Unit 06:

Periodic Trends

Packet Grade:_____________/100pts

- **Reading: 30pts**
- **Worksheets: 10pts each**
- **Review Sheet: 20pts**

Lab 6A Grade:________/100pts

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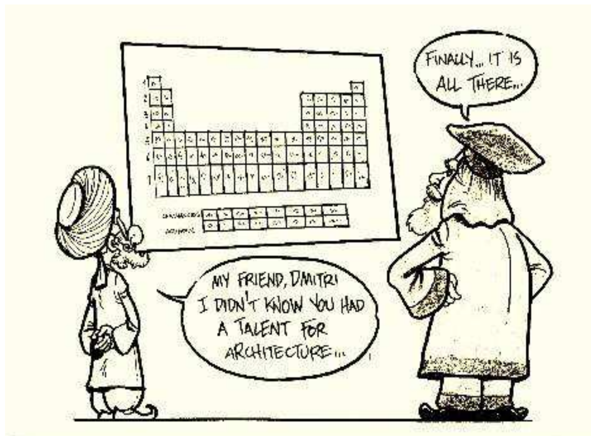


Introductory Resources:

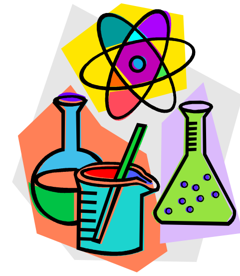
Addison-Wesley v.5 - Chapter 14

Addison-Wesley v.4 - Chapter 12

Addison-Wesley v.3 - Chapter 12

Main Idea Summary:

- Elements that have similar properties also have similar electron configurations and are members of the same group.
- Regular changes in the electron configuration of the elements cause gradual changes in both the physical and chemical properties within a group and within a period.
- Atomic radii decrease as you move from left to right in a given period.
- Ionization energy increases as you move from left to right in a given period.
- Atomic radii increase within a given group because the outer electrons are farther from the nucleus as you go down the group.
- Ionization energy decreases as you move down through a group.



Name:______________________________

Period: 1 2 3 4 5 6 7 8

Chemistry Unit 6 Bell Question/DLO sheet

Date	DLO	Bell Question	Bell Answer

Unit 06 – Periodic Trends NAME: ______________________

Reading Comprehension Sheet

13.1 Development of the Periodic Table

Dmitri Mendeleev

Henry Mosley

13.2 The Modern Period Table

Period

Groups

Periodic Law

Representative Elements

Periodic (use a dictionary not in the book)

Use the definition of dictionary to explain how the periodic table got its name.

13.3 Electron Configuration and Periodicity

Noble Gas

Representative Element

Alkali Metal

Alkaline earth metal

Halogen

Transition Metal

Inner transition metal

How can you tell an element belongs to group 1A

How can you tell an element belongs to group 4A

13.4 Periodic Trends in Atomic Size

Atomic Radii

Atomic Size _____________________ as you move ______________ a group.

Why?

Atomic Size _____________________as you move ________________ a period..

Why?

13.5 Periodic Trends in Ionization Energy

Ionization Energy

Ionization Energy _____________________ as you move ______________ a group.

Why?

Ionization Energy _____________________as you move ________________ a period..

Why?

13.6 Periodic Trends in Ionic Size

Ionization Energy

Ionic Size ____________________ as you move _____________ a group.

Why?

Ionic Size ____________________as you move _______________ a period..

Why?

13.7 Periodic Trends in Electronegativity

Ionization Energy

Electronegativity ____________________ as you move _____________ a group.

Why?

Electronegativity ____________________as you move _______________ a period..

Why?

