AP Chemistry Name \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

MOAR Free Response Practice

**The Periodic Law and Ionic Charge**

1. The PES spectrum for an element can be found below:

Intensity

0.63 0.77 3.24 5.44 39.2 48.5 433

Binding Energy (x 10-19 J)

1. Identify the type of orbital responsible for the peak at:
2. 0.63
3. 0.77
4. 3.24
5. Identify the element:
6. Using evidence from the PES spectrum:
   * 1. From which orbital would electrons be lost most easily? Justify your answer.
     2. Why would a 4+ ion be highly unlikely for this element? Justify your answer.
     3. What would be the most common ionic charges for this element? Justify your answer.
7. What would be another element that would likely form similar ionic charges as this element? Justify your answer.
8. Oxygen and sulfur both form ions with a -2 charge, despite sulfur having 16 electrons to oxygen’s 8.
   1. What is the periodic law and how does it account for this observation?
   2. Based on the periodic law, we might expect the next element to form a -2 ion to have 24 electrons, but instead we must wait for an element with 34 electrons to form a -2 anion. Why?
9. Lead generally forms ions with a +2 and a +4 charge. For both cations:
   1. Which electrons were lost to form each?
   2. What kind of electrons were these (core or valence)? Explain.
   3. What would be the expected ionic charges for:
      1. Ge?
      2. Tl?
10. Both Al and Cl are found in the second period. The binding energies for the PES peak for their 3p electrons can be found below:

Element Binding Energy of 3p electrons (x 10-19 J)

Al 0.58

Cl 1.25

* 1. In terms of orbital stability, why does chlorine tend to gain one electron and aluminum lose three?
  2. Using evidence from the PES binding energies, explain why aluminum tends to form cations and chlorine tends not to?
  3. Scandium (Sc) can form +3 cations just as aluminum can:
     1. Explain how this is possible despite Al and Sc being in different groups:
     2. What charge might we find on a scandium ion that would be unlikely on an aluminum ion? Justify your answer.

1. Atoms within the same group tend to form ions with similar charges whereas ionic charges tend to vary significantly throughout a period. Explain this observation.
2. For each of the following cations, explain from which orbital the electrons were lost and how many. In addition, indicate which orbital electrons would be lost from first by listing that orbital first.
   1. Mn7+
   2. Cr3+
   3. K+
   4. Ga3+
   5. Sn4+
   6. Cd2+
   7. Ag+
3. Sulfur tends to form ionic compounds with copper (II) ions in the 1:1 ratio.
   1. Another ion also forms an ionic compound with copper(II) in a 1:1 ratio. What group is the atom that formed this ion likely in?
   2. Another ion forms an ionic compound with copper(II) in a 3:2 ratio (3 copper cations for every 2 anions). What group is the element that forms this anion likely in?
4. Phosphorus and arsenic can both form ions with a -3 charge.
   1. Explain how this is possible despite arsenic having 18 more electrons than P.
   2. Explain why the next element that may carry a -3 charge will have exactly 18 more electrons than arsenic.

**Periodic Law and Atomic Radii**

* + - 1. For the following statements, decide if you disagree or agree.

1. As atomic number increases down a group, the atomic radius of the atoms will increase. Justify.
2. Flourine has a smaller atomic radii (42 pm) than oxygen (48 pm) principally because there is more shielding between the valence electrons and the nucleus in an oxygen atom. Justify.
   * + 1. Which electrons (core or valence) are responsible for shielding the nuclear charge and how does this shielding influence the atomic radii?
       2. We might normally define the atomic radius as being the distance between the nucleus and the outermost electron(s). Based on our understanding of the quantum model of the atom, why can this not be measured with certainty?

4. The PES spectra for nitrogen and fluorine can be found below:

F

N

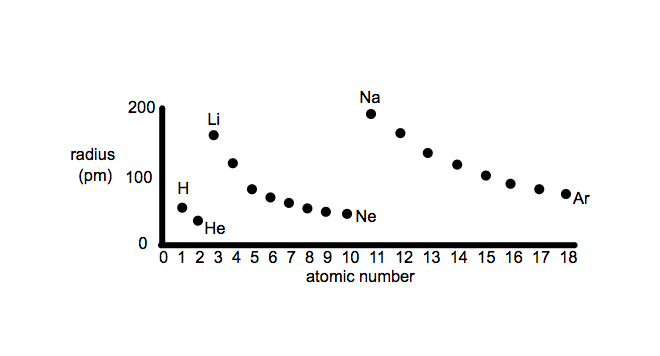
Intensity

1.31 3.04 52.6 1.68 3.88 67.2

Binding Energy (x 10-19 J) Binding Energy (x 10-19 J)

