

Unit 06:

Periodic Trends

Driving Questions

- ❖ How is the periodic table organized according to trends
- ❖ How has the periodic table changed over time?
- ❖ Why are trends useful and what are periodic trends?

Connections to Past/Future Units

- ❖ Understand how science changes over time, with new technology and new ideas
- ❖ Practice of graphing skills and data analysis
- ❖ Conceptualize the atom using past and new information
- ❖ Understand the layout and use of the periodic table

Objectives: *SWBAT*...

- ☐ explain how the periodic table is based on patterns
- ☐ use graphing skills to discover periodic table trends
- ☐ Experimentally determine how density changes within periods and groups.
- ☐ write the electron configurations in short hand using the Noble Gas configuration
- ☐ apply ionization energy, atomic/ionic size, electronegativity, electron affinity trends
- ☐ create and analyze graphs of elemental properties

Essential Vocabulary

Trend
Period
Group
Periodic law
Atomic radius
Ion
Ionization energy
Electronegativity

Electron affinity
Noble gas
Representative elements
Metals
Non-metals
Metalloids
Alkali metals
Alkaline earth metals

Lanthanide series
Actinide series
Rare earth metals
Halogens
S block
P block
D block
F block

Name: _____ Period: _____

Packet Grade: _____/132pts

❖ Objective; reflection /10pts

❖ Reading Charts: 30pts

❖ Worksheets: 10pts per page

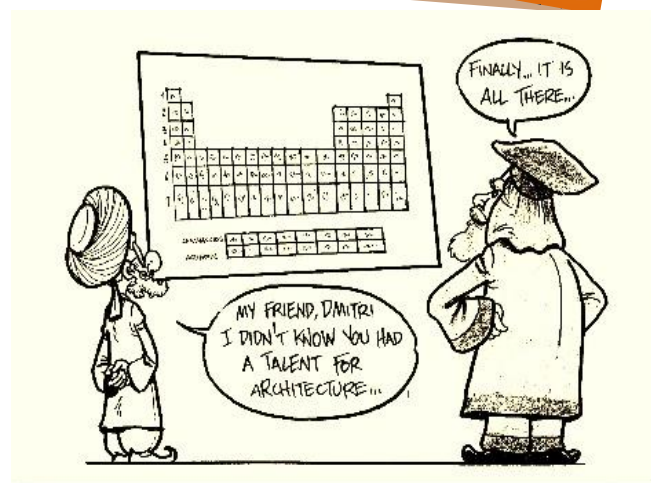
❖ Stamps: 2pts/stamp

Activity 6A: _____/100pts

Activity 6B: _____/35pts

Lab 6A Grade: _____/100pts

Lab 6B Grade: _____/100pts



Personal Objective: Looking at the objectives above, what more do you want to learn this unit?

Reflection: Did you accomplish your personal objective? What further questions do you have about this unit?

Activity 6A:

One of the most important tools for a successful life is being able to predict outcomes based on limited data. Being able to look at data, rearrange it and form a useful opinion is something that we do from childhood through to adulthood. This is generally learned at a young age; through structured discipline, small children are able to predict the consequences of their actions. A good financial analyst is someone who can “predict” market trends, and capitalize on the up and down movements. You have already done some predictions in your math class, if you have graphed a simple equation; depending on the slope of the line you can say that the trend is increasing, decreasing, or constant. Knowing the trend allow you to generalize what will happen for data that is not represented.

Scientists (and the ACT) also have conventions on what good graphs look like. Whenever you graph data in this class, you should make sure to include the following things.

Here is a data set. Organize the data into a table. Graph the data. Then, answer the questions.

Correctly Labeled Axis

-
-
-

Title

-
-

A “KEY”

- .

PRACTICE:

1. If we graph how the number of hours spent studying affects your grade on next class period’s quiz, what would we title our graph?
 - a. What would we measure on our X-axis? **Hours Spent Studying (Time)** or **Grade on Quiz**
 - b. What would we measure on our Y-axis? **Hours Spent Studying (Time)** or **Grade on Quiz**
 - c. So, what would be an appropriate title for our graph?

2. If we graphed how temperature affects the pressure of air at a defined volume, how would we graph that?
 - a. What would we measure on our X-axis? **Temperature** or **Pressure?**
 - b. What would we measure on our Y-axis? **Temperature** or **Pressure?**
 - c. So, what would be an appropriate title for our graph?

