Lecture 3: Periodic Properties of the Elements & Basic Concept of Chemical Bonding (Part I)

Light-emitting diodes *(LEDs) can be made from semiconductors composed of Group 3A and Group 5A elements.*

Group 3A
13 Aluminum Al
31 Gallium Ga
49 Indium In

LEDs made of a mixture of

GaAs gives RED light,

GaP and AlP give GREEN light

GaN and InN give BLUE light.

Group 5A
7 Nitrogen N
15 Phosphorus P
33 Arsenic As

Periodic Properties of the Elements

Outline

- Introduction
- Effective Nuclear Charge
- Property 1: Sizes of Atoms and Ions
- Property 2: Ionization Energy
- Property 3: Electron Affinity
- Property 4: Metals, Metalloids and Nonmetals

Periodic Properties of the Elements

Outline

- Introduction
  - Development of Periodic Table
  - Periodicity of Chemical Elements
- Effective Nuclear Charge
- Property 1: Sizes of Atoms and Ions
- Property 2: Ionization Energy
- Property 3: Electron Affinity
- Property 4: Metals, Metalloids and Nonmetals

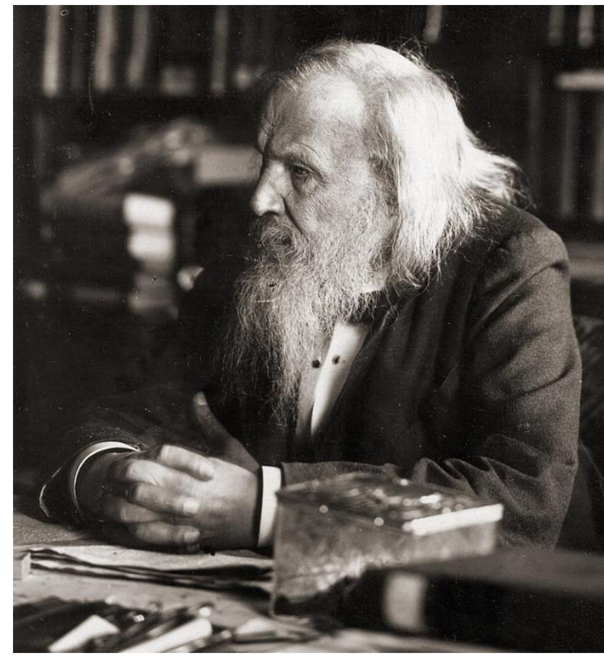
Intro: Development of Periodic Table

- Certain elements, such as Fe/Ag/Au…, appear in nature in elemental form and were thus discovered thousands of years ago.
- However, the majority of elements are not found in nature in their elemental form. Scientists were thus unaware of their existence for centuries.
- From 1735 onwards, advances in chemistry made it easier to isolate elements from their compounds.

H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	[113]	Fl	[115]	Lv	[117]	[118]

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No

Ancient Times	Middle Ages–1700	1735–1843	1843–1886	1894–1918	1923–1961	1965–
(9 elements)	(6 elements)	(42 elements)	(18 elements)	(11 elements)	(17 elements)	(15 elements)



In 1869, **Dmitri Mendeleev** (Russia) formulating the Periodic Law and creating a farsighted version of the periodic table of elements.

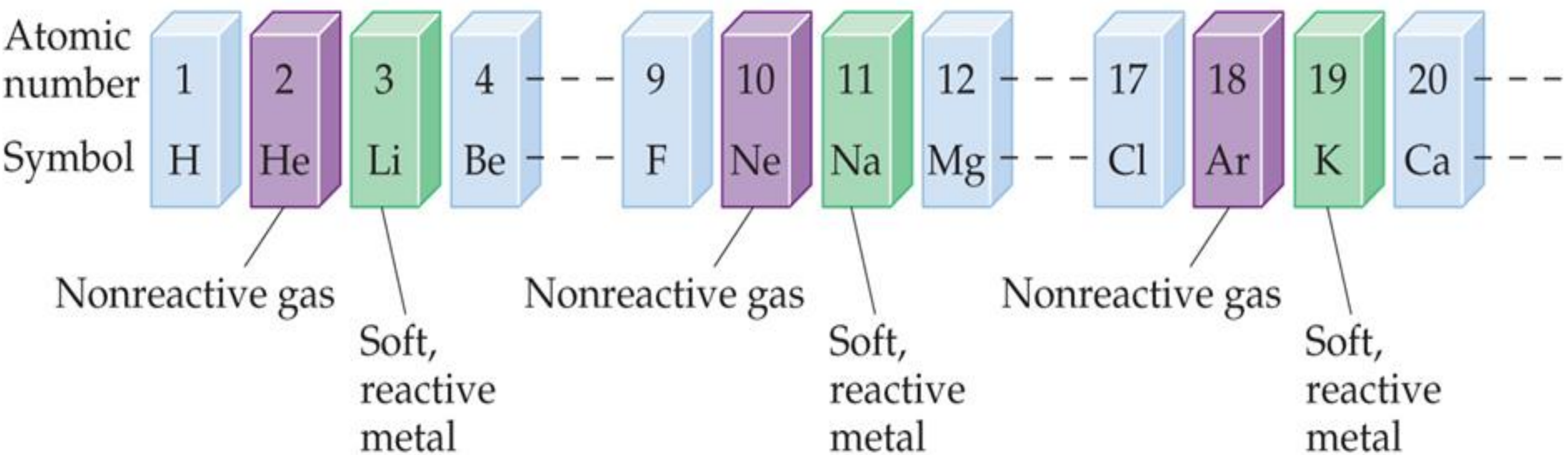
Intro: Development of Periodic Table

- English physicist **Henry Moseley (1887–1915)** developed the concept of ***atomic numbers***.
- Moseley found that **each element produced X rays of a unique frequency** by bombarding different elements with high-energy electrons.
- He arranged the X-ray frequencies in order by assigning a unique whole number, called an ***atomic number***, to each element.

Group → / Period ↓	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La *	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac **	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
			*	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
			**	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Intro: Periodicity of Chemical Elements

- We now know that *atomic number (Z)* is a key property of element.
- It indicates the number of **protons/electrons** of a given **neutral element**.
- When the elements are arranged in order of increasing *Z*, a repeating pattern of chemical properties emerges.



Similar chemical/physical properties recur periodically when the elements are arranged in order of increasing *atomic number*.

Intro: Periodicity of Chemical Elements

To understand this....we need to apply the electron configuration of atoms (from Lecture 2) for a deeper understanding

Element	Full electron configuration	Condensed electron configuration
Lithium $_3Li$	$1s^2 2s^1$	$[He]2s^1$ where [He] represents the electron configuration of helium.
Sulfur $_{16}S$	$1s^2 2s^2 2p^6 3s^2 3p^4$	$[Ne]3s^2 3p^4$ where [Ne] represents the electron configuration of neon.
Bromine $_{35}Br$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$	$[Ar]4s^2 3d^{10} 4p^5$ where [Ar] represents the electron configuration of argon.

We focus on **the valence (outermost shell) electrons** because it is these electrons that **mainly** determine chemical properties as they are responsible for chemical reactivity/chemical bonding.

But **the core (inner shell) electrons** also **minorly** affect atomic properties. They do so indirectly by affecting the extent to which the valence electrons "see" the nucleus.

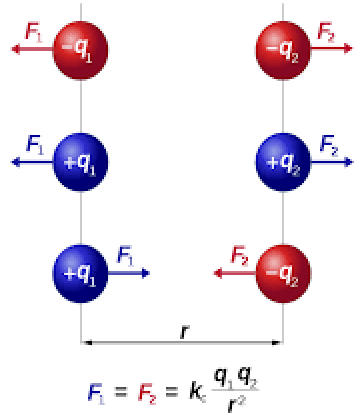
The best way to deal with this is by means of the unifying KEY concept of "Effective Nuclear Charge"

Periodic Properties of the Elements

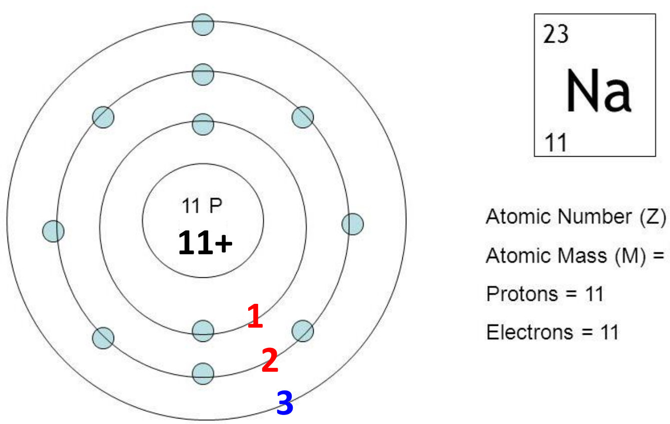
Outline

Effective Nuclear Charge (Z_{eff})

- **Coulomb's law** tells us that the strength of the interaction between two electrical charges depends on **the magnitudes of each charge** and the **distance between them**.
- **For a given many-electron atom**, each **electron (negatively charged)** is attracted to the **nucleus (positively charged)** and repelled by other electrons.



Sodium Atom



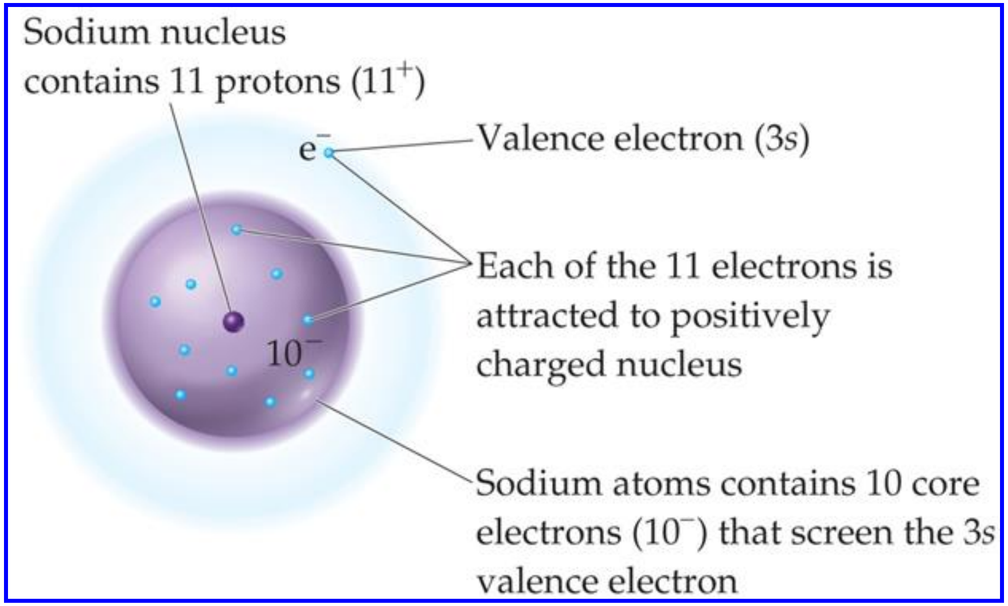
Atomic Number (Z) = 11

Atomic Mass (M) = 23

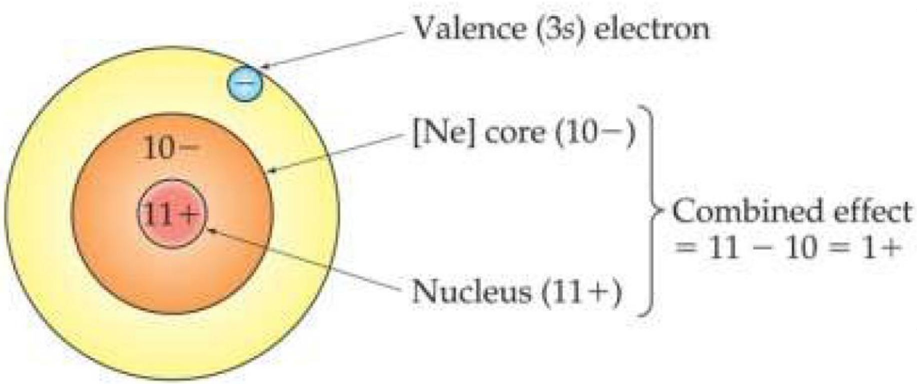
Protons = 11

Electrons = 11

Na: $1s^2 2s^2 2p^6 3s^1$ or $[Ne]3s^1$



Effective Nuclear Charge (Z_{eff})



