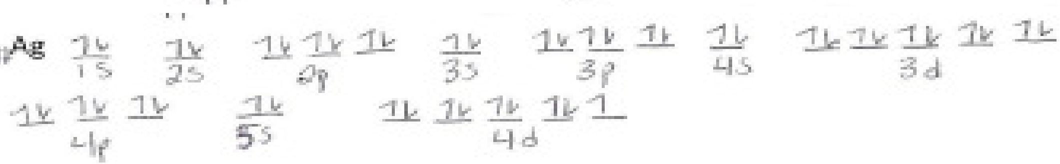
6. Al



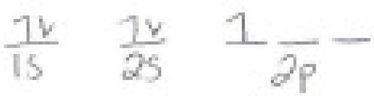
7. $_{74}$W



8. $_{47}$Ag

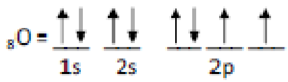


9. B



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:

  $1s^2 2s^2 2p^4$

- Manganese

  $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$

- Rubidium

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

- Aluminum

  $1s^2 2s^2 2p^6 3s^2 3p^1$

- 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

  [Ar] $4s^2 3d^5$

- Rubidium

  [Kr] $5s^1$

- Aluminum

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

- 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
  $[Ne]3s^2 3p^2$ **$3s^2 3p^2$** **4 valence electrons**

- Uranium
  $[Rn]7s^2 5f^4$ **$7s^2$** **2 valence electrons**

- Lead
  $[Xe]6s^2 4f^{14} 5d^{10} 6p^2$ **$6s^2 6p^2$** **4 valence electrons**

Octet Rule and Oxidation States

- The octet rule states the electrons need __**eight**__ 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 ___**ion**____ 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 __**ion**___.
- An ion with a positive charge is a ____**cation (lost electrons)**_____.
- An ion with a negative charge is an ___**anion (gained electrons)**___.

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	17	$3s^23p^5$	G	1	Cl^{-1}	$3s^23p^6$	18
Rb	37	$5s^1$	L	1	Rb^{+1}	$4s^24p^6$	36
Ca	20	$4s^2$	L	2	Ca^{+2}	$3s^23p^6$	18
Na	11	$3s^1$	L	1	Na^{+1}	$2s^22p^6$	10

Element	Symbol	# protons	# electrons	# neutrons	Valence Configuration
Calcium (+2)	Ca^{2+}	20	18	20	$3s^23p^6$
Aluminum (+3)	Al^{+3}	13	10	14	$2s^22p^6$
Barium (0)	Ba	56	56	81	$6s^2$
Sulfur (+2)	S^{-2}	16	18	16	$3s^23p^6$
Potassium (+)	K^+	19	18	20	$3s^23p^6$

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*** **= principle quantum number (energy level quantum number)**
  - 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.
- ℓ = **_azimuthal quantum number (sublevel quantum number)**
  - 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= **magnetic quantum number (orbital quantum number)**
  - 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= **electron spin quantum number**
  - 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:

    n = 2, l = 1, m_l = -1, m_s = -1/2

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

    4 (2 electrons in the s sublevel, 4 in the p sublevel, that's 8 electrons total, half have +1/2 spin and half have -1/2 spin)

  - Which of the following quantum numbers would **NOT** be allowed in an atom

    n = 2, *l* = 2 and m_l = -1 **not allowed (when n=2, l cannot = 2)**
    n = 4, *l* = 2 and m_l = -1 **allowed**
    n = 3, *l* = 1 and m_l = 0 **allowed**
    n = 5, *l* = 0 and m_l = 1 **not allowed (when l = 0, m_l can only = 0)**

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

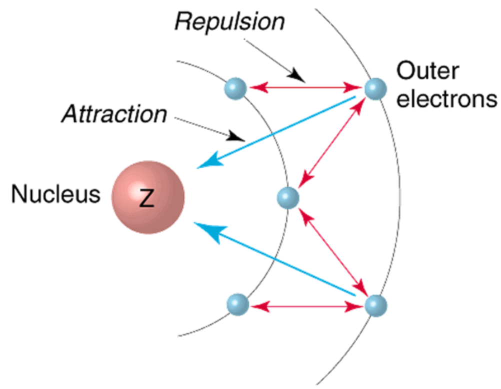
10) $n = 6, \ell = 2, m_l = -2, m_s = +\frac{1}{2}$

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

Periodic Trends- Notes

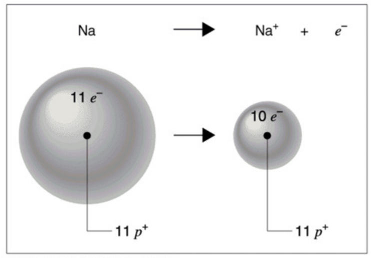
- Effective Nuclear Charge: The *s*, *p*, *d*, and *f* orbitals within a given shell have slightly different energies.
    - The difference in energies between subshells result in electron-electron repulsion which __**SHIELDS**_____outer electrons from the nucleus.
    - The _**NET**_ nuclear charge 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.