Draw Figure 13.10 on page 367

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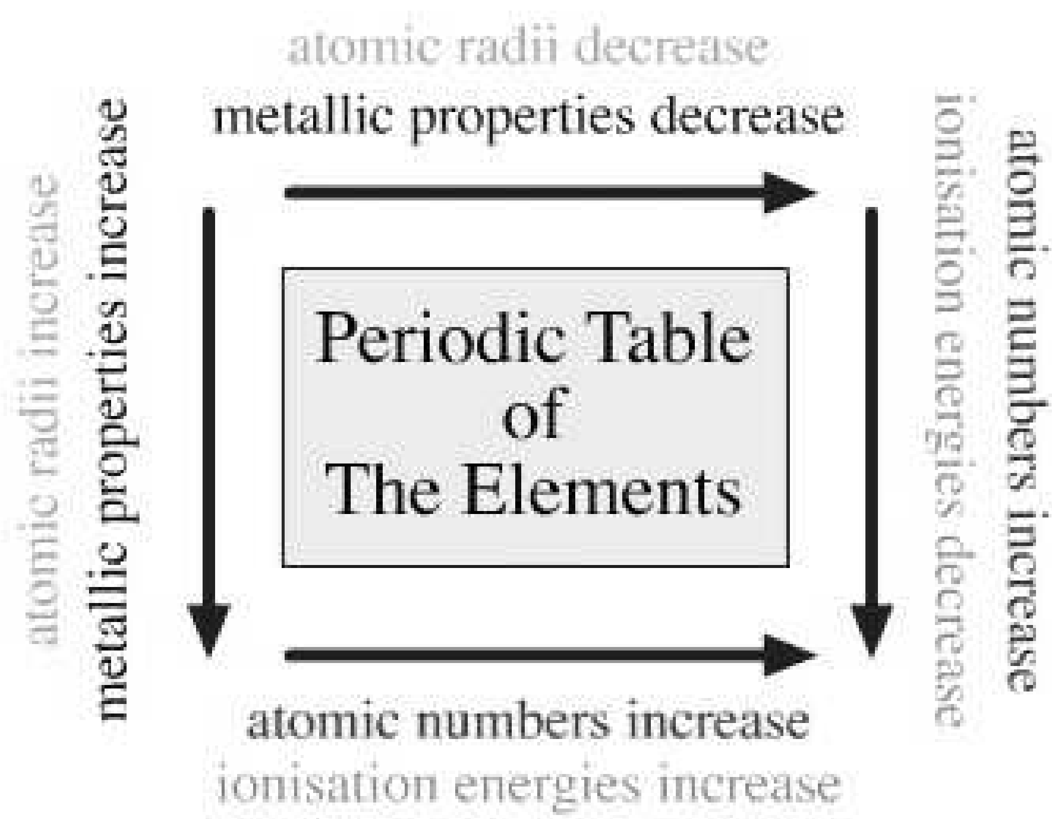
Discussion Sheet 6a – Periodic Trends

The Periodic Table was developed by Dmitri Mendeleev in the mid-1800s. Mendeleev, a Russian, put all of the elements in order based on mass, and noticed a periodic reoccurrence of chemical and physical properties. He arranged the elements in columns. Elements in each column have similar properties. Occasionally, he would find a hole in the table. He could use the surrounding information to predict the properties of a yet to be discovered element. As new elements were discovered, they neatly filled holes in the periodic table.

As scientists began to work with Mendeleev's table, some changes were made. The vertical columns became known as families or groups. Horizontal rows are known as periods. Most importantly, instead of putting elements in order of mass, elements are now arranged in order of atomic number.

Similar electron configurations are found in vertical columns. The elements of the first columns end in s^1. The elements of the second column end in s^2. The similarity of the electron configurations of atoms in vertical columns is why atoms are similar, vertically.

The repetitive nature of the periodic table allows scientists to know something about an element based on where it is located in the periodic table. The following diagram summarizes some of the more important repeating trends found in the periodic table. These trends will be dealt with in more detail on a subsequent Discussion Sheet.



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Worksheet 6.01 – Completion

Use this completion exercise to check your knowledge of the terms and your understanding of the concepts introduced in this chapter. Each blank can be completed with a term, short phrase, or number.

The periodic table organizes the elements into vertical ________________ and horizontal ________________ in order of increasing ________________. The table is constructed so that elements that have similar chemical properties are in the same ________________. The elements in Groups 1A through 7A are called the ________________. The ________________ make up Group 0. The elements in Groups 2A and 3A are interrupted in periods 4 and 5 by the ________________ and in periods 6 and 7 by the ________________.

The atoms of the noble gas elements have their outermost *s* and ________________ sublevels filled. The outermost *s* and *p* sublevels of the representative elements are ________________.

Atomic radii generally ________________ as you move from left to right in a period. Atomic size generally ________________ within a given group because there are more ________________ occupied and an increased ________________ effect, despite an increase in nuclear ________________.

