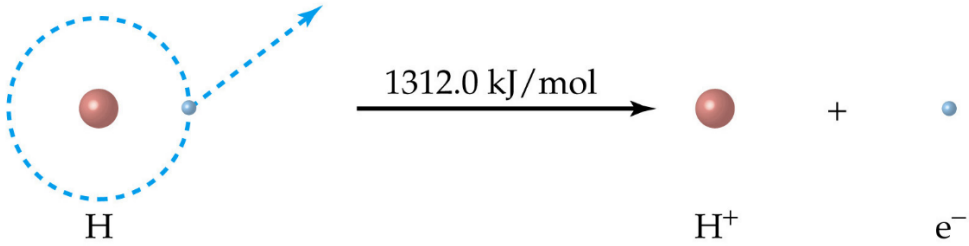
Module-10: Periodic Trends

THE IONIZATION ENERGY

- The **Ionization energy** (IE) is the energy required to remove an electron from an isolated **gaseous** atom.

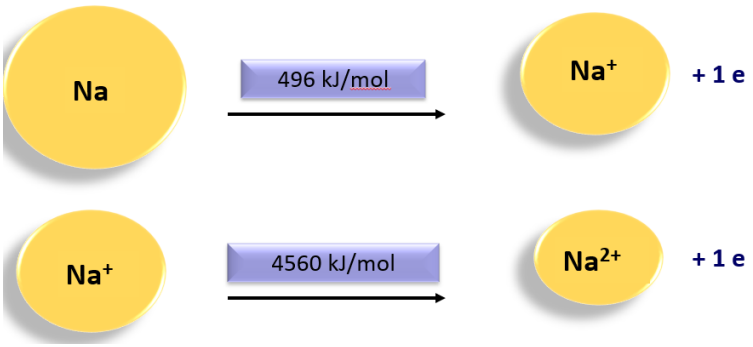


- Ions are charged species and they can be either positively charged or negatively charged. Generally, metallic elements form positively charged ions called cations and non-metals form negatively charged ions called anions.
- The number of charges on ions depend on the location of the element on the periodic table.
- The first group element sodium (Na) after losing an electron, the cation attains the electronic configuration of the nearest noble gas, neon.

$$Na \rightarrow Na^+ + e^-$$

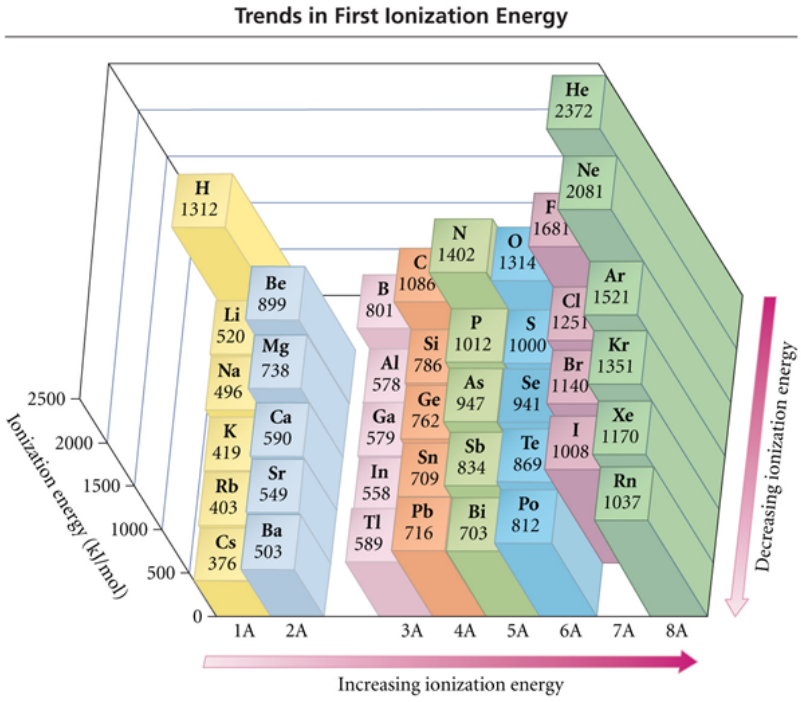
$Na = [Ne]\ 3s^1$ $\qquad$ $Na^+ = [Ne]$

Once a noble gas configuration is reached, removing further electrons would require a very large amount of energy.