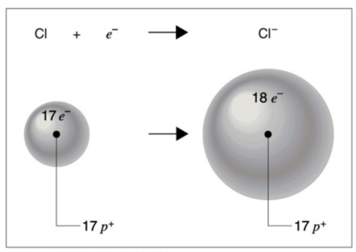


- Shielding: As you go down the periodic table, the number of shells increases which results in greater electron-electron repulsion.
    - The more shells there are, the further from the nucleus the valence electrons are.
    - Therefore, more shielding means the electrons are __**lESS**______ attracted 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 ___**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 atomic radius ____**DECREASES**________.

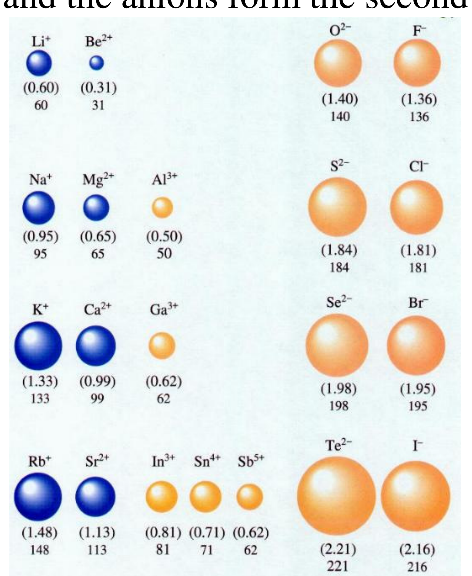
- Ionic Radius is defined as half the distance between adjacent nuclei of the same ion.
  - For __CATIONS___ an electron was lost and therefore the ionic radius is smaller than the atomic radius.



  - For ___ANIONS_____ 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 ___DECREASES__.
    - 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 ___REMOVE_____ 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 ____DECREASES___.
  - As you move left-to-right across a period, the nuclear charge increases, making it harder to remove an electron, thus the ionization energy _____INCREASES______.

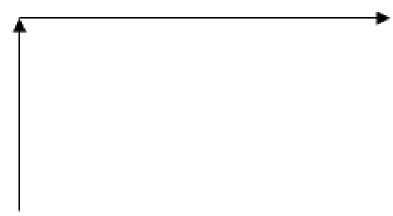
- Electronegativity is defined as the relative ability of an atom to attract electrons in a ___CHEMICAL BOND_____.
  - As you move down a group atomic size increases, causing available electrons to be further from the nucleus, therefore the electronegativity ___DECREASES______.
  - As you move left-to-right across a period, the nuclear charge increases, making it easier to gain an electron, thus the electronegativity _____ INCREASES_______.
- 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 ____INCREASES______.
  - As you move left-to-right across a period, the nuclear charge increases, making it harder to lose electrons, thus the reactivity ___DECREASES_.
- 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 _____DECREASES___.
  - As you move left-to-right across a period, the nuclear charge increases, making it easier to attract electrons, thus the reactivity ___INCREASES_.

Atomic Radius

Ionic Radius

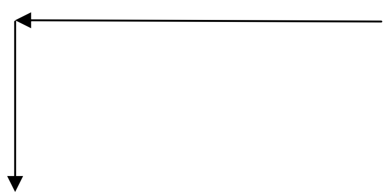
Ionization Energy

Electronegativity



Reactivity

Metals



NonMetals

Periodic Trends

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

   Mg has less energy levels than Ca, so it is smaller.
   Mg has more protons than Na, so the electrons (which are in the same energy level) are held in tighter, making Mg smaller than Na.

2. Would you expect a Cl^- ion to be larger or smaller than a Mg^{2+} ion? Explain.
   Cl^- would be larger than Mg^{+2} because when e- are gained, the protons can't hold them in as tightly, whereas when e- are lost, the remaining e- are held in more tightly by the protons.

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

   The energy levels that electrons are filling have more of an effect than the nuclear charge. When nuclear charge increases, it is only by one proton. When you add an energy level, you are adding at least 8 electrons.

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

   S^{-2} ions have gained 2 electrons, while Cl^- ions have only gained 1 electron, so more e- cause the ion to be larger (more spread out)

5. Compare the ionization energy of sodium to that of potassium and EXPLAIN.
   IE for Na is greater than IE for K. Electrons are harder to remove from Na, so a higher energy is required, because the e- are held in tighter by the nucleus in Na because there are less energy levels

6. Explain the difference in ionization energy between lithium and beryllium.
   The IE for Be is higher than the IE for Li. The electrons in Be are held in tighter by the nucleus because there are more protons in Be than in Li (the electrons are all in the same energy level), so the energy required to remove an electron from Be is higher than the energy required to remove an electron from Li.