1. Which of these two would have the smaller atomic radii? Justify your answer using data from the PES spectra.
2. What is responsible for the difference in atomic radii? (Shielding or nuclear charge) Justify your answer.
3. Why is the F 1s peak of a higher binding energy than the N 1s peak?
4. How would the radii of an F- ion compare to that of a neutral F atom? Justify your answer.
5. Francium has the largest atomic radii of any element (270 pm)on the periodic table. Explain how this is possible considering it has a very high nuclear charge.
6. Rank the following atoms/ions in order of decreasing atomic radii:
7. Ca, Br, Br-, Ca2+
8. Cl, Cl-, Ar

1. Boron, aluminum, and gallium are in group 13. Their atomic radii are (in pm) 85, 143, and 135.
   1. Why is the atomic radii of gallium an unexpected value?
   2. What can you discern about the shielding ability of the 3d orbital based on this data?
      * 1. If one were to graph the atomic radii versus atomic number, the graph would look like the one below:



1. Explain the following:
2. The substantial increase in atomic radii from He to Li:
3. The decrease in atomic radii from Na to Ar
4. The increase in atomic radii from He to Ne to Ar
5. Would you expect the atom with atomic number 19 to have a radii larger or smaller than Na? Explain:
   * + 1. Magnesium tends to form ions with a 2+ charge. Would these ions have an ionic radii greater or less than:
   1. The atomic radii of neutral magnesium. Justify your answer.
   2. The ionic radii of fluoride (F-). Justify your answer.
      * 1. For the following statements, agree or disagree and then justify your answer.
6. Xenon has a larger atomic radii than krpton because of the greater number of protons in the xenon nucleus.
7. Sulfur’s valence electrons feel a stronger nuclear charge than do phosphorus’s valence electrons because there is less shielding from sulfur’s core electrons.
   * + 1. Written below are the electron configurations of five different elements:

Element A: 1s22s22p63s1

Element B: 1s22s22p3

Element C: 1s22s22p63s23p63d10

Element D: 1s22s22p63s23p64s23d10

Element E: 1s22s22p63s23p64s23d104p4

1. Which of these would have the largest atomic radii? Justify your answer.
2. Which would have the smaller atomic radii; element “B” or element “A”? Justify your answer.
3. List all of these elements that would have an atomic radii smaller than germanium (Ge).
   * + 1. Rank the following atoms/ions in order of increasing atomic radii:
          1. Ga, Ge, Si
          2. P, S, Ar, P3-

**Periodic Law and Ionization Energy**

1. What is ionization energy and what are two factors that influence it? For each factor, describe how it influences the ionization energy.

2. Which of the processes below represents the following:

Mg 🡪 Mg+ + e-

Mg+ + e- 🡪 Mg

Mg+ 🡪 Mg2+ + e-

Mg2+ + e- 🡪 Mg+

a. The first ionization energy for magnesium?

b. The second ionization for magnesium?

c. The process requiring the most energy to occur?

3. The following questions pertain to the graph below:



a. Explain why, in general, there is an increase in the first ionization energy from Li to Ne.

b. Explain why the first ionization energy of B is lower than that of Be.

c. The first ionization energy of O is lower than that of N.

d. Predict how the first ionization energy of Na compares to those of Li and of Ne. Explain.

4. Explain the following observations:

a. The first ionization energy of Mn is 717 kJ/mol while the first ionization energy of Cr is 656 kJ/mol.

b. The noble gases have the highest ionization energies for their period.

5. The ionization energies for strontium and magnesium are listed below:

1st IE (kJ/mol) 2nd IE(kJ/mol) 3rd IE (kJ/mol)

Sr 550 1064 4130

Mg 740 1450 7730

a. Explain why the ionization energies of Mg are consistently higher than those of Sr:

b. Why is there such a large difference between the 2nd and 3rd ionization energies for these two elements?

6. The PES spectrum for lithium and helium can be seen below:

Li

He

Intensity

n

0.52 6.26 2.37

Binding Energy (x 10-19 J)

a. Which orbitals are responsible for producing the peaks listed at binding energies:

i. 0.52?

ii. 6.26?

iii. 2.37?

b. Why is less energy required to remove lithium’s 2s electron compared to helium’s 1s electrons?

c. Why is more energy required to remove lithium’s 1s electrons than helium’s 1s electrons?

d. Which element would have the highest:

i. First ionization energy? Justify your answer.

ii. Second ionization energy? Justify your answer.