The energy required to remove an electron from an atom is known as the ________________ energy. This quantity generally ________________ as you move left to right across a period. The ease with which an atom gains an electron, or the ________________, decreases as you move ________________. The ability of a bonded atom to attract electrons to itself is known as ________________, and this quantity ________________ as we move from left to right across a period.

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Discussion Sheet 6b – Electron Shorthand Configuration

The periodic table is organized in such a way that the elements in a group have similar characteristics. The rightmost column, Group 0 is for the Noble Gases. These elements that make up the Noble Gas group have the outermost s and p sublevels filled. Electron configuration can be tedious and take up much time. Therefore, there is a shorthand method available, using the noble gases.

Shorthand method Steps:

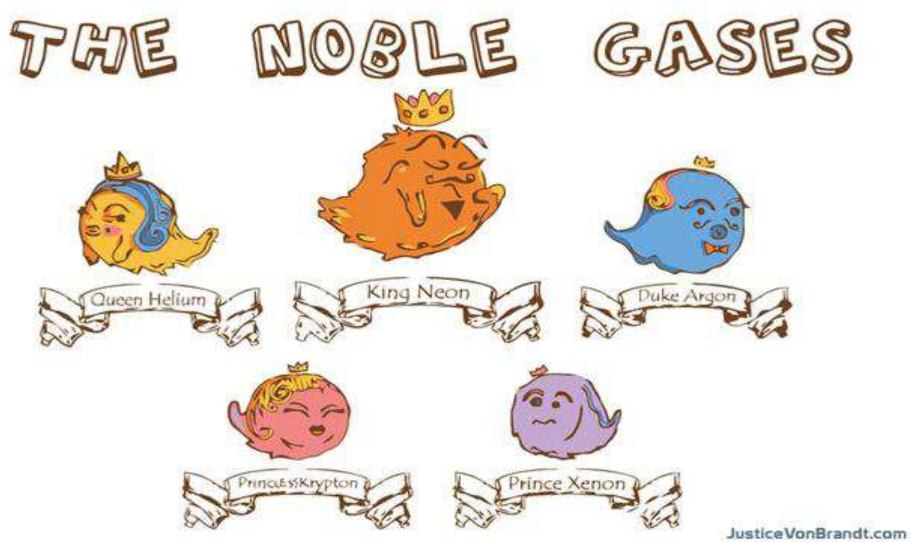
1. Determine the closest noble gas to the element you need the electron configuration for (the noble gas should have a lower atomic number). Example: closest noble gas that comes BEFORE the element zinc (Zn) is Argon (Ar). Zn is atomic number 30, and Ar is atomic number 18.

2. The noble gas symbol is put in brackets: [ ]. Example: [Ar].

3. The remaining electrons are put into electron configuration. Example: Ar has a filled 3p, with a total of 18 electrons. Zn has a total of 30 electrons. 30-18=12. 12 electrons need to be put in subshells. The shorthand electron configuration for Zn would be [Ar] $4s^2 3d^9$.

Noble Gases Analogy:

The noble gases can be thought to live in castles. The only way to get to a castle is for an element (that is not a noble gas) to go back to the nearest castle. This means that the element you want to determine the electron configuration for will have a higher atomic number than the noble gas castle you will use.



SWS 6.01 Short Configurations NAME: ________________________

1. **Write the full electron configuration for vanadium (V).**

2. **What is the nearest noble gas (with a lower atomic number) to V?**

3. **Write the short hand electron configuration for V.**

4. **Write the <u>full and short</u> electron configuration for silicon.**

5. **Write the long and short hand configuration for the following:**

- **Selenium**

- **Tungsten**

- **Strontium**

- **Radium**

6. **Write the short hand electron configuration for the following:**
   - **Germanium**
   - **Xenon**
   - **Gold**
   - **Samarium**
   - **Plutonium**
   - **Iron**
   - **Mercury**

7. **Identify the following elements based on their electron configurations:**
   - **___________[Kr]$5s^2 4d^{10} 5p^2$**
   - **___________ [Ne] $3s^2 3p^4$**
   - **___________ [Kr]$5s^2 4d^6$**
   - **___________ [Ar]$4s^2 3d^8$**
   - **___________ [Xe]$6s^2$**
   - **___________ [Rn]$7s^2 5f^4$**
   - **___________ [Rn]$7s^2 5f^7$**
   - **___________ [Xe]$6s^2 4f^{14} 5d^{10} 6p^1$**
   - **___________ [Ar]$4s^2$**

Unit 06 – Advanced Atomic Theory NAME: ________________________

Worksheet 6.02 – Short Configurations

Fill in the table below. The first two problems have been done for you.