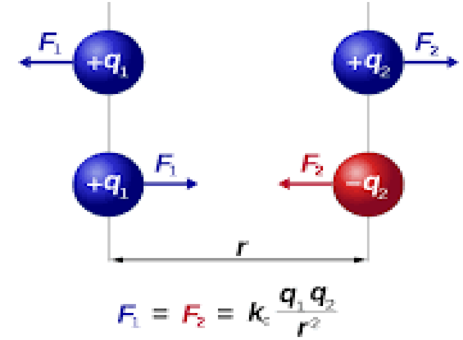
$_{11}$Na ($1s^2 2s^2 2p^6 3s^1$)

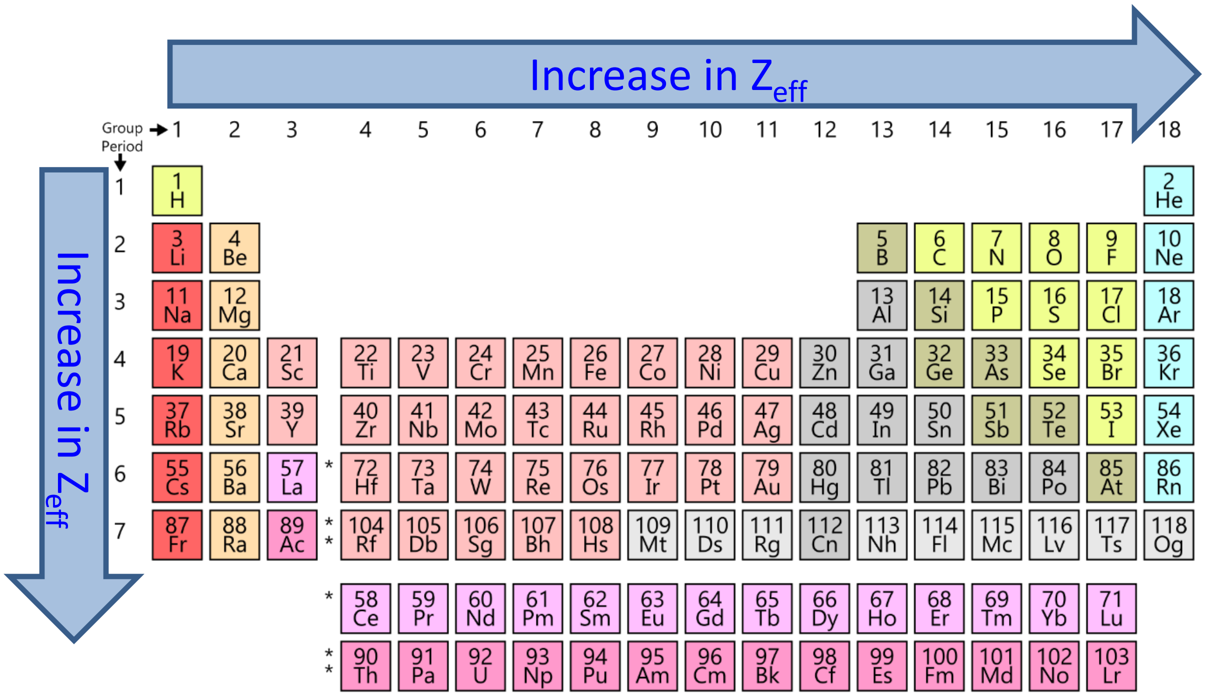
- Considering a sodium atom, if the 10 core electrons completely ***screen*** the actual 11+ nuclear charges, then the valence 3*s* electron will experience a net nuclear charge of *about* 1+.
- This net 1+ charge is the ***effective nuclear charge*** (***Z_{eff}***) for the valence 3s electron of the sodium atom.

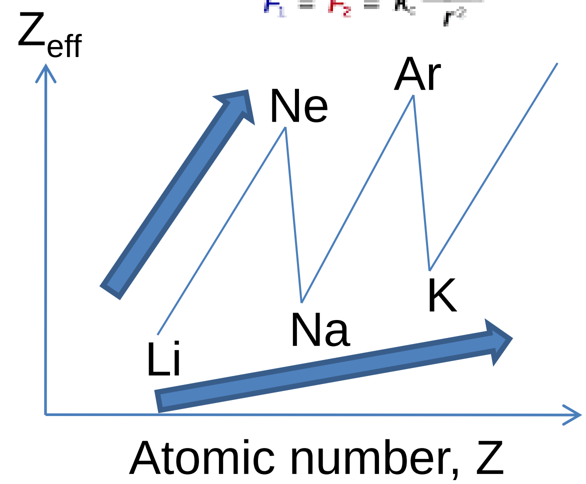
Example 1: $_3$Li ($1s^2 2s^1$)
3+ nucleus & 2*core electrons
The valence electron (i.e. the one in 2s) experiences a ***Z_{eff}*** of *about* 3 - 2 = 1+.

Example 2: $_4$Be ($1s^2 2s^2$)
4+ nucleus & 2*core electrons
The ***Z_{eff}*** experienced by **each** 2*s* electron is *about* 4 - 2 = 2+.

Effective Nuclear Charge (Z_{eff})







- Valence electrons in the same shell screen one another ineffectively

Li & Be

*Therefore Z_{eff} increases with increasing atomic number across any **period***
(because the increase in nucleus charge dominant the Z_{eff} while the increased repulsion between valence electrons is relatively minor).

- Larger electron cores are less able to screen the valence electrons from the nuclear charge.

Li & Na

*Therefore Z_{eff} **slightly** increases with increasing atomic number down any **group***
(because the average nucleus-electron distance is also increasing).

Effective Nuclear Charge (Z_{eff})

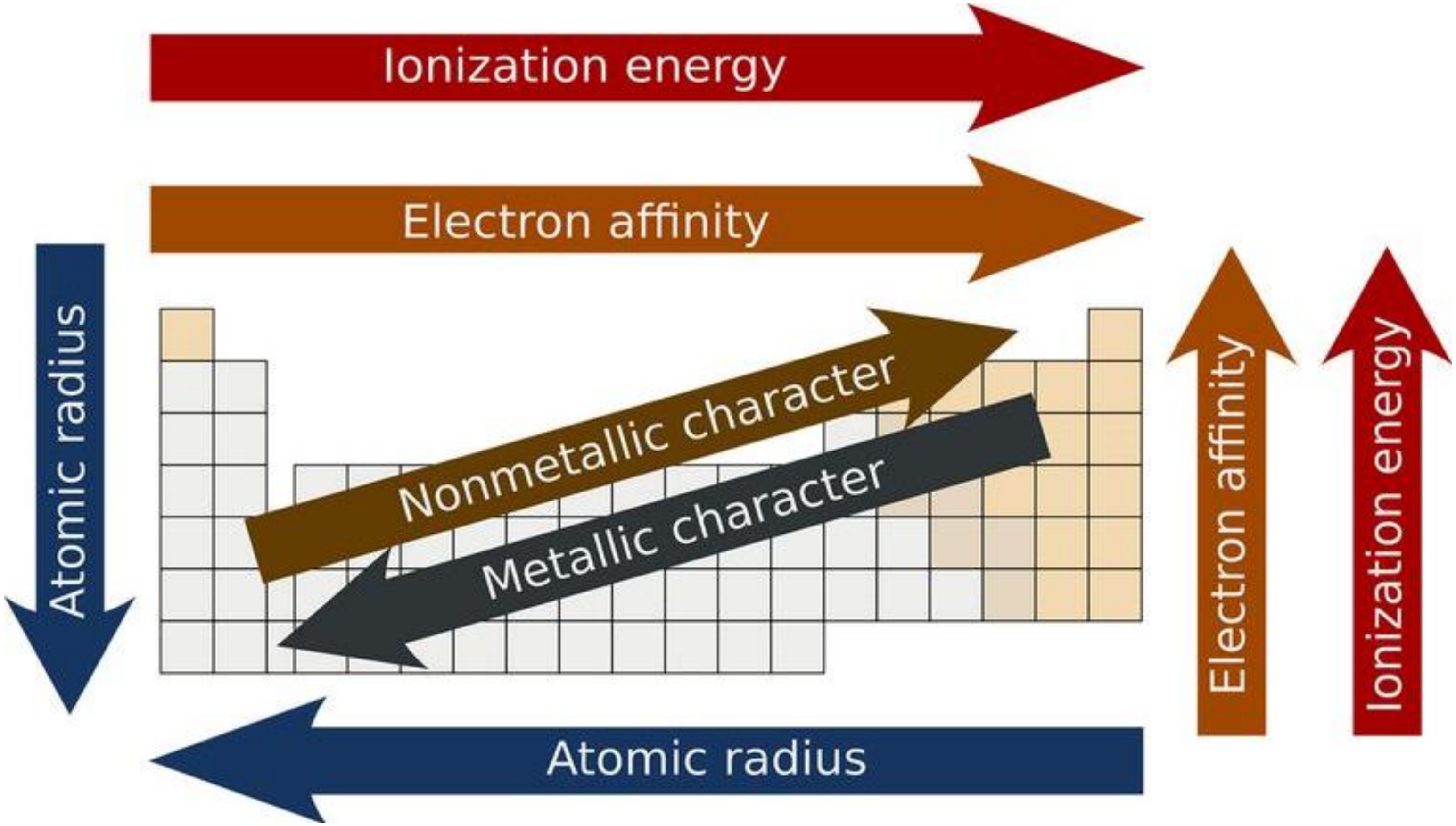
By knowing…

The periodic Z_{eff} trend of elements

We can now rationalize observed periodic properties in

- Sizes of Atoms and Ions,
- Ionization Energy,
- Electron Affinity,
- Metal, Metalloids and Nonmetals

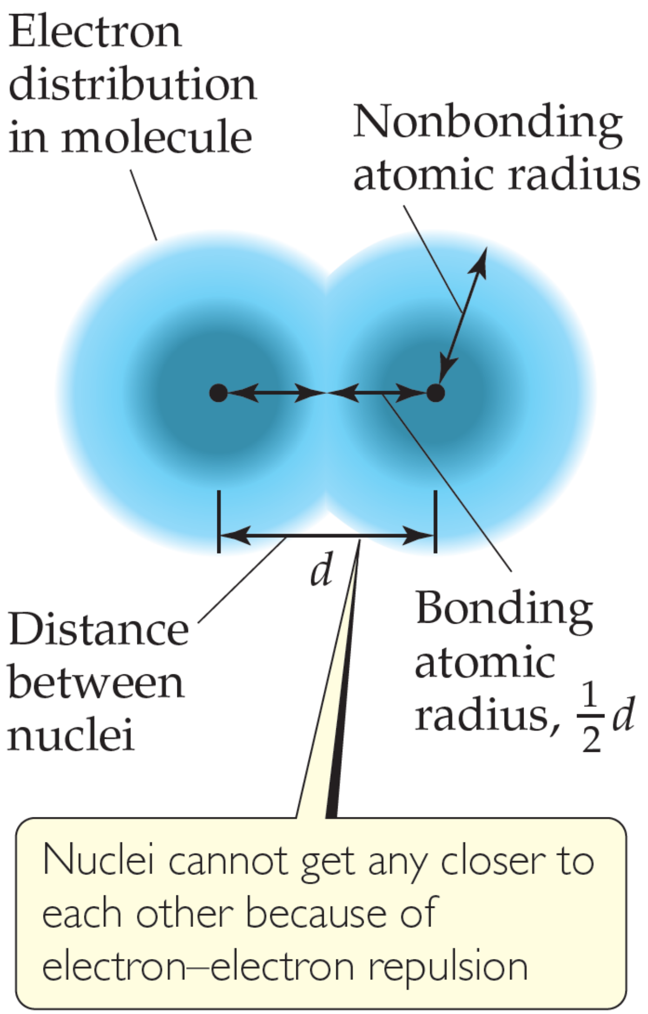
which can reflect the chemical reactivity of the elements.