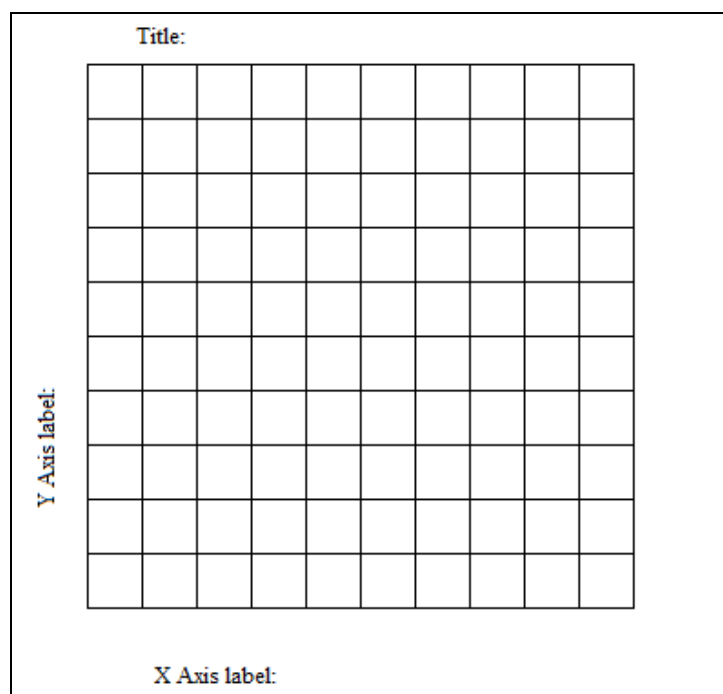
3. If we graphed how amount of sunlight affects plant growth, how would we graph that?
 - a. What would we measure on our X-axis? **Hours of sunlight** or **Plant growth**
 - b. What would we measure on our Y-axis? **Hours of sunlight** or **Plant growth**
 - c. So, what would be an appropriate title for our graph?

Data Set 1: Gasoline prices

You may not be currently concerned about gasoline prices. However, you should be when you start to drive, and even if you are not driving, increased fuel prices will drive up the cost of cab rides, bus, and train fares. Here are some actual prices of gasoline. They represent the average pump price for one gallon of regular gasoline in the Midwest for the given date.

Sept. 02, 1996: \$1.164	Sept. 03, 2001: \$1.707	Sept. 04, 2006: \$2.572	Sept. 05, 2011: \$3.697
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First, put the data into the data table below. Then, graph the data points:



Based on the data table, what is your observation of the historical price of gasoline? _____

Using the graph, how can you predict the price of gasoline on September 04, 2003? _____

What do you predict the price of a gallon of gas will be on September 04, 2003 _____.

You just practiced an important ACT skill called interpolation!

Using the graph, how can you predict the price of gasoline on September 06, 2016? _____

What do you predict the price of a gallon of gas will be on September 06, 2016 _____.

You just practiced an important ACT skill called extrapolation!

Data Set 2: Sand and Water warming

	Starting Temp. (Celsius)	1 min	2 min	3 min	4 min	5 min	6 min	7 min	8 min	9 min	10 min/ Final Temp.
Sand	18	19	21	22	24	25	26	28	30	31	32
Water	20	20	20	21	22	23	23	24	25	26	27



Which substance absorbs heat more quickly?

How do you know? (Use evidence!)

There was a time when the periodic table did not exist. People did experiments with substances they called elements, and they catalogued the elements' properties. However, no one had yet arranged them in any significant way. Also, not all elements had been discovered which would cause "holes" in any arrangement system.

To see how this may have caused problems arrange the following eight symbols in the 3x3 grid provided. You must write the symbol and mass in each square. You should put the symbols in their appropriate boxes by considering both their mass and color. Once you have written in the symbols you will notice that there is an empty box. Based on your arrangement you should determine where the missing symbol belongs.

Symbol	Mass	Color
Se	1	Blue
Te	23	Blue
N	28	Green
Et	4	Red
T	15	Red
He	7	Green
R	26	Red
Pa	12	Blue

Based on your arrangement, where is the missing box? _____

What is the color and approximate mass of the missing symbol? _____

READING ASSIGNMENT: Read and take notes on pages _____ in your textbook.

