1.7 Periodic Trends

Essential knowledge statements from the AP Chemistry CED:

- The organization of the periodic table is based on the recurring properties of the elements and explained by the pattern of electron configurations and the presence of completely or partially filled shells (and subshells) of electrons in atoms.
- Trends in atomic properties within the periodic table (periodicity) can be qualitatively understood through the position of the element in the periodic table, Coulomb's law, the shell model, and the concept of shielding/effective nuclear charge. These properties include the following.
    - atomic and ionic radii
    - ionization energy
    - electronegativity
    - electron affinity
- Periodicity is useful to predict /estimate values of properties in the absence of data.

Coulomb's law describes the force between two charged particles. This equation is useful when studying periodic trends.

$$F_{coulombic} \propto \frac{q_1 q_2}{r^2}$$

When comparing the atoms of two different elements that are located in the same period,

- The valence electrons of each atom are located in the same energy level.
- The element with more protons has a greater nuclear charge, and there is a stronger attraction between the nucleus and the valence electrons.
- According to Coulomb's law, the greater the magnitude of charge, the stronger the attractive force between oppositely charged particles.

When comparing the atoms of two different elements that are located in the same group,

- The valence electrons of each atom are located in different energy levels.
- Electrons located in a higher energy level are farther away from the nucleus.
- Electrons located in a lower energy level are closer to the nucleus.
- According to Coulomb's law, the smaller the distance between oppositely charged particles, the greater the attractive force between them.

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1. Which element, Li or Be, has a smaller atomic radius? Justify your answer in terms of atomic structure and Coulomb's law.

2. Which element, Li or Na, has a smaller atomic radius? Justify your answer in terms of atomic structure and Coulomb's law.

3. Based on your answers to Questions #1 and #2, arrange the atoms Li, Be, and Na in order of increasing atomic radius.

smallest atomic radius	- - - - - - - ->	largest atomic radius

4. The atomic radius of the Na atom is different than the ionic radius of the Na^+ ion.

   (a) Write the complete ground state electron configuration for Na and for Na^+.

   Na _____ Na^+ _____

   (b) Which particle, Na or Na^+, has a larger radius? Justify your answer in terms of atomic structure.

Ion	Ionic Radius (pm)
Fe^{2+}	92
Fe^{3+}	79

5. The ionic radii of two different ions are shown in the table above.

   (a) Write the ground state electron configuration for Fe^{2+} and for Fe^{3+}.

   Fe^{2+} _____ Fe^{3+} _____

5. (b) In terms of atomic structure, explain why the ionic radius of Fe^{2+} is larger than that of Fe^{3+}.

6. The atomic radius of the F atom is different than the ionic radius of the F^- ion.

   (a) Write the complete ground state electron configuration for F and for F^-.

   F _____ F^- _____

   (b) Which particle, F or F^-, has a larger radius? Justify your answer in terms of atomic structure.

K^+	Ca^{2+}	S^{2-}	Cl^-

7. Each of the ions shown in the table above are members of an isoelectronic series. This means that each ion has the same number of electrons.

   (a) Arrange these ions in order of increasing ionic radius.

smallest ionic radius	- - - - - - - ->	- - - - - - - ->	largest ionic radius

   (b) Justify your answer.

Ionization energy is normally expressed in units of kilojoules per mole, and is defined as the energy required to remove one mole of electrons from one mole of gaseous atoms (or ions) in their ground states. Removing the outermost electron from a neutral atom is called the *first ionization energy* (IE_1). Removing the outermost electron from a +1 ion is called the *second ionization energy* (IE_2), etc.

$$Mg(g) \rightarrow Mg^{+}(g) + e^{-} \quad IE_1 = 738 \text{ kJ/mol}$$

$$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-} \quad IE_2 = 1451 \text{ kJ/mol}$$

$$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^{-} \quad IE_3 = 7733 \text{ kJ/mol}$$

8. As you move from left to right across a horizontal row (period) on the periodic table,

   atomic radius values tend to _____ from left to right, and

   first ionization energy values tend to _____ from left to right.

9. As you move from top to bottom down a vertical column (group) on the periodic table,

   atomic radius values tend to _____ from top to bottom, and

   first ionization energy values tend to _____ from top to bottom.

On the AP Exam,

- you will NOT earn credit for simply referring to the relative position of the elements on the periodic table without an explanation.
- you will NOT earn credit for using one trend to explain another trend.

Explain why the first ionization energy value of Mg (738 kJ/mol) is greater than the first ionization energy value of Na (496 kJ/mol).	
Ionization energy increases from left to right across a period. Therefore it requires more energy to remove a valence electron from a Mg atom than it does to remove a valence electron from a Na atom.	Unacceptable response because there is no explanation.
Mg has a smaller atomic radius than Na. Therefore it requires more energy to remove a valence electron from a Mg atom than it does to remove a valence electron from a Na atom.	Unacceptable response because it uses one trend to explain another trend.
The valence electrons in Na and Mg are located in the same energy level (n = 3). Na has 11 protons, and Mg has 12 protons. Since Mg has a greater nuclear charge than Na, there is a stronger attraction between the nucleus and the valence electrons. Therefore it requires more energy to remove a valence electron from a Mg atom than it does to remove a valence electron from a Na atom.	Acceptable response because it uses principles of atomic structure to explain the data.

