

Part I: Summary of Periodic Trends

1. Define the phrase “periodic trend” in your own words. A pattern exhibited by the chemical or physical characteristics of the elements on the periodic table. Increases across and decreases down → IE, EA, EN Decreases across and increases down → AR, IR (sort of), MP, TC

Property	Definition	Trend From Left to Right Across a Period	Trend Down a Group
Atomic radius	Half the distance between two identically bonded atoms.	Decreases	Increases
Ionic radius	The radii of the most commonly found ion of the element.	Decreases across the metals and then gets much larger at the nonmetals and then decreases again.	Increases
First Ionization Energy	The energy required to remove one electron from a neutral atom of an element.	Increases	Decreases
Electron Affinity	The energy change that occurs when an electron is acquired by a neutral atom	Increases	Decreases
Number of Valence Electrons	The electrons that are available to be lost, gained, or shared in the formation of chemical compounds.	Increases	Same
Melting Point	The temperature at which a solid turns into a liquid.	Decreases	Increases
Thermal Conductivity	The ability of a substance to conduct heat.	Decreases	Increases

Electronegativity	The ability of an atom to pull the bonded electrons to itself in a chemical compound.	Increases	Decreases

Part II: Explaining Periodic Trends

1. How is ionization energy related to how easily an electron can be lost from an atom?

The easier it is to lose an electron the smaller the amount of energy needed to remove it (smaller ionization energy).

2. What happens to the amount of positive charge in the nucleus as you go from left to right across a period?

The amount of positive charge increases as the number of protons increase from left to right across a period.

- a. How would this account for the trend you discovered in atomic radius?

Electrons are held more tightly and are pulled in closer by the nucleus as the number of protons increases within a level.

- b. How would this account for the trend you discovered in the first ionization energy?

Because of the additional protons, electrons are held more tightly by the nucleus which increases the amount of energy (ionization energy) needed to remove them.

- c. How would this account for the trend that you discovered in electron affinity?

The additional protons pull the electron cloud in closer which reduces the amount of energy needed to attract its electrons. This in turn increases the amount of potential energy the atom has for gaining an additional electron.

3. What is different about the number of energy levels of electrons as you go down a group?

The number of energy levels increases as you go down a group.

- a. How would this account for the trend you discovered in atomic radius?

The additional energy levels of electrons take up more space. The larger the electron cloud the larger the atomic radius.

- b. How would this account for the trend you discovered in the first ionization energy?

The electrons are not held as tightly by the nucleus and, therefore, can be lost more easily. This means that the ionization energy decreases as the atomic radius increases.

4. What is the relationship between the group (groups 1 & 2 and 13 – 18) and the number of valence electrons?

The ones place in the group number is equal to the number of valence electrons for these groups.

- a. How does this account for the ions that are formed for each group?

If the atom has less than 4 valence electrons (VE), the atom will lose its VE to obtain a full octet. If the number of VE exceeds 4, it's easier for an atom to gain electrons to obtain a full octet. Atoms that have 4 VE can either lose 4 or gain 4 electrons to obtain a full octet.

- b. How does this account for the trend you discovered in the ionic radius?

Metals have less than 4 valence electrons and therefore lose them to obtain a full octet. Losing electrons causes the radius to decrease in size from atom to ion.

Nonmetals have 4 or more valence electrons and therefore gain electrons to obtain a full octet. Gaining electrons causes the radius to increase in size from atom to ion.

- c. What is the relationship between the number of valence electrons and an element's electronegativity?

As the number of valence electrons increases the electronegativity increases.

5. Which properties from the summary of Periodic Trends Table describe metallic character?

Melting Point and Thermal Conductivity.

- a. In general, what happens to the metallic character that elements have as you go from left to right across a period?

Metallic character decreases

- b. In general, what happens to the metallic character that elements have as you go down a group?

Metallic character increases

6. What is Effective Nuclear Charge? (See pages 154 – 155)

The difference between the full nuclear charge (the attraction for the protons), Z , and the screening effect (shielding) of the inner electrons is called the effective nuclear charge, or Z_{eff} .

