The Do's and Don'ts of Periodic Trends

Name________________________Per____

Periodic trend questions require you to know the **why** behind the trend. A question might go like this, "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 is not a complete answer.

Almost all of the properties that are asked rely on the Coulombic (electrostatic, i.e. + / −) attraction (or − / − repulsions) 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 states that the force of attraction between two oppositely charged particles (or repulsion between two same charge particles) is*
>
> ***directly proportional to the magnitude of the charges** and*
>
> ***inversely proportional to the distance** between those charges.*

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

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 effective nuclear charge, *enc* to justify trends **across a period**.
2) Use increased distance (*energy level*) to justify trends **down a column**.

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

Atomic radii decrease as atomic numbers increase across a given period.	
DO	**DON'T**
More protons cause greater effective nuclear charge, *enc*, increasing the attractive force of the nucleus and therefore pulls the outer electrons 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. This would not be complete answer. Do not simply state "more protons" with out commenting that the protons are acting on electrons in the *same* energy level.

Atomic radii increase as atomic number increases down a column or group	
DO	**DON'T**
Increased number of **occupied energy levels** increases the distance between the nucleus and outer electrons.	You should not simply say that radii increase down a column. This would not be complete answer.
Full energy levels (inner core electrons) provide shielding between the nucleus and valence electrons, thus within a column, the effective nuclear charge, *enc*, is nearly constant in a column	Do NOT use the number of protons as an argument since the *enc* is relatively constant in a column, and the valence electrons are not feeling more attractive force.

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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. Ionization energy is *always* endothermic (energy IN).

$X + I.E. \rightarrow X^{+} + e^{-}$

Ionization energy increases as atomic number increases across a given period.	
DO	**DON'T**
More protons cause greater effective nuclear charge, *enc*, increasing the attractive force of the nucleus and therefore holds the valence electrons more tightly, making the ionization energy higher. More protons are acting on electrons in the **same** valence orbital. Be sure and reference Coulomb's Law.	Don't simply state that ionization energy increases from left to right across a period. This would not be complete explanation. Don't simply say "more protons." (It is important to emphasize that the more protons are acting on electrons in the same valence orbital.) Do not argue more electrons, that would be a counter claim.
Exceptions occur for groups III and VI.	Don't think that the trends are without anomalies. Continuing across the period, the increase in *enc* outweighs the cause of the exception and the trend continues.
A drop in IE occurs for group III because the *p* electrons do not penetrate as close to the nucleus as *s* electrons do, and thus *p* electron experiences less nuclear force outweighing the increase in *enc* and are therefore not held as tightly.	Don't state that the *p* orbital is farther away from the nucleus than the *s* orbital. It is not.
A drop in IE occurs for group VI because the increased repulsion created by the first pairing of electrons in the *p*-orbitals outweighs the increase in *enc* and thus less energy is required to remove the electron.	Don't state that the atoms in group V are more stable because they have a half filled sublevel.

Ionization energy decreases as atomic number increases down a column or group.	
DO	**DON'T**
Increased number of occupied energy levels increases the **distance** over which the nucleus must act and therefore reducing the distance over which the *ENC* must act. Be sure and reference Coulomb's Law.	Don't simply saying that IE decreases down a column. This would not be complete explanation.
Full energy levels (core electrons) block some of the nuclear force between the nucleus and valence electrons, thus within a column, the effective nuclear charge, *ENC*, is somewhat constant.	Don't use use ENC as a justification. The ENC is constant in a column. This would not be complete explanation

Electron affinity is NOT the opposite of ionization energy, but involves the addition of an electron to a gaseous *atom not ion*, 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.

$X + e^{-} \rightarrow X^{-}$ (note: The energy was not put into the equation because it can be exothermic or endothermic.)

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Successive Ionization energy refers to the energy needed to remove the "next" electron from a *gaseous* ion, meaning the energy required to remove the 2nd, 3rd, 4th, and more electrons one at a time. Successive Ionization energy *always* increases, but there are differences in the amount of each increase.

$X + \text{more I.E.} \rightarrow X^{2+} + e^-$

There will be one ionization that will be MUCH greater increase over the previous	
DO	**DON'T**
Removing the electron that is one more than the number of valence electrons will be in an entire **energy level closer** to the nucleus and and that electron will be acted upon by a greater *ENC*. Do think about the number of valence electrons.	Don't simply that that the number of protons is more than the number of electrons, because the number of protons is not changing, and the justification must describe a change in force. Do not argue that the increase is due to the difficulty of breaking up a full energy level.
Removing each successive valence electron will require a **slightly** larger amount of energy due to the previous removal of electrons causing less electron-electron repulsion and allowing the valence electrons to skootch close to the nucleus and thus be attracted by a slightly greater force due to the slight decrease in distance between the nucleus and the outermost electrons.	Do not concern yourself with an electron proton ratio.

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 more loosely held electron will result in a more reactive metal. This is directly tied to the same reasons for ionization energy trends. With an increased number of energy levels comes increased **distance** from the nuclear attraction making the electron more loosely held and more 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 for your justification.

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 *ENC* and the least number of energy levels should be the most reactive nonmetal (F) because the nucleus has the greatest *ENC* working over the shortest **distance**.	Don't simply say that nonmetals are most reactive at the top right corner of the table. This is an incomplete justification.

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Ionic radius is the distance from the nucleus to the outer edge of the electron cloud in a charged ion. Ionic radii increase going down all columns for the same reason that atoms increase; because of the additional energy levels present.

Positive ions are smaller than their "parent" neutral atoms.	
DO	**DON'T**
Positive metal ions result from the loss of **all** the valence electrons. This means the outermost electrons in a cation are now in an energy level one level closer to the nucleus than the electrons in the original neutral atom.	Don't stop simply saying that the positive ion is smaller because the atom lost electrons. (The mention of energy levels is essential to providing a complete explanation.)
If the entire set of valence electrons are not removed, there will decreased electron/electron repulsions between the remaining valence electrons allowing the valence 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 "parent" neutral atoms.	
DO	**DON'T**
Negative nonmetal ions result from the addition of valence electrons. Increased electron/electron **repulsions** cause the electron cloud to expand.	Don't say that the ion is bigger simply because it has more electrons.

Isoelectronic ions are different sizes due to the number of protons.	
DO	**DON'T**
State that the number of electrons is the same, but the number of protons is different. An increased number of protons will draw the electron cloud closer causing a smaller ion.	Don't say that a negative isoelectronic ion has more repulsion, the electrons are the same, so the repulsion is the same. The issue is the number of protons.

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.

- Make a claim.
- State the relevant evidence.
  Locate *both* elements on the periodic table and state what is the same and what is different
  - State the energy level and the sub-level containing the valence electrons for *each* element.
  - State the number of protons. State the effective nuclear charge (*ENC*)
  - Don't be afraid to state the obvious.
- Justify with the appropriate reasoning
  Do the atoms have the same or different number of energy levels?
  - If same number of energy levels (same row in the periodic table): justify with *ENC*
  - If different number of energy levels (different row in the periodic table): justify with *the difference in* distances between outmost electrons and the nucleus.
- **AVOID the word "it" at all costs. Never, never use the word "it" in your arguments. Use the word for whatever you are referring to, even if the sentence seems awkward.**