Unit 1, Lesson 04: Answers to Homework on Periodic Trends

1. Read pages 152 - 157 in your text.
2. An atom of argon has 18 electrons while an atom of sodium has only 11 electrons. However, an atom of argon is SMALLER than an atom than sodium. Explain why this is true.

   Both argon and sodium are in period 3, so they have the same shielding effect. As you move across a period, net nuclear attraction (Z_{eff}) increases which means that the valence electrons experience a stronger attraction to the nucleus. Argon has a NNA of 8+, while sodium has a NNA of only 1+. Argon's stronger NNA pulls its valence electrons in closer to the nucleus, which results in a smaller atom.

3. What three factors influence the magnitude of ionization energy?

   Ionization energy depends on:
   - net nuclear attraction (as NNA increases, ionization energy increases)
   - shielding effect (as shielding effect increases, ionization energy decreases)
   - which electron is being removed (if an electron is being removed from a full or half-full sublevel, then ionization energy will be higher because it is disrupting a stable electron arrangement)

4. Explain why lithium has a lower first ionization energy than fluorine.

   Lithium and fluorine are both in the second period, so they both have the same shielding effect. As you move from left to right across a period, net nuclear attraction increases, so the valence electrons are attracted more strongly to the nucleus and are harder to remove. Because fluorine has a higher NNA than lithium, it requires more energy to remove a valence electron from fluorine which results in fluorine having a higher first ionization energy.

5. Explain why potassium has a lower first ionization energy than lithium.

   Potassium and lithium are both in group I, so they both have the same NNA. As you go down a group on the periodic table, shielding effect increases so the valence electrons are further from nucleus. Because the valence electrons are further from the nucleus, they are less strongly attracted by the nucleus and easier to remove. It takes less energy to remove a valence electron from potassium, so its first ionization energy is lower.

6. As a general rule, positive metal ions are smaller than their neutral metal atoms, while negative non-metal ions are larger than their neutral non-metal atoms. Explain these patterns.

   Positive metal ions have fewer electrons than their parent atom, but the same net nuclear attraction and shielding effect. Because there are fewer electrons, positive ions are smaller than their parent atom.

   Negative non-metal ions have more electrons than their parent atom, but the same net nuclear attraction and shielding effect. Because there are more electrons, negative ions are larger than their parent atom.

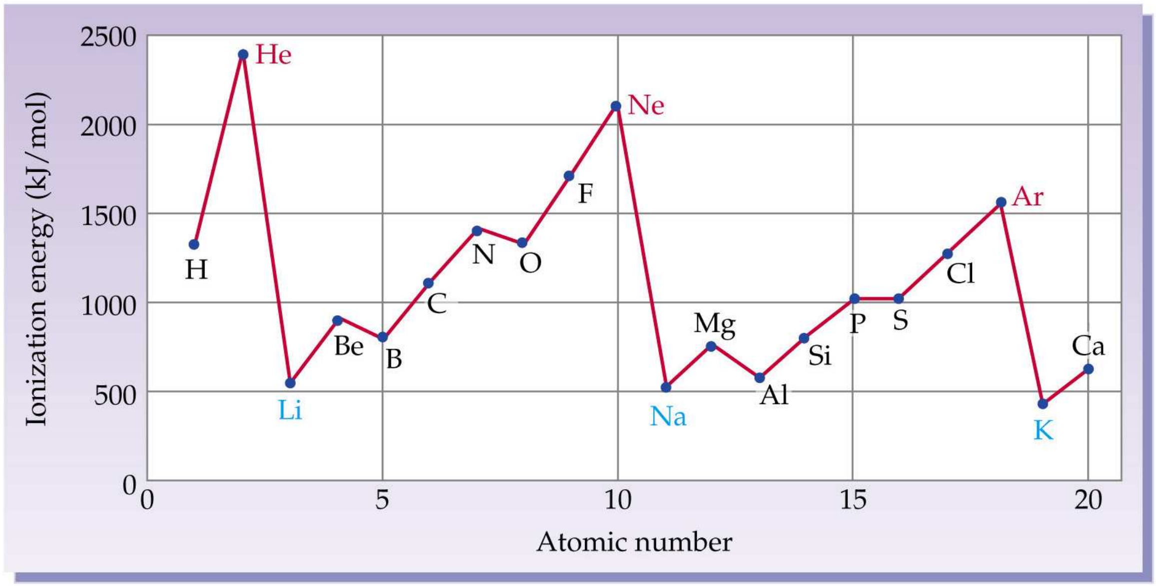
7. Put the following in order of ***increasing*** radius. Explain why you put them in this order:

