SCPS Chemistry Worksheet – Periodicity

A. Periodic table

1. Which are metals? Circle your answers: C, Na, F, Cs, Ba, Ni

Which metal in the list above has the most metallic character? Explain.
Cesium – as the largest atom, the lowest ionization energy and the most reactivity with nonmetals. This can be determined by its position lowest in the alkali metal group.

2. Write the charge that each of the following atoms will have when it has a complete set of valence electrons forming an ion.
O – 2- Na 1+ F 1- N 3- Ca 2+ Ar - none

3. What is the most common oxidation number for calcium? Explain.
The last page of the powerpoint on *Chemistry, Atoms and Ions* provides the definition of oxidation number as the charge on the ion. You were also asked to define this term in your element brochure. Pages 222 and 980 of your text book defines oxidation number as the positive or negative charge of a monatomic (one atom) ion. These numbers may be confirmed on many of the websites you have used

Calcium will have a 2+ oxidation number as it tends to easily lose its 2 valence electrons, the $4s^2$ electrons

4. Name two more elements with that oxidation number and explain your choice.
Other alkaline earth metals (group 2) will also have 2+ oxidation numbers since they all have 2 s valence electrons that they easily lose. This would include Be, Mg, Sr, Ba and Ra.

5. What element in period 3 is a metalloid?
Silicon – Si is the only metalloid in period three.

6. When element with atomic number 118 is discovered, what family will it be in?
Element 118 should be a <u>noble gas</u>, as 118 electrons would arrange to fill 7p sublevel

7. Make an argument for placing hydrogen in the halogen family rather than the alkali metals.
Since Hydrogen – H has 1 valence electron it is usually placed in the alkali metal group, as that electron is found in the s sublevel. However, the 1s sublevel has room for only 2 electrons. This creates a situation where the sublevel needs only 1 more electron to be complete. Members of the halogen family also only need one electron to complete their sublevel (p). Therefore, hydrogen would fit in this group as well – if the group was defined as needing only one more electron to complete a sublevel rather than the s AND p sublevels.

8. The alkali metals have a single electron in the highest energy level.
9. The alkaline earth metals achieve the electron configurations of noble gases by losing 2 e-
10. The transition metals vary in the number of electrons in the highest energy level
11. The halogens achieve the electron configuration of noble gases by gaining one electron.
12. The noble gases have full s and p orbitals in the highest occupied energy levels.
13. The noble gases are stable and un-reactive (inert)
14. The halogens are highly reactive and readily form salts with metals.
15. The alkaline earth metals are metals that are more reactive than the transition elements but less reactive than the alkali metals.

16. Predict the oxidation number based on the electron configuration shown.

$1s^2\ 2s^2\ 2p^6\ 3s^2$ (2+)	$1s^2\ 2s^2\ 2p^6$ 0 (none)
$1s^2\ 2s^2\ 2p^6\ 3s^1$ (1+)	$1s^2\ 2s^2\ 2p^5$ (1-)
	$1s^2\ 2s^2\ 2p^1$ (3+)

B. Ionization Energy

1. Choose the element with the greatest first ionization energy:

Carbon or aluminum	Chlorine or argon
Calcium or strontium	Chlorine or fluorine
Helium or lithium	Sulfur or chlorine

2. Which has the larger ionization energy – sodium or potassium? Why?

   Na has the larger 1st ionization energy (495) because the lone e- in the 3s sublevel feels more of the pull from the nucleus in the Na atom. There are fewer inner core e- shielding the 3s e- so more IE is needed to remove the first e-. In the K atom, the 1st IE is smaller than that of Na (418 to 495). Potassium is located in period 4, the e- is further from the nucleus, does not feel the pull of the nucleus as tightly. The inner core e- are shielding the 4s e- thus requiring less energy to remove the e-.

   When totaling all IE for the electrons of the atoms, K has the greater total IE because more e- are removed thus requiring more total energy.

3. Explain the difference in first ionization energy between lithium and beryllium.

As you travel across the PT, the atomic number increases because the number of p+ are increasing with each consecutive atom. With the increase in p+ comes an increase in nuclear charge. The nucleus of the Be atom has a greater positive charge, thus wants to hold on more tightly to the 2 e- located in the 2s sublevel. No additional e- are added to the inner core so no additional shielding takes place. Therefore, Be requires more ionization energy to remove the first e- in the 2s sublevel.

