The Do's and Don'ts of Periodic Trends

Name________________________Per____

The AP® Chemistry test typically addresses periodic trends in the free response questions. A statement such as "In terms of atomic structure, explain why the first ionization energy of selenium is less than that of bromine." Often these questions are not well answered because students simply state the trend, such as "ionization energy increases as you move left in a period," and fail to address the forces at work that created the trend in the first place. You must address the *forces* that create the trend and avoid the impulse to simply use the location on the table as an explanation of the observed property. If needed, chant over and over again, "A trend is an observation, not an *explanation*!" It is fine to state the trend in your answer, but you must also go further by explaining what causes the observed trend.

In addition to discussing forces, you must mention BOTH of the atoms or ions that are in the question when stating your answer. Addressing one and leaving the other as implied does not usually earn the point for this type of question.

Almost all of the properties that are asked about in exam questions rely on the Coulombic (electrostatic) attraction between the outer electrons and the nucleus. Answers to these questions should always include a statement about how this attraction is affected. The concept of Coulomb's Law is important and you should mention it when appropriate in your explanations.

> *Coulomb's Law shows that the force of attraction between two oppositely charged particles is directly proportional to the magnitude of the charges and inversely proportional to the distance between those charges.*

$$F_{attraction} \ \alpha \ \frac{(Q^+)(Q^-)}{d^2} \quad \text{or} \quad F = k\frac{(Q^+)(Q^-)}{d^2}$$

Chemistry is often defined as "the study of matter and energy". Emphasize the *energy* of attractions and repulsions in your explanations.

Trends in the Periodic Table

Justifying all of the trends on the periodic table can be simplified using these two generalizations:

1) Use Z_{eff} to justify trends across a period.
2) Use increased distance (greater value of n) to justify trends down a group.

Atomic radius refers to the distance between the nucleus and the outer edge of the electron cloud. It is influenced by the nuclear pull and the number of energy levels.

Atomic radii decrease as atomic numbers increase in any given period	
DO	**DON'T**
Greater effective nuclear charge, Z_{eff}, increases the attractive force of the nucleus and therefore pulls the electron cloud closer to the nucleus resulting in a smaller atomic radius.	Don't simply stating that atomic radii decrease from left to right across a period. You will not earn points.

Atomic radii increase as atomic number increases down a column or group	
DO	**DON'T**
Increased number of energy levels (n) increases the distance over which the nucleus must attract and therefore reduces the attraction for electrons.	You must not simply saying that radii increase down a column. You will not earn points.
Full energy levels provide shielding between the nucleus and valence electrons, thus within a column, the effective nuclear charge, Z_{eff}, is somewhat constant.	Don't use shielding for explanations *across* a period. Only full energy levels, not full sublevels, are of concern in a shielding argument.

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Ionization energy refers to the energy needed to remove an electron from a *gaseous* atom or ion, i.e. an isolated one, not part of a solid, liquid or a molecule. It is *always* endothermic.

Ionization energy increases as atomic number increases in any given period	
DO	**DON'T**
Greater effective nuclear charge, Z_{eff}, increases the attractive force of the nucleus and therefore holds the electrons more tightly.	Don't simply stating that ionization energy increases from left to right across a period. You will not earn points
Exceptions occur between groups II and III and V and VI.	Don't think that the trends are without anomalies.
1) A drop in IE occurs between groups II and III because the *p* electrons do not penetrate the nuclear region as greatly as *s* electrons do and are therefore not as tightly held.	1) Don't state that *p* electrons are farther away from the nucleus. You will not earn points.
2) A drop in IE occurs between groups V and VI because the increased repulsion created by the first pairing of electrons in the p-orbitals outweighs the increase in Z_{eff} and thus less energy is required to remove the electron.	2) Don't state that the atoms in group V are more stable because they have a half filled sublevel. This is wrong, wrong, wrong! You will not earn points.

Ionization energy decreases as atomic number increases down a column or group	
DO	**DON'T**
Increased number of energy levels (*n*) increases the distance over which the nucleus must pull and therefore reduces the Coulombic (electrostatic) attraction for electrons.	Don't simply saying that IE decreases down a column. You will not earn points.
Full energy levels provide shielding between the nucleus and valence electrons, thus within a column, the effective nuclear charge, Z_{eff}, is somewhat constant.	Don't use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument. You will not earn points.

