Name: ________________________________ Band: _________ Date: _______________

AP Practice: Periodic Trends

Periodic trends are popular on the AP exam, so you better know them! Here are some things to keep in mind when answering PTE free response questions:

- Trends in atomic radius should be explained in terms of:
  - Number of energy levels,
  - Effective nuclear charge, and/or
  - Electron-electron repulsions.
- Trends in ionization energy, electron affinity, and metallic character (metals easily lose electrons) should be explained in terms of:
  - Orbital energy/distance of electron from nucleus,
  - Effective nuclear charge, and/or
  - Shielding.

1. Consider the atoms of the elements rubidium and strontium.
   a. Write the equation for the first ionization of rubidium and the first ionization of strontium.
   b. Predict which of the two elements has a lower first ionization energy. Justify your prediction.
   c. Predict which element will form ions with the smallest ionic radius. Justify your answer.

2. The table below shows the first three ionization energies for atoms of four elements from the third period of the periodic table. The elements are numbered randomly. Use the information in the table to answer the following questions.

	First Ionization Energy (kJ mol^{-1})	Second Ionization Energy (kJ mol^{-1})	Third Ionization Energy (kJ mol^{-1})
Element 1	1,251	2,300	3,820
Element 2	496	4,560	6,910
Element 3	738	1,450	7,730
Element 4	1,000	2,250	3,360

   a. Which element is most metallic in character? Explain your reasoning.
   b. Identify element 3. Explain your reasoning.
   c. Write the complete electron configuration for an atom of element 3.
   d. What is the expected oxidation state for the most common ion of element 2?
   e. What is the chemical symbol for element 2?
   f. A neutral atom of which of the four elements has the smallest radius?

3. Consider the two elements magnesium and sodium.
   a. Write the equation for the first ionization of atomic sodium.
   b. Explain why the radius of the Na^{+} ion is larger than the radius of the Mg^{2+} ion.
   c. The second ionization energy of Na is 4562 kJ/mol, whereas the second ionization energy of Mg is 1450 kJ/mol. Account for this difference using principles of atomic structure.

4. In the periodic table, as the atomic number increases from 11 to 17, what happens to the atomic radius?
   a. It remains constant.
   b. It increases only.
   c. It increases, then decreases.
   d. It decreases only.
   e. It decreases, then increases.

5. The ionization energies for element X are listed in the table below. On the basis of the data, what is the most likely identity of element X?

Ionization Energies for element X (kJ mol^{-1})				
First	Second	Third	Fourth	Fifth
580	1,815	2,740	11,600	14,800

   a. Na
   b. Mg
   c. Al
   d. Si
   e. P

6. Which species, Zn or Zn^{2+}, has the greater ionization energy? Justify your answer.

7. Answer the following questions related to sulfur.
   a. Consider the two chemical species S and S^{2-}.
      i. Write the electron configuration of each species.
      ii. Explain why the radius of S^{2-} ion is larger than the radius of the S atom.
      iii. Which of the two species would be attracted into a magnetic field? Explain.
   b. The S^{2-} ion is isoelectronic with the Ar atom. From which species, S^{2-} or Ar, is it easier to remove an electron? Explain.

8. Using principles of atomic and molecular structure and the information in the table below, answer the following questions about atomic fluorine, oxygen, and xenon, as well as some of their compounds.

Atom	First Ionization Energy ($kJ\ mol^{-1}$)
F	1,681.0
O	1,313.9
Xe	?

   a. Write the equation for the ionization of atomic fluorine that requires 1,681.0 $kJ\ mol^{-1}$.
   b. Account for the fact that the first ionization energy of atomic fluorine is greater than that of atomic oxygen. You must discuss both atoms in your response.
   c. Predict whether the first ionization energy of atomic xenon is greater than, less than, or equal to the first ionization energy of atomic fluorine. Justify your prediction.

9. Account for each of the following observations in terms of atomic theory and/or quantum theory.
   a. Atomic size decreases from Na to Cl in the periodic table.
   b. The first ionization energy of K is less than that of Na.

10. Suppose that a stable element with atomic number 119, symbol Q, has been discovered.
    a. Write the ground-state electron configuration for Q, showing only the valence-shell electrons.
    b. Would Q be a metal or a nonmetal? Explain in terms of electron configuration.
    c. On the basis of periodic trends, would Q have the largest atomic radius in its group or would it have the smallest? Explain in terms of electronic structure.

11. Use the principles of atomic structure and/or chemical bonding to explain each of the following. In each part, your answer must include references to both substances.
    a. The atomic radius of Li is larger than that of Be.
    b. The second ionization energy of K is greater than the second ionization energy of Ca.

12. As shown in the table below, the first ionization energies of Si, P, and Cl show a trend.

Element	First Ionization Energy ($kJ\ mol^{-1}$)
Si	786
P	1,012
Cl	1,251

a. For each of the three elements, identify the quantum level (e.g., n=1, n=2, etc.) of the valence electrons in the atom.
b. Explain the reasons for the trend in first ionization energies.

13. Account for the following observations using principles of atomic structure and/or chemical bonding. For each part, your answer must include specific information about both substances.
a. The Ca^{2+} and Cl^{-} ions are isoelectronic, but their radii are not the same. Which ion has the larger radius? Explain.
b. Carbon and lead are in the same group of elements, but carbon is classified as a nonmetal and lead is classified as a metal.
c. The first ionization energy of Be is 900 $kJ\ mol^{-1}$, but the first ionization energy of B is 800 $kJ\ mol^{-1}$.