Element	# of Electrons	Long Configuration	Short Configuration
Mg	12	$1s^2\ 2s^2\ 2p^6\ 3s^2$	$[Ne]\ 3s^2$
I	53	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^5$	$[Kr]\ 5s^2\ 4d^{10}\ 5p^5$
Mn			
K			
As			
Sr			
Cs			
W			
Sn			
Br			
Sb			
Al			
Xe			
Po			
Zn			

Lab 6A: Graphing Trends in the Periodic Table

Introduction

Because the periodic table is set up according to electron configuration, the properties of elements change in a predictable way as you move around the periodic table. Today you will learn about four periodic trends: atomic radius, ionization energy, electronegativity, and charge when ion formed and answer the question "Are certain properties of elements periodic functions of the atomic numbers?"

Definitions

Atomic radius: typically defined as half the distance between the nuclei of identical atoms that are bonded together because that is the way it is measured. Basically it is the distance from the center of the atom to the outer edge of the atom, but remember that atoms don't have an exact outer edge. Typically measured in pm.

Ionization energy: the energy required to remove one electron form a neutral atom. Measured in kJ/mol. The lower this energy is the easier it is for that element to become a positive ion.

Electronegativity: the measure of an atom's ability to attract electrons in a chemical bond. Electronegativity is not measured in any units. The most electronegative element, fluorine, is assigned electronegativity of 4.0, and all the other elements are assigned their relative value. The higher the electronegativity, the more likely an element is to become a positive ion.

Charge of ions: the charge ion an element will form can be easily predicted using the periodic table for the s-block and p-block elements. These elements form ions which have electron configurations that will end in p^6 by gaining or losing electrons. The d-block and f-block are not as predictable.

Materials

- 4 sheets of graph paper
- 2 colored pencils
- periodic table of trends
- straight edge

Procedure

1. For elements 1-36, make a graph of **atomic radius** as a function of atomic number. Plot atomic number on the X axis and atomic radius on the Y axis. Make sure each scale is uniform and covers the range of numbers to be plotted. Label the graph. Include a label and unit for each axis.
2. On a separate graph, graph the **atomic radius** information for elements in Group 1, Group 2, Group 17, and Group 18 using a different color. Label the graph.

3. For elements 1-36, make a graph of the energy required to remove the easiest electron, **ionization energy**, as a function of atomic number. Plot **atomic number** on the X axis and **ionization energy** on the Y axis. Label the graph.
4. On a separate graph, graph the **ionization energy** information for elements in Group 1, Group 2, Group 17, and Group 18 using a different color. Label the graph.

5. For elements 1-36, make a graph of the ability of an atom or molecule to attract pairs of electrons in the context of a chemical bond, **electronegativity**, as a function of atomic number. Plot **atomic number** on the X axis and **electronegativity** energy on the Y axis. Label the graph.

6. On a separate graph, graph the **electronegativity** information for elements in Group 1, Group 2, Group 17, and Group 18 using a different color. Label the graph.

Observations

1. What happens to the atomic radius as the atomic number increases across a period?
2. What happens to the atomic radius as the atomic number increases down a group?
3. What happens to the ionization energy as the atomic number increases across a period?
4. What happens to the ionization energy as the atomic number increases down a group?
5. What happens to electronegativity as you move from left to right on the periodic table?
6. What happens to electronegativity as you move from top to bottom of the periodic table?
7. Which groups seem to have a very predictable charge when its elements form ions?
8. Are there exceptions to the periodic trends?
9. Where did most of the exceptions to the periodic trends occur?

Conclusion

1. What properties of the elements are periodic functions of their atomic numbers?
2. Why does atomic radius change as it does?
3. What do you think would happen to the atomic radius of an atom if it became a negative ion?
4. What do you think would happen to the atomic radius of an atom if it became a positive ion?
5. Why does the energy to remove an electron change as it does?