4. The first and second ionization energies of magnesium are both relatively low, but the third ionization energy requirement jumps to five times the previous level. Explain. What is the most likely ion for magnesium to become when it is ionized?

Mg is more likely to become a cation with a 2+ charge by giving up 2 e- from the 3s sublevel. Mg more readily gives up these two e- thus requiring low IE to remove them. However, once the two 3s e- are removed, then Mg has an electron configuration of a noble gas (Ne). Therefore it is unwilling to give up the next e- as easily thus requiring a larger 3rd IE to remove the 3rd e-.

5. Compare the first ionization energies for the noble gases.

He has the highest 1st IE of any element on the PT. The remaining noble gases have the highest 1st IE of any element in their respective periods. However, 1st IE values decrease as you travel down the Nobel gas family to Ra with the lowest IE of the Noble gas family. While still the the highest IE in the 7th row, Ra is a larger atom with more inner core electrons of any atom I the family, thus the IE for the Ra e- is the lowest.

6. Compare the first ionization energies for a noble gas with that of a halogen in the same period. Support your comparison with an orbital diagram.

1st IE of each halogen is lower than that of the noble gas element next to it in the same family. The orbital diagram will show one less e- for each halogen family and each noble gas family with full valence electron shell. A full shell is what the elements strive for so noble gas members require larger IE in order to lose their e-.

7. Where would the largest jump in ionization energies be for oxygen? (with the loss of how many electrons?)

The largest ionization jump will occur following the loss of the 6th e-. The 7th IE is the largest number because at that point, you are trying to remove the $2s^2$ electrons from a full valence shell.

8. How can you tell from a list of ionization energies for an element where a kernel (non valence) electron has been removed?

Because this is where the highest jump in IE occurs, following the removal of the valence e-. Attempting to remove inner core e- (kernel e-) requires larger amounts of energy.

C. Electronegativity and Electron Affinity

1. Arrange the following elements in order of increasing electronegativity.

a. gallium, aluminum, indium indium, gallium, aluminum

b. calcium, selenium, arsenic Selenium, Arsenic, Calcium

c. oxygen, fluroine, sulfur Fluorine, Oxygen, Sulfur

d. phosphorus, oxygen, germanium Oxygen, Phosphorus, Germanium

2. Will the electronegativity of barium be larger or smaller than that of strontium? Explain.

EN of Ba (0.89) is smaller than EN of Sr (0.95). Because of larger size and more inner core e- shielding, Barium will more readily share its valence electron in the 6s sublevel than Sr will share it valence e- in the 5s sublevel.

3. Compare the electronegativity of tellurium to that of antimony. Explain your reasoning.

As you move across the PT EN increases so that the EN of Te (2.1) is greater than Sb (2.05). Te has a greater demand for e- in its desire to form an octect, a full valence shell. Therefore in a chemical bond, Te will pull harder on e- shared with another atom. Te will have agreater attraction for the e- in the bond.

4. The family within any period with the greatest negative electron affinity is usually the ___.

a. alkali metals b. transition metal c. halogens d. noble gases

5. Contrast ionization energy and electron affinity. In general, what can you say about these values for metals and non-metals?

Ionization energy is the energy required to remove an e- from a gaseous neutral atom. It is the energy utilized in the formation of a cation. Ionization energies for metals will be lower than ionization energies for non-metals because IE increases as you travel across the PT.

Electron Affinity is an atom's attraction for electrons. Energy is either given off or force onto an atom when an e- is added to a neutral atom. Electron affinity is a negative measurement of increasing value as you move across the PT. It is the energy released or gained in forming an anion. Nonmetals normally form anion therefore, the EA of non metals will be greater (more negative).

6. What is the difference between electron affinity and electronegativity.

Electron affinity is the atoms attraction for e- the measurement of energy when an e- is added to an atom. EN is a relative value reported on the Pauling scale that tells us how much an atom pulls or attracts an e- while in a chemical bond with another atom.

7. Why is it difficult to determine electron affinities for metals?

Most metals, especially those in the alkali and alkaline groups, want to form cations by giving away e- rather than forming anions by assuming an e- from another atom. For this reason, it is difficult to measure the EA of metals.