SEE

THINK

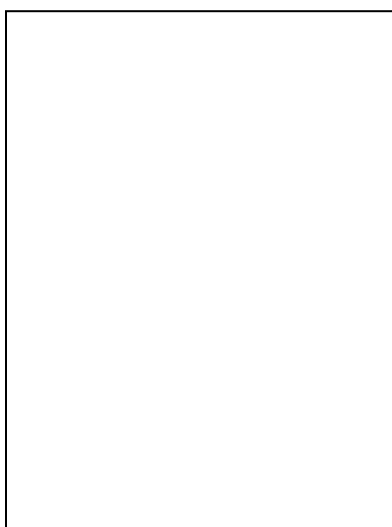
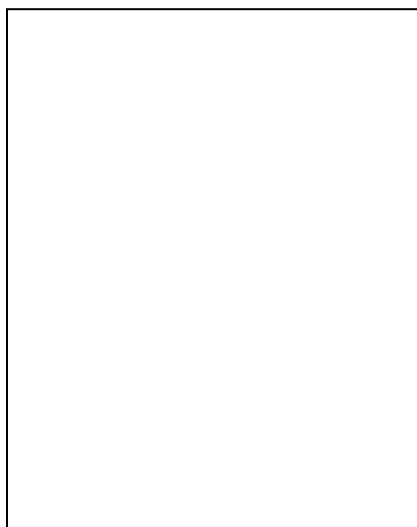
Unit 06 – Periodic Trends

NAME: _____

Activity 6B: Periodic Table Cards**Instructions**

Spread out all the cards and look at their properties. If you arrange them correctly you be able to determine all the characteristics of your missing card.

1. What are the six properties that you can determine about the cards?
 - a.
 - b.
 - c.
 - d.
 - e.
 - f.
2. What properties do all the cards have?
3. Arrange the cards in the way that makes the most sense to you using all the properties. When all the cards are arranged correctly, you should be able to determine the properties of your two missing cards.
4. How did you arrange them?
5. Draw your missing cards below:



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NAME: _____

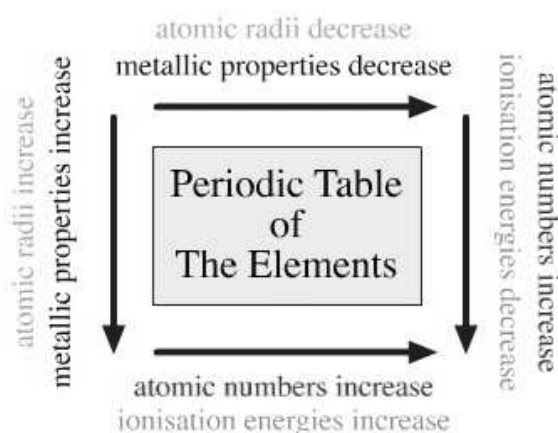
Discussion Sheet 6a – Periodic Trends

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.



READING ASSIGNMENT: Read and take notes on pages _____ in your textbook.

SEE**THINK**

Unit 06 – Periodic Trends

NAME: _____

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 $[\text{Ar}] 4s^2 3d^{10}$.

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.

THE NOBLE GASES



Unit 05 – Advanced Atomic Theory

NAME: _____

WS 6.01 Short Configurations

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 full and short 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:

- _____ $[\text{Kr}]5s^24d^{10}5p^2$
- _____ $[\text{Ne}]3s^23p^4$
- _____ $[\text{Kr}]5s^24d^6$
- _____ $[\text{Ar}]4s^23d^8$
- _____ $[\text{Xe}]6s^2$
- _____ $[\text{Rn}]7s^25f^4$
- _____ $[\text{Rn}]7s^25f^7$
- _____ $[\text{Xe}]6s^24f^{14}5d^{10}6p^1$
- _____ $[\text{Ar}]4s^2$

Unit 06 – Advanced Atomic Theory
Worksheet 6.02 – Short Configurations

NAME: _____

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

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

NAME: _____

Purpose

To investigate how density changes as you move across a period and down a group.

SWBAT experimentally determine how density changes within periods and groups.

Materials

Carbon	Lead	Tin	Copper
Aluminum	Colored pencils	Iron	Balance
Graph paper	Graduated cylinder	Ruler	Calculator

Procedure

1. Determine the mass and volume of each sample. Record in your data table.
2. Calculate the density of each sample. (density=mass/volume)
3. Order the sample from lowest to highest period.
4. Graph density vs. element period number for your samples.