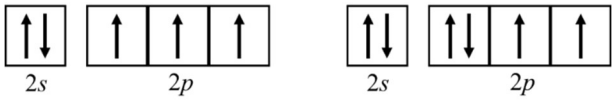
Explain why the first ionization energy value of K (419 kJ/mol) is less than the first ionization energy value of Na (496 kJ/mol).	
Ionization energy decreases from top to bottom down a group. Therefore it requires less energy to remove a valence electron from a K atom than it does to remove a valence electron from a Na atom.	Unacceptable response because there is no explanation.
K has a larger atomic radius than Na. Therefore it requires less energy to remove a valence electron from a K atom than it does to remove a valence electron from a Na atom.	Unacceptable response because it uses one trend to explain another trend.
Na has three occupied energy shells, and K has four occupied energy shells. The valence electron in Na is located in a 3s orbital, whereas the valence electron in K is located in a 4s orbital. Since the valence electron in K is farther away from the nucleus than the valence electron in Na, there is a weaker attraction between the nucleus and the valence electron. Therefore it requires less energy to remove a valence electron from a K atom than it does to remove a valence electron from a Na atom.	Acceptable response because it uses principles of atomic structure to explain the data.

Two Anomalies in the Horizontal Trend for First Ionization Energy

Element	Li	Be	B	C
Electron Configuration	$1s^2 2s^1$	$1s^2 2s^2$	$1s^2 2s^2 2p^1$	$1s^2 2s^2 2p^2$
Ionization Energy (kJ/mol)	520	899	801	1086

Although B has one more proton than Be, the ionization energy of B is slightly less than that of Be. This decrease in ionization energy can be explained as follows. The outermost electron for B is located in the 2*p* subshell, whereas the outermost electron for Be is located in the 2*s* subshell. The 2*p* subshell is slightly higher in energy than the 2*s* subshell. It requires slightly less energy to remove an electron from the 2*p* subshell than it does to remove an electron from the 2*s* subshell.

Element	C	N	O	F
Electron Configuration	$1s^2 2s^2 2p^2$	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^4$	$1s^2 2s^2 2p^5$
Ionization Energy (kJ/mol)	1086	1402	1314	1681



Nitrogen Oxygen

Although O has one more proton than N, the ionization energy of O is slightly less than that of N. This decrease in ionization energy can be explained as follows. There is slightly more electron-electron repulsion between the paired electrons in the p^4 configuration of O as compared to the p^3 configuration of N. This electron repulsion in the p^4 configuration explains why it requires slightly less energy to remove an electron from an atom of O than it does to remove an electron from an atom of N.

Element	1st IE	2nd IE	3rd IE	4th IE	5th IE	6th IE	7th IE
Na	496	4562	6910	9543	13,354	16,613	20,117
Mg	738	1451	7733	10,543	13,630	18,020	21,711
Al	578	1817	2745	11,577	14,842	18,379	23,326
Si	786	1577	3232	4356	16,091	19.805	23,780
P	1012	1907	2914	4964	6274	21,267	25,431
S	1000	2252	3357	4556	7004	8496	27,107
Cl	1251	2298	3822	5159	6542	9362	11,018

10. Consider the data for successive ionization energy in the table above.

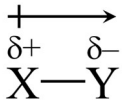
    (a) In terms of atomic structure and Coulomb's law, explain why the ionization energy values increase as successive electrons are removed from an atom.

    (b) In terms of atomic structure and Coulomb's law, explain why the 2nd IE for Na is much higher than the 2nd IE for Mg.

Element	1st IE	2nd IE	3rd IE	4th IE	5th IE
X	1087	2353	4621	6223	37,831

11. Based on the information in the table above, how many valence electrons does element X have? Justify your answer.

Electronegativity is defined as the tendency of an atom to attract electrons to itself in a chemical bond. The higher the electronegativity value is, the greater the attraction for electrons. Electronegativity values are used when determining if a particular chemical bond is classified as nonpolar covalent, polar covalent, or ionic. The greater the difference in electronegativity between two atoms, the more polar the bond is. Suppose that a polar covalent bond is formed between two atoms X and Y as shown below.



If atom X is less electronegative than atom Y, there is a partial positive charge (δ+) on atom X and a partial negative charge (δ–) on atom Y. The arrow above the polar covalent bond represents the dipole, which is generated whenever two electrical charges of opposite sign are separated by a distance. The arrow always points toward the atom that has the higher electronegativity value. The measure of the magnitude of the dipole is called the dipole moment. In general, the greater the difference in electronegativity, the greater the magnitude of the dipole moment.

Electronegativity Values

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5

Notice that the noble gases (He, Ne, Ar, etc.) are not included in the data table above. This is because the atoms of the noble gases ordinarily do not form chemical bonds or share electrons with other atoms.