D. Definitions - match

Atomic radius	First ionization energy	Noble gases
Decrease	Increase	Nonmetals
Electron affinity	Ionization energy	Semimetal
electronegativity	Metals	Shielding effect
	Noble gas configuration	

1. Ionization energy is the energy required to remove an electron from an atom.
2. The attraction of an atom for an additional electron is called electron affinity.
3. The energy needed to remove the most loosely held electron from a neutral atom is called 1^{st} ionization energy.
4. When they have a(n) noble gas configuration, ions have a stable, filled outer electron level.

5. Along with the increased distance of the outer electrons from the nucleus, the shielding effect of the inner electrons causes ionization energy to decrease going down a column of the periodic table.

6. A low ionization energy is characteristic of a(n) metal.
7. Ionization energies tend to increase across periods of the periodic table.
8. An element with a high ionization energy is classified as a (n) Non-metal (will also take Nobel gas as an answer).

9. The attraction an atom has for shared electrons is called electronegativity.

10. The distance from the nucleus to the outer most electron is known as atomic radius.

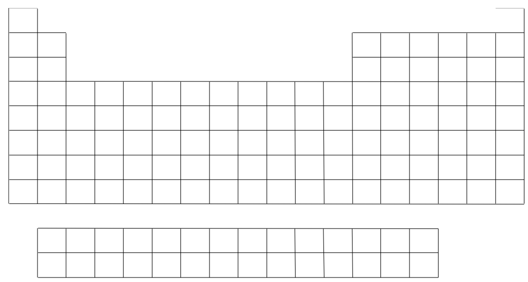
11. The Noble gases do not have measured electronegativites since they do not commonly form compounds.

12. The electron arrangement with a complete outermost s and p sublevel is known as Noble gas configuration.

E. Trend Chart

Draw in the trends on the periodic table: Look in your book for these trends

Ionization energy   electronegativity   atomic radius   electron affinity   shielding effect



F. Atomic Radius

1. Circle the atom in each pair with the larger atomic radius?

Li or K	Cl or Br
Ca or Ni	Be or Ba
Ga or B	Si or S
O or C	Fe or Cu

2. Chlorine selenium, and bromine are located near each other on the periodic table. Which of these elements is the smallest atom and which has the highest ionization energy?

   Chlorine has the smallest atomic radius and the highest ionization energy.

3. Which of the following atoms is smallest: nitrogen, phosphorus, or arsenic? Which of these atoms has the most negative electron affinity? Nitrogen is the smallest atom and the most negative EA.

4. Which of the following is the largest: a potassium atom, a potassium ion with a charge of 1+ or a rubidium atom? LOOK at the charts in your BOOK – pages 163 and 166.

5. Which of the following is the largest: a chlorine atom, a chlorine ion with a charge of 1- or a bromine atom?

6. Which of the following is the smallest: a lithium atom, a lithium ion with a charge of 1+ or a sodium atom?

7. Use the atomic theory to explain why within a family such as the halogens, the ionic radius increases as the atomic number increases.

The ionic radius increases as you travel down the family of halogens because the number of protons becomes increasing larger with increasing atomic number. Also, each element below the previous element has one additional energy level. The electrons in the higher energy level are further from the nucleus and thus feel less of the charge of the nucleus.

8. In terms of electron configuration and shielding, why is the atomic radius of sodium smaller than that of potassium?

Potassium has more inner core electrons and thus more shielding of the valence electrons. The inner core e- shield the valence electrons and do not pull as tightly on the valence. Also, the valence e- of Potassium is in a higher principle energy level which is further away from the nucleus of the atom thus allowing the atom to be larger.

9. In terms of electron configurations and shielding, Why do atoms get smaller as you move across a period? As you move across the PT, e- are added in the same valence shell so the nuclear charge can pull the e- closer to the nucleus. The inner core electrons do not increase as you move across the PT, you are only adding e- to the valence shell.