   a) Mn, Mn^{2+}, Mn^{4+} in order of increasing radius would be: Mn^{4+}, Mn^{2+}, Mn
   - Explanation: the manganese ions and atom have the same number of protons and same net nuclear attraction, but different numbers of electrons. The smaller the number of electrons, the smaller the species so Mn^{4+} (21 electrons), followed by Mn^{2+} (23 electrons) and Mn (25 electrons) is the largest

b) P^{3+} , P^{3-} , P , P^{5+} in order of increasing radius would be: P^{5+}, P^{3+}, P, P^{3-}

- Explanation: all of these ions and atom have the same number of protons and same net nuclear attraction, but different numbers of electrons. The smaller the number of electrons, the smaller the species so P^{5+} (10 electrons), followed by P^{3+} (12 electrons), P (15 electrons) and P^{3-} (18 electrons)

First Ionization Energy vs. Atomic Number for the First 20 Elements



8. Refer to the graph above to answer the following questions:

a) What is significant about the IE_1 for the Noble gas elements? Explain.

The Noble gas elements have the highest IE_1 of any element in the same period. This is because the first ionization for a Noble gas element is removing an electron from a stable octet and must overcome an extremely strong net nuclear attraction of 8+. Both of these factors result in a very high IE_1.

b) What is significant about the IE_1 for the Alkali metal elements? Explain.

The Alkali metal elements have the lowest IE_1 of any element in the same period. This is because the first ionization for an Alkali metal is removing an valence electron to achieve a stable octet and must only overcome a weak net nuclear attraction of 1+. Both of these factors result in a very low IE_1.

c) How do the IE_1 for the Period 2 elements compare to those of Period 3? Explain.

IE_1 for the elements of Period 3 follow the same general pattern as those of Period 2; however, they are slightly lower than their corresponding element in Period 2. This is because the Period 3 elements have a higher shielding effect. Their valence electrons are further from the nucleus and easier to remove which results in lower IE_1 than for the corresponding atom in Period 2.

d) The IE to remove a $2s^2$ electron is higher than the IE to remove a $2p^1$ electron. Explain.

When you remove a s^2 electron, you are removing an electron from a full sub-level and disrupting a stable electron arrangement, which requires slightly more energy. Removing a p^1 electron does not disrupt a full or half-full sub-level, so it does not require as much energy. Also, because of the spherical shape of s orbitals, they essential shield the p electrons from the nucleus and slightly reduce the amount of attraction experienced by the p electrons for the nucleus. This will also result in a lower IE_1 for the p^1 electron.

e) The IE to remove a $2p^3$ electron is higher than the IE to remove a $2p^4$ electron. Explain.

When you remove a p^3 electron, you are removing an electron from a half-full sub-level and disrupting a stable electron arrangement, which requires slightly more energy. Removing a p^4 electron does not disrupt a full or half-full sub-level, so it does not require as much energy. This results in a lower IE_1 for the p^4 electron.

9. The first eight ionization energies for phosphorus are shown in the table below:

ELEMENT	IONIZATION ENERGIES (eV, Electron Volts)							
	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8
Phosphorus	11.0	19.7	30.2	51.4	65.0	220.4	263.3	309.3

a) Write the first six ionization reactions for phosphorus. Include the ionization energy required for each reaction.

