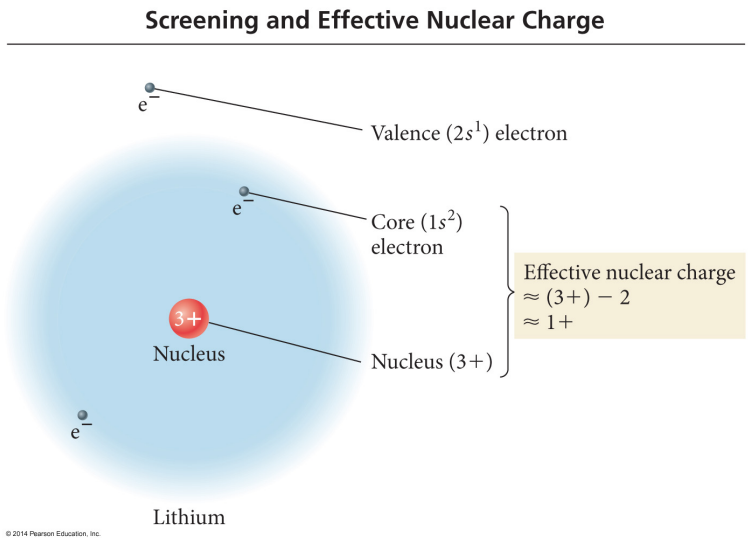
Chapter 8:
Periodic Properties of the Elements

- We will now explore how electron configurations affect various properties of atoms, and the trends for these properties moving across the periodic table.
- The properties we will study are:
  - Size of atoms
  - Size of ions
  - Ionization energies
  - Electron affinities
  - Metallic character

8.6 Periodic Trends in the Size of Atoms

- In general, properties of elements depend on the strength of the attraction between outer electrons and the nucleus.
- According to Coulomb's Law, the attraction is stronger as the **charge on the nucleus (Z) increases**, and as the electron gets **closer to the nucleus**.
- The charge of the nucleus increases as Z increases, but the electrons do not always "feel" all of the charge due to **shielding** by the inner electrons (Figure 8.11).



Effective Nuclear Charge

- The charge that an electron actually experiences is called the **effective nuclear charge,** $Z_{\text{efffective}}$.

$$Z_{\text{effective}} = Z - \text{shielding}$$

- Outer electrons are shielded by inner electrons. Electrons in the same shell do not shield very well.
- Comparing elements in the same period of the periodic table: As you go **left to right** across the row, since the number of protons in the nucleus is increasing.:

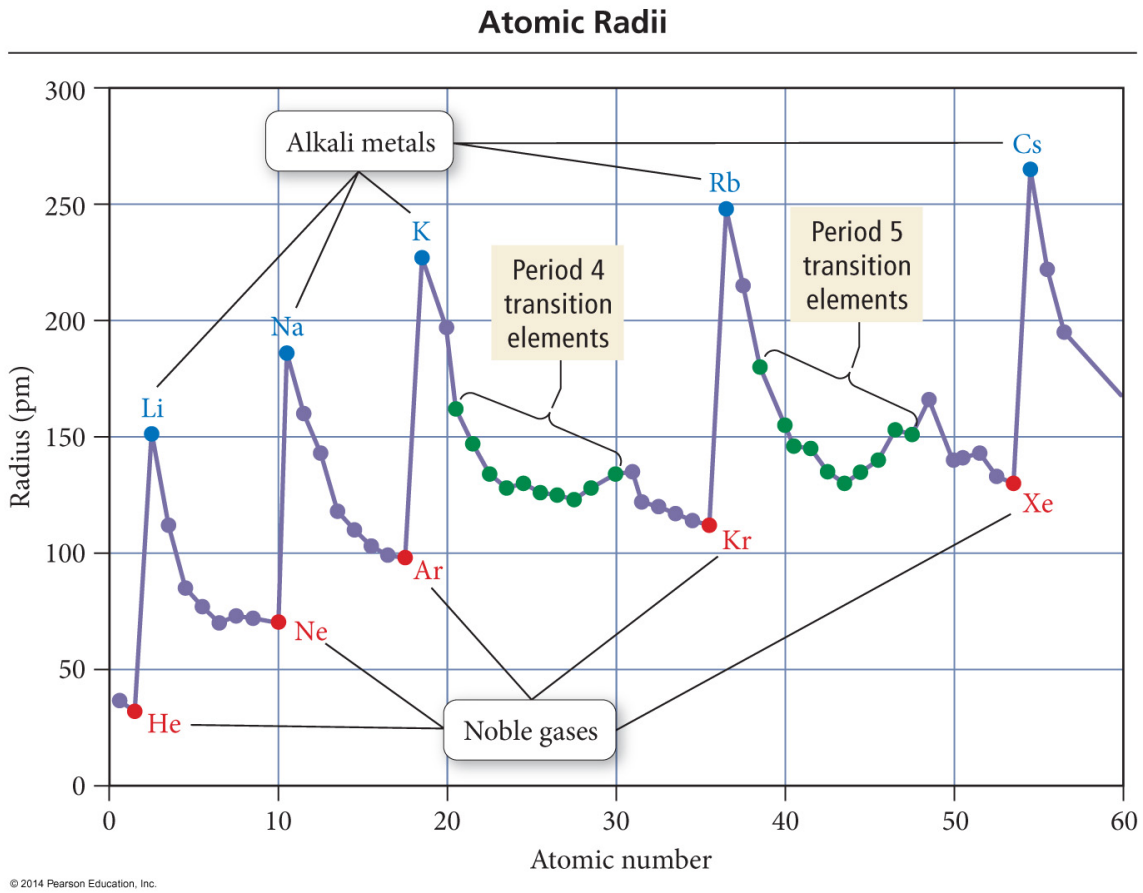
$$Z_{\text{effective}} \text{ increases}$$