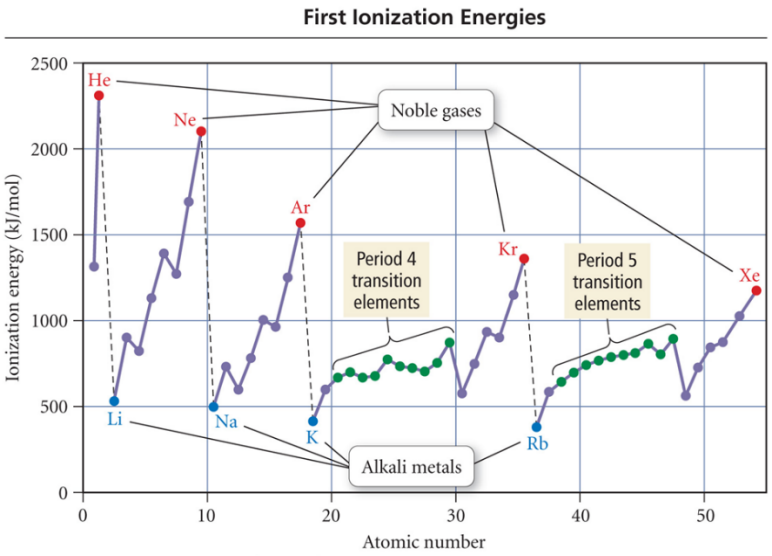


Due to this extremely high second ionization energy, divalent compounds of sodium are unknown. For example, stable $NaCl_2$ has never been made.

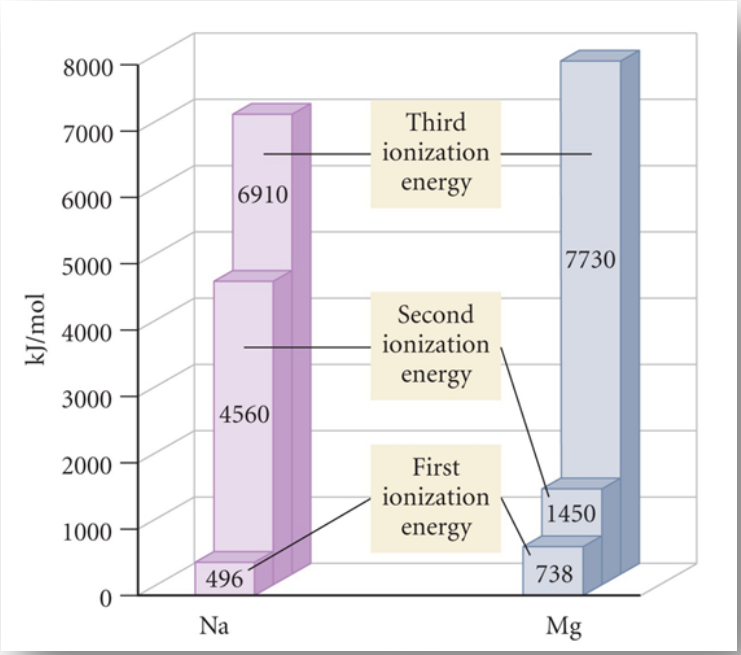
- The 1st Ionization energy (IE) **decreases** down the group in the periodic table because the valence electrons are farther away from nucleus.



- The 1st IE generally **increases** across the period, left to right, due to the fact that the effective nuclear charge increases.



- However, there are exceptions, even though oxygen is to the right of nitrogen in periodic table, the 1st IE of oxygen is lower than that of nitrogen.
- The reason for this discrepancy is attributed to the fact that, in one of the 2*p* orbitals in oxygen there are two electrons, and the repulsion between these two electrons decrease the IE.
- Trends in Second and Successive Ionization Energies: After removal of the 1st electron the size of the atom decreases, the decrease in size of the atom (ion) is due to having more protons than electrons.
- The decrease in size brings the outer electrons closer to the nucleus, therefore it is harder to remove more electrons.