$_{15}P \quad 1s^22s^22p^63s^23p^3 \rightarrow {}_{15}P^{1+} \quad 1s^22s^22p^63s^23p^2 + 1\ e^- \qquad IE_1 = 11.0\ eV$

$_{15}P^{1+} \quad 1s^22s^22p^63s^23p^2 \rightarrow {}_{15}P^{2+} \quad 1s^22s^22p^63s^23p^1 + 1\ e^- \qquad IE_2 = 19.7\ eV$

$_{15}P^{2+} \quad 1s^22s^22p^63s^23p^1 \rightarrow {}_{15}P^{3+} \quad 1s^22s^22p^63s^2 + 1\ e^- \qquad IE_3 = 30.2\ eV$

$_{15}P^{3+} \quad 1s^22s^22p^63s^2 \rightarrow {}_{15}P^{4+} \quad 1s^22s^22p^63s^1 + 1\ e^- \qquad IE_4 = 51.4\ eV$

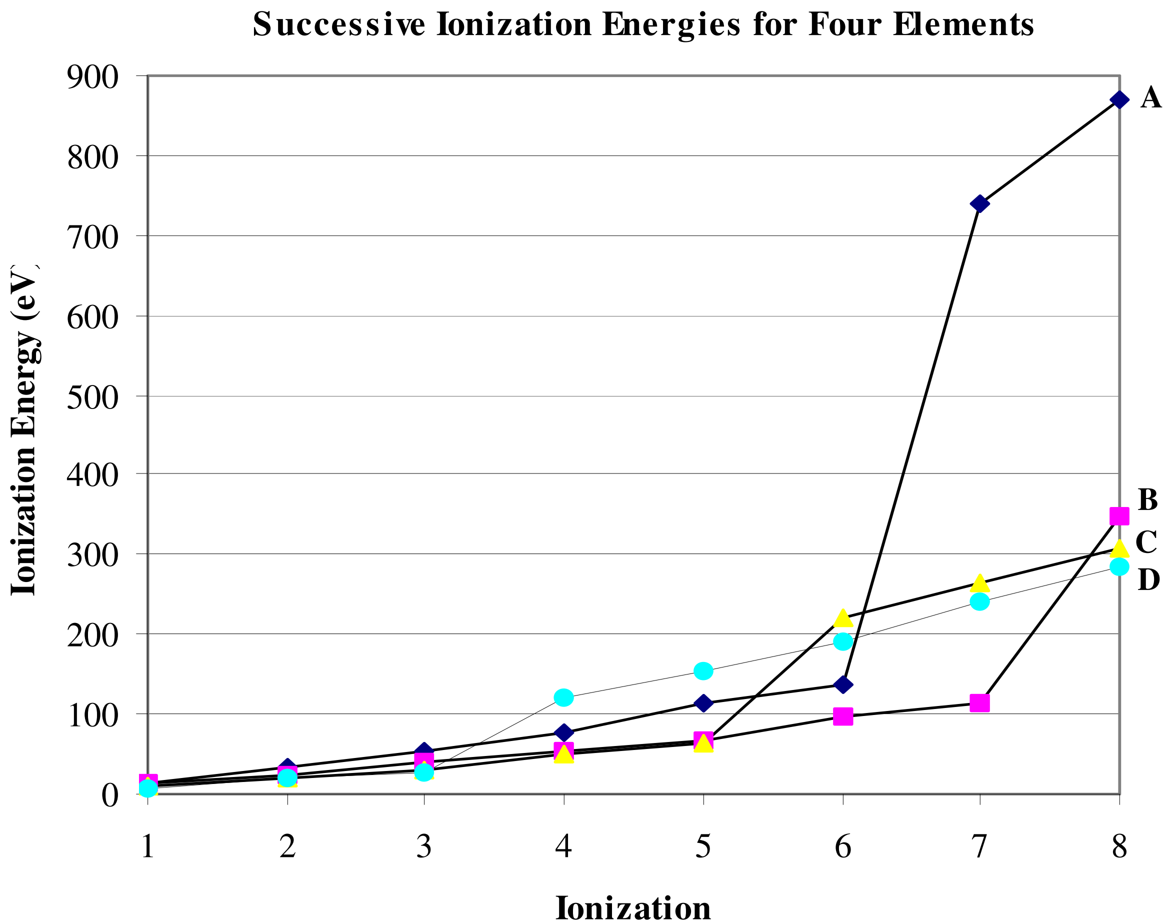
$_{15}P^{4+} \quad 1s^22s^22p^63s^1 \rightarrow {}_{15}P^{5+} \quad 1s^22s^22p^6 + 1\ e^- \qquad IE_5 = 65.0\ eV$

$_{15}P^{5+} \quad 1s^22s^22p^6 \rightarrow {}_{15}P^{6+} \quad 1s^22s^22p^5 + 1\ e^- \qquad IE_6 = 220.4\ eV$

b) Explain why IE_6 is so much higher than IE_5.