Periodic Properties of the Elements

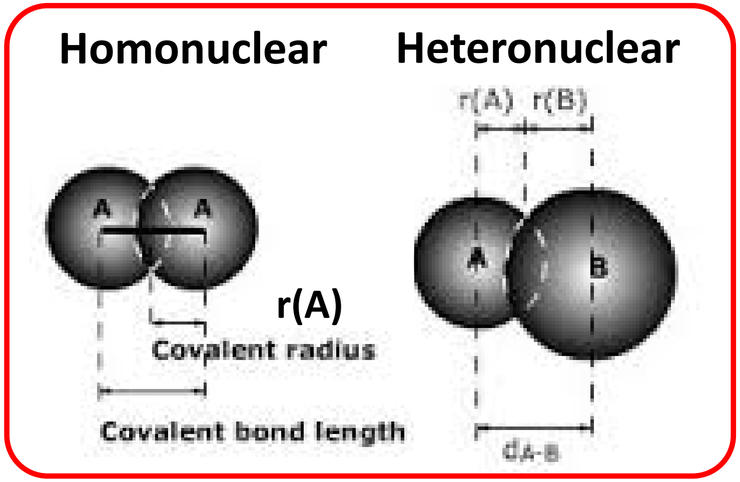
Outline

Sizes of Atoms: Bonding Atomic Radius

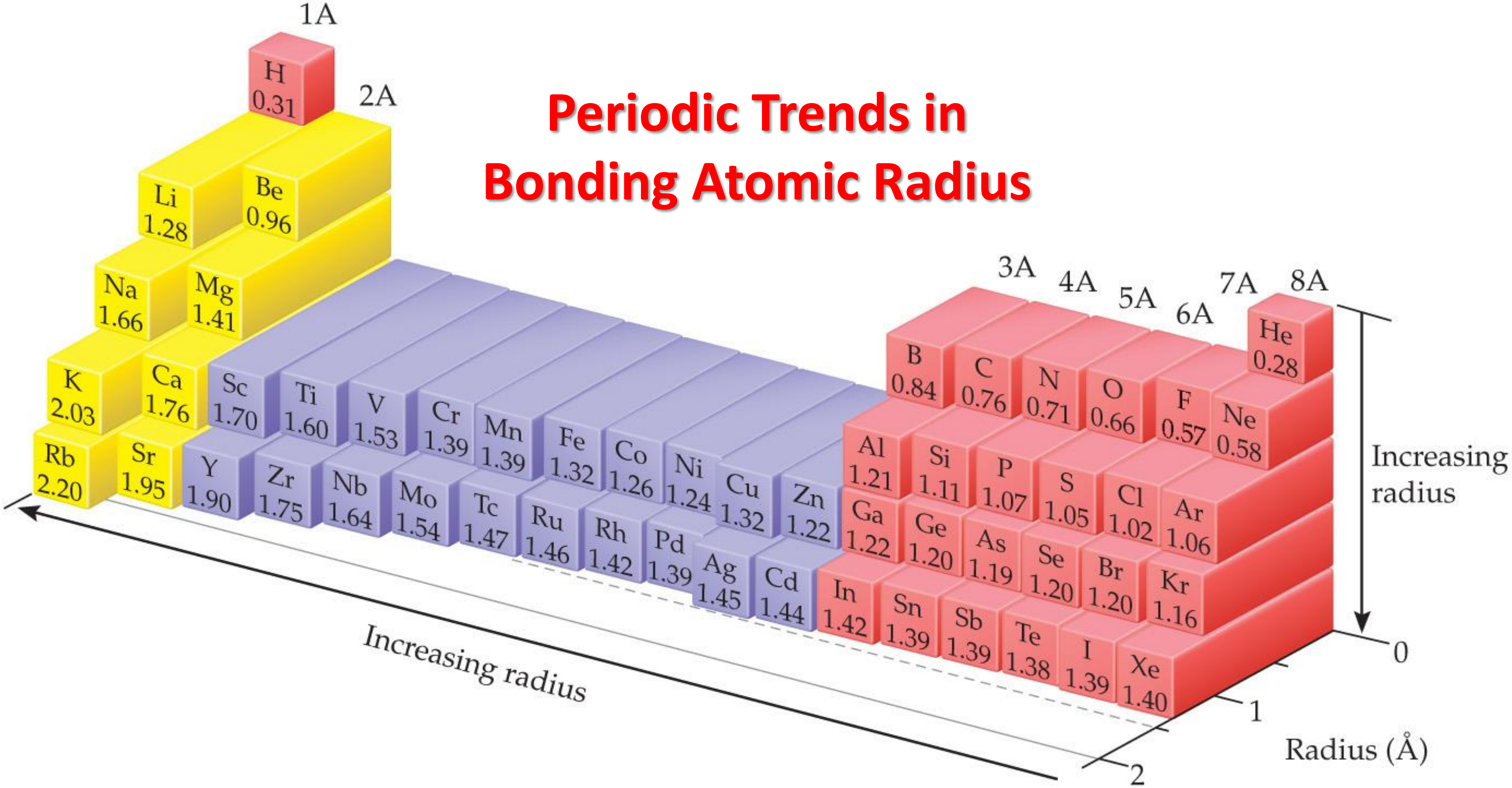


For homonuclear molecule…

The **bonding atomic radius** is defined as one-half of the distance (d) between covalently bonded nuclei.



Sizes of Atoms: Bonding Atomic Radius

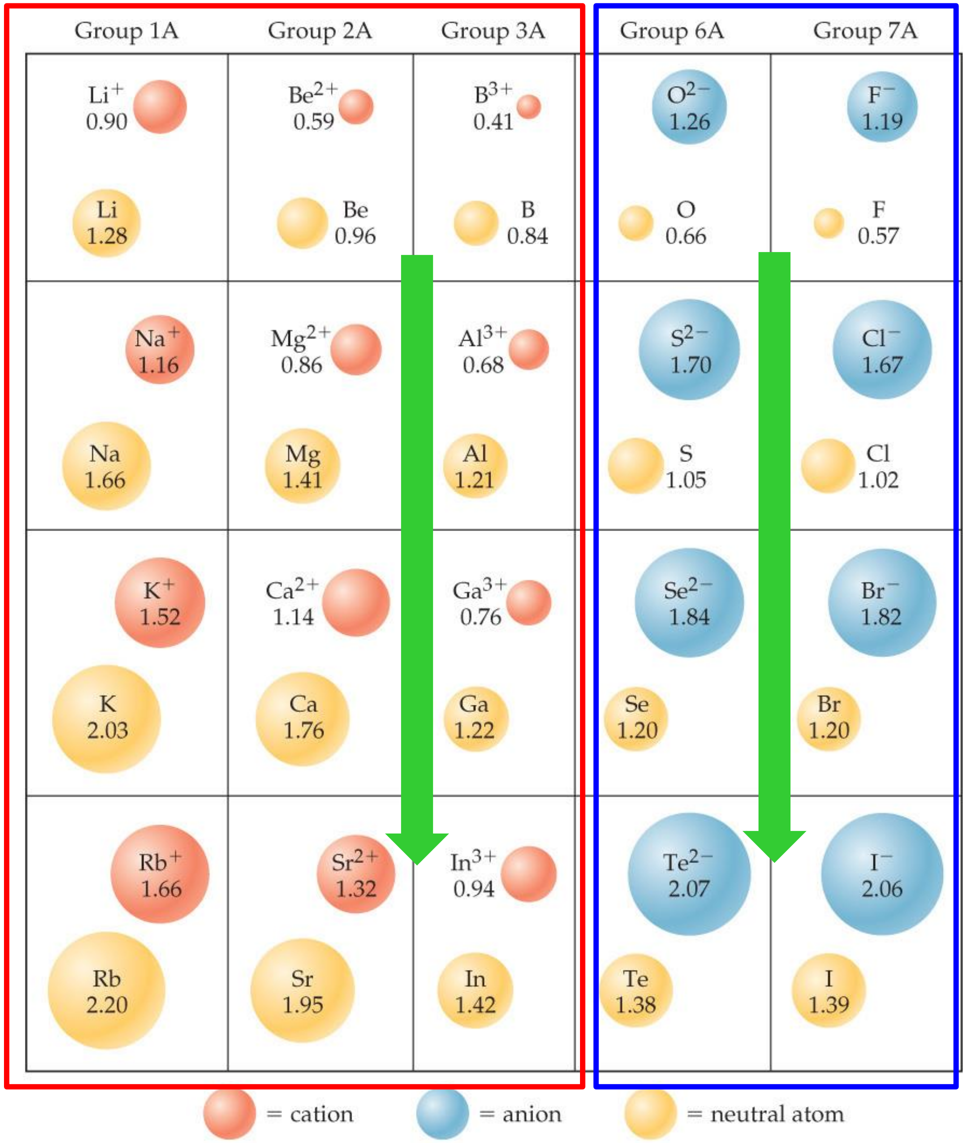


Bonding atomic radius tends to:

- decrease with increasing atomic number across ***a period*** (due to <u>increasing Z_{eff}</u>) .
- increase with increasing atomic number down ***a group*** (due to increasing value of the <u>principal quantum number *n*</u> of the outer electrons, which are farther away from the nucleus).

Sizes of Ions: Ionic Radius

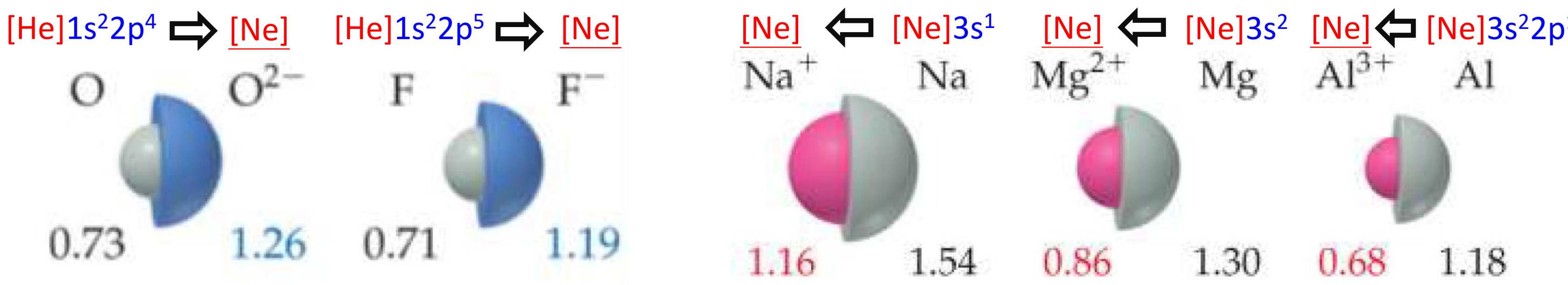
- Cations are smaller than their parent atoms because the outermost electrons are removed so that the repulsions between electrons are reduced.
- Anions are larger than their parent atoms because electrons are added so that the repulsions between electrons are increased.
- For ions carrying the same charge, Ionic radius increase in size as you go down a group because of the increasing value of *n*.



Sizes of Ions: Ionic Radius

Isoelectronic series

e.g. ions with [^{10}Ne] electron configuration for ^{8}O, ^{9}F, ^{11}Na, ^{12}Mg, ^{13}Al



—Increasing nuclear charge ⟶

O^{2-}	F^-	Na^+	Mg^{2+}	Al^{3+}
1.26 Å	1.19 Å	1.16 Å	0.86 Å	0.68 Å

—Decreasing ionic radius ⟶

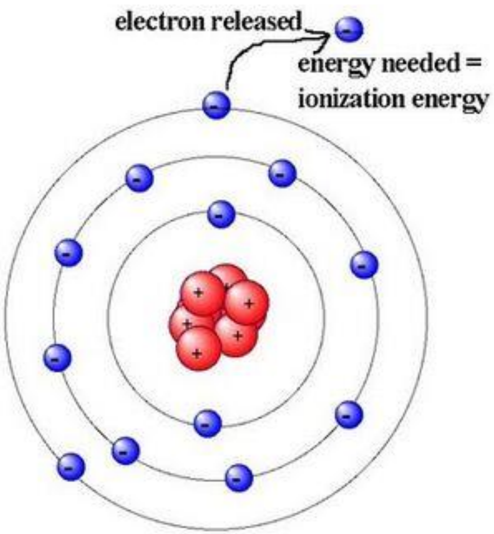
Because the number of electrons remains constant, ionic radius decreases with increasing nuclear charge (Z_{eff}) as the electrons are more strongly attracted to the nucleus.

Periodic Properties of the Elements

Outline

Ionization Energy

Ionization energy of an atom or ion is the *minimum* energy required to ***remove an electron (i.e. the outmost electron with highest energy)*** from the ground state of the gaseous atom or ion.



