The Do's and Don'ts of Teaching Periodic Trends

The AP® Chemistry test typically addresses periodic trends in the free response questions. Students are often given 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. Your first order of business when teaching this unit is to emphasize that students must address the forces that create the trend and avoid the impulse to 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 their answer, but they must also go further by explaining what causes the observed trend.

In addition to discussing forces, train your students to mention BOTH of the atoms or ions in the question when stating their answer. Addressing one and leaving the other 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 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 should be taught and students should be encouraged to mention it in their 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.

$$F_{attraction} \propto \frac{(q^+)(q^-)}{d^2}$$

Chemistry is often defined as "the study of matter and energy". Emphasize the *energy* of attractions and repulsions throughout this unit of study.

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**
Teach students that the effective nuclear charge, Z_{eff}, increases the attraction of the nucleus and therefore pulls the electron cloud closer to the nucleus resulting in a smaller atomic radius.	Don't' let students get away with simply stating that atomic radii decrease from left to right across a period.

Atomic radii increase as atomic number increases down a column or group	
DO	**DON'T**
Teach students that the increased number of energy levels (*n*) increases the distance over which the nucleus must pull and therefore reduces the attraction for electrons.	Don't let students get away with simply saying that radii increase down a column.
Teach students that full energy levels provide some shielding between the nucleus and valence electrons.	Don't let students use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument.

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**
Teach students that the effective nuclear charge, Z_{eff}, increases the attraction of the nucleus and therefore holds the electrons more tightly.	Don't' let them get away with simply stating that ionization energy increases from left to right across a period.
Teach students the exceptions that occur between groups II and III and V and VI.	Don't let them think that the trend is unwavering.
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 let them state that *p* electrons are farther away from the nucleus.
2) A drop in IE occurs between groups V and VI because the increased repulsion created by the first pairing of electrons outweighs the increase in Z_{eff} and thus less energy is required to remove the electron.	2) Don't let them state that the atoms in group V are more stable because they have a half filled sublevel. This is wrong, wrong, wrong!

Ionization energy decreases as atomic number increases down a column or group	
DO	**DON'T**
Teach students that the increased number of energy levels (*n*) increases the distance over which the nucleus must pull and therefore reduces the attraction for electrons.	Don't let students get away with simply saying that IE decreases down a column.
Teach students that full energy levels provide some shielding between the nucleus and valence electrons.	Don't let students use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument.

Electron affinity is NOT the opposite of ionization energy, but involves the addition of an electron to a gaseous atom or ion, which can be exothermic or endothermic. The exothermic values can be confusing for students since − 500 kJ represents a higher electron affinity than −100 kJ. You can tell students to consider the absolute value of the energy term since the negative sign is simply indicating the direction of energy flow (out of the system).

Electronegativity is a property (there are several scales) which measures 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**
Teach students that the effective nuclear charge, Z_{eff}, increases the attraction of the nucleus and therefore it strengthens the attraction for the electrons.	Don't' let them get away with simply stating that electronegativity increases from left to right across a period.

Electronegativity decreases as atomic number increases down a column or group	
DO	**DON'T**
Teach students that the increased number of energy levels (n) increases the distance over which the nucleus must pull and therefore reduces the attraction for electrons.	Don't let students get away with simply saying that electronegativity decreases down a column.
Teach students that full energy levels provide some shielding between the nucleus and valence electrons.	Don't let students use shielding for explanations across a period. Only full energy levels, not full sublevels, are of concern in a shielding argument.

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**
Teach students that positive metal ions result from the loss of valence electrons. In many cases this means the farthest electrons are now in a smaller principal energy level (n) than the original neutral atom.	Don't let students stop at 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.
Teach students to address the ratio of protons to electrons. As electrons are lost the ratio of $p+/e-$ increases and thus the electrons are held closer and with more strength.	Don't let students 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 smaller than their respective neutral atoms	
DO	**DON'T**
Teach students that negative nonmetal ions result from the addition of valence electrons. The primary explanation is the change in the proton to electron ratio. As electrons are added the $p+/e-$ ratio decreases and the electrons are not as closely held.	Don't let students say that the ion is bigger simply because it has more electrons. They must address the $p+/e-$ ratio and the $e-/e-$ repulsions to earn maximum credit.
Teach students that increased electron/electron repulsions also play a role in expanding the electron cloud.	Don't let students neglect this important effect. Electron repulsions are a powerful force within the atom.

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**
Teach students that 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 let students simply say that metals are more reactive at the bottom left corner of the table.

Non-metals are more reactive as you move up a column	
DO	**DON'T**
Teach students that 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 let students simply say that nonmetals are most reactive at the top right corner of the table.

Final thoughts

Students often have trouble immediately recognizing the difference between the two species given. Teach them to follow these three steps EVERY time they answer a periodicity question and their scores are sure to increase.

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