1. How does ionization energy relate to atomic radii (inversely or directly)? Explain.

2. Explain the difference between Ca and K in regard to

a. Their first ionization energies,

b. Their second ionization energies.

3. It is observed that the first ionization energy of Mg is 738 kilojoules per mole and that of Al is 578 kilojoules per mole.

a. Account for this difference and justify your answer.

b. How much energy would be required to remove the first single electron from a magnesium atom?

c. What wavelength of light would be required to remove this electron?

d. Which element (Mg or Al) would be expected to have the “2p” peak in their PES spectra with the highest binding energy? Justify your answer.

e. Would the binding energy of aluminum’s 3p electron be higher or lower than that of magnesium’s 3s electrons? Justify your answer.

4. The PES spectrum for sulfur is shown below:

Intensity

1.0 2.05 16.5 22.7 239

Binding Energy (x 10-19 kJ)

a. Which peak corresponds to sulfur’s valence “p” orbital electrons?

b. Would the same peak in Cl be expected to have a higher or lower binding energy compared to S? Justify your answer.

c. Would the same peak in P be expected to have a higher or lower binding energy than that of sulfur? Justify your answer.

d. Why the binding energy of helium’s 1s electrons is so much lower (2.37) than that of sulfur’s 1s electrons?

5. Rank the following elements in order of increasing first ionization energy?

a. Na, K, Sc

b. I, Xe, Te

**Periodic Law, Electronegativity and Metallic Character**

1. Do metallic elements typically have high or low ionization energies? Explain your answer.

2. Lead and carbon reside in the same period yet lead demonstrates metallic properties while carbon does not. Explain.

3. Define electronegativity and provide an explanation as to why the noble gases do not have measurable electronegativity values.

4. Draw a graph that shows the general trend in electronegativity when plotted vs. the atomic number for the first 18 elements. Do not use specific electronegativity values but do ensure your points are internally consistent.

5. For the following compounds, indicate which element would have the strongest pull on electrons and explain why:

a. SF6

b. H2O

c. N2O

d. CH3Cl

1. Both bromine and copper are in the same period:

a. Explain why bromine is a non-conductor of electricity while copper is an excellent conductor of electricity.

b. Silver is in the same group as copper. Is it more or less metallic than copper? Explain.

2. Explain the general trend in electronegativity as the atomic number increases down a group AND provide a rationale for this trend.

3. Chlorine has 8 more protons in it’s nucleus than does fluorine. Provide a reason for it’s observed electronegativity (3.2) being lower than that of fluorine (4.0).

4. Draw a graph that shows the trend in metallic character for group 16 when plotted vs. atomic number. Do not use quantitative values for metallic character but make sure your data is internally consistent.

5. Rank the following elements in order of decreasing electronegativity

a. C,Si, S, N

b. Ca, Al, Mg, P

**Periodic Table and Specific Groups**

1. The alkali metals are more reactive than the alkaline earth metals:

a. Provide an example of an alkali metal and an alkaline earth metal.

b. In terms of ionization energies, provide a rationale for this difference in reactivity.

b. Which alkali metal would be expected to be more reactive (K or Na)? Justify your answer.

2. An atom forms an ion which combines with the aluminum ion in a 3:1 ratio. Which group must this element belong to on the periodic table?

3. Using the following clues, identify the element:

a. Halogen with an atomic radii smaller than barium but larger than chlorine.

b. Has three unpaired electrons in it’s 4d orbital (two options - list both)

c. Transition metal with the highest ionization energy.

d. Alkali metal with the smallest atomic radii.

1. The halogens are the most reactive group of non-metals. Propose a reason for this.

2. Group 16 elements are sometimes referred to as the chalcogens.

a. Which element would be most metallic within this group? Justify your answer.

b. When oxygen is combined with fluorine, which element would have the:

i. Larger ionization energy? Justify your answer.

ii. Smaller atomic radii? Justify your answer

3. Using the following clues, identify the following element:

a. Transition metal with the largest atomic radii with a partially filled 4d orbital. Justify your answer.

b. Period 2 element with 2 unpaired electrons (two choices - list both). Justify your answer.

c. Alkaline earth metal with the lowest 3rd ionization energy. Justify your answer.