For phosphorus, IE_6 is removing the 6^{th} valence electron and disrupting a stable octet electron arrangement, which requires a tremendous amount of energy. IE_5 is not disrupting a full or half-full sub-level, so it requires relatively little energy.

10. The successive ionization energies for four unknown elements are reported below. Prepare a graph showing ionization energy vs. successive ionization. (You can do one single graph or four separate graphs).



10. For each element:

a) Identify which ionization reaction is breaking a stable octet electron arrangement.

Unknown A: IE_7 is breaking a stable octet

Unknown B: IE_8 is breaking a stable octet

Unknown C: IE_6 is breaking a stable octet

Unknown D: IE_4 is breaking a stable octet

b) How many valence electrons does each neutral atom have?

Unknown A has 6 valence electrons (6 electrons come off easily)

Unknown B has 7 valence electrons (7 electrons come off easily)

Unknown C has 5 valence electrons (5 electrons come off easily)

Unknown D has 3 valence electrons (3 electrons come off easily)

c) Identify which group on the Periodic Table each element is found in.

Unknown A has 6 valence electrons so it is in Group VI (16)

Unknown B has 7 valence electrons so it is in Group VII (17)

Unknown C has 5 valence electrons so it is in Group V (15)

Unknown D has 3 valence electrons so it is in Group III (3 or 13)

ELEMENT	IONIZATION ENERGIES (eV, Electron Volts)							
	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8
Unknown A	13.6	35.1	54.9	77.4	113.9	138.1	739.1	871.1
Unknown B	13.0	23.8	39.9	53.5	67.8	96.7	114.3	348.3
Unknown C	11.0	19.7	30.2	51.4	65.0	220.4	263.3	309.3
Unknown D	6.0	18.8	28.4	120.0	153.8	190.4	241.9	285.1

11. Answer questions 3, 4, 5, 6b,c,g,h, 7 and 8 on pages 157- 158.

Question 3 (p. 157): Identify the group number, period number and orbital block of the following elements:
a) [Kr] $5s^2$ $4d^1$: Group 3 (IIIA), period 5, "d" orbital block (element Y, transition metals)
b) [Ar] $4s^2$ $3d^{10}$ $4p^3$: Group 15 (VB), period 4, "p" orbital block (element As)
c) [He] $2s^2$ $2p^6$: Group 18 (VIII), period 2, "p" orbital block (element Ne, Noble gases)
d) [Ne] $3s^2$ $3p^1$: Group 13 (IIIB), period 3, "p" orbital block (element Al)

Question 4 (p. 157): Which requires more energy: removing a valence electron from its atom or removing an electron from an inner energy level? Explain.
Removing an inner electron will require more energy for two reasons:
- an inner electron is closer to the nucleus and has lower shielding effect, so it will be harder to remove
- an inner electron will be part of a full sub-level, which will be very stable electron arrangement so it will require more energy to disrupt

Question 5 (p. 157): On which side of the periodic table would you find an element whose atom is likely to form a cation? Which atomic property is related to this question?
Elements on the left hand side of the periodic table (metals) are more likely to form cations. This is because these elements generally have low ionization energies (will easily lose electrons) and low electron affinity (will not gain electrons).

Question 6b (p. 158):
Electron configurations and orbital diagrams are both representations of the locations of the electrons in an atom. Orbital diagrams include slightly more information, because they show the distribution of the electrons in the various sublevels and include the spin on each electron.

