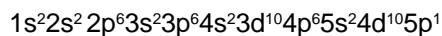


Chemistry I Practice Test Unit 3 "The Periodic Table"

Give complete explanations with your answers to the following questions:

1. An element has the following electron configuration:



- What period and group is it in?
 - How many valence electrons does it have and where are they located? (Energy level and orbital.)
 - Explain how to use the electron configuration to identify the period and group that it is located in.
 - What type of ion will this element form? Be specific about the charge and explain your reasoning.
 - Compare the radius of an atom of this element with the radius of an ion of this element and explain your reasoning.
2. Summarize the trends in ionic radius, atomic radius, first ionization energy and electronegativity as you move down a group in the periodic table. Explain why each trend occurs in terms of electron configurations, energy levels and nuclear charge.
3. Summarize the same trends as in #2 as you move across a period in the periodic table.
4. Rank the following atoms and ions in size from least to greatest. Explain your answer.
- Se Se⁻² Br
 - Se⁻² Br I
 - Al Al⁺³ C
5. Rank the following atoms in electronegativity from least to greatest. Define electronegativity. Explain your answer.
- F Ne Cl
 - Na Li Cs
 - K Br Rb
6. Rank the atoms in question #5 from least to greatest ionization energy. Define ionization energy. Explain your answer.
7. Find the number of valence electrons in the following elements. Explain why valence electrons are important.
- Cs
 - Cl
 - P
 - Sn
 - Xe
 - Sr
8. On the blank periodic table found on the back of this paper, identify the locations of the following groups of elements. Draw any necessary boundaries. Identify all of the group and period numbers.

metals nonmetals s-block p-block semimetals metals alkali metals

transition metals d-block halogens noble gases alkaline earth metals

Solutions to Unit 3 Practice Test Chemistry I

1. **a.** period 5 group 13 **b.** 3 valence electrons in the 5s and 5p orbitals.
- c.** The highest energy level occupied by electrons corresponds to the period number and the number of valence electrons is the ones digit of the group number. (e.g. group 13 = 3 valence electrons.)
- d.** This element will form a +3 ion because by losing 3 electrons, the 5th energy level will be empty leaving the 4th energy level completely filled which is chemically stable.
- e.** The radius of an atom of this element is larger than the radius of an ion because when the atom loses 3 electrons it will also lose an energy level.

2. **ionic radius** - increases as you move down a group because each successive element moving down a group has 1 more energy level than the previous element.

atomic radius - same trend as ionic radius.

first ionization energy - decreases as you go down a group because the added energy levels shield the positive charge of the nucleus which holds the electrons to the atom thus weakening this force.

electronegativity - decreases as you go down a group because the added energy levels shield the positive charge of the nucleus which attracts the electrons from outside the atom.

3. **ionic radius** - decreases as you move across a period because the added positive charge pulls the electrons closer to the nucleus. The ionic radius increases however, when you get to elements which form negative ions (nonmetals) because the negative ions have an additional energy level. The negative ions then continue to decrease in radius as you move across the period.

atomic radius - decreases as you move across a period because additional protons in the nucleus pull the outer energy level in closer to the nucleus resulting in a smaller radius.

ionization energy - increases as you move across a period because additional protons create more positive charge which holds the electrons to the atom with greater force.

electronegativity - increases as you move across a period because the additional protons create more positive charges in the nucleus which results in a greater attractive force towards electrons from outside the atoms.

4. **a.** $\text{Br} < \text{Se} < \text{Se}^{2-}$ Br is smallest because it is furthest to the right in the period. Se^{2-} is larger than Se because the additional electrons take up more space.

b. $\text{Br} < \text{I} < \text{Se}^{2-}$ Br is smaller than Se^{2-} because it is further to the right in the same period. Since I and Se^{2-} are not in the same group we cannot use general trends to determine their relative sizes. Literature reveals that Se^{2-} has a radius of 184 pm while I has a radius of 138 pm.

c. $\text{Al}^{3+} < \text{C} < \text{Al}$ Both C and Al^{3+} are smaller than Al because they have one less energy level. Al^{3+} is smaller than C because it has a greater nuclear charge which pulls the second energy level closer to the nucleus.

5. Electronegativity is the attraction that an atom has for other atom's electrons.

a. $\text{Ne} < \text{Cl} < \text{F}$ Neon has no electronegativity because it does not want any more electrons because of its completely filled energy levels. F is more electronegative than Cl because it has one less energy level to shield the positive nucleus from the outside of the atom.

b. $\text{Cs} < \text{Na} < \text{Li}$ Cs is the least electronegative because it is the furthest down the group. Na is the second least because it is below Li which is the most electronegative because it is the highest in group 1 of the three atoms.

5 c. $Rb < K < Br$ Rb and K are both on the left side of the table, Br is on the right, making Br the most electronegative. Rb is below K in group 1, making it less electronegative than K.

6. Ionization energy is the energy required to remove an electron from an atom.

a. $Cl < F < Ne$ Cl is in a higher period than F and Ne and will have a lower ionization energy because of the shielding effect that the extra energy level has on the attractive force from the nucleus that holds the electrons to the atom. Ne has a higher ionization energy because it lies to the right of F in period 2 and therefore has higher nuclear charge, which has a stronger hold on the outer electrons .

b. $Cs < Na < Li$ Since the ionization energy decreases as you move down a group and Cs is the furthest down the group, it has the lowest ionization energy followed by Na which is further down group 1 than Li.

c. $Rb < K < Br$ Rb and K both have lower ionization energies than Br because they are on the far left side of the periodic table which has lower ionization energies than the right side. Rb has a lower ionization energy because it lies below K in group 1.

7. Valence electrons are important because they are the electrons that are involved in any chemical change that the element undergoes. We can use the valence electrons to predict the chemical behavior of elements.

a. 1

b. 7

c. 5

d. 4

e. 8

f. 2

8.