12. As you move from left to right across a horizontal row (period) on the periodic table,

    electronegativity values tend to _____ from left to right.

    As you move from top to bottom down a vertical column (group) on the periodic table,

    electronegativity values tend to _____ from top to bottom.

13. The smaller the atomic radius is, the _____ the electronegativity value is.

    The larger the atomic radius is, the _____ the electronegativity value is.

    The most electronegative element on the periodic table is _____.

Electron affinity is a periodic trend is a bit confusing to understand. Electron affinity is defined as the energy change that occurs when an electron is added to a gaseous atom to form a negatively charged anion. Consider the following examples.

$$F(g) + e^- \rightarrow F^-(g) \qquad \Delta E = -328 \text{ kJ/mol}$$

$$Li(g) + e^- \rightarrow Li^-(g) \qquad \Delta E = -60 \text{ kJ/mol}$$

If ΔE is negative, energy is released. If ΔE is positive, energy is absorbed. The greater the attraction is between an atom and an added electron, the more negative the value of ΔE is. The more negative the value of ΔE is, the greater the electron affinity is. As you can see in the table below, the trends in electron affinity are not necessarily clear and predictable.

In general, more energy is released when a nonmetal atom gains an electron than when a metal atom gains an electron. For the noble gases, the electron affinity has a positive value. This indicates that the $X^-(g)$ ion is less stable than the $X(g)$ atom.

Electron Affinity (kJ/mol)

H –73							He +48
Li –60	Be +48	B –27	C –122	N +7	O –141	F –328	Ne +116
Na –53	Mg +40	Al –42	Si –134	P –72	S –200	Cl –349	Ar +96
K –48	Ca –2	Ga –29	Ge –119	As –78	Se –195	Br –325	Kr +96

1.8 Valence Electrons and Ionic Compounds

Essential knowledge statements from the AP Chemistry CED:

- The likelihood that two elements will form a chemical bond is determined by the interactions between the valence electrons and nuclei of elements.
- Elements in the same column of the periodic table tend to form analogous compounds.
- Typical charges of atoms in ionic compounds are governed by their location on the periodic table and the number of valence electrons.

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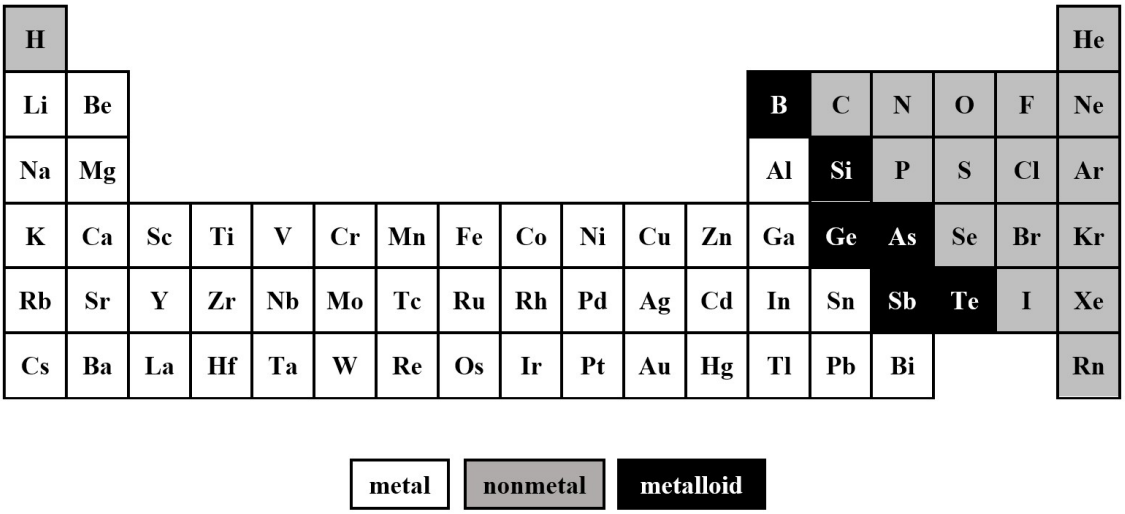
14. Write the correct number of valence electrons for each of the following elements.

Element	Li	Be	B	C	N	O	F	Ne
Valence Electrons								

15. Write the correct charge (e.g., 1+, 2+, 1–, 2–, etc.) that each of the following elements has when it forms a stable monoatomic ion.

Element	Li	Be	B	C	N	O	F	Ne
Charge				N/A				



Binary ionic compounds (e.g., NaCl) normally consist of a metal and a nonmetal. The chemical formula of a binary ionic compound can be determined be examining the charges on each ion. The formula is written as an empirical formula, and should have an overall charge of zero.

16. Write the correct chemical formula for the binary ionic compound that is formed from the combination of each of the following pairs of elements.

Elements	Chemical Formula of the Binary Ionic Compound
Li and F	
Na and S	
Mg and Cl	
Al and O	
Ca and P	

Elements in the same group (column) of the periodic table have the same number of valence electrons. This explains why elements in the same group tend to form analogous compounds.