G. Concept Mastery Questions

1. The shielding effect increases with increasing atomic number within a ___.
a. period b. group c. both d. neither

2. In any ___, the number of electrons between the nucleus and the outer energy level is the same.
a. period b. group c. both d. neither

3. Within a ____, the nucleus has a stronger ability to pull on the outermost (valence) electrons in elements of high atomic number.
a. period b. group c. both d. neither

4. In a ____, electron affinity values become more negative as atomic number increases.
a. period b. group c. both d. neither

5. The halogens are considered a ____.
a. period b. group c. both d. neither

6. Which atom has the greater nuclear charge?
a. Na b. Al c. P d. Ar

7. Which atom demonstrates the greatest shielding effect?
a. Na b. Al c. P d. Ar

8. The atoms Na, Al, P, and Ar all have the same
a. shielding b. size atomic radius c. number of valence electrons d. number of kernel electrons

9. Which element on the periodic table has
a. lowest ionization energy - Cesium
b. highest second ionization energy – Lithium
c. highest electronegativity - Fluorine
d. highest ionization energy – Helium – followed by Fluorine
e. largest atomic radius – Cesium

10. Explain the relationship between the relative size of an ion to its atom and the charge on the ion.

Cation: has lost an e-, therefore electrostatic repulsion b/w e- is decreased and the size of the ion shrinks. The cation of an element is smaller than the neutral atom of the same element.

Anion: has gained an e-, therefore electrostatic repulsion b/w e- is increased and the size of the ion increases. The anion of an element is larger than the neutral atom of the same element.

11. Explain why noble gases are inert and do not form ions.

Inert means un-reactive. The noble gases have a full octect of valence electrons, 2s and 6 p electrons in its outer valence shell. The noble gases do not need to nor do they have to accept or lose electrons. The number of valence electrons determines an element's chemical properties, therefore with 8 valence e-, the Noble gases do not form ions.

12. Will the shielding effect be more noticeable in metals or non metals? Explain your answer.

When comparing metals and nonmetals across the same period, the metals will have the more noticeable shielding affect. As you move across the PT, the number of p+ in the nucleus increases thus increasing the nuclear charge of the atom. The shielding effect of the inner core e- becomes less noticeable with the nonmetals because the valence e- are pulled more tightly towards the nucleus.

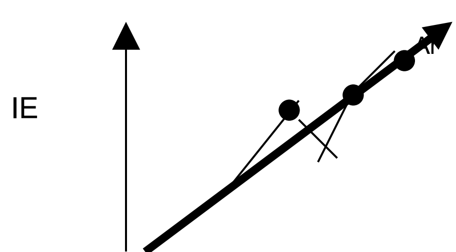
13. Why do elements in the same family generally have similar properties? Choose one as an example to support your reasoning.

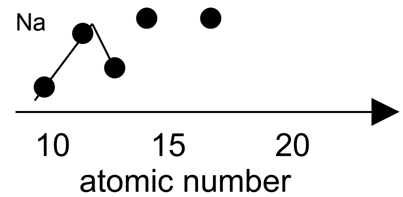
Halogens Family – The number of valence e- determine the element's chemical properties such as reactivity. Each member of the halogen family has 7 e- in its valence shell. Therefore, the members of the halogen family all have similar chemical properties.

14. Arrange each of the following in order of increasing ionization energy and explain your reasoning: Calcium, iron, copper, bromine and krypton.

Calcium, copper, iron, bromine, krypton.

15. Factors affecting ionization energy include nuclear charge, the shielding effect, the atomic radius and the electron arrangements in a sublevel. Use the appropriate factors to explain the overall trend indicated by the dark line and the exceptions to it in the graph (light lines).





The trend chart above is very similar to the 1st Ionization Energy trends located in your text book. The trend shows the 1st IE for each element in the 3rd period of the PT. Starting the Na and ending with Ar, the IE values in kJ/mol are, Na (495), Mg (737), Al (577), Si (786), P (1011), S (999), Cl (1251), Ar (1520). As the nuclear charge increases moving across the periodic table, the atomic radius decreases as valence e- are pulled closer to the nucleus. Moving across the PT, there is an increase in valence e- but not an increase in inner core electrons. The shielding effect of the inner core electrons is increasingly overcome by the increasing nuclear charge. Fore these reasons, the energy required to remove the 1st electron increases as you move across the PT. The above trend demonstrates this increase.

