Trends in the Periodic Table

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Effective nuclear charge:

- effective nuclear charge is the attraction felt by the valence electrons from the nucleus

- increases across a period :
  increases across because the number of protons is increasing in the nucleus, while the electrons being added are the same distance away

- increases down a group:
  increases down because more protons are being added.
  However the core electrons shield the valence electrons from the effect of the nucleus, hence decreasing effective nuclear charge down the group.

Atomic Radius:

- refers to the distance from the center of the nucleus to the outer most shell.

- decreases across a period :
  decreases across because positive charge, (i.e. Effective nuclear charge) is increasing in the nucleus; since # of protons increases, electrostatic attraction of the nucleus for electrons increases;
  Since the added electrons are the same distance away, so the electrons are pulled closer towards the nucleus

- increases down a group:
  Atomic radius increases going down a group because more energy levels are being added

Practice Problems

1. Of silicon, Si, magnesium, Mg, or sulphur, S, which has the largest atomic radius? Why?

2. Of the elements aluminium , Al, sodium, Na, or chlorine, Cl, which has one has the largest atomic radius?

3. Of the elements lead, Pb, barium, Ba, or astatine, At, which one has the smallest atomic radius?

4. Which atom in the pair, silicon, Si, and lead, Pb, has the larger atomic radius? Why?

5. Which atom in the pair phosphorus, P, and antimony, Sb, has the larger atomic radius?

6. Which of the following elements has the smallest atomic radius? S, Se, O, Te
7. The following is a block of elements from the periodic table, use it to answer the following questions:

A	B	C	D
E	F	G	H
I	J	K	L

Which element has the (i) largest atomic radius?
(ii) smallest atomic radius?

Ionic radii: Size of cations and anions:

- a **cation (+)** has a smaller ionic radius than a neutral atom of the same element, (i.e. Cations are smaller than the parent atom):

  because the nuclear charge which pulls in electrons is greater, making the radius smaller.
  In addition, an energy level may have been removed in forming the cation.

- an **anion (-)** is always greater than a neutral atom of the same element, (i.e. Anions are larger than the parent atom) :

  because when you add an electron, there is more repulsion in the electron cloud and it spreads out, increasing the ionic radius.

- Compare electron– proton attraction and electron– electron repulsion, example: Na^{+} and F^{-1}

Practice Problems

Which ion in each of the following pairs has the larger ionic radius?

a. Na, Na^{+1} b. Cl, Cl^{-1} c. Al, Al^{+3} d. O, O^{-2} e. Mg, Mg^{+2} f. O^{-1}, O^{-2} g. P^{-2}, P^{-3}

Ionization Energy:

- Defined as the energy required to remove an electron from a gaseous atom in its ground state to from a gaseous cation.

- $X_{(g)} + \Delta E \longrightarrow X^{+1}_{(g)} + e^{-}$ ΔE = Ionization Energy

- The value of the first ionization energy depends upon:

  (1) the effective nuclear charge,
  (2) the distance between the electron and the nucleus, and
  (3) the 'shielding' produced by lower energy levels.

- generally increases across a period:

  increases across a period due to the increasing nuclear charge and decreasing atomic radius; thus stronger attraction between the electrons and the nucleus and hence more energy is required to remove the electrons.

- generally decreases down a group:

  decreases down a group because effective nuclear charge is decreasing and because the atomic radius is increasing; thus weaker hold on the electrons and hence less energy is required to remove the electrons.

<u>Metallic character: decreases across a period, increases down a group</u>

Practice Problems

1. Of the elements calcium, Ca, beryllium, Be, or magnesium, Mg, which has the highest ionization energy? Why?

2. Of the elements boron, B, aluminium, Al, or gallium, Ga, which one has the highest ionization energy?

3. Of the elements argon, Ar, krypton, Kr, or xenon, Xe, which has the highest ionization energy?

4. Which of the following elements has the smallest ionization energy? Sr, Ca, Ba, Mg

5. Which of the following elements has the largest ionization energy? Br, K, As, Ca

6. Elements X, Y, and Z are found in the same group of the periodic table, with X on top and Z on the bottom. Which element will have
(a) the largest atomic radius
(b) the largest ionization energy
(c) the most metallic character?

Electron Affinity:

- The electron affinity is a measure of the attraction an atom has for other electrons apart from its own.

- Defined as the energy released when an electron is added to a gaseous atom in its ground state to form a gaseous anion

- Electron affinity is a useful measure of electron attraction by atoms.

- electron affinity is the energy change when an electrons is added to an atom in the gas phase:

- $X_{(g)} + e^- \longrightarrow X^{-1}_{(g)} + \Delta E$ ΔE = electron affinity

- increases across a period:

  increases across because elements go from metals to non-metals.
  Non-metals want electrons to bring them closer to their stable octet.
  Atoms release energy because as atoms become more stable they have less energy.