- **First ionization energy I_1:**
(remove an e- from neutral atom)

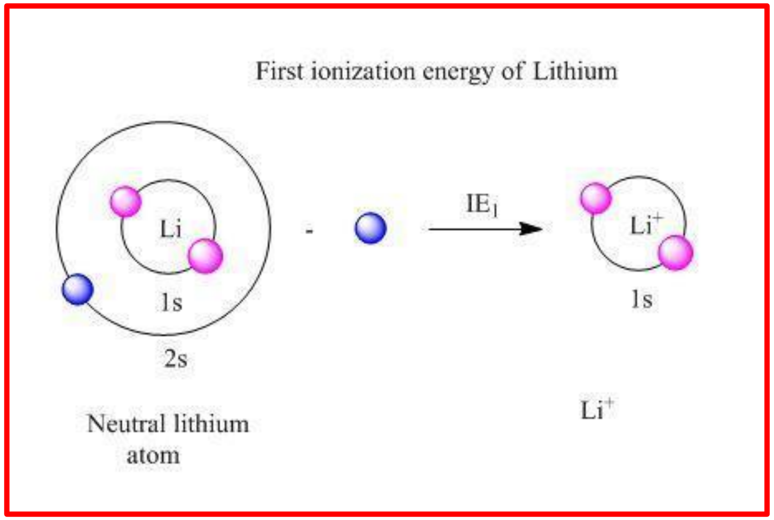
Li $(1s^2 2s^1) \longrightarrow$ Li$^+$ $(1s^2)$ + e^-

Cl $([Ne]3s^2 3p^5) \longrightarrow$ Cl$^+$ $([Ne]3s^2 3p^4)$ + e^-

- **Second ionization energy I_2:**

Li$^+$ $(1s^2) \longrightarrow$ Li^{2+} $(1s^1)$ + e^-

Cl$^+$ $([Ne]3s^2 3p^4) \longrightarrow$ Cl^{2+} $([Ne]3s^2 3p^3)$ + e^-



Ionization Energy: Elements in the Same Period

Table 7.2 Successive Values of Ionization Energies, I, for the Elements Sodium through Argon (kJ/mol) **Period 3!**

Element	I_1	I_2	I_3	I_4	I_5	I_6	I_7
Na	496	4562			(inner-shell electrons)		
Mg	738	1451	7733				
Al	578	1817	2745	11,577			
Si	786	1577	3232	4356	16,091		
P	1012	1907	2914	4964	6274	21,267	
S	1000	2252	3357	4556	7004	8496	27,107
Cl	1251	2298	3822	5159	6542	9362	11,018
Ar	1521	2666	3931	5771	7238	8781	11,995

(valence electrons)

Three trends can be obtained from this table!

- The first ionization energy generally increases with atomic number across a period (?)

  Valence electrons screen one another ineffectively, while Z_{eff} increases with increasing atomic number!

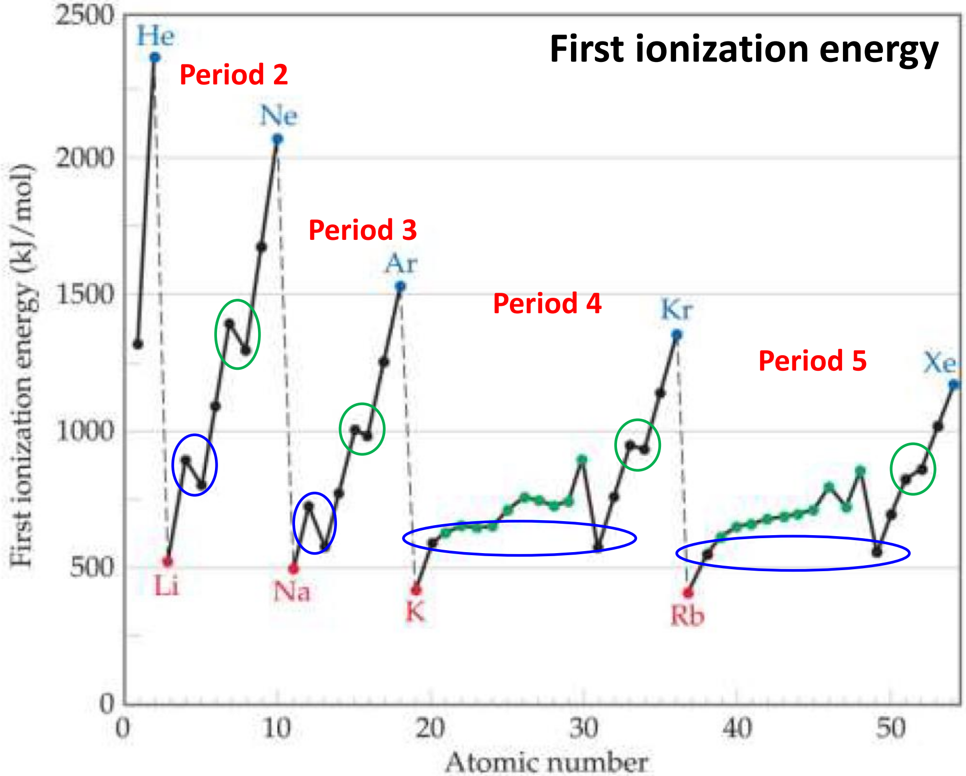
- Ionization energy increases with each successive removal of electron for an element (?)

  $I_1 < I_2 < I_3$…. Electron is being removed from an increasingly more positive ion (i.e. Z_{eff} is increased since the target electron is less screened now)

- A sharp increase in ionization energy occurs when a *core electron* is removed (?)

  Core electrons experience greater Z_{eff} than the valence electrons. (for Na^+/Mg^{2+}/Al^{3+}…possesses the electron configuration of [Ne])

Ionization Energy: Elements in the Same Period



In general, as one goes across a row (period), it gets harder to remove an electron.

–As you go from left to right, atomic number & Z_{eff} increases.

However, there are "two apparent discontinuities" in this trend

–The first occurs between group IIA and IIIA.

–The second occurs between group VA and VIA.

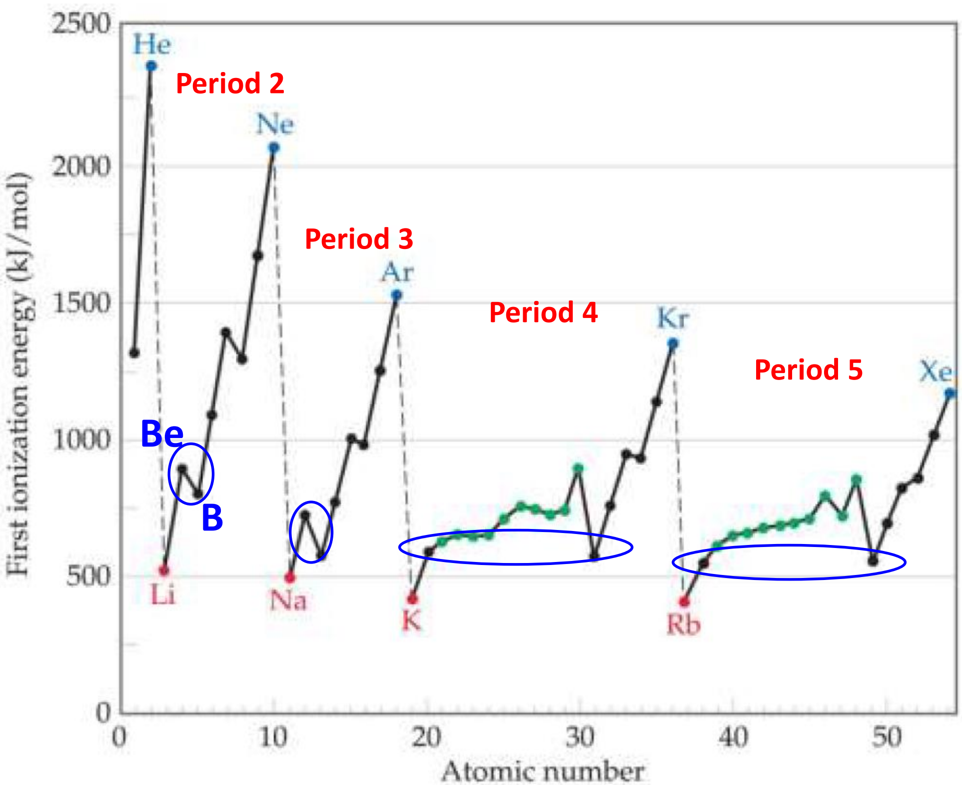
Ionization Energy: Elements in the Same Period

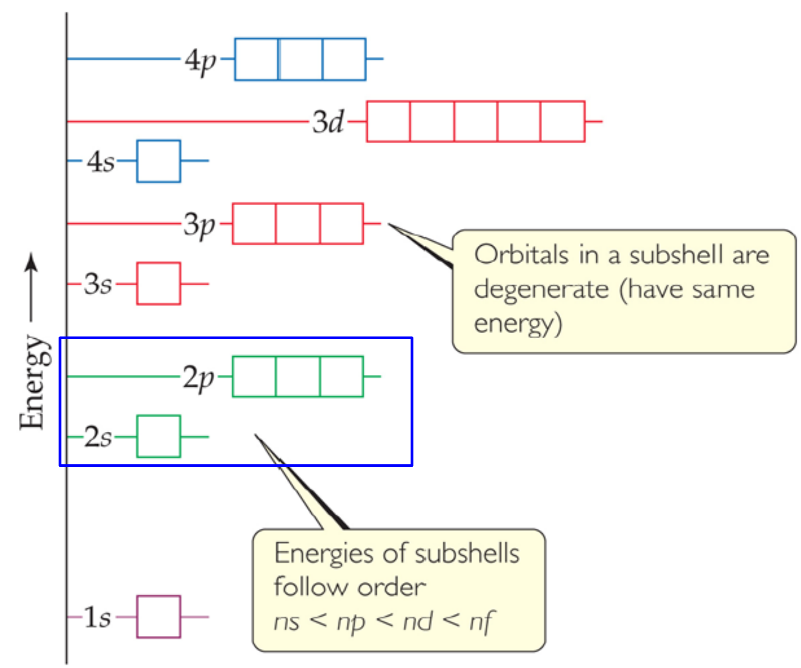
–The first discontinuity between group IIA and IIIA.

Take period 2 as an example,

Ionization energy: Be ([He]$2s^2$) > B ([He]$2s^2 2p^1$)

This is due to the third valence electron of B must occupy the 2p subshell (which is higher in energy than 2s subshell & hence not favored).





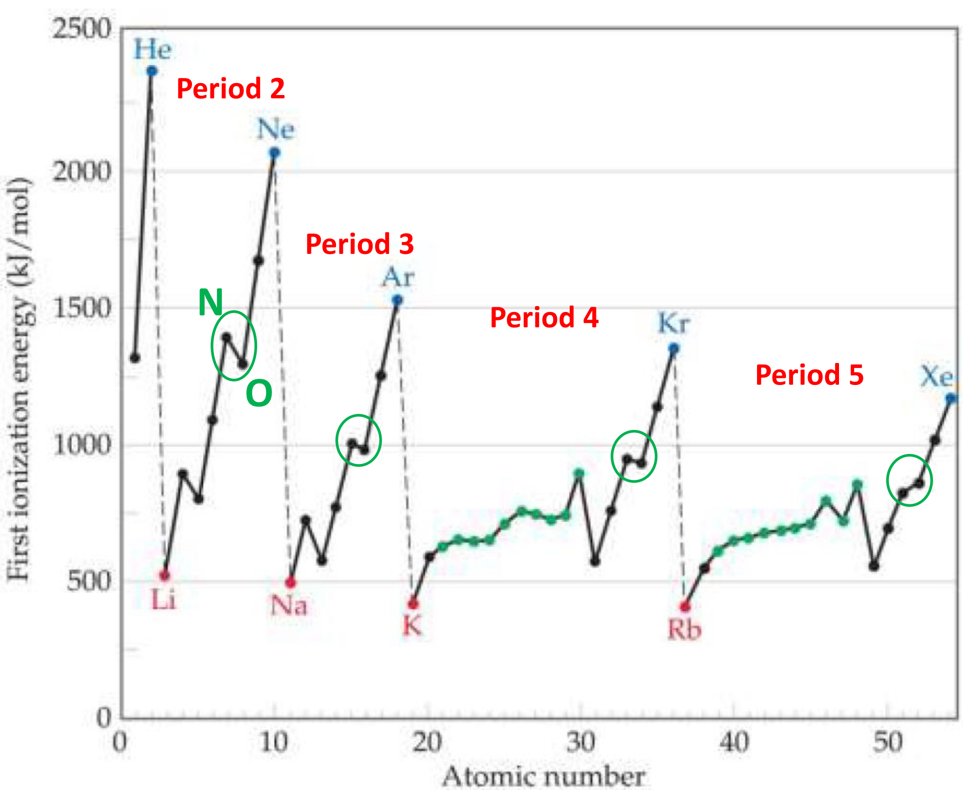
Ionization Energy: Elements in the Same Period

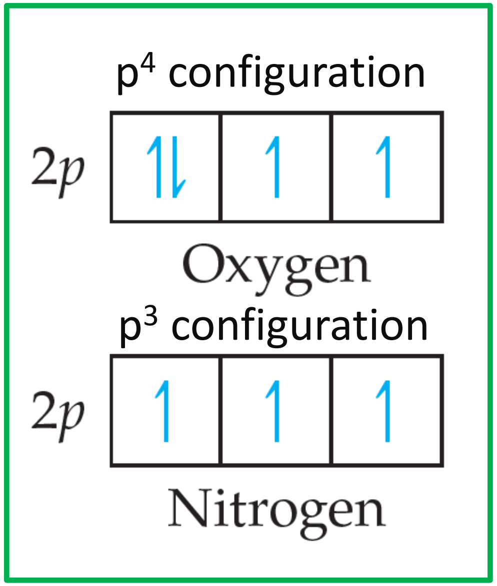
–The second discontinuity between group V and VIA.