Teacher Key

**The Periodic Law and Ionic Charge**

Classwork:

1. The PES spectrum for an element can be found below:

Intensity

0.63 0.77 3.24 5.44 39.2 48.5 433

Binding Energy (x 10-19 J)

1. Identify the type of orbital responsible for the peak at:
2. 0.63 = ***4s2***
3. 0.77 = ***3d1***
4. 3.24 = ***3p6***
5. Identify the element: ***Sc***
6. Using evidence from the PES spectrum:
   * 1. From which orbital would electrons be lost most easily? Justify your answer.

***The 4s orbital as it has the lowest binding energy.***

* + 1. Why would a 4+ ion be highly unlikely for this element? Justify your answer.

***After losing both the 4s and 3d electrons (3 total), the binding energy of the 3p orbital electrons is much higher due to the full valence shell and less shielding compared to the 4s and 3d orbitals. As a result they would be difficult to remove.***

* + 1. What would be the most common ionic charges for this element? Justify your answer.

***+2 and +3 – the +2 from losing just the 4s electrons, the +3 from losing both the 4s and 3d electrons.***

1. What would be another element that would likely form similar ionic charges as this element? Justify your answer.

***Y – it is in the same group so has the same valence/outer shell electron configuration.***

1. Oxygen and sulfur both form ions with a -2 charge, despite sulfur having 16 electrons to oxygen’s 8.
   1. What is the periodic law and how does it account for this observation?

***The periodic law states that the properties of the elements repeat as a function of their atomic number. This is due to the nature of principal energy levels filling and then new ones starting again.***

* 1. Based on the periodic law, we might expect the next element to form a -2 ion to have 24 electrons, but instead we must wait for an element with 34 electrons to form a -2 anion. Why?

***The sulfur atom is two shy of a full 3rd energy level, so in addition to the two additional electrons needed to fill the 3rd energy level, all new electrons get added to the 4s(2), the 3d(10), and then 4 more to get to the same valence configuration as O and S. (18 total)***

1. Lead generally forms ions with a +2 and a +4 charge. For both cations:
   1. Which electrons were lost to form each?

***Pb2+ = lost 6p2 electrons***

***Pb4+ = lost 6p2 and 6s2 electrons***

* 1. What kind of electrons were these (core or valence)? Explain.

***Valence electrons as they are outermost s and p electrons***

* 1. What would be the expected ionic charges for:
     1. Ge? ***+2 and +4***
     2. Tl? ***+2 and +3***

Homework

1. Both Al and Cl are found in the second period. The binding energies for the PES peak for their 3p electrons can be found below:

Element Binding Energy of 3p electrons (x 10-19 J)

Al 0.58

Cl 1.25

* 1. In terms of orbital stability, why does chlorine tend to gain one electron and aluminum lose three?

***Cl needs to gain 1 electron to reach a stable valence shell (full s and p)***

***Al needs to lose 3 electrons to reach a stable valence shell, a more energetically favorable option than gaining 5 electrons.***

* 1. Using evidence from the PES binding energies, explain why aluminum tends to form cations and chlorine tends not to?

***Chlorine has much higher binding energies for it’s valence electrons than aluminum making it much more plausible it would lose electrons.***

* 1. Scandium (Sc) can form +3 cations just as aluminum can:
     1. Explain how this is possible despite Al and Sc being in different groups:

***Al loses its outermost 3p and 3s electrons***

***Sc loses its outermost 4s then 3d electrons***

* + 1. What charge might we find on a scandium ion that would be unlikely on an aluminum ion? Justify your answer.

***+2 , as scandium loses its 4s electrons first, aluminum loses it’s 3p1 electron first, then its 3s electrons.***

1. Atoms within the same group tend to form ions with similar charges whereas ionic charges tend to vary significantly throughout a period. Explain this observation.

***Atoms within the same group have the same valence shell configuration whereas atoms in the same period all have different valence shell configurations.***

1. For each of the following cations, explain from which orbital the electrons were lost and how many. In addition, indicate which orbital electrons would be lost from first by listing that orbital first.
   1. Mn7+***4s2 then 3d5***
   2. Cr3+ ***4s2 then 3d1***
   3. K+ ***4s1***
   4. Ga3+ ***4p1 then 4s2***
   5. Sn4+ ***5p2 and 5s2***
   6. Cd2+ ***5s2***
   7. Ag+ ***5s1***
2. Sulfur tends to form ionic compounds with copper (II) ions in the 1:1 ratio.
   1. Another ion also forms an ionic compound with copper(II) in a 1:1 ratio. What group is the atom that formed this ion likely in?

***Same as sulfur – group 16***

* 1. Another ion forms an ionic compound with copper(II) in a 3:2 ratio (3 copper cations for every 2 anions). What group is the element that forms this anion likely in?