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

$S^{-2} > Cl^{-} > K^{+} > Ca^{+2}$ Elements that gain e- are larger; the more e- that is gained, the larger the ion. Elements that lost e- are smaller; the more e- lost, the smaller the ion.

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

a. I, I^{-} $I^{-} > I$ because when you add e-, the atom becomes larger, since there are more electrons than protons
b. K, K^{+} $K > K^{+}$ because when you remove e-, the atom becomes smaller, because the remaining e- are held in tighter by the protons.
c. Al, Al^{+3} $Al > Al^{+3}$ same explanation as above

9. 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.**
O would have a greater electron affinity (electronegativity) because the net nuclear charge (Zeff) is higher when the e- are held in tighter (because of more protons), so it's easier to attract other electrons.

Review

Give the Orbital Diagram for the following elements:

1. Chromium (Exception!) 24e-

↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑↓ 3s ↑↓ ↑↓ ↑↓ 3p ↑ 4s ↑ ↑ ↑ ↑ ↑ 3d

2. Nitrogen 7e-

↑↓ 1s ↑↓ 2s ↑ ↑ ↑ 2p

3. Strontium 38e-

↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑↓ 3s ↑↓ ↑↓ ↑↓ 3p ↑↓ 4s ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ 3d ↑↓ ↑↓ ↑↓ 4p ↑↓ 5s

4. Aluminum 13e-

↑↓ 1s ↑↓ 2s ↑↓ ↑↓ ↑↓ 2p ↑↓ 3s ↑ __ __ 3p

Give the COMPLETE electron configuration for the following elements:

5. Argon 18e-

$1s^2 2s^2 2p^6 3s^2 3p^6$

6. Phosphorous 15e-

$1s^2 2s^2 2p^6 3s^2 3p^3$

7. Iron 26e-

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

8. Uranium 92e-

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 6p^6 7s^2 5f^4$

Complete the table.

Element	Total # of electrons	Valence Configuration	Gain or Lose e⁻	How Many?	Ion Symbol	New Valence Configuration	Total # of e⁻
Phosphorous	15	$3s^2 3p^3$	gain	3	P^{-3}	$3s^2 3p^6$	18
Chlorine	17	$3s^2 3p^5$	gain	1	Cl^{-1}	$3s^2 3p^6$	18
Cesium	55	$6s^1$	lose	1	Cs^{+1}	$5s^2 5p^6$	54
Lithium	3	$2s^1$	lose	1	Li^{+1}	$1s^2$	2

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

9. Phosphorous

1 1 (1) 3p   n = 3   l = 1 m_l = +1 m_s = +1/2

10. Manganese

1 1 1 1 (1) 3d   n = 3   l = 2 m_l = +2 m_s = +1/2

11. Silver Exception!

1↓ 1↓ 1↓ 1↓ (1↓) 4d   n = 4   l = 2 m_l = +2 m_s = -1/2

12. Promethium

1 1 1 (1) _ _ _ 4f   n = 4   l = 3 m_l = 0 m_s = +1/2

13. Iodine

1↓ (1↓) 1 5p   n = 5   l = 1 m_l = 0 m_s = -1/2

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 = $+\frac{1}{2}$ valid! (middle orbital) (up arrow) 2nd E level, p sublevel

1 1 _ 2p Carbon

15. n = 4, *l* = 0, m_l = 2, m_s = $-\frac{1}{2}$ not valid!
m_l cannot equal 2 when l = 0!

16. n = 1, *l* = 1, m_l = 0, m_s = $+\frac{1}{2}$ not valid!
l cannot equal 1 when n = 1!

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

ELEMENTS	ATOMIC RADIUS	IONIZATION ENERGY	ELECTRONEGATIVITY	REACTIVITY
Sodium and Aluminum	Na	Al	Al	Na
Chlorine and Iodine	I	Cl	Cl	Cl
Oxygen and Fluorine	O	F	F	F
Magnesium and Calcium	Ca	Mg	Mg	Ca

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

- 2nd most reactive group of metals
2 valence electrons
react with nonmetals
group 2

24. Alkali Metals

most reactive group of metals
1 valence electron
react with nonmetals (esp. halogens)
- found combined w/ other elements in nature

25. Halogens

- most reactive group of nonmetals
- 7 valence electrons
- reacts with alkalis and other metals

26. Noble Gases

- nonreactive group
- 8 valence electrons
- found uncombined in nature