Take period 2 as an example,

Ionization energy: N ([He]$2s^2 2p^3$) > O ([He]$2s^2 2p^4$)

This is due to the repulsion of paired electrons in the p^4 configuration of O atom. While No electron repulsion in the p^3 configuration of N atom according to Hund's rule.

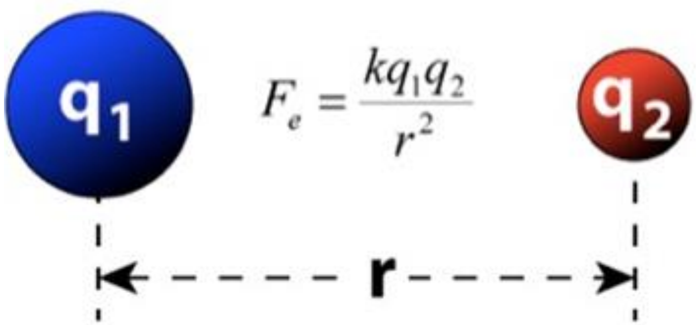


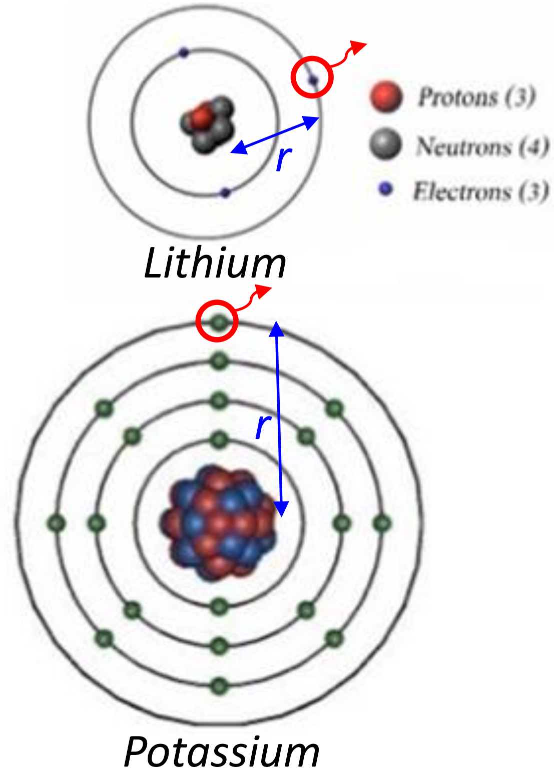


Ionization Energy: Elements in the Same Group

Group 1	1^{st} IE (kJ/mol)
Li	520
Na	496
K	419

decrease slightly!





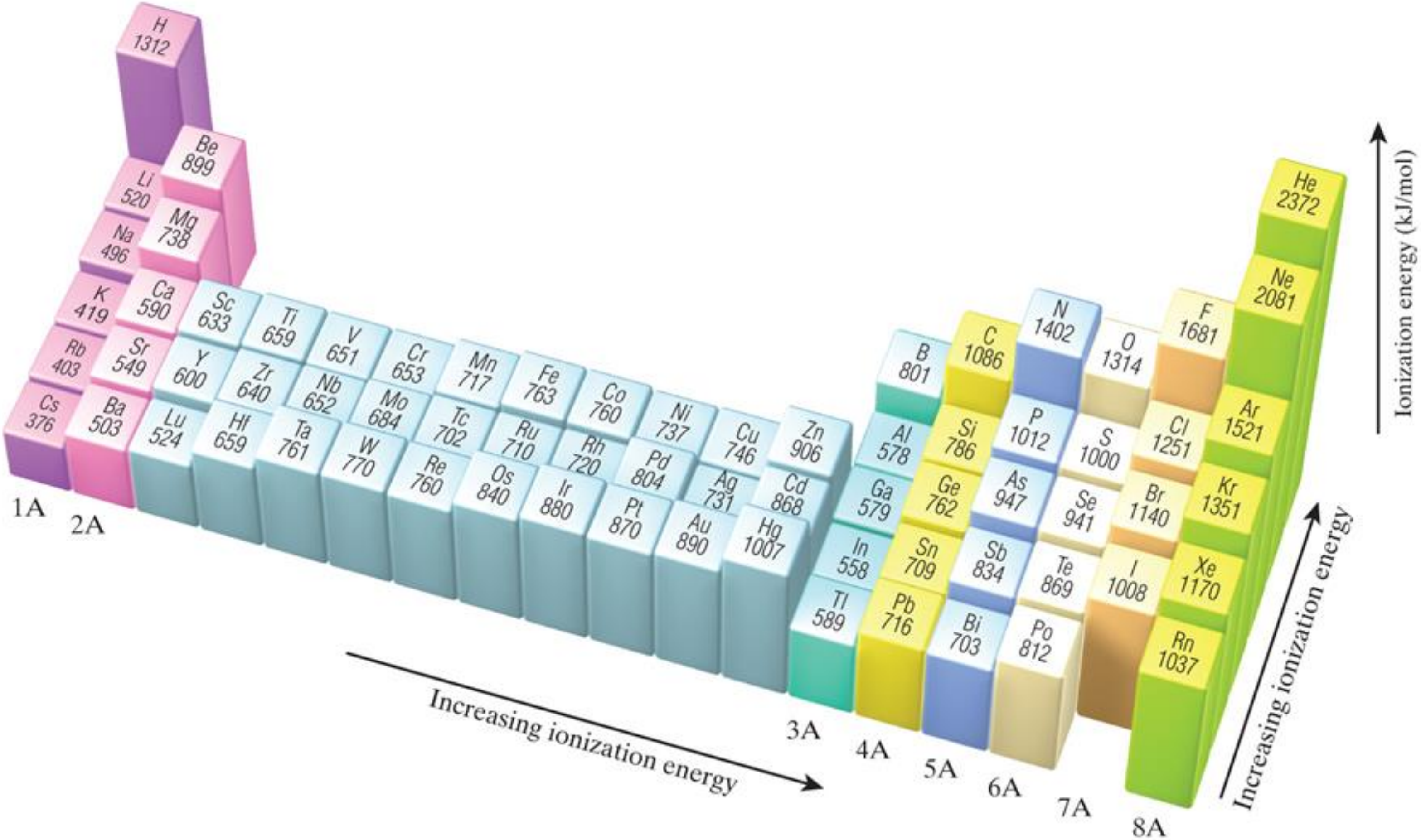
The Z_{eff} is about 1 for them…while

- Going from Li to K, the outmost electron is further away from the nucleus and the fore of attraction decreases *(Coulombic force is inversely proportional to r^2)*.

The decreased Coulombic force for bigger element in a group making it easier to remove electron and hence lower in ionization energy (e.g. potassium here).

Periodic Trends in the First Ionization Energy (I_1)

- I_1 tends to increase with increasing atomic number **across each period** because Z_{eff} is increasing across a period (two discontinuities!).
- I_1 tends to decrease with increasing atomic number **down each "A" group** because the valence electrons are farther from the nucleus, although Z_{eff} is slightly increasing down a group.



Periodic Properties of the Elements

Outline

Ionization Energy vs. Electron Affinity

Ionization energy describes the energy change when *an atom **loses** an electron*, whereas **electron affinity** describes the energy change when *an atom **gains** an electron*.

Take chlorine (Cl) as an example…

The first **ionization energy** of Cl(*g*): 1251 kJ/mol, the ***positive value*** of the ionization energy means that energy must be put into the atom to remove the electron.

$$\textit{Ionization energy:}\quad \underset{[\mathrm{Ne}]3s^2 3p^5}{\mathrm{Cl}(g)} \longrightarrow \underset{[\mathrm{Ne}]3s^2 3p^4}{\mathrm{Cl}^+(g)} + \mathrm{e}^- \qquad \Delta E = 1251\ \mathrm{kJ/mol}$$

*All ionization energies for atoms are positive: **Energy must be given** to remove e^-!*

The energy change that occurs when an electron is added to a gaseous atom is called the **electron affinity** because it measures the attraction, or *affinity,* of the atom for the added electron

$$\textit{Electron affinity:}\quad \underset{[\mathrm{Ne}]3s^2 3p^5}{\mathrm{Cl}(g)} + e^- \longrightarrow \underset{[\mathrm{Ne}]3s^2 3p^6}{\mathrm{Cl}^-(g)} \qquad \Delta E = -349\ \mathrm{kJ/mol}$$

*The negative sign indicating that **energy is released** during the process!*

Electron Affinity

- ***Electron affinity*** of an atom is the energy change that occurs when an electron is added to the atom.
- *Most atoms have negative electron affinity* (i.e. releasing energy; Only for elements in Group 8A and Be/Mg/N are positive).

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

Electron Affinity: Elements in the Same Period

- Generally, It tends to become more negative across a period (WHY?).

  The greater the attraction between an atom and an added electron (i.e. Z_{eff}), the more negative the atom's electron affinity (i.e. tend to gain an electron).

- However, there are three apparent discontinuities in this trend

  Between groups (1) 1A & 2A; (2) 4A & 5A; (3) 7A & 8A

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

Electron Affinity: Elements in the Same Period

1st Discontinuity between group

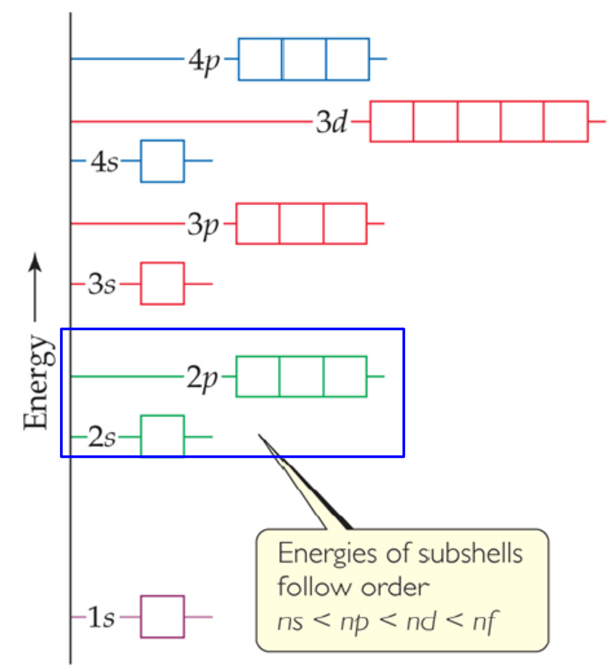
(1) 1A & 2A

Li ([He]$2s^1$) → Li^- ([He]$2s^2$); Be ([He]$2s^2$) → Be^- ([He]$2s^2 2p^1$)

Because Be^- involves the occupying of 2p with higher-energy is energetically unfavorable, the electron affinity is positive.

Similar explanation applies to discontinuity between 1A & 2A in other periods

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



Electron Affinity: Elements in the Same Period

2nd Discontinuity between group

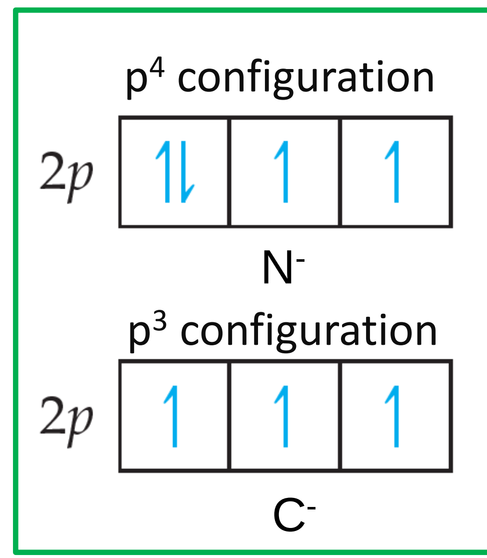
(2) 4A & 5A

C ([He]$2s^2 2p^2$) → C^- ([He]$2s^2 2p^3$); N ([He]$2s^2 2p^3$) → N^- ([He]$2s^2 2p^4$)

Since neutral N have half-filled *p* subshells already, the added electron generates extra electron–electron repulsions.

Similar explanation applies to discontinuity between 4A & 5A in other periods

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



Noted that a discontinuity in first ionization energy for the same reason!