- Electrons are also being added to the same shell so shielding does not increase much.

Atomic Radius

- **Atomic radius** is a measure of the size of an atom. It can be defined as half of the distance between two atoms of the same element bonded together.

Atomic Radius (contd.)



- Notice the trends in atomic radius:
  - **Moving left to right across a row, radius decreases**
  - **Moving down a group, radius increases**

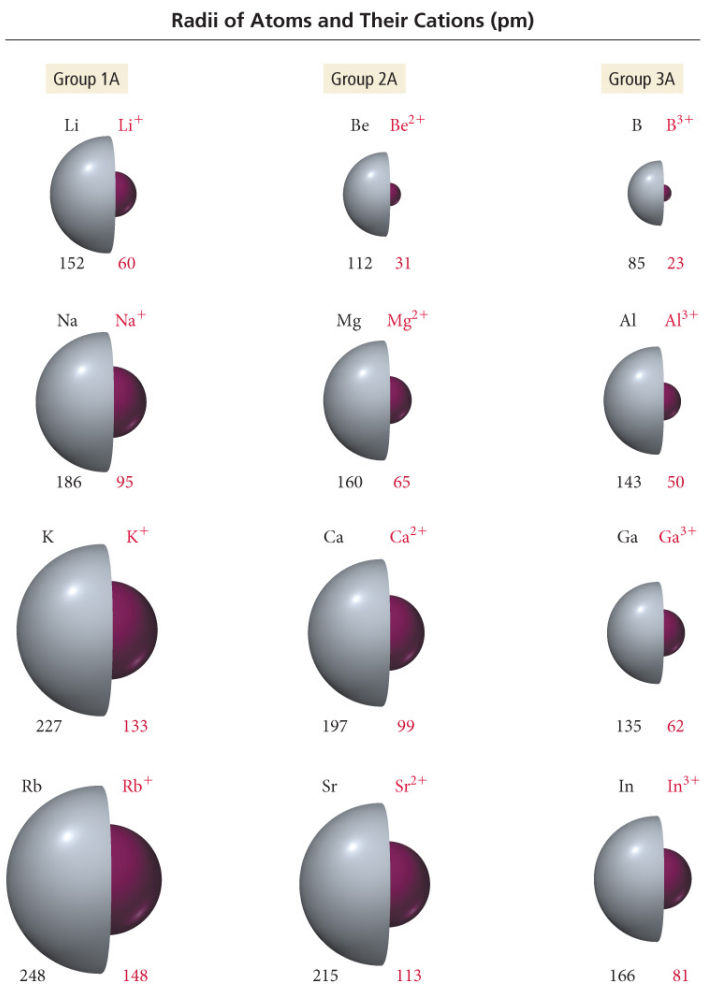
Atomic Radius (contd.)

How can we explain the trends in atomic radius?

- As we have seen, moving left to right across a period increases $Z_{\text{effective}}$.
- A larger $Z_{\text{effective}}$ means that the outer electrons are attracted closer to the nucleus, so the atom is smaller.
- **As you go down a group:** The size of atoms increase since electrons are being added to orbitals with **higher** $\boldsymbol{n}$.

