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# Pre-Health Post-Baccalaureate Program
# Study Guide and Practice Problems

Course: CHM 2045

Textbook Chapter: 8.3 – 9.3 (Silberberg 6e)

Topics Covered: Periodic Trends
Chemical Bonding
Models

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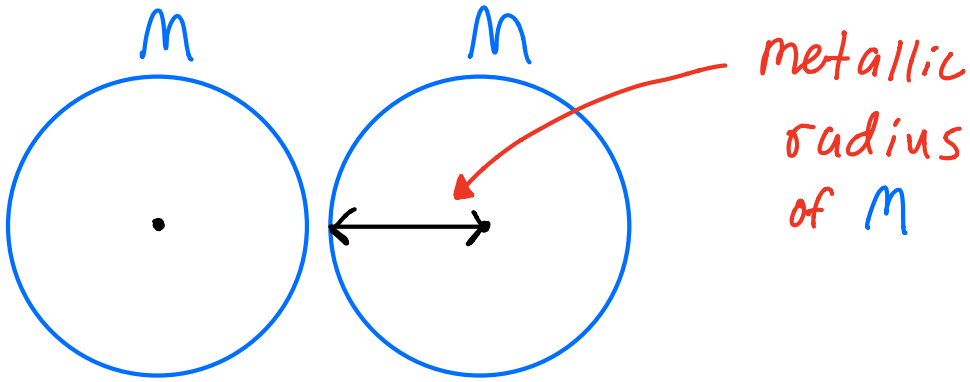
# Periodic Trends

- Periodic trends are patterns within a period or group of the periodic table which inform us of an element's properties in relation to their location on the table, and relative to other elements.
- It is important to not only know the periodic trends, but to understand conceptually why these trends exist.

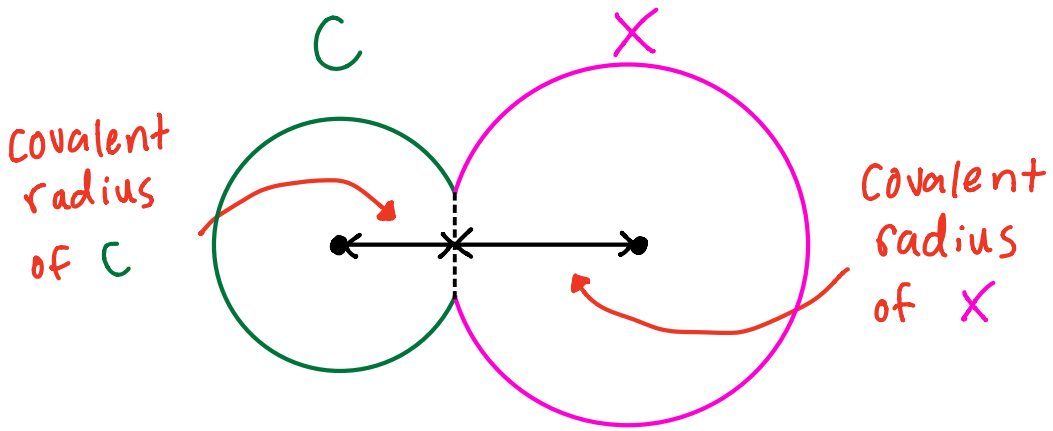
## ① Atomic Size

- There are two ways we can think about atomic size: the metallic radius (used primarily for, you guessed it, metals) and the covalent radius (used primarily for molecules
- metallic radius:



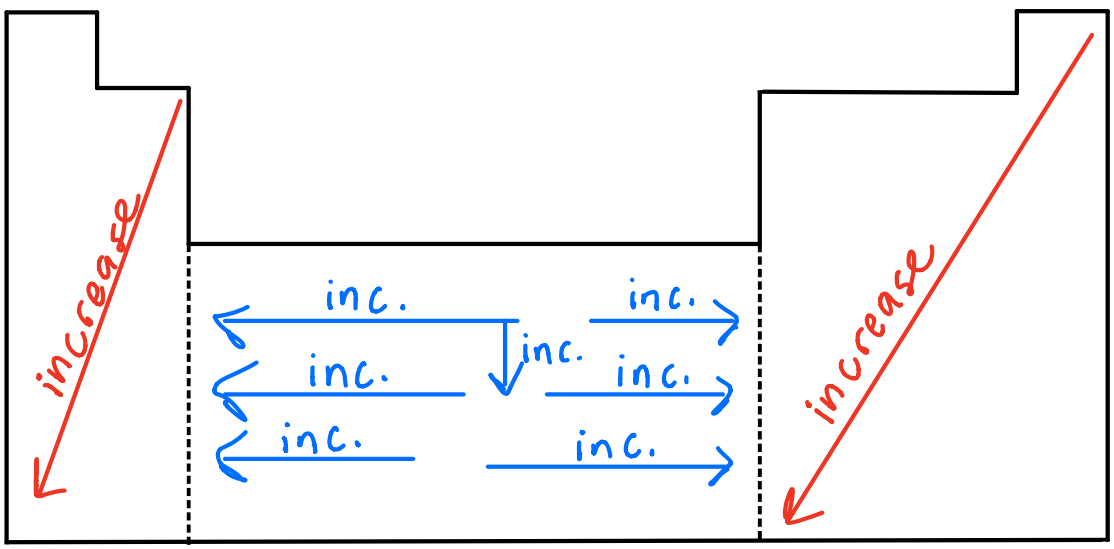

It is equal to half of the distance between two metal atoms within a crystal structure.

– Covalent radius:




It is equal to half of the distance between to non-metal atoms which are covalently bonded.

– Trend:



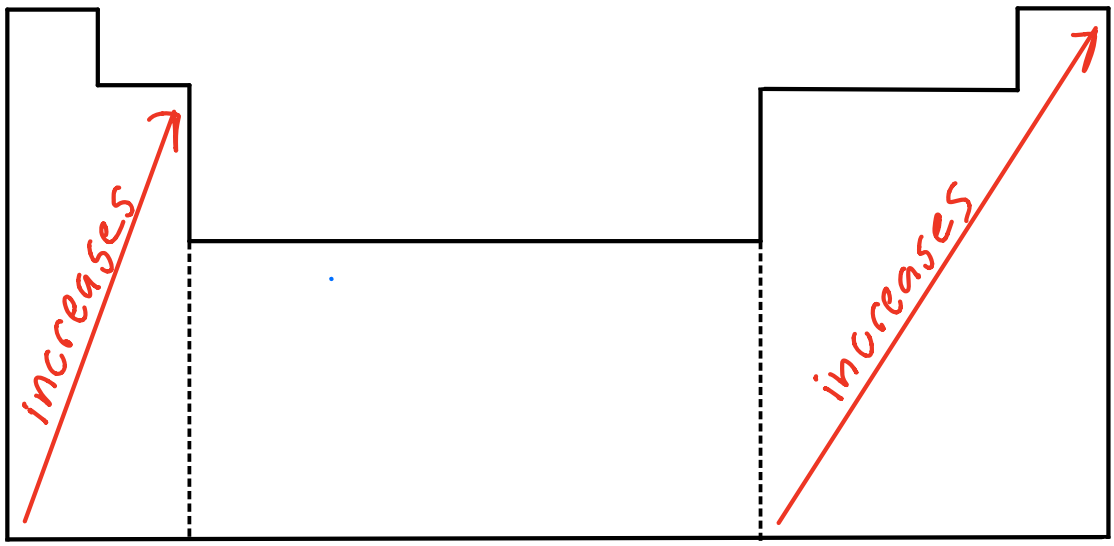
We will focus on the trend within the main group elements (the trend is less consistent with transition metals).
Down a group, atomic radius increases due to increased n value.
Across a period, atomic radius decreases due to increased $Z_{eff}$.

(2) Ionization energy

— Ionization energy is the amount of energy required to "strip off" one mole of electrons from one mole of gas atoms (or ions)

you can think about it another way: how hard is it to pull an electron off of an atom of a certain element?