Electron Affinity: Elements in the Same Period

3rd Discontinuity between group

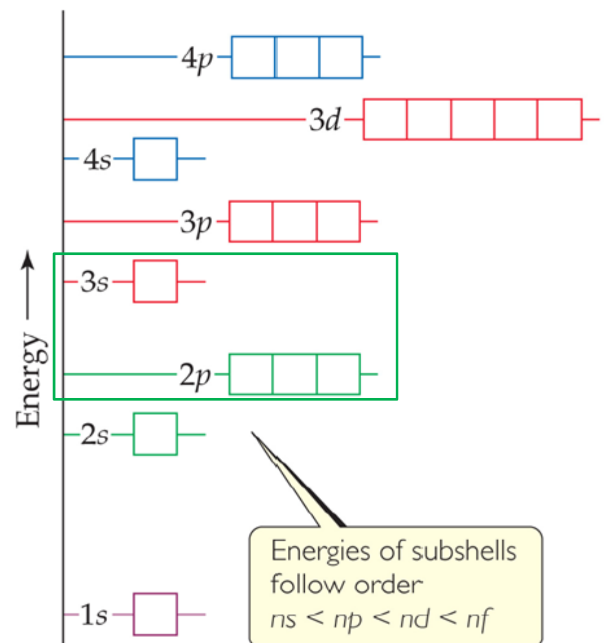
(3) 7A & 8A

F ([He]$2s^2 2p^5$) → F^- ([He]$2s^2 2p^6$); Ne ([He]$2s^2 2p^6$) → Ne^- ([He]$2s^2 2p^6 3s^1$)

By gaining an electron, a halogen atom forms a stable anion that has a noble-gas configuration. The addition of an electron to a noble gas is energetically unfavorable due to the involvement of a higher-energy shell.

Similar explanation applies to discontinuity between 7A & 8A in other periods

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



Electron Affinity: Elements in the <u>Same Group</u>

- ***Electron affinity*** decreases ***relatively little*** down a group (cf. a period)(WHY?).

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

When we proceed down a group the average distance between the added electron and the nucleus increases, causing the **electron–nucleus attraction** to <u>**decrease**</u> ($\propto 1/r^2$) and hence <u>**decrease**</u> in **electron affinity**.

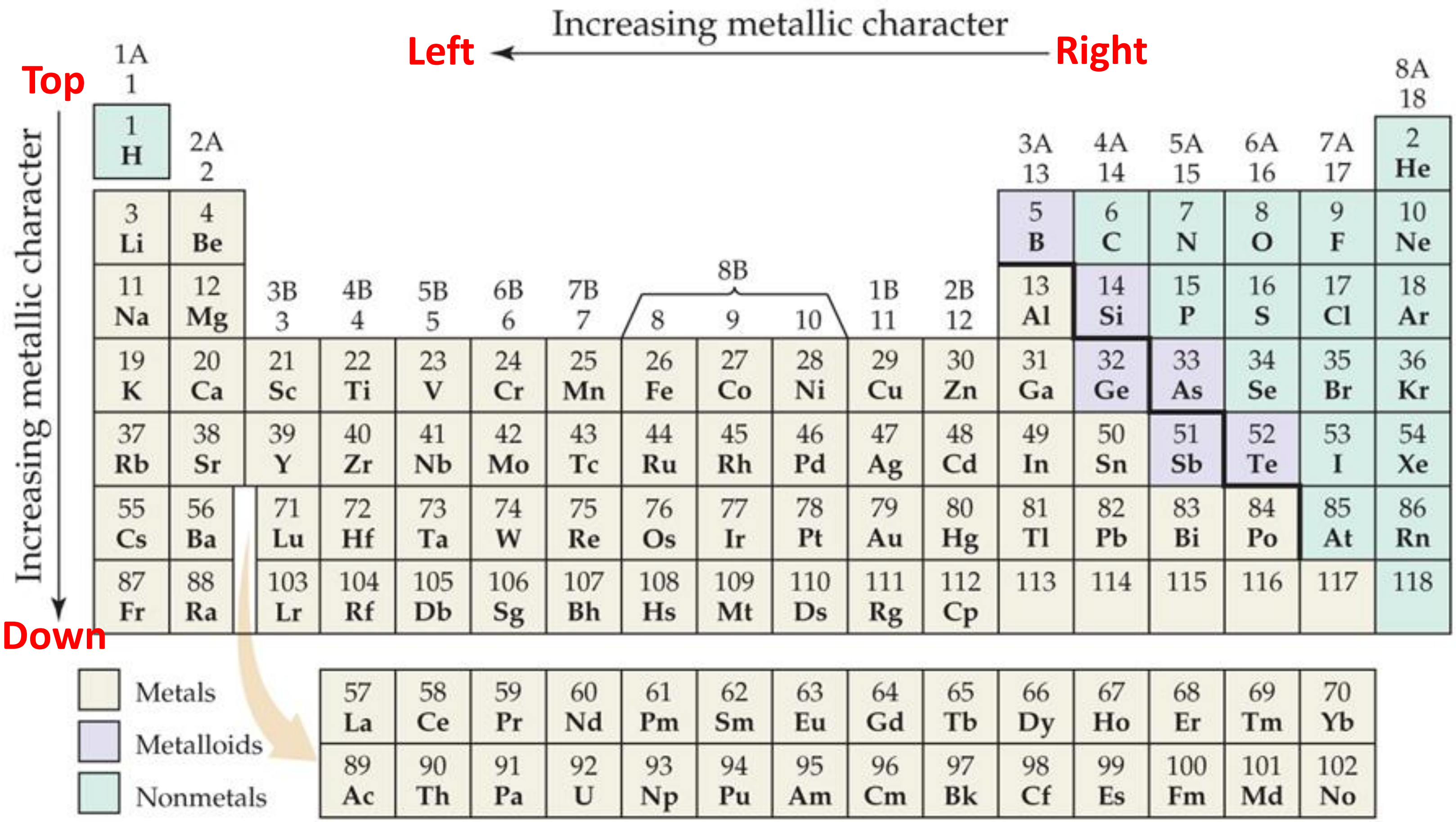
Periodic Properties of the Elements

Outline

- Introduction
- Effective Nuclear Charge
- Property 1: Sizes of Atoms and Ions
- Property 2: Ionization Energy
- Property 3: Electron Affinity
- Property 4: Metals, Metalloids and Nonmetals

Metals, Metalloids and Nonmetals

The distribution of metal/metalloid/nonmetal of elements in periodic table



Metal (87/112; 80%); Metalloid (6/112; 5%); Nonmetal (19/112; 15%)

Metals, Metalloids and Nonmetals

Metals vs. Nonmetals

Table 7.3 Characteristic Properties of Metals and Nonmetals

Metals	Nonmetals
Have a shiny luster; various colors, although most are silvery	Do not have a luster; various colors
Solids are malleable and ductile	Solids are usually brittle; some are hard, and some are soft
Good conductors of heat and electricity	Poor conductors of heat and electricity
Most metal oxides are ionic solids that are basic	Most nonmetal oxides are molecular substances that form acidic solutions
Tend to form cations in aqueous solution	Tend to form anions or oxyanions in aqueous solution

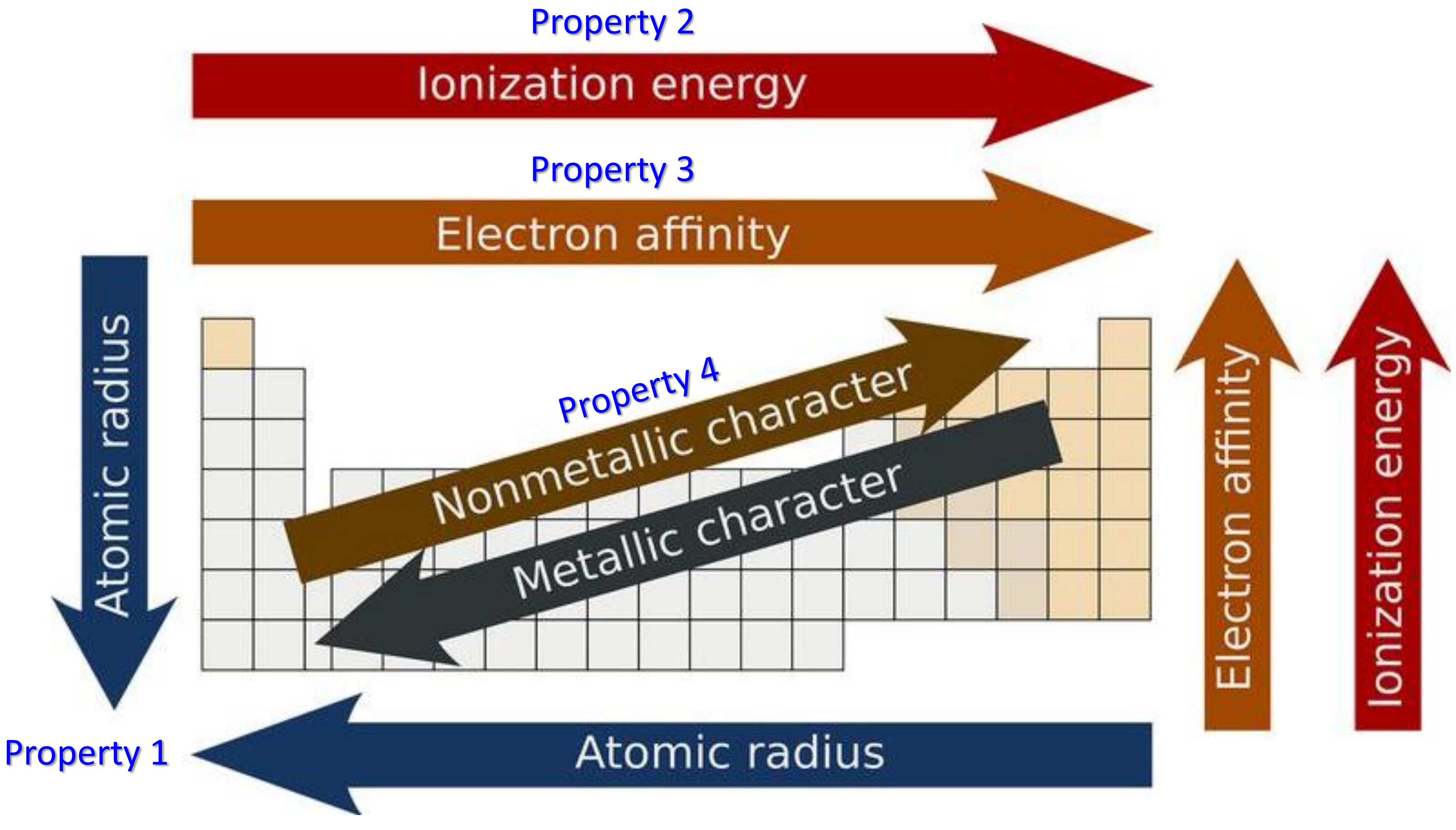
Metalloids

- **B, Si, Ge, As, Sb, Te**
- They exhibit properties between metal & nonmetals.
- For instance, silicon looks shiny, but is brittle and fairly poor conductor.

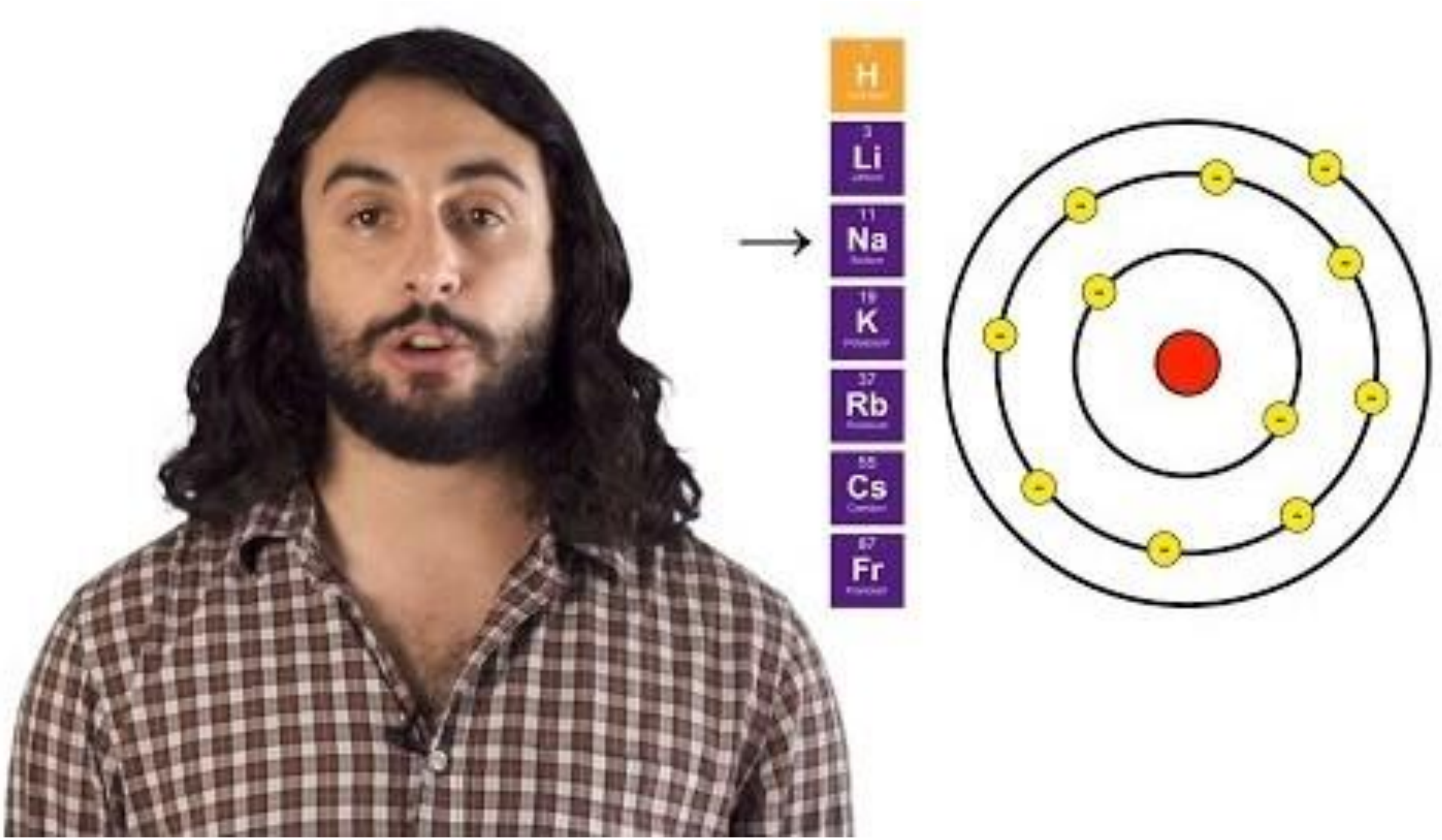
For detailed properties of some common elements pls see the textbook (Chapter 7)!