Data

Element	Mass (g)		Volume (mL)		Calculated Density
Carbon		/		=	
Lead		/		=	
Tin		/		=	
Copper		/		=	
Aluminum		/		=	
Iron		/		=	

[illegible]

Key

- ☐ Carbon
- ☐ Lead
- ☐ Tin
- ☐ Copper
- ☐ Aluminum
- ☐ Iron

Analysis and Conclusions

1. Which elements are in the same group?
2. How does density change as you move down this group in the periodic table?
3. Which elements are in the same period?
4. How does density change as you move from left to right across the period?
5. Use your graph to predict the density of germanium and silicon.
6. Calculate the percent error of your densities you determined as compared with accepted values.
 - a. Carbon 2.266 g/cm^3
 - b. Lead 11.342 g/cm^3
 - c. tin 7.265 g/cm^3
 - d. copper 8.961 g/cm^3
 - e. aluminum 2.701 g/cm^3
 - f. iron 7.871 g/cm^3
7. What are some possible sources of error for your experiment?

Lab 6B: Graphing Trends in the Periodic Table

SWBAT create and analyze graphs of elemental properties

**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 from 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 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** 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)	Electronegativity
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

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Discussion Sheet 6c – Atomic Radius and Ionic Radius

SWBAT compare the size of atoms of different elements based on their position in the periodic table.

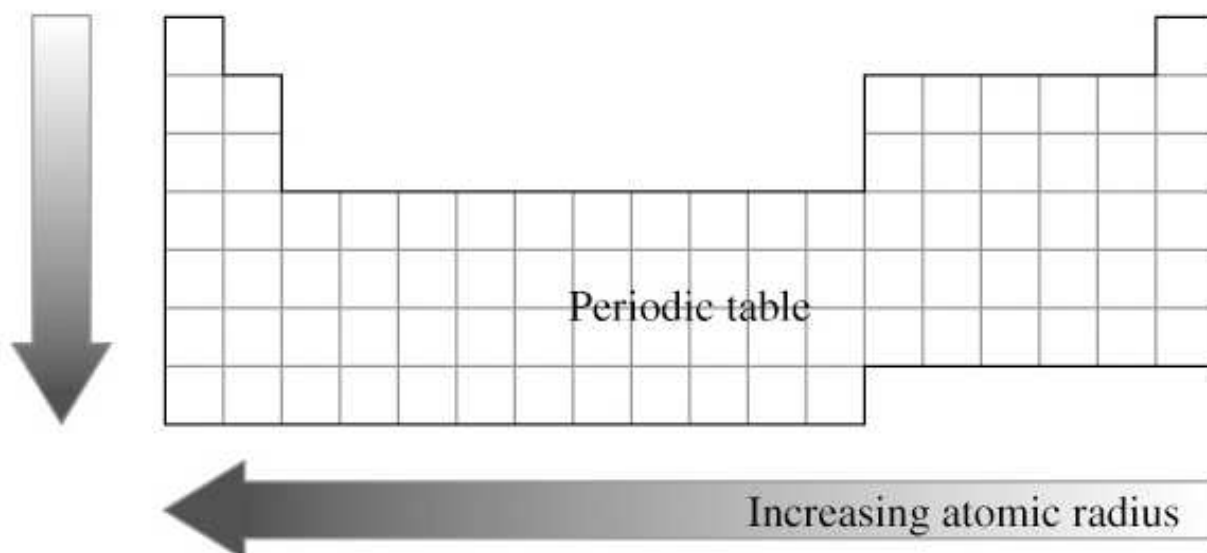
The relative size of an atom or ion can be determined based on its location in the periodic table. The radius of an atom is half the distance between the nuclei of two like atoms. Therefore, radius is directly proportional to size.

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.**

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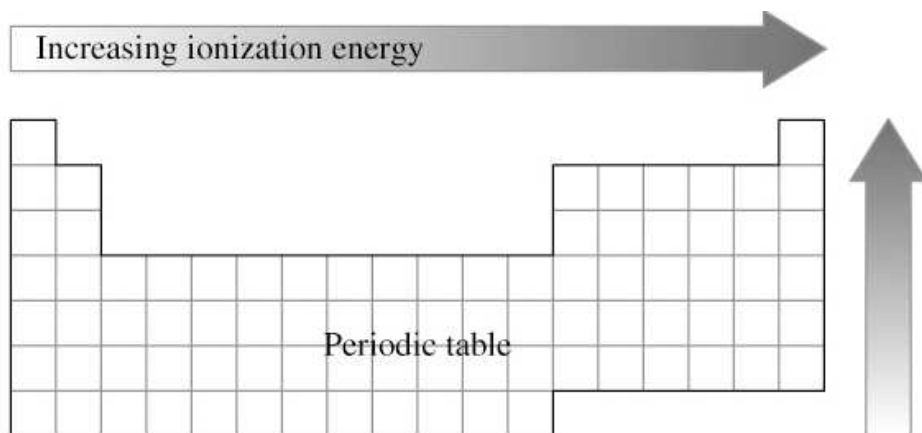
Discussion Sheet 6d – Ionization Energy, Electronegativity, and Electron Affinity

SWBAT compare the ionization energy and electronegativity of elements based on their position in the periodic table.