Part III: Applying Periodic Trends

- Which group on the periodic table has the largest atomic radii? Group 1
- Which group on the periodic table has the largest first ionization energy? Group 17
- What element on the periodic table has the highest/lowest electron affinity? Cl Fr
- What element on the periodic table has the largest/smallest atomic radius? Fr He
- What element on the periodic table has the largest/smallest ionic radius? Si⁻⁴ C⁺⁴
(Ignoring H)
- What element on the periodic table has the highest/lowest ionization energy? F Fr
- What element on the periodic table has the highest/lowest electronegativity? F Fr
- Compare the element Fe and Cu by answering the next three questions.
 - Which element has the smaller radius? Cu
 - Which element has the larger first ionization energy? Cu
 - Which element has the least metallic character? Cu
- Predict the number of valence electrons for the next four elements.
 - Ba 2
 - Pb 4
 - Po 6
 - Rn 8
- In each pair of elements, circle the one that has the larger atomic radius.
 - Cs or **Fr**
 - Sn or **Pb**
 - Ag** or Cd
- In each pair of elements, circle the one that has the larger first ionization energy.
 - Cs or **Ba**
 - Sb** or Bi
 - Au or **Hg**
 - O or **N**
- In each pair of elements, circle the one that has the most metallic character.
 - Ba or **Ra**
 - Pb or **Pt**
 - Po** or Te
- Place the following elements / ions in order of increasing radius.
 - Na, K, Li Li < Na < K
 - Te, I, Xe Xe < I < Te
 - Fe, P, O O < P < Fe
 - O, O⁻¹, O⁻² O < O⁻¹ < O⁻²
 - O⁺¹, O, O⁻¹ O⁺¹ < O < O⁻¹
 - Se⁻², Ca⁺², Br⁻¹, K⁺¹ Ca⁺² < K⁺¹ < Br⁻¹ < Se⁻²

14. In each pair of elements, circle the one that has the larger electron affinity.

- a. F or **Cl** b. **Si** or P c. **Cu** or Zn d. Mn or **Fe**

15. In each pair of elements, circle the one that has the larger second ionization energy.

- b. **Li** or Be b. Be or **B** c. **S** or P

16. Determine which element in the following compounds has the highest electronegativity. Then determine if the compound is polar or nonpolar. If it is polar, label the positive and negative ends of the compound.

- a. $(+)\text{MgO}^{(-)}$ polar b. $(-)\text{N}_2^{()}$ nonpolar c. $(+)\text{RbF}^{(-)}$ polar

17. What is an isoelectronic series? Give an example. Is a group of ions that have the same

number of electrons and the same electron configuration. For example: O^{-2} , F^{-1} , Na^{+1} , Mg^{+2} ,

Al^{+3} (10 electrons each and size decreases \rightarrow) (Sr^{+2} , As^{-3} , Se^{-2} , Rb^{+1} , Br^{-1})

18. Explain how the atomic structure, including subatomic particles and energy levels of electrons, relate to how easily valence electrons are removed from or added to an atom. Since chemical reactivity of metals is based on losing electrons, explain how this accounts for the increased chemical reactivity of metals as you go down a group on the periodic table. Since chemical reactivity of nonmetals is based on gaining electrons, explain how this accounts for the increased chemical reactivity of nonmetals as you go up a group on the periodic table.

As the pull of the nucleus on the valance electrons decreases, it becomes easier for the valence electrons to be lost from an atom. The decrease in pull on the valence electrons occurs as the atomic radius increases (as energy levels are added). An increase in energy levels of electrons block the pull of the nucleus from fully reaching the valence electrons (nuclear charge – shielding = Effective nuclear charge), making it easier to lose electrons.

Metals have a tendency to lose electrons when they react because they typically have less than 4 valence electrons and it takes less energy to lose a few electrons than gain a lot of them. As you go down a group on the periodic table there are more energy levels, the valence electrons are further from the nucleus, and the inner electrons block or shield some of the nuclear pull (nuclear charge – shielding = Effective nuclear charge). The shielding that occurs makes it easier for electrons to be lost which means that the metals that are closer to the bottom of the periodic table are more reactive. This means that the energy needed to

remove a valence electron (ionization energy) decreases as the atomic radius increases. As you go across a period, the number of protons increase and the number of valence electrons also increase. The more valence electrons a metal has the more energy is needed to lose them. The total energy needed to remove all of the valence electrons in a metal is the addition of all of the ionization energies. (1^{st} IE + 2^{nd} IE + 3^{rd} IE ...).

When nonmetals react they have a tendency to gain electrons because they typically have 4 or more valence electrons and it takes less energy to gain a few electrons than lose a lot of them. The atomic radius of nonmetals decreases as you go across the periodic table and increase as you go down a group. The atomic radius decreases, as you go across, because the number of protons increase which increases the nuclear charge (the attraction the protons have for the electrons). The smaller the atomic radius the more reactive the nonmetal. The increase of atomic radius as you go down a group of nonmetals means that the reactivity of the nonmetals decreases as you go down the group. Larger radius...less attraction for the valence electrons (further away and shielding)...less attraction for gaining an additional electron.