Electron affinity is NOT the opposite of ionization energy, but involves the addition of an electron to a gaseous atom, which can be exothermic or endothermic. An exothermic value indicates that the resulting negative ion is more stable than the original atom. Thus a higher negative value indicates that the nuclear force is stronger and thus attracts the electron more effectively. Positive electron affinities indicate that the nucleus is not as effective at attracting the electron, and it must be forced into the atom.

Electronegativity is an assigned property which indicates the attraction of an atom for the *pair* of outer shell electrons in a covalent bond with another atom. Electronegativity patterns are the same as electron affinity patterns for the same reasons. Both of these properties focus on the attraction that the nucleus has for electrons.

Electronegativity increases as atomic numbers increase in any given period	
DO	**DON'T**
Greater effective nuclear charge, Z_{eff}, increases the attractive force of the nucleus and therefore it strengthens the attraction for the electrons.	Don't simply state that electronegativity increases from left to right across a period. You will not earn points.

Electronegativity decreases as atomic number increases down a column or group	
DO	**DON'T**
Increased number of energy levels (n) increases the distance over which the nucleus must pull and therefore reduces the attraction for electrons.	Don't simply saying that electronegativity decreases down a column. You will not earn points.
Full energy levels provide shielding between the nucleus and valence electrons, thus within a column, the effective nuclear charge, Z_{eff}, is somewhat constant.	Don't use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument. You will not earn points.

Ionic radius is the distance from the nucleus to the outer edge of the electron cloud in a charged ion. The same radii trends apply once you divide the table into the metal and non-metal sections. Within the metal section the positive ionic radii decrease from left to right with only minor changes in the transition metals. Once you get to the nonmetal section and the ions are now negative and larger they will again decrease in radii from left to right. Ionic radii increase going down all columns because of the additional energy levels present (n).

Positive ions are smaller than their respective neutral atoms	
DO	**DON'T**
Positive metal ions result from the loss of valence electrons. In many cases this means the outermost electrons are now in a smaller principal energy level (n) and are thus much closer than the electrons in original neutral atom.	Don't stop simply saying that the positive ion is smaller because it lost electrons. The mention of energy levels (n) is essential to earning the point on this type of question.
If the entire set of valence electrons are not removed, there will decreased electron/electron repulsions between the remaining electrons allowing the electron cloud to contract.	Don't neglect this important effect. This is especially useful when comparing ionic radii that do not involve a complete loss of a valence energy level.

Negative ions are larger than their respective neutral atoms	
DO	**DON'T**
Negative nonmetal ions result from the addition of valence electrons. Increased electron/electron repulsions also cause the electron cloud to expand.	Don't say that the ion is bigger simply because it has more electrons. Electron repulsions are a powerful force within the atom.

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Reactivity depends on whether the element reacts by losing electrons (metals) or gaining electrons (non-metals).

Metals are more reactive as you move down a column	
DO	**DON'T**
Because metals react by losing electrons, a loosely held electron will result in a more reactive metal. This is directly tied to ionization energy. With an increased number of energy levels (*n*) comes increased distance from the nuclear attraction and thus a more loosely held electron available for reacting.	Don't simply say that metals are more reactive at the bottom left corner of the table. You will not earn points.

Non-metals are more reactive as you move up a column	
DO	**DON'T**
Because nonmetals tend to gain electrons, a strong nuclear attraction will result in a more reactive non-metal. This means that an atom with the highest Z_{eff} and the least number of energy levels should be the most reactive nonmetal (F) because its nucleus exerts the strongest pull.	Don't simply say that nonmetals are most reactive at the top right corner of the table. You will not earn points.

Final thoughts

It is so important to recognize the difference between the two species given. Follow the following steps EVERY time you answer a periodicity question and you will make good scores.

- Locate *both* elements on the periodic table and state the principal energy level (*n*) and the sub-level containing the valence electrons for *each* element.
- Do they have the same or different *n* values?
  - If same *n*: justify with Z_{eff} ;
  - If different *n*: justify with *n* vs. *n* distances.