AP CHEMISTRY

TOPIC 4: ATOMIC STRUCTURE & THE PERIODIC TABLE, PART B

Day 47:

- Isoelectronic
- Thomson's Experiments
- Rutherford's Experiment
- Heisenberg Uncertainty Principle
- Periodic Trends
- Atomic Radius
- Ionization Energy
- Electron Affinity
- Electronegativity

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1) Use your knowledge of the periodic table of the elements to answer the following questions.
   a) Explain the trend in electronegativity from P to S to Cl.

Electronegativity is the pull from the nucleus of one atom on the electrons of other atoms. The trend is that as one moves from left to right along a period (row), the electronegativity increases. Therefore, from Phosphorus to Sulfur to Chlorine the electronegativity increases.

Electronegativity is a measure of the ability of an atom or molecule to attract electrons from another atom (in the context of a chemical bond). The type of bond formed is largely determined by the difference in electronegativity between the atoms involved. Atoms with similar electronegativities will share an electron with each other and form a covalent bond. However, if the difference is too great, the electron will be permanently transferred to one atom to the other atom and an ionic bond will form. Furthermore, in a covalent bond, if one atom pulls slightly harder than the other, a polar covalent bond will form.

   b) Explain the trend in electronegativity from Cl to Br to I.

The trend is that as one moves from the top toward the bottom along a group (column), the electronegativity decreases. Therefore, from Chlorine to Bromine to Iodine the electronegativity decreases.

Each element has a characteristic electronegativity ranging from 0 to 4 on the Pauling scale. The most strongly electronegative element, fluorine, has an electronegativity of 3.98 while weakly electronegative elements, such as lithium, have values close to 1. The least electronegative element is Francium at 0.7. In general, the degree of electronegativity decreases down each group and increases across the periods.

Across a period, non-metals tend to gain electrons and metals tend to lose them due to the atom striving to achieve a stable octet (full outer-most energy-level "s" and "p" orbitals). Down a group, the nuclear charge has less effect on the outermost shells. Therefore, the most electronegative atoms can be found in the upper, right hand side of the periodic table, and the least electronegative elements can be found at the bottom left.

   c) Explain the trend in atomic radius from Li to Na to K.

The trend is that as one moves from the top toward the bottom along a group (column), the atomic radius (for neutral atoms) increases. Since energy levels (orbitals) are filled with electrons, therefore causing the radius to increase. The Effective nuclear charge is reduced as more orbitals are filled. The effective nuclear charge is the net positive charge experienced by an electron in a multi-electron atom. The term "effective" is used because the shielding effect of negative electrons prevents higher orbital electrons from experiencing the full nuclear charge which does allow the outer electrons to be "pulled" as close to the nucleus as electrons that are closer to the nucleus.

d) Explain the trend in atomic radius from Al to Mg to Na.

The trend is that as one moves from left to right along a period (row), that atomic radius (neutral atoms) decreases. However, this question is asking us to move right to left and when this happens the neutral radius increases.

2) Write the electron configuration (long form, using “arrows”) for neutral silver in its ground state? Is it diamagnetic or paramagnetic?

1s	2s	2p			3s	3p			4s	3d					4p			5s	4d				
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑↓	↑↓	↑↓	↑↓	↑↓

Did you learn “special stabilities” in general chemistry ??? Now would be a good time to learn them…

Paramagnetic, the “4d” orbital does NOT have all orbitals paired

3) Explain each of the following in terms of atomic and molecular structures and / or forces.

a) The first ionization energy for magnesium is greater than the first ionization energy for calcium.

Magnesium is in group 2A, period 3. Calcium is in group 2A, period 4. Due to calcium residing in a higher energy level (period 4), the Effective Nuclear Charge (protons pull on the electrons around the nucleus) for calcium is not as high (strong) as it is for magnesium. In other words, the nuclear pull on calcium’s own valence electrons are not as strong as the nuclear pull on magnesium’s valence electrons. Since calcium’s effective nuclear charge is not as strong, it is “easier” to remove calcium’s valence electrons.

b) The first and second ionization energies for calcium are comparable, but the third ionization energy is much greater.

The first and second ionization energies for calcium are comparable (low), but the third ionization energy is much greater. This is because removing two electrons (the valence) electrons from calcium is “easy” while trying to remove CORE electrons are difficult. When an atom’s “s” and “p” orbitals are filled with electrons, the atom no longer will “easily” allow electrons to be removed.

4) Write the electron configuration (notation form, using “exponents”) for barium (with a +2 charge) in its ground state? Is it diamagnetic or paramagnetic? Also, what is barium (with a +2 charge) isoelectronic with?

$$Ba{:}\ \ 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2\, 3d^{10}\, 4p^6\, 5s^2\, 4d^{10}\, 5p^6\, 6s^2$$

$$Ba^{+2}:\ \ 1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^6\, 4s^2\, 3d^{10}\, 4p^6\, 5s^2\, 4d^{10}\, 5p^6\, 6s^0$$

Diamagnetic (for BOTH neutral and +2), the orbital does have all its orbitals paired.

Isoelectronic with Xenon