***Must have -3 charge, so group 15.***

1. Phosphorus and arsenic can both form ions with a -3 charge.
   1. Explain how this is possible despite arsenic having 18 more electrons than P.

***Despite arsenic having more electrons, the bulk of them are in the next highest energy level yielding the same valence shell configuration as P – ending in p3.***

* 1. Explain why the next element that may carry a -3 charge will have exactly 18 more electrons than arsenic.

***3 to fill the 4p, 2 to fill the 5s, 10 to fill the 4d, and 3 in the 5p to have the same valence configuration as P and As***

**Periodic Law and Atomic Radii**

Classwork

* + - 1. For the following statements, decide if you disagree or agree.

1. As atomic number increases down a group, the atomic radius of the atoms will increase. Justify.

***Agree - As more energy levels are filled, there becomes increased shielding of the nuclear charge allowing the valence electrons to expland outward expanding the atomic radii***

1. Flourine has a smaller atomic radii (42 pm) than oxygen (48 pm) principally because there is more shielding between the valence electrons and the nucleus in an oxygen atom. Justify.

***Disagree –The smaller atomic radii is due to a higher nuclear charge due to F’s extra proton while the shielding is essentially the same between the two.***

* + - 1. Which electrons (core or valence) are responsible for shielding the nuclear charge and how does this shielding influence the atomic radii?

***Core – they cancel out the nuclear charge so the valence electrons feel less coulombic attraction and therefore expand outward.***

* + - 1. We might normally define the atomic radius as being the distance between the nucleus and the outermost electron(s). Based on our understanding of the quantum model of the atom, why can this not be measured with certainty?

***We cannot know the precise location of an electron so we cannot know the precise boundary of an atom***

4. The PES spectra for nitrogen and fluorine can be found below:

F

N

Intensity

1.31 3.04 52.6 1.68 3.88 67.2

Binding Energy (x 10-19 J) Binding Energy (x 10-19 J)

1. Which of these two would have the smaller atomic radii? Justify your answer using data from the PES spectra.

***F, the outermost 2p electrons are held tighter than those of N as evidenced by the higher binding energy – so the nuclear charge is stronger and the electrons are pulled inward more.***

1. What is responsible for the difference in atomic radii? (Shielding or nuclear charge) Justify your answer.

***Nuclear charge – the shielding for both atoms is essentially the same as they are in the same period.***

1. Why is the F 1s peak of a higher binding energy than the N 1s peak?

***F has the higher nuclear charge and therefore exerts a stronger coulombic attraction.***

1. How would the radii of an F- ion compare to that of a neutral F atom? Justify your answer.

***Larger, the extra electron does not feel a strong nuclear charge as it has been diminished by the existing electrons, therefore increasing the radii.***

5. Francium has the largest atomic radii of any element (270 pm) on the periodic table. Explain how this is possible considering it has a very high nuclear charge.

***It also has 6 full energy levels of shielding between the nucleus and valence electrons. All other elements – such as Ra or U have more protons than Fr for it’s period.***

1. Rank the following atoms/ions in order of decreasing atomic radii:
2. Ca, Br, Br-, Ca2+ ***Ca > Br- > Br > Ca2+***
3. Cl, Cl-, Ar ***Cl- > Cl >Ar***

1. Boron, aluminum, and gallium are in group 13. Their atomic radii are (in pm) 85, 143, and 135.
   1. Why is the atomic radii of gallium an unexpected value?

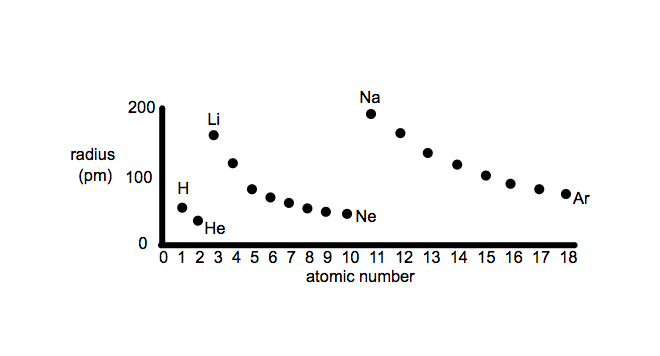
***It has more shielding so we would expect the radii to be larger.***

* 1. What can you discern about the shielding ability of the 3d orbital based on this data?