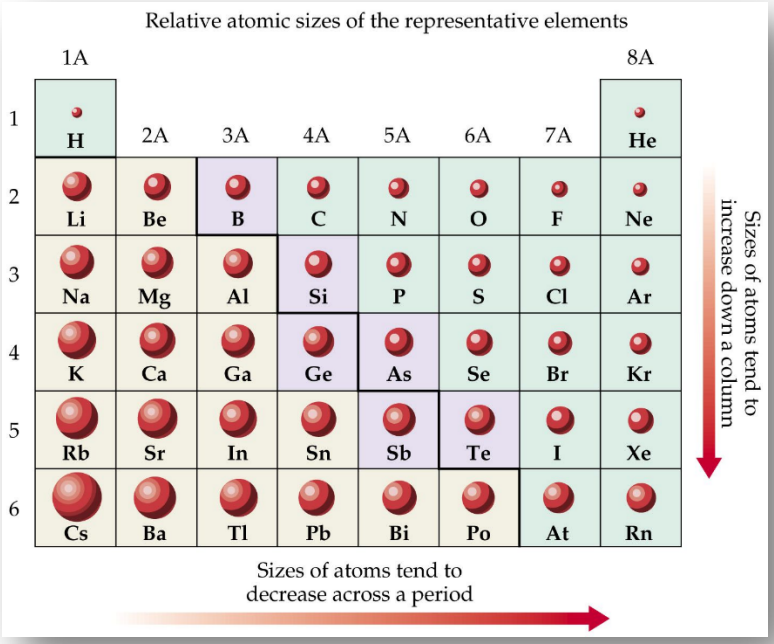


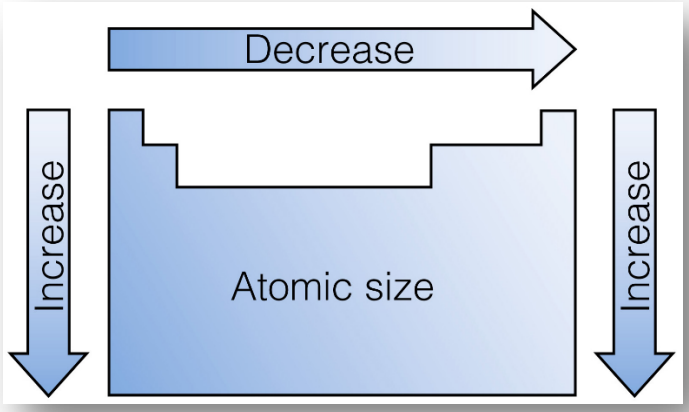
- After removing two electrons from the 2nd group element magnesium, we get the Mg^{2+} ion. The Mg^{2+} ion has the electronic configuration of the nearest noble gas neon (Ne). The energy needed to remove the 3rd electron from a stable Mg^{2+} is incredibly high (7730 kJ/mol) that Mg^{3+} ion is virtually unknown.
- Like Na^{+} and Mg^{2+} ions, any ion with noble gas electronic configuration cannot be ionized further, because the energy required to overcome a stable noble gas electronic configuration is very high.

RELATIVE SIZES OF ATOMS AND IONS

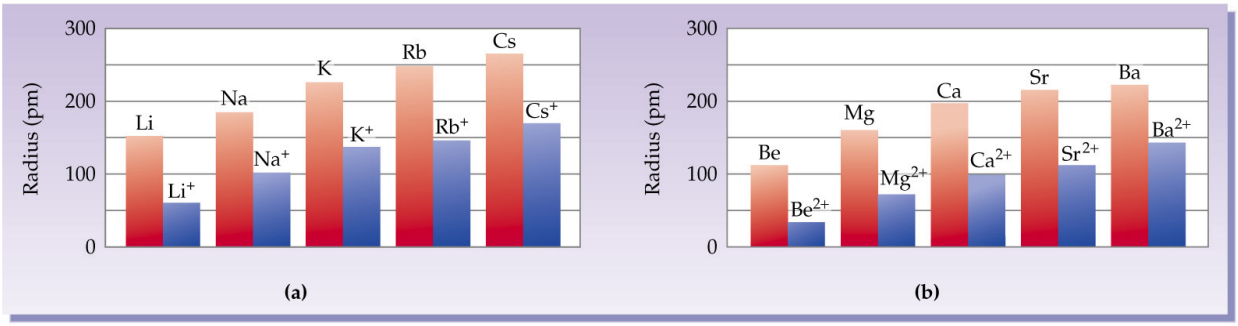
- Atomic radii decrease across period (left to right), this is due to the fact that, we are adding electrons to same valence shell, effective nuclear charge increases and valence shell held closer to the nucleus.

