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.

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 negative 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.

Graph the first 20 elements (atomic number 1-20) atomic radius relative to their atomic number.

Table of data

Element	Atomic Radius (pm)	Ionization Energy (kJ/mol)	Electronegativity	Charge when ion formed
Hydrogen	37	1312	2.1	+ or -
Helium	31	2372	NA	No ion formed
Lithium	152	520	1.0	+
Beryllium	112	900	1.5	2+
Boron	85	801	2.0	3+
Carbon	77	1086	2.5	4-
Nitrogen	75	1402	3.0	3-
Oxygen	73	1314	3.5	2-
Fluorine	72	1681	4.0	-
Neon	71	2081	NA	No ion formed
Sodium	186	496	0.9	+
Magnesium	160	738	1.2	2+
Aluminum	143	578	1.5	3+
Silicon	118	787	1.8	4+
Phosphorous	110	1012	2.1	3-
Sulfur	103	1000	2.5	2-
Chlorine	100	1251	3.0	-
Argon	98	1521	NA	No ion formed
Potassium	227	419	0.8	+
Calcium	197	590	1.0	2+
Scandium	162	633	1.3	3+
Titanium	147	659	1.5	4+
Vanadium	134	651	1.6	3+ or 5+
Chromium	128	653	1.6	6+ or 3+ or 2+
Manganese	127	717	1.5	7, 6, 4, 3, or 2+
Iron	126	762	1.8	2+ or 3+
Cobalt	125	760	1.8	2+ or 3+
Nickel	124	737	1.8	2+ or 3+
Copper	128	746	1.9	+ or 2+
Zinc	134	906	1.6	2+
Gallium	135	579	1.6	3+
Germanium	122	762	1.8	4+
Arsenic	120	947	2.0	3-
Selenium	119	941	2.4	2-
Bromine	114	1140	2.8	-
Krypton	112	1351	NA	No ion formed

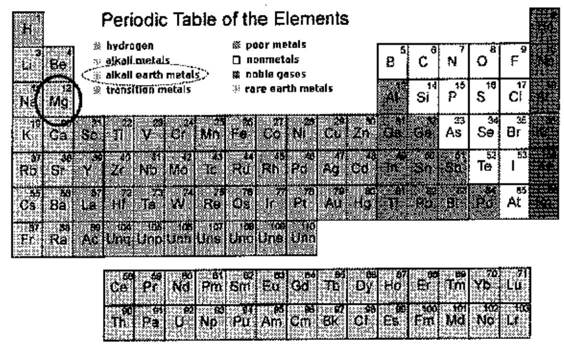
Name: ____________________ Unit 3: Periodic Table

POGIL: Periodic Table Trends

(Adapted from Rush Henrietta CSD)

Why:

The Periodic Table is one of the greatest inventions in the history of man. It allows scientists to predict physical and chemical properties of the elements. Dimitri Mendeleev (a Russian scientist) and Robert Mosley (a British chemist) put together this table in the late 1800's based on properties of the elements known at the time. The trends in these properties as you go across periods and down groups is the subject of this Chem POGIL



Success Criteria:

- Understand the meaning of
  atomic radius, reactivity, electronegativity, ionization energy
- Recognize trends in atomic radius, reactivity, ionization energy and electronegativity as you go across periods and down groups.

Resources:

Periodic Table, Unit 3 Notes, & Chapter 5 in *Modern Chemistry* textbook

Prerequisites:

Answer the following questions using the resources above.