Question 6c (p. 158):
Ionization energy and atomic radius are inversely related and both are determined by shielding effect and net nuclear attraction:
- as shielding effect increases, atomic radius increases and ionization energy decreases
- as net nuclear attraction increases, atomic radius decreases and ionization energy increases

Question 6g (p. 158):
Electron configurations, in general, determine the arrangement of elements on the periodic table because electron arrangement determines the physical and chemical properties of the elements:
- the total number of valence electrons determines which group an element belongs to
- the principal quantum number of the last "s" term determines which period an element belongs to
- all elements whose electron configurations end in s^1 or s^2 are found in the "s" block
- all elements whose electron configurations end in p^1 to p^6 are found in the "p" block
- all elements whose electron configurations end in d^1 to d^{10} are found in the "d" block

- all elements whose electron configurations end in f^1 to f^{14} are found in the "f" block

Question 6h (p. 158):
Electron configurations and quantum numbers are both representations of the locations of electrons in an atom:

- both use "n" to represent the principal quantum number or the distance of the electron from the nucleus
- electron configurations use a letter (s, p, d, f etc) to represent the shape of an orbital while quantum numbers use a number to represent the shape (l = 0, 1, 2, 3 etc, respectively)
- quantum numbers include another two numbers: the magnetic quantum number (m_l) to describe the orientation of the orbital in space and the magnetic spin quantum number (m_s) to indicate the spin of the electron, either + ½ or - ½

Question 7 (p. 158):

Ion	Electron configuration	Condensed e- configuration	Pattern
a) Na^{1+}	$1s^2\ 2s^2\ 2p^6$	[Ne]	isoelectronic with a Noble gas
b) Ca^{2+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	[Ar]	isoelectronic with a Noble gas
c) Cl^{1-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	[Ne] $3s^2\ 3p^6$	isoelectronic with a Noble gas
d) S^{2-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	[Ne] $3s^2\ 3p^6$	isoelectronic with a Noble gas

Careful with question 7. To be completely correct, the condensed electron configuration is for the ***preceding*** Noble gas plus valence electrons, which is why we must use neon as the Noble gas for chlorine and sulfur.

Question 8 (p. 158):
Given IE_1 = 738 kJ/mol, IE_2 = 1451 kJ/mol and IE_3 = 7733 kJ/mol, predict the valence electron configuration for this atom:

- the first two ionization energies are quite low so these must be valence electrons
- there is a huge jump to IE_3 so this ionization is breaking a stable octet
- because there are two valence electrons, the electron configuration for this element will end in s^2

12. Explain the difference between electron affinity and electronegativity.

    Electron affinity measures the amount of energy released or absorbed when a neutral atom gains an electron:

    - if an atom will be more stable with the added electron (eg. a non-metal, but not a Noble gas) then it will release energy when the electron is added an electron affinity will be negative
    - if an atom will be less stable with the added electron (eg. a metal or Noble gas) then it will require energy to make the electron "stick" and electron affinity will be positive

    Electronegativity measures the relative attraction that an atom has for the electrons in a bond:

    - all electronegativity values are zero or greater (positive)
    - if an atom will be more stable with the added electron (eg. a non-metal, but not a Noble gas) then it will have a large electronegativity
    - if an atom will be less stable with the added electron (eg. a metal or Noble gas) then it will have a small electronegativity

13. An atom has a large negative value for electron affinity. What type of atom would this be?

A large negative electron affinity indicates that energy is released when the atom gains an electron so it is more stable with the added electron. This atom would be a non-metal, particularly from the far right of the periodic table, but not a Noble gas.

14. An atom has a large positive value for electron affinity. What type atom would this be?

    A large positive electron affinity indicates that energy must be added to make the electron "stick" to the atom. The atom is more stable without the added electron so it would be a metal, particularly from the far left of the periodic table.

15. Summarize the trends on the Periodic Table for electronegativity, first ionization energy and atom radius in a way that is meaningful to you. Be able to explain each of these trends.

    You must do this question yourself. Know how each of these trends is related to the other:
    - as electronegativity increases, atomic radius decreases
    - as atomic radius increases, first ionization energy decreases
    - as first ionization energy increases, electronegativity increases etc.