As the atomic number increases, the atomic radii increase within a group.

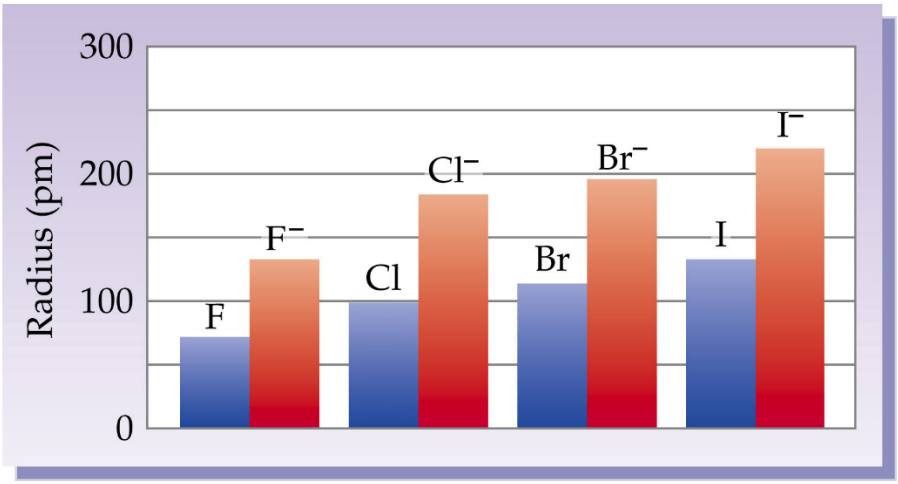




The ionic radii of cations ions are always smaller than the atoms from which they are formed.



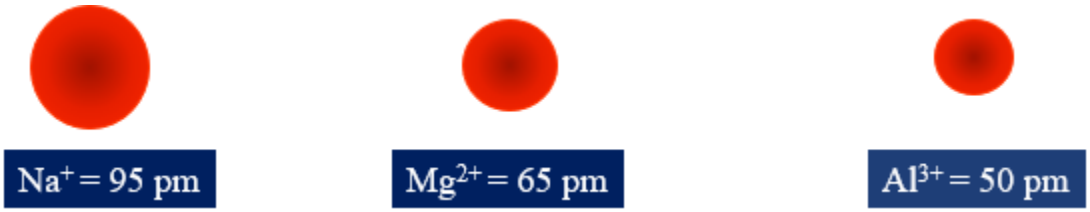
The anions are larger than the neutral atoms because of additional electron-electron repulsions and a decrease in effective nuclear charge.



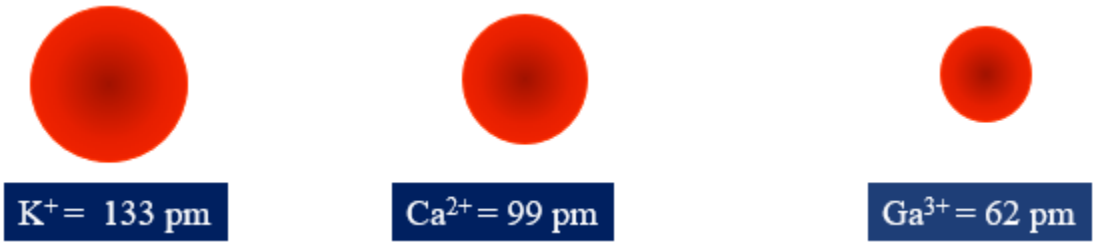
Isoelectronic ions: Theses are ions with the same number of electrons but with differing number of charges. For example, Na^{+}, Mg^{2+} and Al^{3+} ions have ten electrons each.