1. What happens to the number of valence electrons as you go down a group on the periodic table?

2. What happens to the number of valence electrons as you go from left to right across a period?

3. What happens to the number of energy levels as you go down a group on the periodic table?

4. What happens to the number of energy levels as you go from left to right across a period?

5. Draw a Lewis structure for a Cl *atom.*

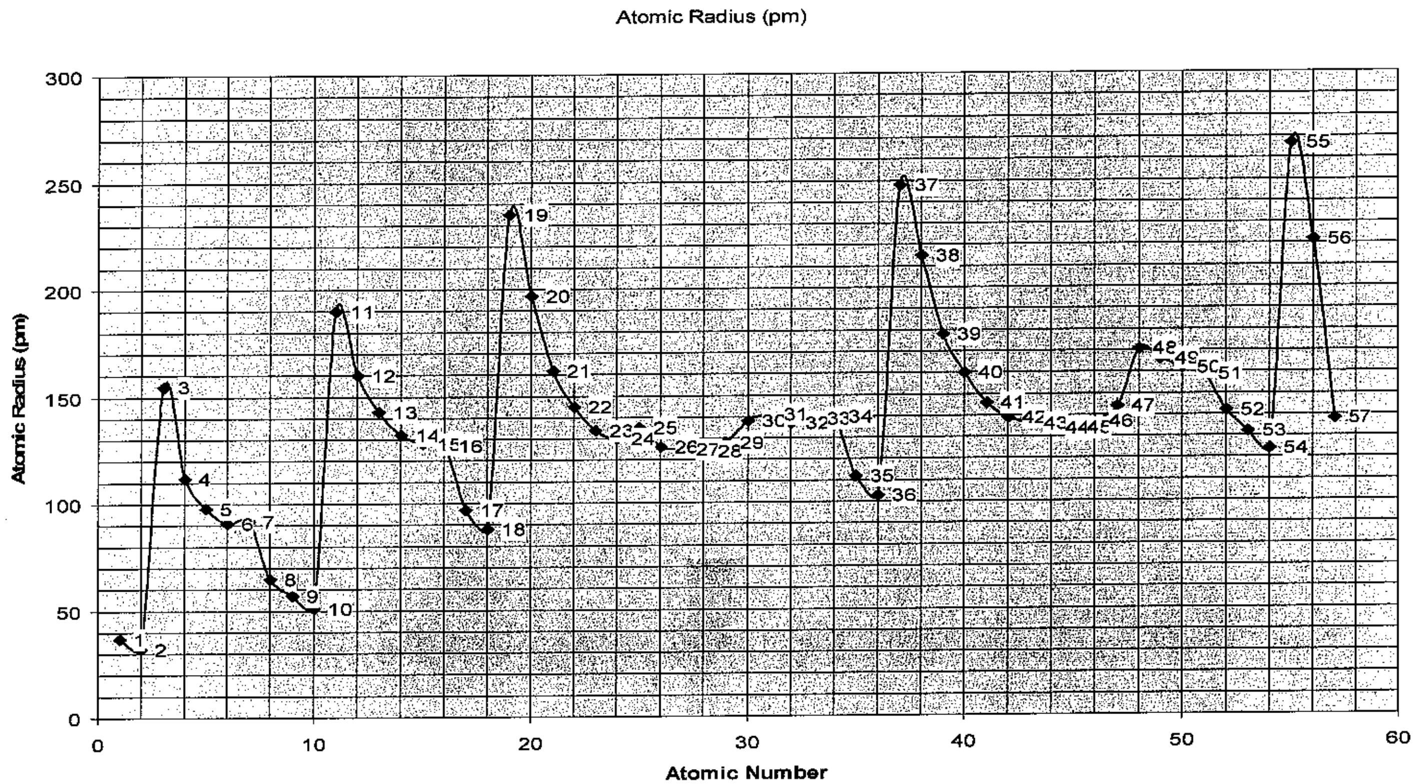
6. Draw a Bohr diagram for the Cl *atom.* ***(Draw above)***

7. In this lab we will be talking about atomic radius.
   a. Look up the atomic radius of Cl on pg. 141: ___________
   b. Why is it possible to show the atomic radius using a Bohr diagram but not by using a Lewis structure?
   c. On your Bohr diagram draw an arrow to represent the radius of the Cl atom.

8. Define *periodic:*

9. Define *trend:*

Model #1: Trends in Atomic Radius



Atomic radius is defined as the distance between the center of the nucleus of an atom and the outermost shell of electrons.

1. What is the unit used for atomic radius? ___________

2. How many of these are in a meter? (hint: pico is 10^{-12}) _________

3. Look at the graph above. Compare it to the periodic table. What do you notice about the location on the periodic table of the elements that represent the high points?

4. What do you notice about the location on the periodic table of the elements that represent the low points?

5. Look up the elements on the graph that make up period 3 of the periodic table. What do you notice about atomic radius as you move from left to right across a period?

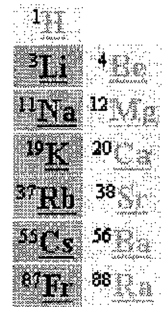
So… state the trend…

ATOMIC RADIUS ______________________________ as you go down a group

and ________________________ as you go from left to right across a period.

<u>Going Further:</u> Explain, in terms of atomic structure and forces, *why* atomic radius decreases from left to right across a period.