However, there are exceptions in that the 1st IE dips at elements Al and again at S. The reason for this can be seen in the electron configurations. The electron configuration for Al shows 13 electrons, $1s^2 2s^2 2p^6 3s^2 3p^1$. The energy to remove the last e- from the 3p sublevel is less than the energy to remove the electrons from the 3s sublevel because the 3s is a full sublevel and the atom wants to hold onto the electrons in the full sublevels. Similar reasoning may be used when explaining the dip between Phosphorus and Sulfur. P has an electron configuration that has a half-filled 3p sublevel. The next element, S, has a 3p sublevel where one of the three orbitals contains two e- while the other two orbitals only contain one e-. It is easier to remove the 4th e- in Sulfur's 3p sublevel than it is to remove the 3rd e- from Phosphorus 3p sublevel because half-filled sublevels are more stable. Draw the orbital diagrams to prove this to yourself.

16. I am an element. I have a high electron affinity, (highly negative value), and my atomic number is X. The element with atomic number X-1 has a lower ionization energy and a lower electron affinity. The element with atomic number N+1 has a higher ionization energy and basically no electron affinity (positive value). I am toxic in my elemental state, but I am very commonly found in my nontoxic ionic form. Within my group, I have the second highest ionization energy. Who am I? (support each step of your reasoning). Chlorine

17. Write a chemical equation that shows the process or events in the formation of an anion.

A + e-   A-

18. What do transition metals have in common with respect to their electron configurations?

All transition metals are located in the d block. Their electron configurations begin with a filled s sublevel of the energy level equal to the row number of the period they are located on. The s sublevel is filled before the d sublevel (s before d). Their valence shells all have 2 electrons in the outermost shell.

19. Consider the table of the first four ionization energies for an element we will call A.

Ionization energy in kJ/mol	1st	2nd	3rd	4th
	578	1817	2745	11580

a. In which group does A appear on the periodic table? 4A

b. What is the most likely oxidation number for element A? (3+)

c. What is the minimum number of electrons that A must have? (13)

d. Write the valence electron sublevel configuration for this element. (sublevel and number of electrons in them) [Ne] $3s^2 3p^1$

20. Can anions of two different elements have the same valence electron arrangement? If so, give examples and discuss. If no, explain why not. Yes

Cl^- and S^{2-} Both anions have the electron configuration of [Ne]$3s^2 3p^6$

21. When an atom loses an electron to become an ion, what happens to its electric charge? To its size? Write a chemical equation that shows the process or events in the formation of this ion. What energy value is associated with this process?

When an atom loses an e- to become an ion, the atom becomes a cation and has a positive charge. The energy value associated with this process is called ionization energy, the amount of energy required to remove an e- form a neutral atom.

A – e- A+

Simplier Short answer questions

Look at questions in Section A, Section D and the first 9 questions of Section G.

Pick out 3-4 questions from the sections above you are confident you know the answers to for the discussion Tuesday. Turn in a written answer to one of the discussion questions for Wednesday.

Section A

8. The alkali metals have a single electron in the highest energy level.
9. The alkaline earth metals achieve the electron configurations of noble gases by losing 2 e-
10. The transition metals vary in the number of electrons in the highest energy level
11. The halogens achieve the electron configuration of noble gases by gaining one electron.
12. The noble gases have full s and p orbitals in the highest occupied energy levels.
13. The noble gases are stable and un-reactive (inert)
14. The halogens are highly reactive and readily form salts with metals.
15. The alkaline earth metals are metals that are more reactive than the transition elements but less reactive than the alkali metals

Section D

Atomic radius	First ionization energy	Noble gases
Decrease	Increase	Nonmetals
Electron affinity	Ionization energy	Semimetal
electronegativity	Metals	Shielding effect
	Noble gas configuration	

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4. When they have a(n) noble gas configuration, ions have a stable, filled outer electron level.
5. Along with the increased distance of the outer electrons from the nucleus, the shielding effect of the inner electrons causes ionization energy to decrease going down a column of the periodic table.
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g. highest second ionization energy – Lithium
h. highest electronegativity - Fluorine
i. highest ionization energy – Helium – followed by Fluorine
j. largest atomic radius – Cesium

Pick a <u>discussion</u> question from sections B, C, F, or >G #9 and turn in a written copy on Wednesday at beginning of class. We will still discuss complex questions on Wednesday