Periodic Properties of the Elements

Effective Nuclear Charge (Z_{eff})



Summary Video



https://www.youtube.com/watch?v=hePb00CqvP0

Basic Concept of Chemical Bonding

Outline (Part I)

- Lewis Symbols & Octet Rule
- Types of Chemical Bonding
- Ionic Bonding
- Covalent Bonding
- Bond Polarity and Electronegativity

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Lewis Symbols & Octet Rule

- The **Lewis symbol** for an element consists of *the element's chemical symbol* plus *a dot for each valence electron*.
- The dots are placed on the four sides of the symbol—top, bottom, left, and right—and each side can accommodate up to two electrons (all four sides are equivalent).

Table 8.1 Lewis Symbols

Group	Period 2 Element	Electron Configuration	Lewis Symbol	Period 3 Element	Electron Configuration	Lewis Symbol
1A	Li	$[He]2s^1$	Li·	Na	$[Ne]3s^1$	Na·
2A	Be	$[He]2s^2$	·Be·	Mg	$[Ne]3s^2$	·Mg·
3A	B	$[He]2s^22p^1$	·Ḃ·	Al	$[Ne]3s^23p^1$	·Ȧl·
4A	C	$[He]2s^22p^2$	·Ċ̣·	Si	$[Ne]3s^23p^2$	·Ṩi·
5A	N	$[He]2s^22p^3$	·Ṇ̇:	P	$[Ne]3s^23p^3$	·Ṗ̣:
6A	O	$[He]2s^22p^4$	:Ọ̇:	S	$[Ne]3s^23p^4$	:Ṩ:
7A	F	$[He]2s^22p^5$	·F̤̈:	Cl	$[Ne]3s^23p^5$	·C̤̈l:
8A	Ne	$[He]2s^22p^6$	:N̤̈e:	Ar	$[Ne]3s^23p^6$	:Ä̤r:

Lewis Symbols & Octet Rule

Octet Rule

Fact: Noble gases (such as He, Ne, Ar) show *high ionization energies* and *low electron affinities (positive value)*, and they are chemically inert (i.e. very stable).

Group	Element	Electron Configuration	Lewis Symbol	Element	Electron Configuration	Lewis Symbol
1A	Li	$[He]2s^1$	$\mathrm{Li}\cdot$	Na	$[Ne]3s^1$	$\mathrm{Na}\cdot$
		...			...	
7A	F	$[He]2s^2 2p^5$	$\cdot\ddot{\underset{\cdot\cdot}{\mathrm{F}}}:$	Cl	$[Ne]3s^2 3p^5$	$\cdot\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}:$
8A	Ne	$[He]2s^2 2p^6$	$:\ddot{\underset{\cdot\cdot}{\mathrm{Ne}}}:$	Ar	$[Ne]3s^2 3p^6$	$:\ddot{\underset{\cdot\cdot}{\mathrm{Ar}}}:$

Low ionization energy

High electron affinity (very negative value)

Theory: The inertness of noble gases results from their stable electron configurations; each (except helium) has **EIGHT** electrons in its outermost shell.

Deduction: Elements should become less reactive when they alter their electron structures to that of a noble gas.

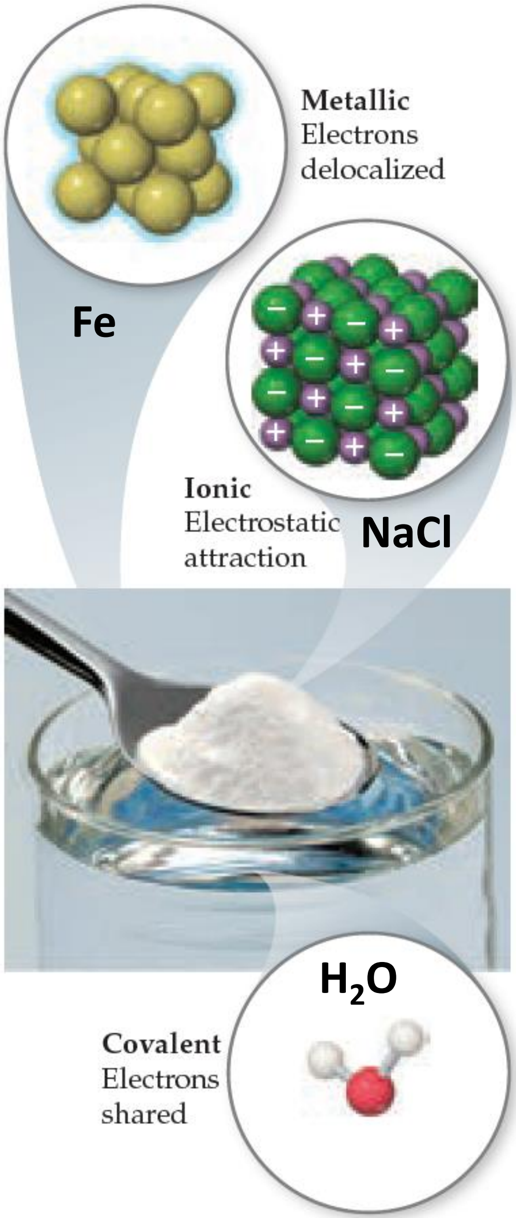
Atoms tend to <u>gain</u>, <u>lose</u>, or <u>share</u> electrons until they acquire *<u>EIGHT</u>* electrons in their valence shells.

Basic Concept of Chemical Bonding

Outline (Part I)

- Lewis Symbols & Octet Rule
- Types of Chemical Bonding
- Ionic Bonding
- Covalent Bonding
- Bond Polarity and Electronegativity

Types of Chemical Bonding



There are three major types of chemical bonds: (Lecture 1)

- **Metallic Bond:**

Delocalization of valence electrons throughout metal atoms in the three-dimensional lattice.

- **Ionic Bond:**

Electrostatic forces between oppositely charged ions.

cations: positively charged ions

anions: negatively charged ions

- **Covalent Bond:**

Sharing of valence electrons between two atoms.

Two or more atoms of different elements joined together by chemical bonds form ***chemical compounds*****.**

Basic Concept of Chemical Bonding

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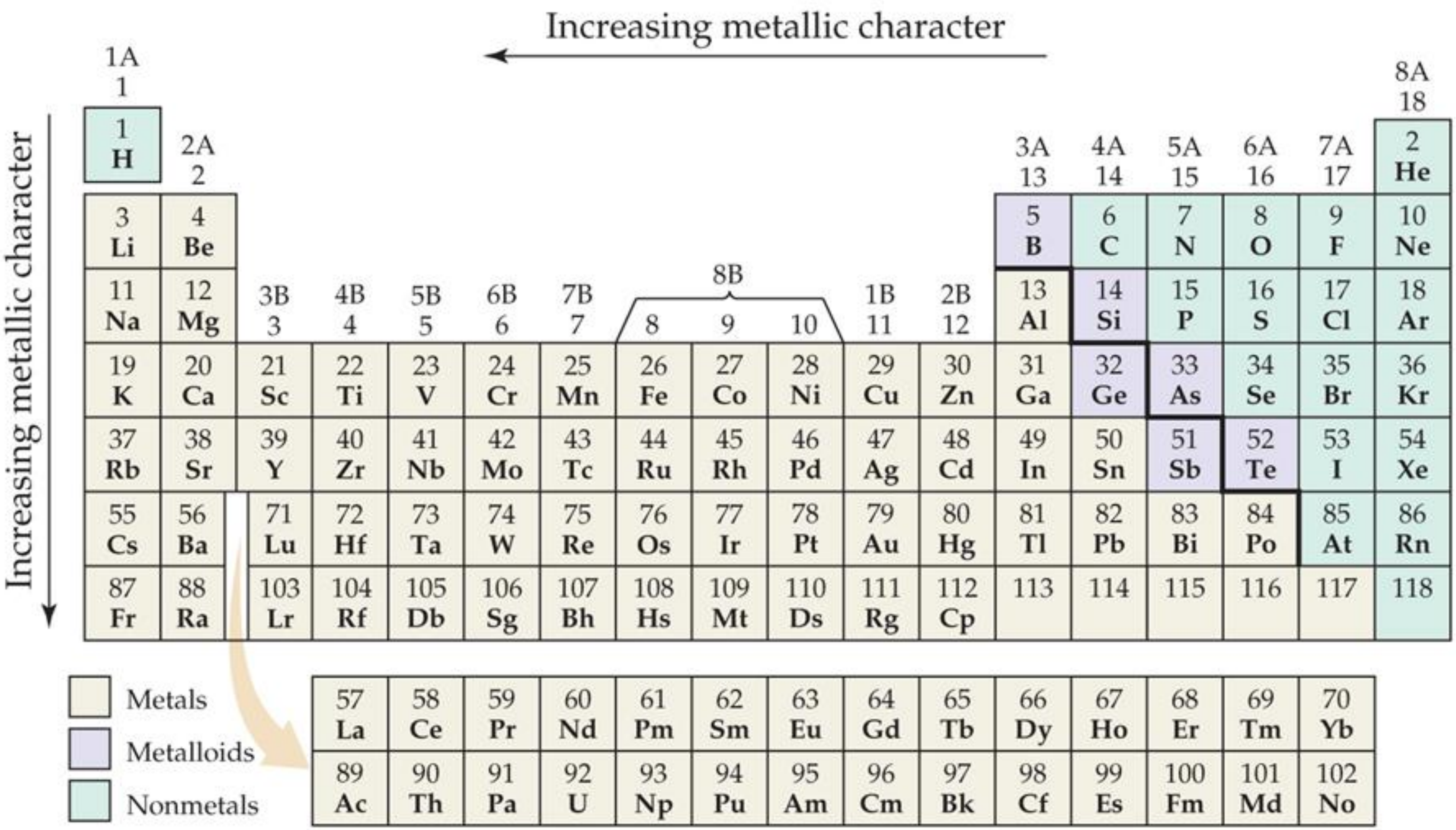
Ionic Bonding

Cation (from Metals) + Anion (from Nonmetals)



Ionic Bonding

Cation (from Metals) + Anion (from Nonmetals)

- **Metals tend to form cations** (the Octet rule!)
  - because they have low ionization energies.
- **Nonmetals tend to form anions** (the Octet rule!)
  - because their high electron affinities (more negative value).