Model #2: Trends in Reactivity



When <u>METALS</u> react they _______________ (gain/lose) electrons. As you go down any group containing metals, it becomes easier to lose electrons because they are further away from the attractive force of the positive nucleus.

1. Look at group 1 on the Periodic Table, the Alkali Metals.
   a. What is similar about their atomic structure?

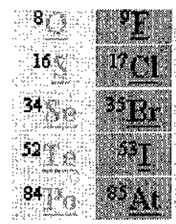
   b. Why does Hydrogen fit into this group?

   c. Why does Hydrogen NOT fit into this group?
   (Hint: Why does it make sense that European Periodic Tables show H in both Group 1 and 17?)

   d. Which metal is the most reactive of the group (that is, loses electrons most easily)? Why?

   e. State the relationship between reactivity and size for metals.

   f. Group 2, the Alkaline Earth Metals, follow similar trends. So which element is more reactive, barium or magnesium? Explain.



When <u>NONMETALS</u> react they ______________ (gain/lose) electrons. As you go down a group the tendency to gain an electron decreases because the attractive force of the nucleus is more "shielded" by the many other layers of electrons in between.

2. Look at group 17 on the Periodic Table, the Halogens…
   a. What is similar about their atomic structure?

   b. Which nonmetal is the most reactive of the group (ie. gains electrons most easily)? Why?

   c. State the relationship between reactivity and size for nonmetals.

3. Group 18, the **<u>NOBLE GASES,</u>** are not reactive at all. Why?

So… state the trend…

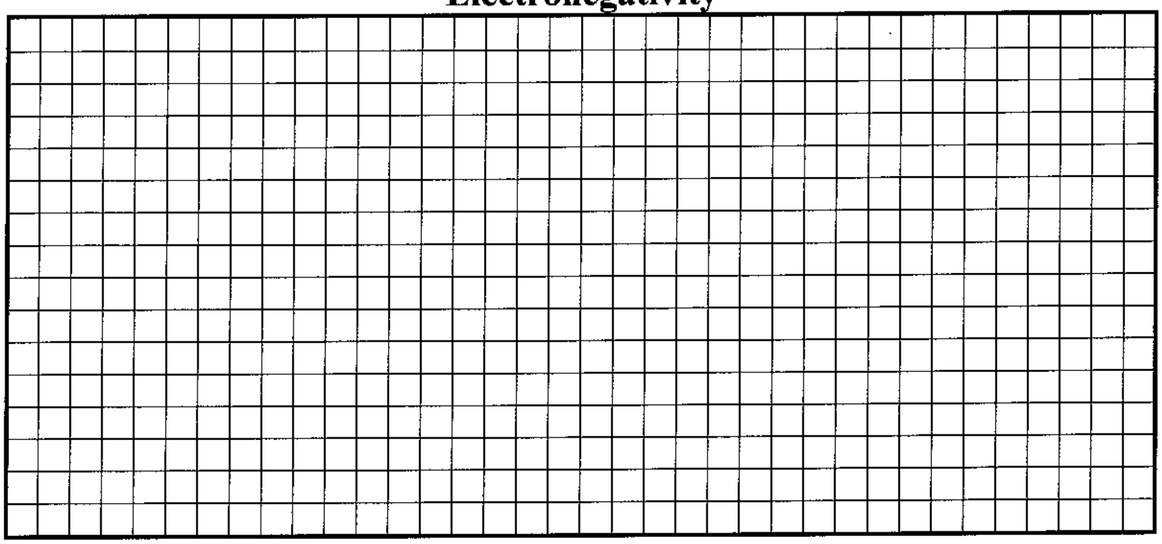
For METALS, REACTIVITY _____________________as you go down a group.
For NONMETALS, REACTIVITY _________________ as you go down a group.

Model #3: Trends in Electronegativity

> ***Electronegativity is the tendency of an atom to attract electrons when it is bonded to another atom. Often, although not always, the difference in electronegativity between two bonding atoms is an indication of whether the compound is ionic, covalent or "in between". (More on this concept, called polarity, in the next Unit.)***

For this graph, plot the atomic number on the x-axis and the "electronegativity" value on the y-axis, for elements **1 - 18**. Use pg. 151 to find these values.