- Trend:




Up a group, the atomic size decreases, and the nucleus is closer to the outermost electron, causing stronger attraction. Because the electron is "held

onto" more tightly, more energy is required to pull it off. Across a period, atomic size decreases while $Z_{eff}$ increases, causing stronger attraction. Once again, because the electron is "held onto" more tightly, more energy is required to pull it off.

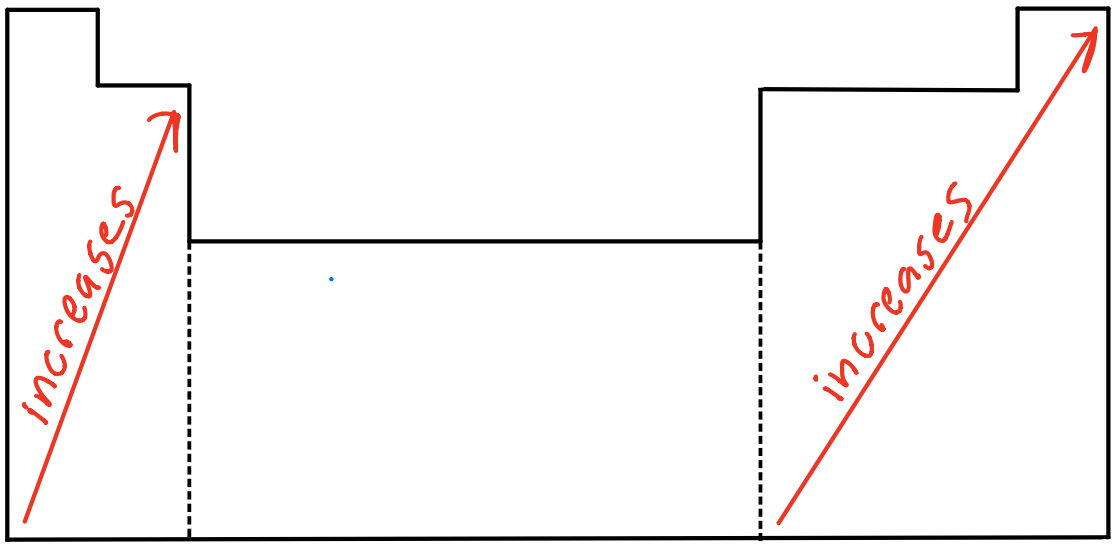
- With each additional electron removed, the successive IE's increase ($IE_1 < IE_2 < IE_3$)

③ Electron affinity

- Electron affinity is the opposite of IE: how much energy does it take to add one mole

of electrons to one mol of gas atoms (or ions).

- Trend:




This trend is not as consistent as the previous two. Just know that this is the general pattern.

# Models of Bonding

## ① Ionic:

- Occurs between a metal ion (⊕ charged due to loss of electron(s)) and a non-metal (⊖ charged due to gain of electron(s))
- The attraction between ⊕ and ⊖ create an ionic solid arranged as a structured lattice
- General formula:

$$M_{(g)} \rightarrow M^+_{(g)} + \cancel{e^-}$$

$$X_{(g)} + \cancel{e^-} \rightarrow X^-_{(g)}$$

$$M_{(g)} + X_{(g)} \rightarrow M^+_{(g)} + X^-_{(g)}$$

- Ion formation costs energy, while solid formation by separate ions releases energy

② Covalent bonding

- Covalent bonding arises through atoms' desire to reach stability through an octet of electrons. To do this, two atoms share a pair of electrons (called a "bonding pair").
- If more of these bonds form, a molecule can have double or triple bonds

- Triple bonds are shorter than double bonds, which are shorter than single bonds
- Triple bonds are stronger than double bonds, which are stronger than single bonds

# Practice Problems

(1) Rank the following in terms of decreasing ionization energy:

I. Alkaline earth metals

II. Halogens

III. Alkali metals

a) I > II > III

b) II = I > III

c) II > III > I

d) II > I > III

(2) Explain, in your own words, why successive IEs are larger than $IE_1$.

(3) Which of the following bonds between a group 1A element and a halogen is an ionic bond?