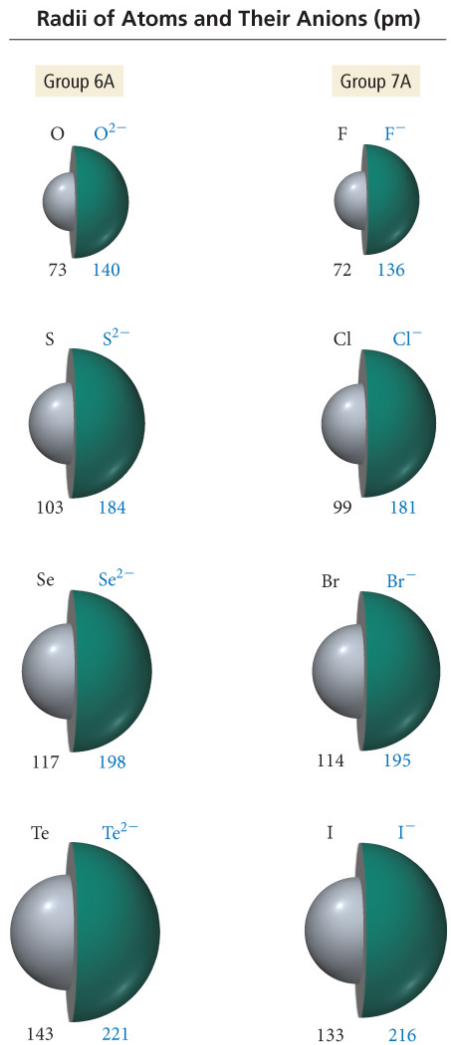
Ionic Radius

- When an atom loses an electron, the resulting cation is always smaller than the neutral atom (Figure 8.12):



Ionic Radius (contd.)

- Anions are larger than the neutral atom (Figure 8.13).



Ionic Radius (contd.)

Trends in ionic radius:

- As you go down a group, ionic radius gets larger, since atomic radius is increasing and charge is constant.
- As you go left to right across a period, change in size is less meaningful since charge is changing.
- Compare a set of ions that are all isoelectronic with Ar:

	S^{2-}	Cl^{-}	K^{+}	Ca^{2+}
Radius in pm:	184	181	133	99

So for isoelectronic ions:

- Ionic radius decreases as the charge becomes more positive
- Ionic radius increases as the charge becomes more negative