Electronegativity



1. Why would an atom want to gain electrons?

2. When looking at the elements in the **same period**, from which group does the element come from that has the highest electronegativity value in every case? _________.... The Lowest? _________

3. Which has higher electronegativity, metals or nonmetals? Why?

4. What is the electronegativity trend from top to bottom in the same group?

5. What is the electronegativity trend from left to right across a period?

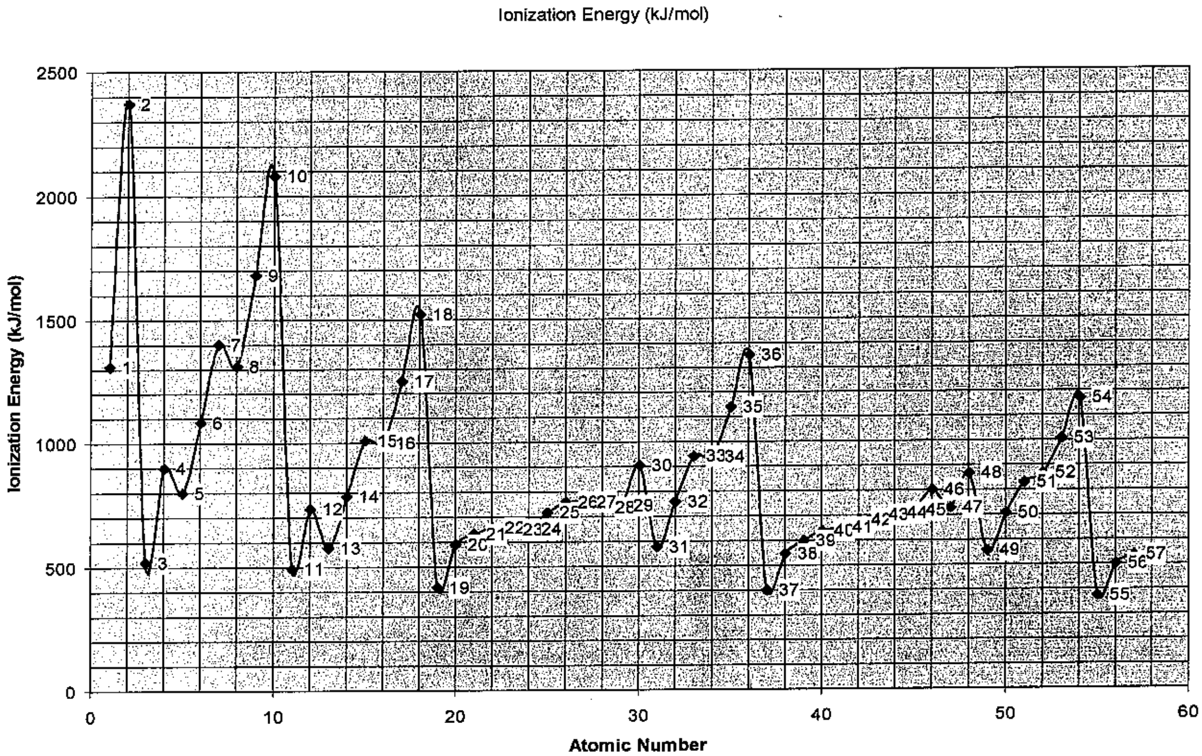
6. Based on the definition of electronegativity, why don't the noble gases have any electronegativity values?

7. Which element has the highest electronegativity of all? Why?

ELECTRONEGATIVITY ___________________________ as you go down a group

and ____________________ as you go from left to right across a period.

Model #4: Trends in Ionization Energy



Ionization energy is the energy required to remove an electron from any atom.

1. Why is it easier to remove an electron from Na than it is from Cl?

2. Why is ionization energy related to, but not the same as, electronegativity?

3. Describe the trend in ionization energy as you go left to right across a period.

4. Describe the trend in ionization energy as you go down a group. What are the numbers of the elements you are comparing?

5. Are these trends similar to other trends. Explain

6. Why do noble gases have the highest ionization energy values?

7. Why do metals have low ionization energy values?

8. Why do nonmetals have high ionization energy values?

So… state the trend…

IONIZATION ENERGY ____________________________ as you go down a group and ________________________ as you go from left to right across a period.

Name ____________________

PERIODIC TABLE WRAP UP

Think about these questions and answer them thoughtfully ON YOUR OWN.

1. What does chemical reactivity mean for metals?

2. Which is the more reactive alkaline earth metal, magnesium or calcium? Explain why.

3. Why does atomic radius increase as you go down a group?

4. Why does ionization decrease as you go down a group?

5. Why is the electronegativity for non-metals much greater than the electronegativity for metals?

Name____________________ **Unit 3: Periodic Table**

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