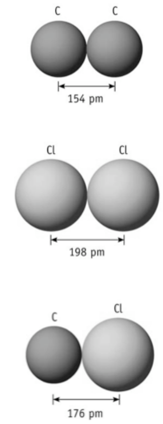
Periodic Properties and Trends

Chapter 8

Atomic Radius (Size)

- The atomic radius of an element can be defined as half the distance between the nuclei of a diatomic molecule of the element.
- *For example, the atomic radius of carbon is 77 picometers (7.7×10^{-11} m).*

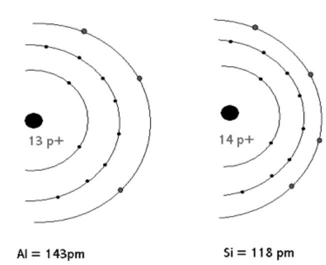


We might expect the atomic radii of elements to increase as we move from LEFT to RIGHTacross the periodic table. However, experimentally, this is not observed. ***Why not?***

Effective nuclear charge (Z_{eff}) **and Shielding**

- ***$Z_{effective}$*** is the average shielded charge felt by an electron.
- ***Shielding*** in an atom is the “blocking” or canceling of positive charge by core electrons.

- As we move from **left → right** on the periodic table, shielding of the nucleus by inner (core) electrons remains constant. *(Shielding by other outer electrons is minimal.)*
- However, as the number of protons increases, the electrons in the valence (outer) shell experience a greater attraction for the nucleus – a higher $Z_{effective}$.
- Because of the greater pull on the outer electrons as we move across the periodic table, atomic radius generally DECREASES from **left →right** within a period.



Atomic Radii (pm)

As we move from **top → bottom** on the periodic table, we would expect the atomic radius to INCREASE, and it does, because the energy level of the outer e^{-}'s increases.

Ionic Radius

- Reported Ionic radii are determined from the measured distances between nuclei in ionic compounds.
- ***Cations*** are always ***smaller*** than the atom they are derived from. In many cases, the outer energy level of the cation is one energy level lower than that of the atom.
- ***Anions*** are always ***larger*** than the atom they are derived from, primarily due to repulsive forces of electrons in the outer energy level.

Ionic Radii (pm)

- The **L→R** trend for ionic radius is DECREASING *for the same type of ion* – cation or anion.
- The **T→B** trend is INCREASING.

Examples: Atomic and Ionic Radii

- Put the following species into order of DECREASING particle size:

1. P P^{3-} S S^{2-}
2. Rb Rb^{+} Sr   I I^{-}
3. B   C   Al

Ionization Energy

- *First Ionization Energy* – the energy required to remove the highest energy electron from an atom.
- *Second Ionization Energy* – the energy required to remove the second highest energy electron (from the (+1) ion.
- *Third, Fourth, Fifth, etc.*
- Ionization energies are positive, as they refer to the energy going into the atom (system) to remove an electron.

1st Ionization Energy

- **L→R** trend: 1st ionization energy INCREASES, because the pull on the outermost electron increases.
- **T→B** trend: 1st ionization energy DECREASES, because the outermost electron is at a higher energy level.

2nd Ionization Energy and higher…

A "jump" between subsequent ionization energies is observed when a STABLE electron configuration is reached.

Beryllium 1st ionization:

Beryllium 2nd ionization:

Element	First	Second	Third	Fourth	Fifth	Sixth	Seventh
H	1312						
He	2372	5250					
Li	520	7298	11,815				
Be	899	1757	14,848	21,006			
B	801	2427	3660	25,025	32,826		
C	1086	2353	4620	6222	37,829	47,276	
N	1402	2857	4578	7475	9445	53,265	64,358
O	1314	3388	5300	7469	10,989	13,326	71,333
F	1681	3374	6020	8407	11,022	15,164	17,867
Ne	2081	3952	6122	9370	12,177	15,238	19,998

Examples – Ionization energy

- Put the following in order of increasing 1st ionization energy:

  Li Be Na

- Where would magnesium fit in?

- Which of the following will have the largest jump in ionization energy between its 4th and 5th ionization energies?

  Na Al Si S

Electron Affinity

- The energy released (negative value) when an atom of an element in the gas phase takes on an electron to become a negative ion.
- ***L→R trend:*** Generally increasing, because the attraction for the outermost electrons increases from left to right.
- ***T→B trend:*** No reliable trend. Generally decreasing, but trend is not highly predictable from top to bottom.

Table of Electron Affinities

Electron Affinities of the Main-Group Elements (kJ/mol)*

Period	IA	IIIA	IVA	VA	VIA	VIIA
1	**H** −73					
2	**Li** −60	**B** −27	**C** −122	**N** 0	**O** −141	**F** −328
3	**Na** −53	**Al** −44	**Si** −134	**P** −72	**S** −200	**Cl** −349
4	**K** −48	**Ga** −30	**Ge** −120	**As** −77	**Se** −195	**Br** −325
5	**Rb** −47	**In** −30	**Sn** −121	**Sb** −101	**Te** −190	**I** −295
6	**Cs** −45	**Tl** −30	**Pb** −110	**Bi** −110	**Po** −180	**At** −270

*Atoms of the alkaline earth metals (Group IIA) and the noble gases (Group VIIIA) do not form stable negative ions.

Electronegativity

- A measure of the relative attraction that a nucleus has on electrons in a chemical bond - the larger the number, the greater the attraction.

  L→R trend: INCREASES

  T→B trend: DECREASES

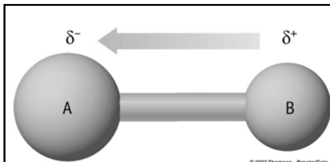
- If two atoms are in a bond, the difference in their electronegativities will determine if the electrons are shared evenly or unevenly.

Electronegativity and Bond Polarity

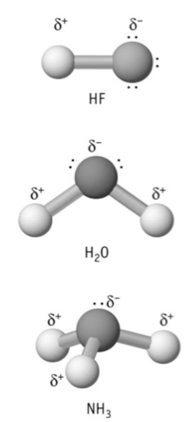
- The greater the electronegativity difference (ΔEN) between two atoms in a bond, the more polar the bond.
- If the difference is very great, then one atom essentially pulls the electrons away from the other making an ionic bond.
- There is a continuum of non-polar to polar to ionic compounds, with no clear breaks.
- However, we will define somewhat arbitrary cut-off points:

ΔEN ≤ 0.4	Non-polar covalent
0.5 ≤ ΔEN ≤ 1.6	Polar covalent
ΔEN ≥ 1.7	Ionic

The cutoff for "ionic" is the point at which a bond has 50% ionic character.



- If the electrons are shared evenly, the bond is ***non-polar***.
- If the electrons are shared unevenly, the bond is considered ***polar*** – **it has a positive end and a negative end.**



Summary of Periodic Properties

Property	***L→R***	***T→B***
$Z_{effective}$		
Shielding		
Electron Energy Level		
Atomic Radius		
Ionic Radius		
Ionization Energy		
Electron Affinity		
Electronegativity		