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 when they are chemically bonded to atoms of another element.
2. The *ionization energy* is the amount of energy required to remove an electron from an atom to make a positive ion.
3. The *electron affinity* is the tendency for an atom to take hold of an electron in order to form a negative ion.

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.**

READING ASSIGNMENT: Read and take notes on pages _____ in your textbook.

SEE

THINK

Unit 06 – Periodic Trends

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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

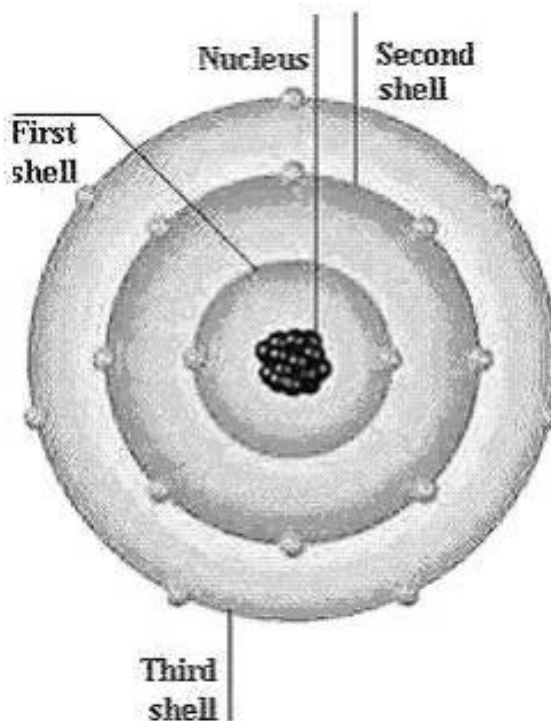
A number of periodic trends exist among the elements and these trends can be explained by looking at variations in atomic structure. Remember, trends occur within groups and within periods. The following are several other trends that also must be considered.

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.**



Unit 6—Periodic Trends

Review Sheet

Name: _____

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

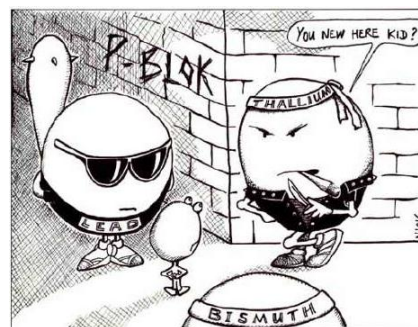
c. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

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

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

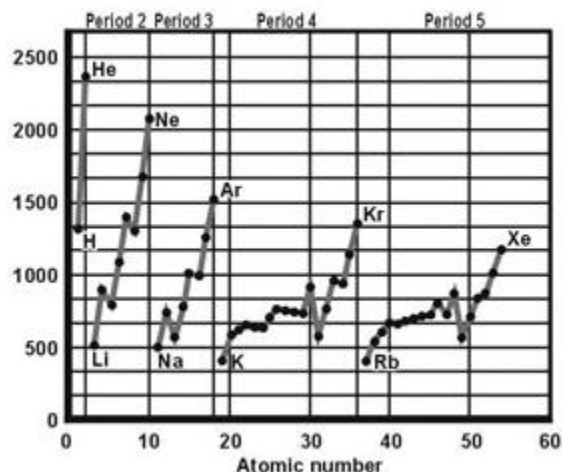
f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^5$

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



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.

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



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

- Metals
- Non-metals
- Stair Step
- Metalloids
- Alkali Metals
- Alkaline Earth Metals
- Lanthanide series
- Actinide series
- Rare Earth Metals
- Halogens
- Noble Gases
- s, p, d, f blocks

Use arrows to label and describe the following:

- 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	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	
87 Fr	88 Ra	89 Ac	104 Unq	105 Unp	106 Unh	107 Uns	108 Uno	109 Une	110 Uun								