***The d orbitals do not shield as well as either the s or p orbital electrons so, the higher nuclear charge of gallium actually pulls the electrons in tighter.***

Homework

* + - 1. If one were to graph the atomic radii versus atomic number, the graph would look like the one below:



1. Explain the following:
2. The substantial increase in atomic radii from He to Li:

***Lithium’s valence electron is in the 2s orbital so the 1s orbital acts to shield the nuclear charge thereby lessening the coulombic attractions and allowing the atom to expand. With helium, there is no shielding between the valence electrons and the nucleus so the attractions are strong and the radii is small.***

1. The decrease in atomic radii from Na to Ar

***All of the elements in this period are shielded by the first two energy levels so it is the nuclear charge that factors. As we move across the period, the atomic number and nuclear charge increase thereby increasing the attractive forces and shrinking the radii.***

1. The increase in atomic radii from He to Ne to Ar

***As we move down this group, each element has new energy level of core electrons shielding the nuclear charge. Due to this, there is less attraction and thereby a greater radii.***

1. Would you expect the atom with atomic number 19 to have a radii larger or smaller than Na? Explain:

***Larger, it now has an extra level of shielding compared to sodium.***

* + - 1. Magnesium tends to form ions with a 2+ charge. Would these ions have an ionic radii greater or less than:
  1. The atomic radii of neutral magnesium. Justify your answer.

***Less than, atoms become smaller as they lose electrons as the coulombic attraction by the nucleus on the remaining electrons is greater.***

* 1. The ionic radii of fluoride (F-). Justify your answer.

***Less than, both ions are isoelectronic with neon and thereby see the same amount of shielding. The magnesium ion however has a much higher nuclear charge and thereby makes the ion smaller.***

* + - 1. For the following statements, agree or disagree and then justify your answer.

1. Xenon has a larger atomic radii than krypton because of the greater number of protons in the xenon nucleus.

***Disagree, the radii is larger due to the increased shielding of xenon’s nuclear charge by additional levels of core electrons.***

1. Sulfur’s valence electrons feel a stronger nuclear charge than do phosphorus’s valence electrons because there is less shielding from sulfur’s core electrons.

***Disagree – the shielding is essentially the same, sulfur’s valence electrons feel as stronger nuclear charge because sulfur has an additional proton in the nucleus.***

* + - 1. Written below are the electron configurations of five different elements:

Element A: 1s22s22p63s1

Element B: 1s22s22p3

Element C: 1s22s22p63s23p63d10

Element D: 1s22s22p63s23p64s23d10

Element E: 1s22s22p63s23p64s23d104p4

1. Which of these would have the largest atomic radii? Justify your answer.

***Element D – it has the most shielding (3 levels of core electrons) and a smaller nuclear charge than Element E***

1. Which would have the smaller atomic radii; element “B” or element “A”? Justify your answer.

***Element A as it has less shielding between it’s valence electrons and the nucleus.***

1. List all of these elements that would have an atomic radii smaller than germanium (Ge).

***Elements E, A, B, C***

* + - 1. Rank the following atoms/ions in order of increasing atomic radii:
         1. Ga, Ge, Si  ***Si < Ge < Ga***
         2. P, S, Ar, P3- ***Ar < S < P < P3-***

**Periodic Law and Ionization Energy**

**Classwork**

* + 1. What is ionization energy and what are three factors that influence it? For each factor, describe how it influences the ionization energy.

***The energy required to remove an electron from an atom. It is influenced by:***

***- Nuclear charge – the higher it is, the higher the IE***

- ***Shielding – the higher it is, the lower the IE***

***- Orbital effects – full or half-full orbitals render stability to the atom and increase the IE.***

2. Which of the processes below represents the following:

I. Mg 🡪 Mg+ + e-

II. Mg+ + e- 🡪 Mg

III. Mg+ 🡪 Mg2+ + e-

IV. Mg2+ + e- 🡪 Mg+

1. The first ionization energy for magnesium?

***I***

1. The second ionization for magnesium?

***III***

1. The process requiring the most energy to occur?

***III***

3. The following questions pertain to the graph below:



1. Explain why, in general, there is an increase in the first ionization energy from Li to Ne.

***The shielding is constant across the period but nuclear charge increases thereby increasing the IE***

1. Explain why the first ionization energy of B is lower than that of Be.

***Beryllium’s IE is unexpectedly higher due to its full 2s orbital.***

1. The first ionization energy of O is lower than that of N.

***Nitrogen’s IE is unexpectedly higher due to its half-full 2p orbital***

1. Predict how the first ionization energy of Na compares to those of Li and of Ne. Explain.

***It will be less than that of both Li and Ne due to the increased shielding of the nucleus by the now full 2nd energy level.***

4. Explain the following observations:

a. The first ionization energy of Mn is 717 kJ/mol while the first ionization energy of Cr is 656 kJ/mol.

