How to Answer Periodic Trends Free Response Questions

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

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

How to Earn Full Points on Periodic Trends Problems

Follow these three steps EVERY time you answer a periodicity question!

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 number of protons; if different *n*: argue with *n* vs. *n* = distance (between valence e⁻ and nucleus).

REMEMBER: a trend is not an explanation!

Simply identifying a trend (atomic radius decreases as you move from left to right across a period, electronegativity decreases as you move down a column, etc) earns ∅ points!

Avoid Losing Easy Points

1. When explaining, you must refer to **ALL** species (atoms, ions) referenced in the question, or you will not get full credit.
2. Read the question: justify with "principles of atomic structure" or "Coulomb's Law" (it will always be one or the other 😊).

Specific Question Types

1. Comparisons between isoelectronic species: explain with **number of p^+**
   a. Isoelectronic species with more protons are **SMALLER** because the valence electrons are more attracted to and thus **CLOSER** to the nucleus.
   b. Isoelectronic species with less protons are **LARGER** because the valence electrons are less attracted to and thus **FARTHER** from the nucleus.
2. Comparisons between an atom and its ion/ions of the same atom, Same n: explain with **e^-/e^- repulsion**
   a. Positively charged cations are **SMALLER** than the neutral atom because of less e^-/e^- repulsion, thus valence electrons are **CLOSER** to the nucleus.
   b. Negatively charged anions are **LARGER** than the neutral atom because of more e^-/e^- repulsion, thus valence electrons are **FARTHER** from the nucleus.
3. Comparisons between an atom and its ion/ions of the same atom, different n: explain with **distance**
   a. If a species has their outermost electrons on a lower energy level (n), their valence electrons are closer to and thus more attracted to the nucleus.

Guided Practice: How will an answer be different when explained in terms of "atomic structure" vs "Coulomb's Law"?

1. Which ion would require more energy to remove an electron: Mn^{3+} or Mn^{4+}? Explain using principles of atomic structure.

Mn^{4+} would require more energy to remove an e^- from than M^{3+}. Removing an e^- from M^{3+} decreases e^-/e^- repulsion, so the e^- cloud shrinks, and valence e^- in Mn^{4+} are closer to (and more attracted to) the nucleus.

2. Which ion would require more energy to remove an electron: Mn^{3+} or Mn^{4+}? Be sure to reference Coulomb's Law in your explanation.

Mn^{4+} would require more energy to remove an e^- from than M^{3+}. Removing an e^- from M^{3+} decreases e^-/e^- repulsion, so the e^- cloud shrinks, + valence e^- in Mn^{4+} are closer to the nucleus. Coulomb's Law states that the attraction between valence e^- and the nucleus increases as distance between them decreases, so more energy is required to overcome that attraction.

Independent Practice

1. Using principles of atomic structure, explain why the Na^+ ion is larger than the Li^+ ion.

Na^+'s valence e^- are in a higher principal energy level than Li^+'s valence e^-, + e^- in higher principal energy levels are further from the nucleus.

2. Which has a higher metallic character, an atom of lithium or an atom of boron? Explain your answer in terms of Coulomb's Law.

An atom of lithium has a higher metallic character than an atom of boron. Both Li's and B's valence e^- are found in the 2^{nd} energy level, but Li^+ has fewer p^+ in its nucleus than B. Coulomb's law states that the attraction between valence e^- and the nucleus decreases when the magnitude of the charges decreases, so it requires less energy to remove a valence e^- from Li than B.

3. Which would be greater for the aluminum atom, its first ionization energy or its second ionization energy? Explain using principles of atomic structure.

Al's second ionization energy (IE_2) is greater than its first ionization energy (IE_1). After removing Al's first e^- (IE_1), e^-/e^- repulsion decreases + so the e^- cloud shrinks. Thus, the remaining valence e^- are closer to the nucleus + so it takes more energy to remove a second e^- (IE_2).

4. Which element, Cl or Br, has a greater first ionization energy? Justify your answer using Coulomb's Law.

Cl has a greater first ionization than Br. Cl's valence e^- are in a lower principal energy level than those Br. Coulomb's law states that the energy of attraction between valence e^- and the nucleus increases as distance between them decreases, so more energy is needed to remove an e^- from Cl than Br.