- decreases down a group;

  decreases down because metallic character is increasing, thus addition of an electron will 'destabilize the metallic atom', therefore the atom is not likely to release energy, (energy is released so atom will be more stable), hence the addition of an electron to an atom going down a group is not likely to be an exothermic process.

- **To explain why, in general, ionization energy and electron affinity follow the same trends throughout the periodic table.**

<u>Answer</u>

Ionization energy is the energy required to remove an electron from an atom.
Electron affinity is the energy released when an atom accepts an electron.

They are both dependent on the same thing, atomic radius.
When the atomic radius is small, it is more difficult to remove an electron, so ionization energy is high.

At the same time, it is easier for the atom to accept another electron, so it releases energy when an electron is added and electron affinity is high.

Practice Problems

1. Which element should have the higher electron affinity, selenium, Se or tellurium, Te? Why?

2. Which element should have the higher electron affinity, arsenic, As, or antimony, Sb? Why?

3. Which of the following elements has the highest electron affinity? Cl, Si, Na, P

4. Which element in the following sets has the largest electron affinity, and why?
(a) B, Li, or F (b) Br, Cl, or I (c) Si, S, Se (d) P, S, N, O

Electronegativity:

- This is the relative measure of the attraction an atom has for electrons in a covalent bond between it and another atom. (Recall: covalent bond = non metal + non metal)

- Linus Pauling assigned values in the range of 0 - 4

Li	Be	B	C	N	O	F	Ne
1.0	1.5	2	2.5	3	3.5	4	

  - Note: noble gases are assigned a value of zero. (WHY?)

  - Use of the Pauling scale to predict the type of bonding in a compound i.e. polarity of a bond:

  - if the difference between electronegativities is:
    - less than 0.7 = purely covalent bond
    - 0.7 – 1.7 = polar covalent bond
    - Greater than 1.7 = ionic bond

- Electronegativity increases across a period and decreases down a group of the periodic table.

- The greater the electronegativity difference between the two atoms involved in a polar covalent bond , the more polar the bond will be.

Practice Problems

1. Predict, and explain, the nature of the bond formed, (i.e. pure covalent, polar covalent or ionic), between the following pairs of elements:

a. Li + Br b. Cl + Cl c. P + Cl d. Mg + S

e. Ga + S f. H + O g. Sr + Cl h. Ca + N

2. Which of the bonds in each of the following is more polar:
a. Br – Br or C– Br b. N – F or N – Cl c. H – Cl or Br — Cl

d. Si – Cl or P – Cl e. O – F or S – O f. Se – Cl or H – Se

Assignment

Multiple Choice:

1. Consider the equation $X_{(g)} + \text{energy} \rightarrow X^{+} + e^{-}$. The "energy" term in the equation represents

a. electron affinity
b. heat of sublimation
c. ionization energy
d. sublimation energy
e. heat of vaporization

2. Which of the following elements requires the least amount of energy to remove an electron from an atom to form an ion?

a. O
b. He
c. K
d. H
e. Fr

3. Why does atomic radius increase from top to bottom in a chemical family?

a. Nuclear charge increases from top to bottom in a chemical family.
b. Nuclear charge decreases from top to bottom in a chemical family.
c. The number of energy levels increases from top to bottom in a chemical family.
d. The number of energy levels decreases from top to bottom in a chemical family.
e. The number of electrons decreases from top to bottom in a chemical family.

4. Why does ionization energy increase from left to right in a period?

a. Nuclear charge increases from left to right in a period.
b. Nuclear charge decreases from left to right in a period.
c. The number of energy levels increases from left to right in a period.
d. The number of energy levels decreases from left to right in a period.
e. The number of electrons decreases from left to right in a period.

5. Why is it easier to remove an electron from potassium than it is to remove an electron from calcium?

a. Potassium has a higher electron affinity.
b. Potassium has a higher ionization energy.
c. Potassium has a lower nuclear charge.
d. Calcium has a lower electron affinity.
e. Calcium is a metal, but potassium is not.

6. Consider the equation $X_{(g)} + e^{-} \rightarrow X^{-} + \text{energy}$. The "energy" term in the equation represents

a. electronegativity
b. electron affinity
c. ionization energy
d. nuclear charge
e. none of the above

7. Elements A, B, C, and D (found in Groups 1–17) have atomic radii of 265 μm, 160 μm, 185 μm, and 175 μm, respectively. Which element will most likely have the highest ionization energy?

a. A
b. B
c. C
d. D
e. not enough information

8. Which one of the following has pure covalent bonding:

a. F – F b. H– F c. P – F d . Na–F