Atomic Number	Element	Atomic Radius (pm)	Ionization Energy (kJ/mol)	Electronega tivity
1	Hydrogen	37	1312	2.1
2	Helium	31	2372	NA
3	Lithium	152	520	1.0
4	Beryllium	112	900	1.5
5	Boron	85	801	2.0
6	Carbon	77	1086	2.5
7	Nitrogen	75	1402	3.0
8	Oxygen	73	1314	3.5
9	Fluorine	72	1681	4.0
10	Neon	71	2081	NA
11	Sodium	186	496	0.9
12	Magnesium	160	738	1.2
13	Aluminum	143	578	1.5
14	Silicon	118	787	1.8
15	Phosphorous	110	1012	2.1
16	Sulfur	103	1000	2.5
17	Chlorine	100	1251	3.0
18	Argon	98	1521	NA
19	Potassium	227	419	0.8
20	Calcium	197	590	1.0
21	Scandium	162	633	1.3
22	Titanium	147	659	1.5
23	Vanadium	134	651	1.6
24	Chromium	128	653	1.6
25	Manganese	127	717	1.5
26	Iron	126	762	1.8
27	Cobalt	125	760	1.8

28	Nickel	124	737	1.8
29	Copper	128	746	1.9
30	Zinc	134	906	1.6
31	Gallium	135	579	1.6
32	Germanium	122	762	1.8
33	Arsenic	120	947	2.0
34	Selenium	119	941	2.4
35	Bromine	114	1140	2.8
36	Krypton	112	1351	NA

Unit 06 – Periodic Trends NAME: ______________________

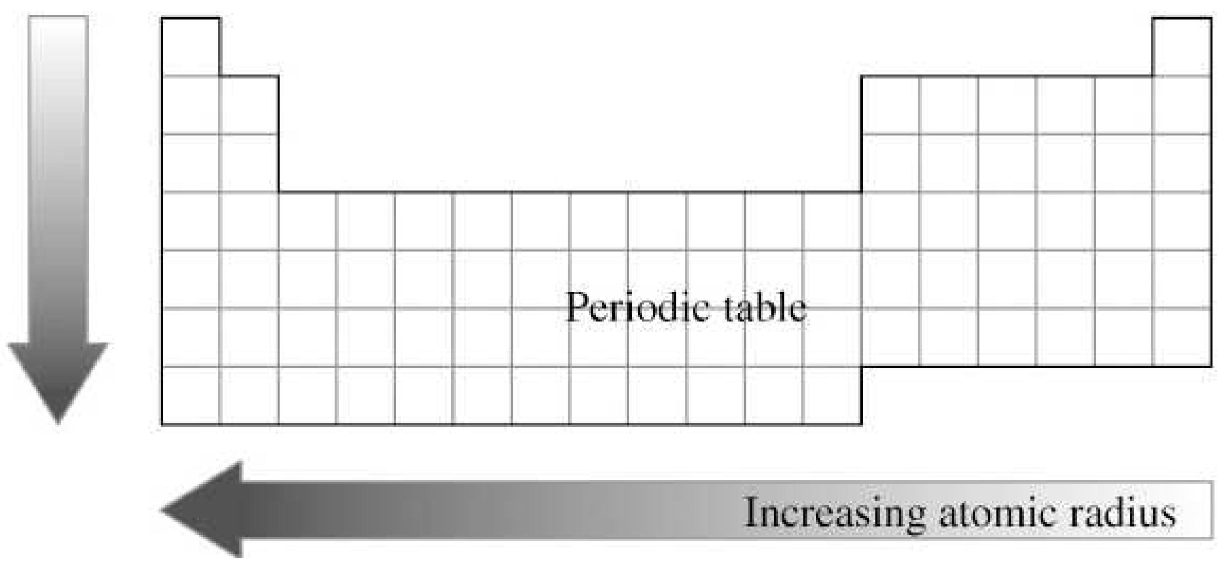
Discussion Sheet 6c – Atomic Radius and Ionic Radius

The relative size of an atom or ion can be determined based on its location in the periodic table. The radius of an atom is half the distance between the nuclei of two like atoms. Therefore, radius is directly proportional to size.