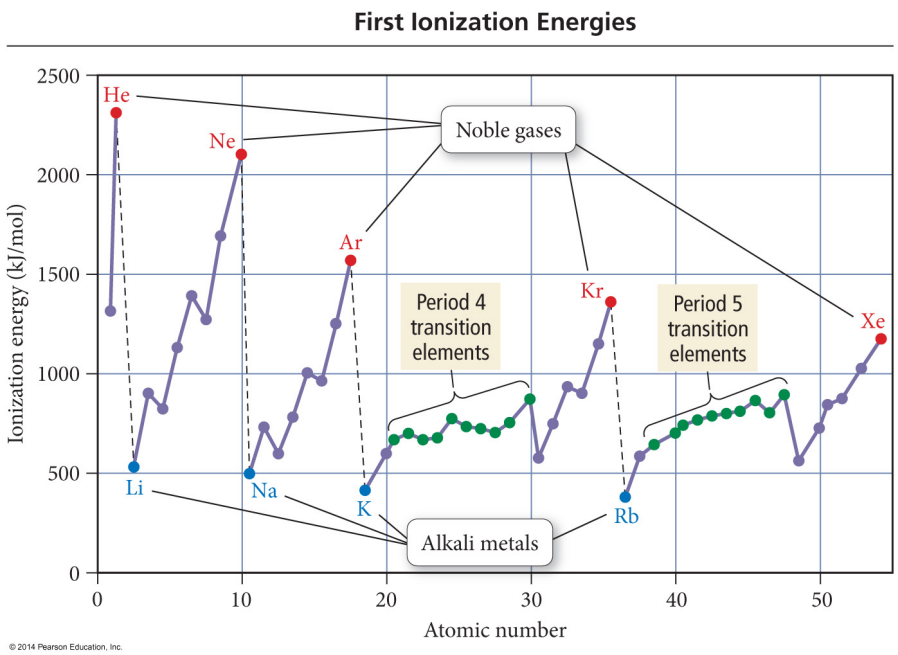
Ionization Energy

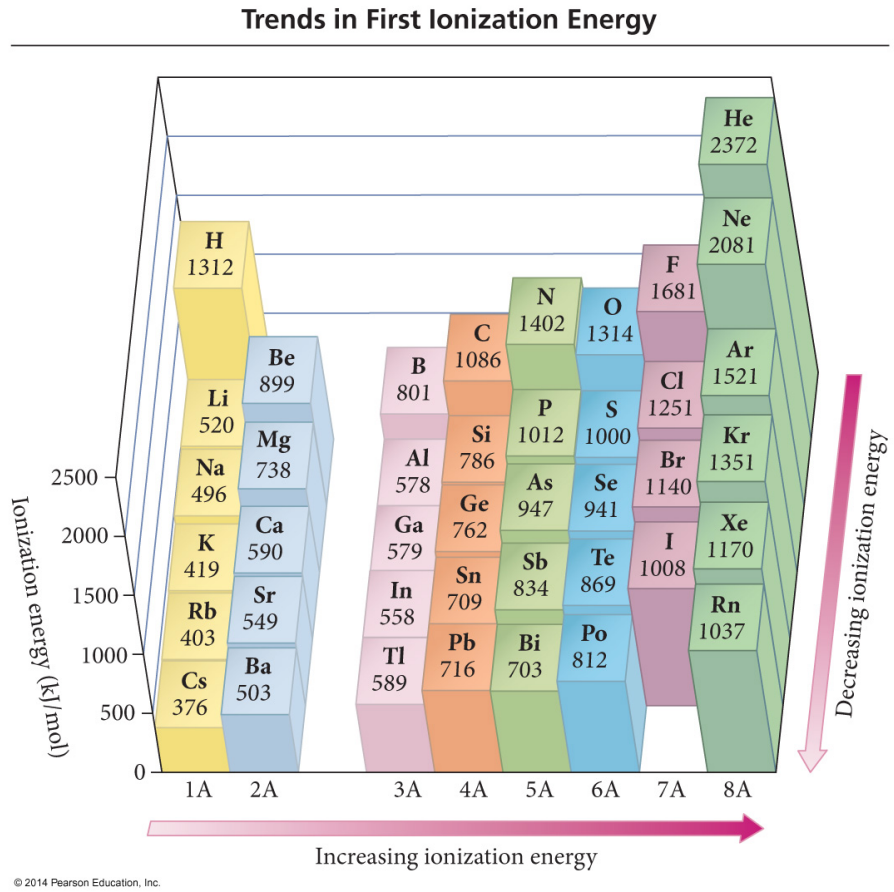
Ionization energy: The energy required to remove an electron from a gaseous atom in its ground state:

$$1^{st}\ IE:\ A\ (g) \rightarrow A^{+}\ (g) + e^{-}$$

- IE has opposite trend to atomic radius (Figures 8.14-15).
- It is easier (takes less energy) to remove an electron from a larger atom than a smaller atom.

Ionization Energy





Ionization Energy

In summary:

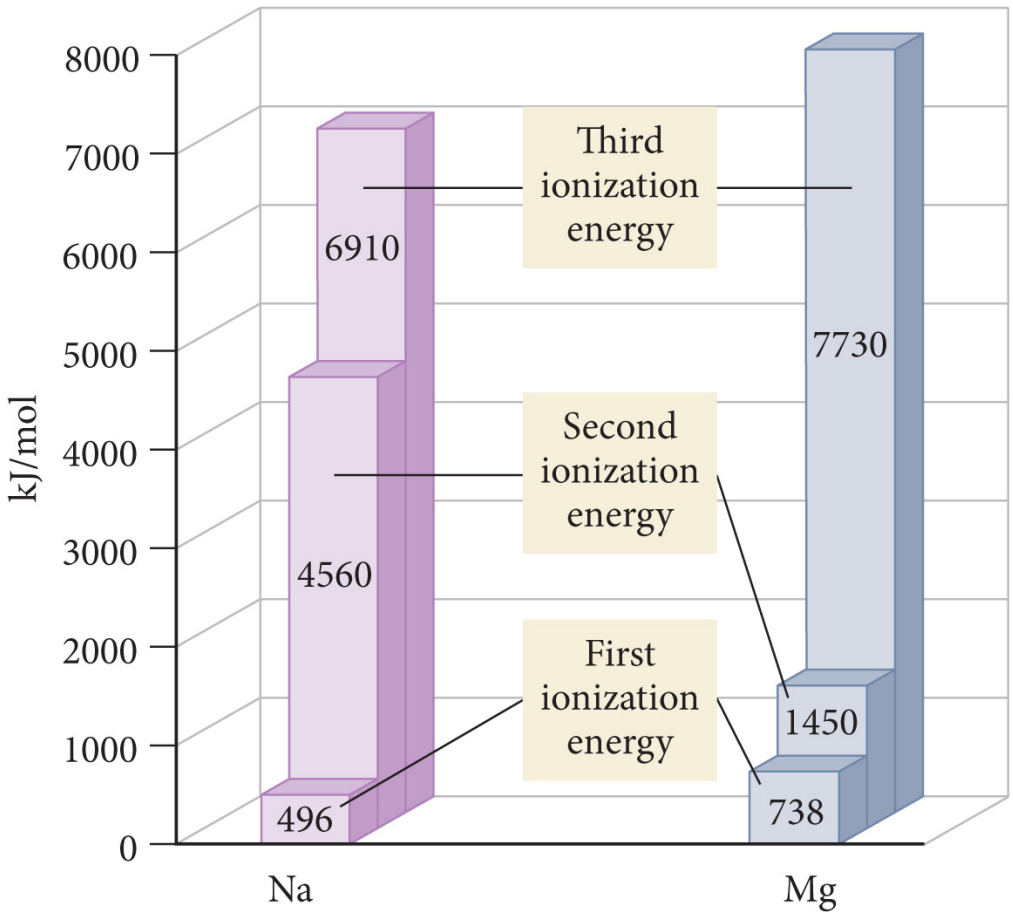
- **Moving left to right, *IE* increases.**
  - $Z_{\text{efffective}}$ is increasing, so the electron is more strongly attracted to the nucleus and it takes more energy to remove it.
- **Going down a group, *IE* decreases.**
- **Second ionization energy (IE_2):**

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

- Successive *IE*s are defined in a similar manner.
- Comparing *IE*s: $IE_1 < IE_2 < IE_3$ etc…
  - As charge increases, it is harder to remove an electron.

Ionization Energy (contd.)

- Compare ionization energies of Na and Mg:



Ionization Energy (contd.)

- IE_1 is smaller for Na, as we would predict from periodic trends. But Na has a *much bigger* IE_2.
- Configurations:

  Na: $1s^2 2s^2 2p^6 3s^1$

  Mg: $1s^2 2s^2 2p^6 3s^2$
- Mg can lose two $3s$ electrons. orbital. For Na, the 2nd electron comes from inner $2p$ orbital – *much* harder to remove.
- **In general:** there is an ***especially large jump*** in IE after an element's <u>last outer electron</u> has been lost.

8.8 Electron Affinity

- When an atom adds an electron to form an anion, usually energy is released (exothermic process, $\Delta H < 0$).
- This energy is the **electron affinity:**

$$1^{st}\ EA = \Delta H \text{ for: } A\,(g) + e^- \rightarrow A^-\,(g)$$

- So the more negative the EA, the "more badly" an atom wants to add an electron.
  - Sometimes EAs are written as a **positive number** (the negative of the ΔH). In this case a more positive EA means adding an electron is more favorable.

Trends in Electron Affinity

- The trends are less regular (Figure 8.16), but in general, **moving right, *EA* increases** (becomes more negative).
- **Going down a group *EA* decreases**.

Electron Affinities (kJ/mol)

1A	2A	3A	4A	5A	6A	7A	8A
H −73							He >0
Li −60	Be >0	B −27	C −122	N >0	O −141	F −328	Ne >0
Na −53	Mg >0	Al −43	Si −134	P −72	S −200	Cl −349	Ar >0
K −48	Ca −2	Ga −30	Ge −119	As −78	Se −195	Br −325	Kr >0
Rb −47	Sr −5	In −30	Sn −107	Sb −103	Te −190	I −295	Xe >0

- Notice that *EA* reaches a maximum in Group **7A** (not 8A)
- Noble gases have very low *EA* – they do not want to add electrons since they have full shells.

Trends in Metallic Character

- Elements become **more** metallic as you go down a group, and **less** metallic as you go left to right across a period (Figure 8.18).
- Since metals tend to lose electrons, elements with lower IE will be more metallic.