I. H - Cl

II. H - Br

III. Na - Cl

IV. K - Br

a) III only

b) IV only

c) III and IV only

d) All

(4) Rank the following based on increasing radius:

I. $Cl$

II. $Cl^-$

III. $Cl^+$

a) III < I < II

b) II < I < III

c) All have the same radius

d) It is impossible to determine without experimental data

# Solutions

① Rank the following in terms of decreasing ionization energy:

I. Alkaline earth metals

II. Halogens

III. Alkali metals

a) I > II > III

b) II = I > III

c) II > III > I

d) II > I > III

IE increases towards the top right of the P.T.

(2) Explain, in your own words, why successive IEs are larger than $IE_1$.

With each electron stripped away, the ion becomes more positively charged and holds onto remaining electrons more tightly. More energy is required to break this attraction.

(3) Which of the following bonds between a group 1A element and a halogen is an ionic bond?

I. H - Cl (g)

II. H - Br (g)

III. Na - Cl

IV. K - Br

a) III only

b) IV only

c) III and IV only

d) All

Though hydrogen is group 1A, it is a nonmetal and therefore forms covalent bonds when bound to Cl and Br.

④ Rank the following based on increasing radius:

I. $Cl$

II. $Cl^-$

III. $Cl^+$

a) III < I < II

b) II < I < III

c) All have the same radius

d) It is impossible to determine without experimental data

$Cl \rightarrow 17\,P,\ 17\,E$

$Cl^- \rightarrow 17\,P,\ 18\,E$

$Cl^+ \rightarrow 17\,P,\ 16\,E$

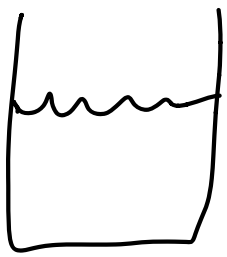
Same # of P... therefore more electrons means larger radius

NaCl

Na — Cl

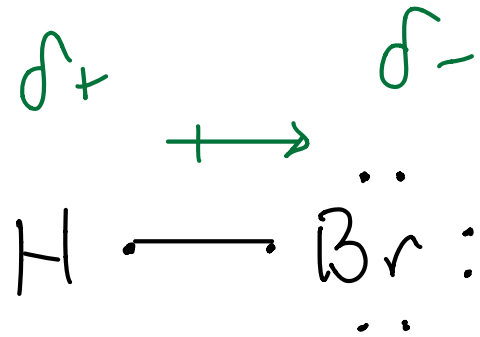
$Na^{\circ}$ + $:\ddot{\underset{\cdot\cdot}{Cl}}:$ (g)

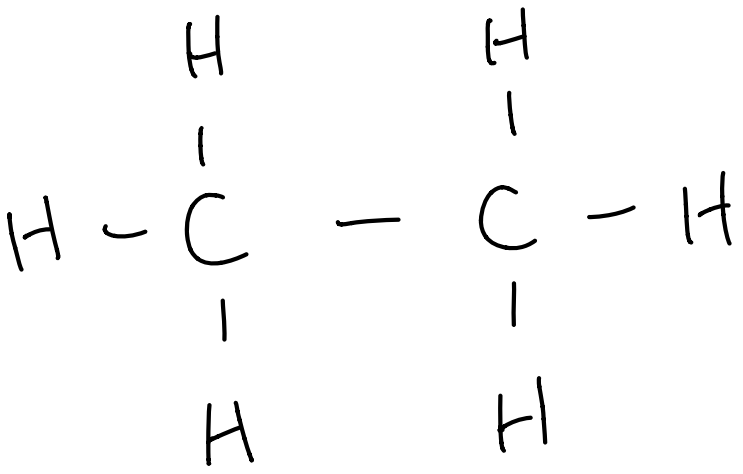
$Na^+$ $Cl^-$



$Na^+$ ---------- $Cl^-$

$Na^+$ ---------- $Cl^-$ ---------- $Na^+$ ---------- $Cl^-$


δ+
δ−
H
Br


H
H
H
C
C
H
H
H
H