Atomic size generally increases as you move down a group of the periodic table. As you descend, electrons are added to successively higher principal energy levels and the nuclear charge increases. The outermost orbital is larger as you move downward. **Atoms at the bottom of the table are bigger than atoms at the top.**

Atomic size generally decreases as you move from left to right across a period. As you go across a period, the principal energy level remains the same. Each element has one more proton and one more electron than the preceding element. The electrons are added to the same principal energy level. The effect of the increasing nuclear charge on the outermost electrons is to pull them closer to the nucleus. **Atoms on the left of a period are bigger than atoms on the right.**

This trend is more pronounced as you move through a group (up or down) than through a period (right or left), because of the addition of new orbitals as you move down the table. **An up/down move in the periodic table is a much more important change in atomic size than a left/right move.**



The trend described above is also true for positive and negative ions. However, there is one fundamental difference caused by the formation of the ions. Positive ions are formed when an atom loses electrons. Therefore, the relative nuclear charge of the ion is more than that of a neutral atom. **Positive ions are always smaller than neutral atoms.** In a similar fashion, negative ions are formed when an atom gains electrons. The relative hold of the nucleus is less in a negative ion than it is for a neutral atom. **Therefore, negative ions are always bigger than neutral atoms.**

<u>Unit 06 – Periodic Trends</u> NAME: ________________________

Discussion Sheet 6d – Ionization Energy, Electronegativity, and Electron Affinity

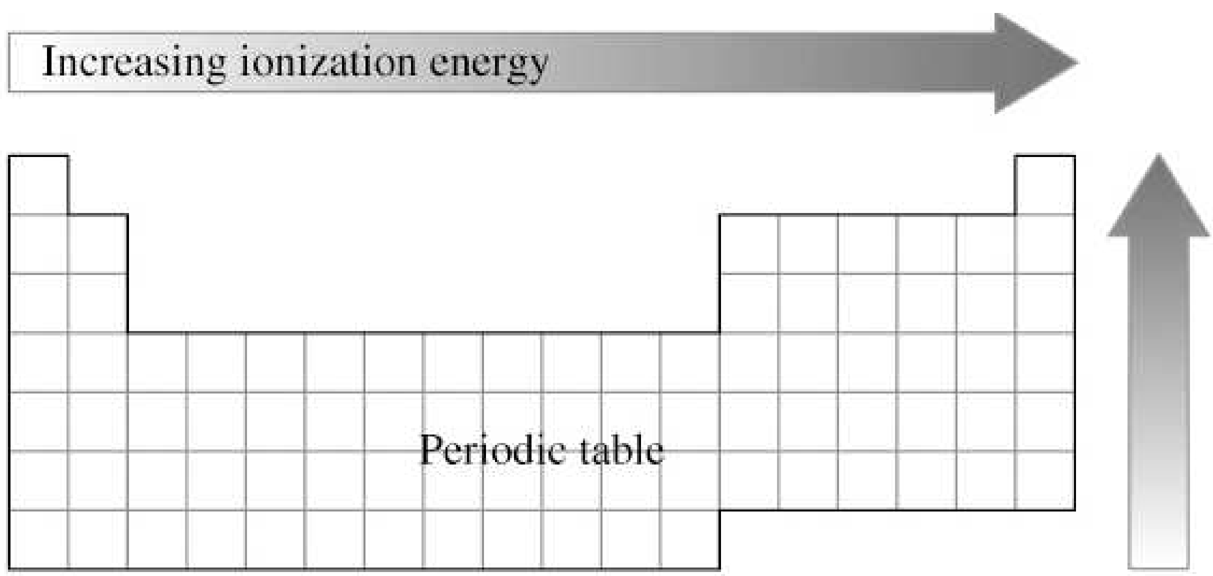
The degree to which an atom holds onto its outermost electrons can also be determined based on its location in the periodic table. This hold a nucleus has on its outermost electrons can be expressed in three ways:

1. The *electronegativity* of an atom is the tendency for atoms of an element to attract electrons <u>when they are chemically bonded to atoms of another element</u>.

2. The *ionization energy* is the amount of energy required to <u>remove an electron from an atom to make a positive ion</u>.

3. The *electron affinity* is the tendency for an atom to <u>take hold of an electron in order to form a negative ion</u>.

The differences between these three things is nuanced, depending on what the atom is, and what type of chemical bond it will form (or has formed). As should be expected, this trend is exactly the opposite of the radius trend, as summarized by the following table.