General trend for **isoelectronic cations**: higher the charge smaller the ionic radii.

(A) Following three cations have 10-electrons each, Al^{3+} is the smallest because it has the highest number of positive charges.



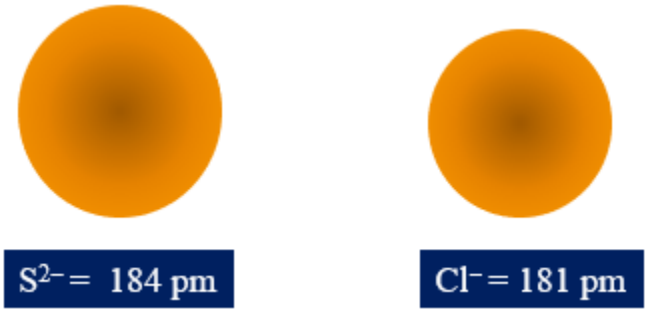
(B) All of the following cations have 18-electrons each, again we see the trend where ion with the highest charge has the smallest radius.



General trend for isoelectronic anions: higher the charge larger the ionic radii.

(A) The following two anions have 10-electrons each, O^{2-} ion with two negative charges is larger than F^- ion which has only one negative charge.

(B) The following two anions have 18-electrons each, we see a similar trend, however, the differences in anion radii are not significant among 3rd period elements.



Practice Problems

1. Which of the following elements has the largest second ionization energy?

(A) Na (B) Cl (C) Mg (D) Si

2. Which of the following processes corresponds to the first ionization energy of calcium?

(A) $Ca(g) \rightarrow Ca^{+}(g) + e^{-}$ (B) $Ca(s) \rightarrow Ca^{+}(s) + e^{-}$

(C) $Ca(s) \rightarrow Ca^{+}(g) + e^{-}$ (D) $Ca(s) + e^{-} \rightarrow Ca^{-}(g)$

3. Which of the following properties, in general, increases from left to right across a period in the periodic table?

(A) cation radii (B) atomic radii

(C) ionization energy (D) metallic character

4. Which of the following elements has the largest atomic radius?

(A) Ge (B) K (C) Ca (D) Cl

5. A section of the p-block elements in the periodic table with all identification features removed is shown below.

V	W	X
	Y	Z

Which element has the smallest atomic radius?

(A) W (B) Y (C) X (D) V

6. Below are data on the first four ionization energies for a fictitious element X.

First ionization energy = 500 kJ/mol Second ionization energy = 2000 kJ/mol

Third ionization energy = 3500 kJ/mol Fourth ionization energy = 18,000 kJ/mol

(A) X could belong to Group IIIA (B) X could belong to Group IIA.

(C) X could belong to Group IVA. (D) X could belong to Group IA.

7. Which of the following elements has the smallest ionization energy?

(A) I (B) Sb (C) Sn (D) Rb

8. Rank the following ions in order of increasing first ionization energy: O^{2-}, Mg^{2+}, F^-, Na^+.

(A) $Mg^{2+} < Na^+ < F^- < O^{2-}$ (B) $Mg^{2+} < O^{2-} < Na^+ < F^-$

(C) $O^{2-} < F^- < Mg^{2+} < Na^+$ (D) $O^{2-} < F^- < Na^+ < Mg^{2+}$

9. Which of the following processes corresponds to the second ionization energy of barium?

(A) $Ba(g) \rightarrow Ba^{2+}(g) + 2\ e^-$ (B) $Ba(g) \rightarrow Ba^+(g) + e^-$

(C) $Ba^+(g) \rightarrow Ba^{2+}(g) + e^-$ (D) $Ba^+(g) + e^- \rightarrow Ba(g)$

10. The change in energy for which of the following processes corresponds to the electron affinity of iodine?

(A) $I_2(g) \rightarrow 2\ I(g)$ (B) $I(g) + e^- \rightarrow I^-(g)$

(C) $I^+(g) + I^-(g) \rightarrow I_2(g)$ (D) $I(g) \rightarrow I^+(g) + e^-$