***The shielding is constant but Mn has the extra proton in the nucleus thereby increasing the nuclear charge increasing the IE***

b. The noble gases have the highest ionization energies for their period.

***They have the highest nuclear charge for the level of shielding experienced by that period.***

5. The ionization energies for strontium and magnesium are listed below:

1st IE (kJ/mol) 2nd IE(kJ/mol) 3rd IE (kJ/mol)

Sr 550 1064 4130

Mg 740 1450 7730

1. Explain why the ionization energies of Mg are consistently higher than those of Sr:

***It has less shielding so the electrons experience a greater nuclear charge and attraction thereby increasing the IE***

1. Why is there such a large difference between the 2nd and 3rd ionization energies for these two elements?

***After both of these elements lose two electrons, they become isoelectronic with the preceding noble gas and lose a level of shielding between the valence electrons and the nucleus thereby creating much greater attractions and a much higher IE.***

6. The PES spectrum for lithium and helium can be seen below:

Intensity

n

0.52 6.26 2.37

Binding Energy (x 10-19 J)

a. Which orbitals are responsible for producing the peaks listed at binding energies:

i. 0.52? ***2s1***

ii. 6.26? ***1s2***

iii. 2.37? ***1s2***

b. Why is less energy required to remove lithium’s 2s electron compared to helium’s 1s electrons?

***The 2s electron is shielded by the 1s orbital***

1. Why is more energy required to remove lithium’s 1s electrons than helium’s 1s electrons?

***Lithium has the higher nuclear charge and the shielding is the same for both elements 1s orbital***

d. Which element would have the highest:

i. First ionization energy? Justify your answer.

***Helium due to less shielding and a full valence shell***

ii.Second ionization energy? Justify your answer.

***Lithium as it is now isoelectronic with helium with a full valence shell but also has the greater nuclear charge for the same amount of shielding.***

Homework

1. How does ionization energy relate to atomic radii (inversely or directly)? Explain.

***Inversely, the greater the atomic radii, the greater the shielding or less nuclear charge – either way – less coulombic attractions, the lower the IE.***

2. Explain the difference between Ca and K in regard to

a. Their first ionization energies.

***Calcium’s will be greater due to the higher nuclear charge for a given level of shielding.***

b. Their second ionization energies.

***Potassium’s will be greater due to its’ being isoelectronic with argon thereby losing a level of shielding and enjoying the stability of a full valence shell.***

3. It is observed that the first ionization energy of Mg is 738 kilojoules per mole and that of Al is 578 kilojoules per mole.

a. Account for this difference and justify your answer.

***Magnesium has a full 2s orbital rendering extra stability and increasing the IE.***

b***.*** How much energy would be required to remove the first single electron from a magnesium atom?

***1.23 x 10-18 J***

c. What wavelength of light would be required to remove this electron?

***162 nm***

1. Which element (Mg or Al) would be expected to have the “2p” peak in their PES spectra with the highest binding energy? Justify your answer.

***Al, due to the higher nuclear charge for the same amount of shielding.***

1. Would the binding energy of aluminum’s 3p electron be higher or lower than that of magnesium’s 3s electrons? Justify your answer.

***Lower, again Mg’s 3s electrons fill their orbital thereby rendering extra stability and therefore the higher binding energy.***

4. The PES spectrum for sulfur is shown below:

Intensity

1.0 2.05 16.5 22.7 239

Binding Energy (x 10-19 kJ)

1. Which peak corresponds to sulfur’s valence “p” orbital electrons?

***The peak with binding energy of 1.0***

1. Would the same peak in Cl be expected to have a higher or lower binding energy compared to S? Justify your answer.

***Higher due to the higher nuclear charge due to the extra proton chlorine has in it’s nucleus.***

1. Would the same peak in P be expected to have a higher or lower binding energy than that of sulfur? Justify your answer.

***Higher, despite having a smaller nuclear charge, the extra stability of the half-full p orbital gives it the higher binding energy.***

1. Why the binding energy of helium’s 1s electrons is so much lower (2.37) than that of sulfur’s 1s electrons?

***For the same shielding, sulfur has a nuclear charge that is 8x that of helium so therefore a much higher binding energy of those electrons.***

5. Rank the following elements in order of increasing first ionization energy?

a. Na, K, Sc  ***K <Na <Sc***

b. I, Xe, Te ***Te < I < Xe***

**Periodic Law, Electronegativity and Metallic Character**

Classwork

1. Do metallic elements typically have high or low ionization energies? Explain your answer.

***Low, metals elements that lose electrons easily so they require low IE.***

2. Lead and carbon reside in the same period yet lead demonstrates metallic properties while carbon does not. Explain.

***Lead’s valence electrons are far more shielded from the nucleus than carbon’s so they are lost much more easily.***

3. Define electronegativity and provide an explanation as to why the noble gases do not have measurable electronegativity values.