As with atomic radius, **an up/down move in the periodic table is a much more important change in ionization energy, electronegativity, or electron affinity than a left/right move.**

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SWS 6.02 – Atomic Radius and Ionic Radius/ Ionization energy/ Electronegativity

1. **Circle the element that has the greatest atomic size:**

- Strontium or chlorine
- Sc or Si
- Cs or Br
- P or Cl
- P or F
- P or N
- Rb or Mg
- K or Cl

2. Circle the element or ion that has the greatest ionic size:

- O^{2-} or S^{2-}
- N^{3-} or P^{3-}
- Na^{+} or Na
- O or S
- F or F^{-}
- P or P^{3-}
- Ba^{2+} or Ba
- Ca^{2+} or Be^{2+}

3. Circle the element with the greatest electronegativity:

- Na or Mg
- K or Mg
- Mg or Sr
- Ge or B
- In or I
- Tl or Bi
- Cl or Se

4. Circle the element with the greatest ionization energy:

- O or S
- N or P
- O or F
- Ba or Fr

5. Circle the element with the greatest electron affinity:

- F or Cl
- Cl or I
- I or Br
- Se or I
- Te or Cl

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Worksheet 6.03 – Atomic Radius and Ionic Radius/ Ionization energy

For each of the following pairs, **circle the atom or ion that is larger.** Use the periodic table to assist you.

Magnesium atom or Sodium atom	Cs^{1+} ion or Ca^{2+} ion
Mn^{2+} ion or Fe^{2+} ion	Strontium atom or Lithium atom
Nitrogen atom or Phosphorus atom	O^{2-} ion or P^{3-} ion
Iodine atom or Chlorine atom	Silicon atom or Fluorine atom
Positive Sr^{2+} ion or Neutral Sr atom	Selenium atom or Gold atom

For each of the following pairs, **circle the element that has the larger ionization energy**. Use the periodic table to assist you.

Li or N	Cl or Se
Li or K	Cl or B
Li or Sc	Br or Pd
Mg or Rb	F or Fe
Mg or C	F or Na

For each of the following pairs, **circle the element that is more electronegative**. Use the periodic table to assist you.

P or As	Ga or Al
P or S	Ga or Cd
P or Si	Au or Mo

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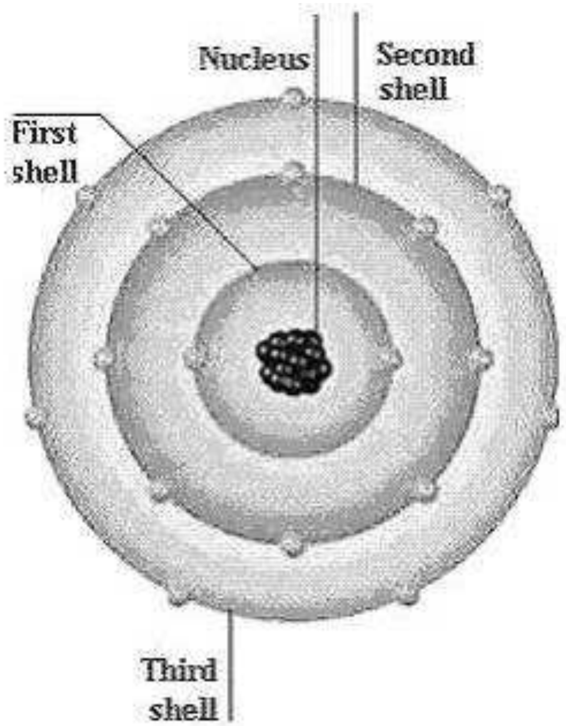
Discussion Sheet 6e – Other Trends

A number of periodic trends exist among the elements and these trends can be explained by looking at variations in atomic structure. Remember, trends occur within groups and within periods. The following are several other trends that also must be considered.



Valence electrons are an atom's outermost electrons. *Shielding* is the collective action of all the non-valence electrons to weaken the attraction between the protons and the valence electrons. **Shielding increases as you move down the periodic table. Shielding does not change as you move from left to right.** Shielding increases as you move down because a downward move is associated with an increased number of shells between the nucleus and the valence electrons. A left/right move has no such increase in shells.

