WLS Periodicity

1. Atomic radius - ½ the distance between the two nuclei in 2 adjacent metal atoms or in a diatomic molecule.

Ionic radius - radius of a cation or anion.

Ionization energy - min amount of nrg (in kJ/mol) required to remove an e- from a gaseous atom in its ground state (i.e., measure of how "tightly" the e- is held in the atom).

Electron affinity - the negative of the nrg change that occurs when an e- is accepted by an atom in the gaseous state to form an anion (i.e., nrg that must be supplied to remove an e- from an anion).

2. When an atom becomes an anion its size (or radius) increases because the nuclear charge remains the same but the repulsion resulting from the additional e-(s) enlarges the e- cloud. Removing one or more e-(s) from an atom reduces e- - e- repulsion but the nuclear charge remains the same; so the e- cloud shrinks & the cation is smaller.

* For isoelectronic ions the # of e- are the same, but the anions have a smaller nuclear charge so the e- cloud is more spread out.
3. $\text{Na} > \text{Mg} > \text{Al} > \text{P} > \text{Cl}$

4. \( F \)

5. The effective nuclear charge that the outermost e\(^{-}\) feels inc across the period.

6. $\text{Mg}^{2+} < \text{Na}^{+} < \text{F}^- < \text{O}^{2-} < \text{N}^{3-}$

7. Both Se and Te are Group 6A elements. Since atomic radius inc going down a column in the periodic table, it follows that $\text{Te}^{2-}$ must be larger than $\text{Se}^{2-}$.

8. H\(^-\) is larger. Both H\(^-\) and He have 2 e\(^-\) (they are isoelectronic) but they have diff # of protons in nucleus. Since H\(^-\) has only 1 proton, the 2 e\(^-\) will experience less nuclear charge, & H\(^-\) will be larger than He.

9. When an e\(^-\) is removed from an atom, the repulsion among the remaining e\(^-\) dec. B/c the nuclear charge remains constant, more nrg is needed to remove another e\(^-\) from the par. charged ion.

10. **Highest - Noble gas**
    We tend to associate full valence-shell e\(^-\) config w/ an inherent degree of chemical stability. The high
Ionization energies of the noble gases, stemming from their low effective nuclear charge, comprise one of the reasons for their stability.

Lowest - group 1A (alkali metals).
Each of these metals has 1 valence electron, which is effectively shielded by the completely filled inner shells. It's energetically easy to remove an electron from the atom of an alkali metal to form an ion.

11. K > Ca > Li > F > Ne

12. Group 3A elements (like Al) all have a single electron in the outermost p subshell, which is well shielded from the nuclear charge by the inner ns² electrons. Therefore, less energy is needed to remove a single p electron than to remove a paired s electron from the same principle energy level (such as for Mg).

\[
\text{Al: } [\text{Ne}] 1s^2 2s^2 2p^1 \\
\quad \quad \quad \quad \quad 3s^1 3p^1
\]

\[
\text{Mg: } [\text{Ne}] 1s^2 2s^2 2p^6 \\
\quad \quad \quad \quad \quad 3s^2 \\
\quad \quad \quad \quad \quad \text{easier to remove than}
\]

\[
\text{than}
\]
13. $\text{ls}^2 \text{ls}^2 \text{2p}^4 \underline{\text{complete}}$ 2080 KJ/mol
$\text{ls}^2 \text{ls}^2 \text{2p}^6 \text{3s}^1 \underline{\text{valence e}^- = 1}$ low ionization energy

14. Ionization is always an endothermic process. By convention, the energy absorbed by atoms (or ions) in the ionization process has a positive value. Thus, ionization energies are all positive quantities. Electron affinity is positive if the reaction is exothermic, and negative if the reaction is endothermic.

15. a) K < Na < Li  b) I < Br < F < Cl
   c) Ca < Ba < P < Si < O

16. greatest: Cl  least: He

17. Alkali metals have a valence e- config of ns, so they can accept another e- in the ns orbital. On the other hand, alkaline earth metals have a valence e- config of ns2. Alkaline Earth metals have little tendency to accept another e- as it would have to go into a higher energy p orbital.