***Attraction for electrons when in a compound. Noble gases do not readily form compounds.***

* + - 1. Draw a graph that shows the general trend in electronegativity when plotted vs. the atomic number for the first 18 elements. Do not use specific electronegativity values but do ensure your points are internally consistent.

Electronegativity

Atomic Number

5. For the following compounds, indicate which element would have the strongest pull on electrons and explain why:

a. SF6 ***F – less shielding***

b. H2O  ***O – much higher nuclear charge***

c. N2O ***O – higher nuclear charge – same shielding***

d. CH3Cl ***Cl – much higher nuclear charge even though the nucleus is shielded more.***

Homework

1. Both bromine and copper are in the same period:

a. Explain why bromine is a non-conductor of electricity while copper is an excellent conductor of electricity.

***Bromine has a much higher nuclear charge for the given amount of shielding so it’s electrons are not lost easily making it a poor conductor.***

* 1. Silver is in the same group as copper. Is it more or less metallic than copper? Explain.

***More, the extra shielding within Ag allows its electrons to be lost more easily.***

2. Explain the general trend in electronegativity as the atomic number increases down a group AND provide a rationale for this trend.

***It decreases due to the extra shielding each element down the group has.***

3. Chlorine has 8 more protons in it’s nucleus than does fluorine. Provide a reason for it’s observed electronegativity (3.2) being lower than that of fluorine (4.0).

***The extra shielding diminishes the nuclear charge felt by chlorine’s valence electrons compared to flourine’s valence electrons.***

1. Draw a graph that shows the trend in metallic character for group 16 when plotted vs. atomic number. Do not use quantitative values for metallic character but make sure your data is internally consistent.

Metallic

Character

Atomic Number

5. Rank the following elements in order of decreasing electronegativity

a. C,Si, S, N ***N > S > C > Si***

b. Ca, Al, Mg, P ***P > Al > Mg > Ca***

**Periodic Table and Specific Groups**

Classwork

1. The alkali metals are more reactive than the alkaline earth metals:

a. Provide an example of an alkali metal and an alkaline earth metal.

***Na and Mg***

b. In terms of ionization energies, provide a rationale for this difference in reactivity.

***For a given period, alkali metals have the smaller nuclear charge and therefore the lower IE so they lose electrons more easily making them more reactive.***

***c.*** Which alkali metal would be expected to be more reactive (K or Na)? Justify your answer.

***K, the higher shielding makes the electrons more readily lost.***

2. An atom forms an ion which combines with the aluminum ion in a 3:1 ratio. Which group must this element belong to on the periodic table?

***Must have -1 charge so the halogens***

3. Using the following clues, identify the element:

a. Halogen with an atomic radii smaller than barium but larger than chlorine. ***Br or I***

b. Has three unpaired electrons in its 4d orbital (two options - list both) ***Nb, Rh***

c. Transition metal with the highest ionization energy. ***Zn***

d. Alkali metal with the smallest atomic radii. ***Li***

Homework

1. The halogens are the most reactive group of non-metals. Propose a reason for this.

***They need gain only 1 electron to have a full valence shell***

2. Group 16 elements are sometimes referred to as the chalcogens.

a. Which element would be most metallic within this group? Justify your answer. ***Po, it has the most shielding so loses electrons easiest***

b. When oxygen is combined with fluorine, which element would have the:

i. Larger ionization energy? Justify your answer.

***F, higher nuclear charge for given amount of shielding***

ii. Smaller atomic radii? Justify your answer

***F, higher nuclear charge, stronger attractions, smaller radii***

3. Using the following clues, identify the following element:

a. Transition metal with the largest atomic radii with a partially filled 4d orbital. Justify your answer. ***Y – least amount of nuclear charge for period***

b. Period 2 element with 2 unpaired electrons (two choices - list both). Justify your answer. ***C and O – carbon is 2p2 and oxygen is 2p4 –***

c. Alkaline earth metal with the lowest 3rd ionization energy. Justify your answer.

***Radium – most shielding of them all –even after losing 2 electrons.***