Nuclear charge is the amount of positive charge exerted by the nucleus. Nuclear charge increases with the addition of more protons. **Therefore, nuclear charge increases as atomic number increases.**

Metallic character is the degree to which an atom behaves as a metal. Metals lose electrons to form positive ions, and are malleable and ductile when in their elemental form. **Metallic character increases as you move down and to the left, just as atomic size does.**

Nonmetallic character is the degree to which an atom behaves as a nonmetal. Nonmetals gain electrons to form negative ions, and are brittle when in their elemental form. **Nonmetallic character increases as you move right and up, just as ionization energy, electronegativity, and electron affinity all do.**

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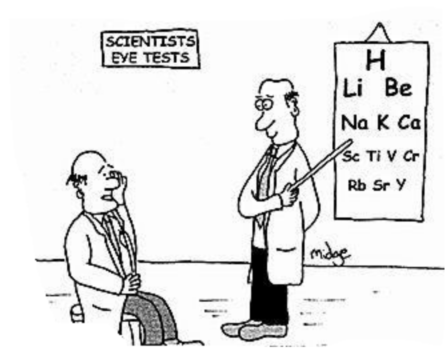
Review Sheet

Complete the following questions as a review for your test.

1. **What aspects of the nature of science and scientific inquiry did you use during the Periodic Cards Activity?**



2. **In terms of the nature of science, how was the periodic table arranged? What changes has it undergone? Why?**

3. **How do you write electron configuration shorthand?**

4. **In general the electronegativity ____________________ as you move from left to right across the periodic table and ______________________ as you move from top to bottom.**

5. **The electronegativity of an atom is defined as**

6. **What is an atomic radius?**

7. **In general the atomic radius of an element ____________________ as you move from left to right across the periodic table and ______________________ as you move from top to bottom.**

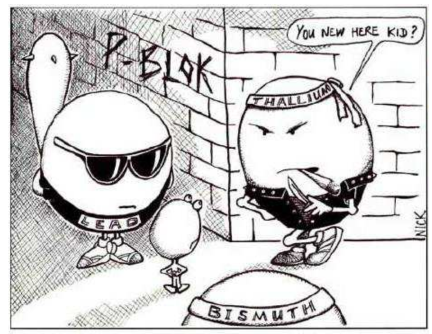
8. **In general the ionization energy ____________________ as you move from left to right across the periodic table and ______________________ as you move from top to bottom.**

9. Ionization energy is defined as the energy required to…?

10. Elements in the same group have the same number of ________________ _________________ and similar properties.

11. Identify the highest energy level, number of valence electrons, and write the shorthand electron configuration for the following elements.

a. Sn



Unwittingly, and against his mother's advice, Vince the first-row Transition Metal had been lured far away from home, and now found himself surrounded by heavier elements of the P-Block.

b. O

c. W

d. Hg

e. Ar

f. Ra

g. Br

12. Determine the period, block, and group of the elements having the following electron configuration.

a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^9$

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

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

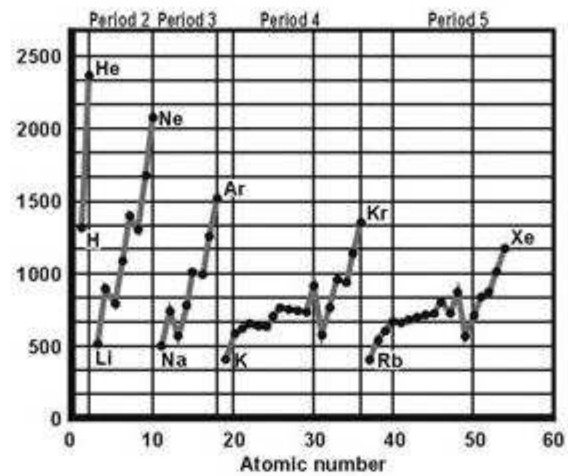
d. $1s^2 2s^2 2p^6 3s^2 3p^3$

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

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

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

13. Explain how you can use this diagram to explain the periodic trend of ionization energy.



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14. Identify all the following on the periodic table below.

a. Metals
b. Non-metals
c. Stair Step
d. Metalloids
e. Alkali Metals
f. Alkaline Earth Metals
g. Lanthanide series
h. Actinide series
i. Rare Earth Metals
j. Halogens
k. Noble Gases
l. s, p, d, f blocks

<u>Use arrows to label and describe the following:</u>

m. how the atomic radius changes across a period and group
n. how the electronegativity changes across a group and period
o. how the ionization energy changes across a period and group

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une	110 Unn								
			58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
			90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	