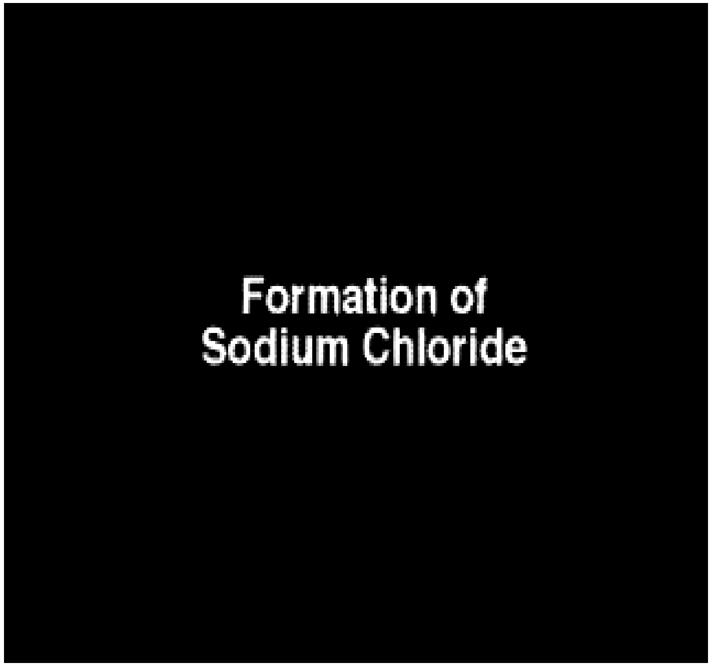
1A	2A	Transition metals										3A	4A	5A	6A	7A	8A
H^+																H^-	NOBLE GASES
Li^+														N^{3-}	O^{2-}	F^-	
Na^+	Mg^{2+}											Al^{3+}		P^{3-}	S^{2-}	Cl^-	
K^+	Ca^{2+}	Sc^{3+}	Ti^{4+}	V^{5+} V^{4+}	Cr^{3+}	Mn^{2+} Mn^{4+}	Fe^{2+} Fe^{3+}	Co^{2+} Co^{3+}	Ni^{2+}	Cu^+ Cu^{2+}	Zn^{2+}				Se^{2-}	Br^-	
Rb^+	Sr^{2+}								Pd^{2+}	Ag^+	Cd^{2+}		Sn^{2+} Sn^{4+}	Sb^{3+} Sb^{5+}	Te^{2-}	I^-	
Cs^+	Ba^{2+}								Pt^{2+}	Au^+ Au^{3+}	Hg_2^{2+} Hg^{2+}		Pb^{2+} Pb^{4+}	Bi^{3+} Bi^{5+}			

Ionic Bonding

Formation of Sodium Chloride (NaCl)

Chlorine gas: $Cl_2(g)$ **$Na(s)$ reacts readily with $Cl_2(g)$ …**

Sodium metal $Na(s)$ **… to produce $NaCl(s)$ (i.e. table salt).**



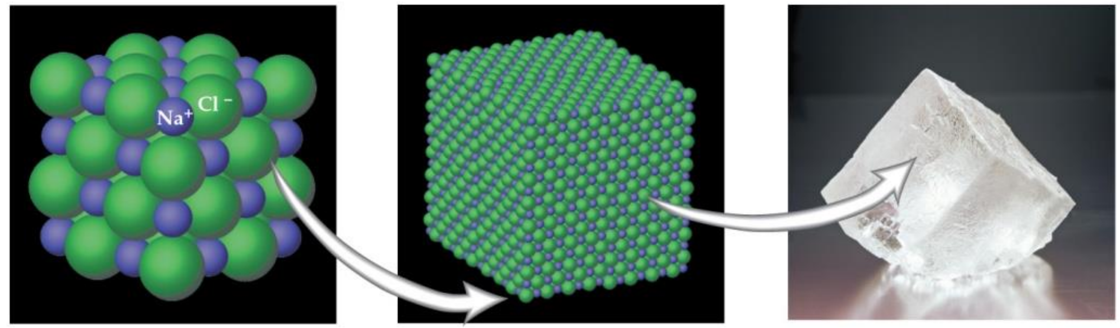
$$\mathrm{Na}\cdot + \cdot\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}: \longrightarrow \mathrm{Na}^{+} + [:\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}:]^{-}$$

$[Ne]3s^1$ $[Ne]3s^23p^5$ $[Ne]$ $[Ne]3s^23p^6$

Ionic Bonding

How do we evaluate the strength of ionic bonding?

Lattice energy (kJ/mol):
The energy required to completely separate *one mole* of a solid ionic compound into its gaseous ions.



Lattice energy is governed by the Coulomb's law (E_{el} = electrostatic potential energy):

$$E_{\text{el}} \propto \frac{q_1 q_2}{d}$$

It increases with:

- increasing charge of the ions (q_i), and
- decreasing size of the ions (i.e. the distance between the two ions d).

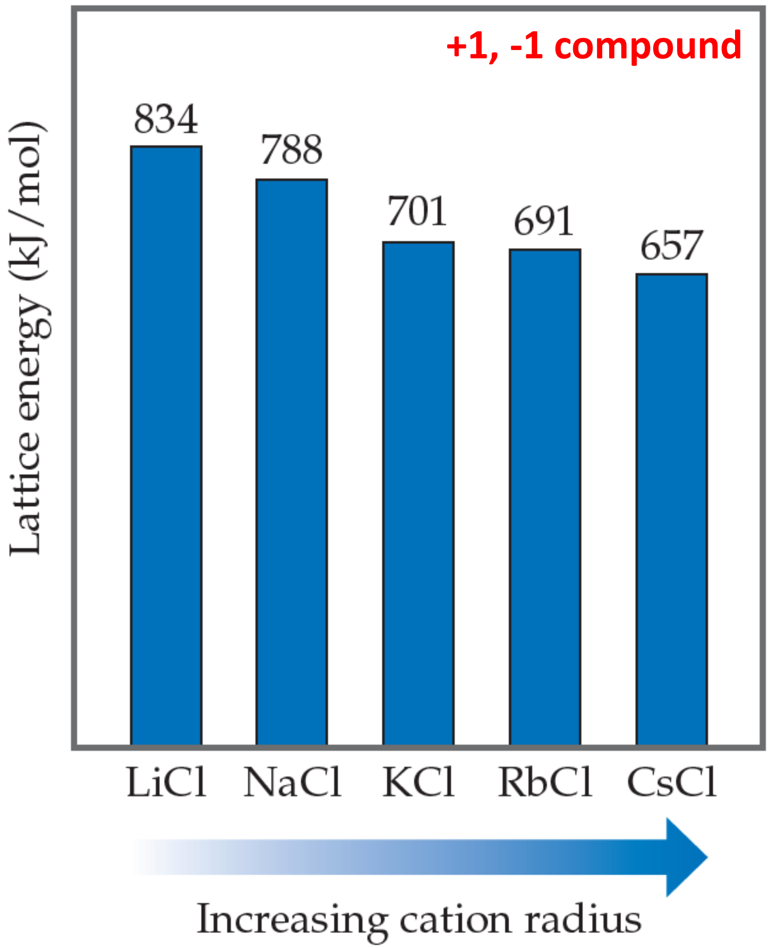
Table 8.2 Lattice Energies for Some Ionic Compounds charge effect

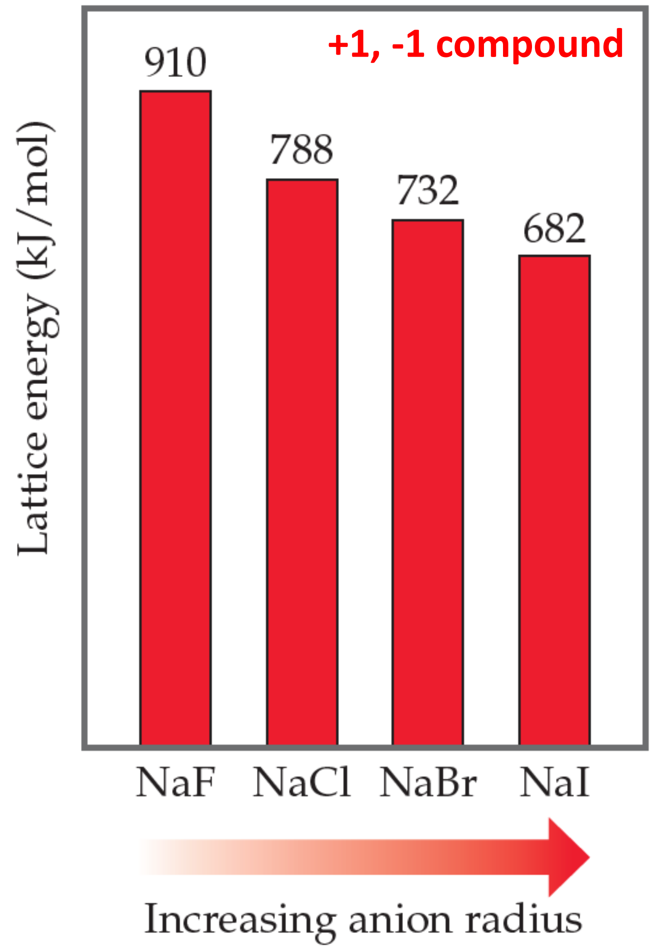
Compound		Lattice Energy (kJ/mol)	Compound		Lattice Energy (kJ/mol)
LiF	+1, -1	1030	$MgCl_2$	+2, -1*2	2326
LiCl		834	$SrCl_2$		2127
LiI		730			
NaF		910	MgO	+2, -2	3795
NaCl		788	CaO		3414
NaBr		732	SrO		3217
NaI		682			
KF		808	ScN	+3, -3	7547
KCl		701			
KBr		671			
CsCl		657			
CsI		600			

Ionic Bonding

Radius effect: lattice energy **decreases with increasing size of the ions** (i.e. the distance between the two ions d).

$$E_{\mathrm{el}} \propto \frac{q_1 q_2}{d}$$





Compared to the charge effect, radius effect is relatively minor!

Basic Concept of Chemical Bonding

Outline (Part I)

- Lewis Symbols & Octet Rule
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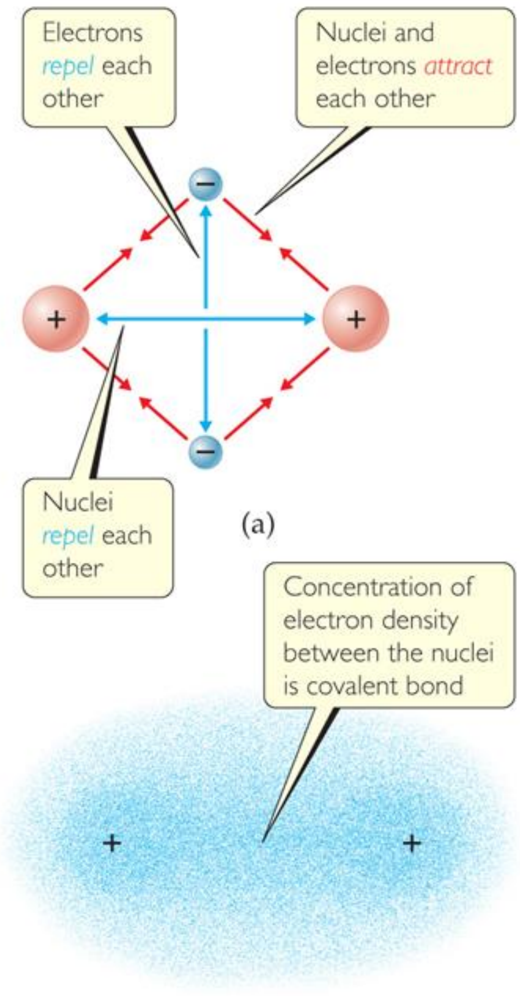
Covalent Bonding

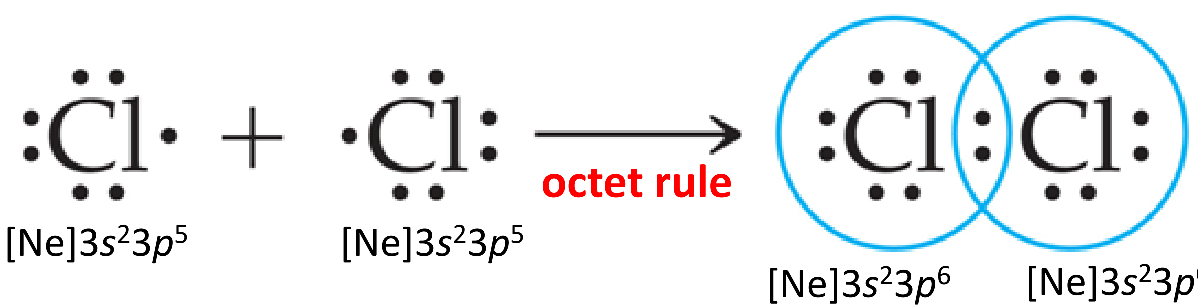
Covalent bonds are formed by ***sharing of electron pairs*** between two atoms (mostly nonmetallic)



(a)

(b)

- Covalent bond is governed by *electrostatic interactions among electrons and nuclei*:
  - electron-electron repulsion
  - nucleus-nucleus repulsion
  - nucleus-electron attraction
- These interactions result in concentrating the electron density between the nuclei.
- Therefore, the nuclei are attracted together by the electron density between them which acts like a "glue".

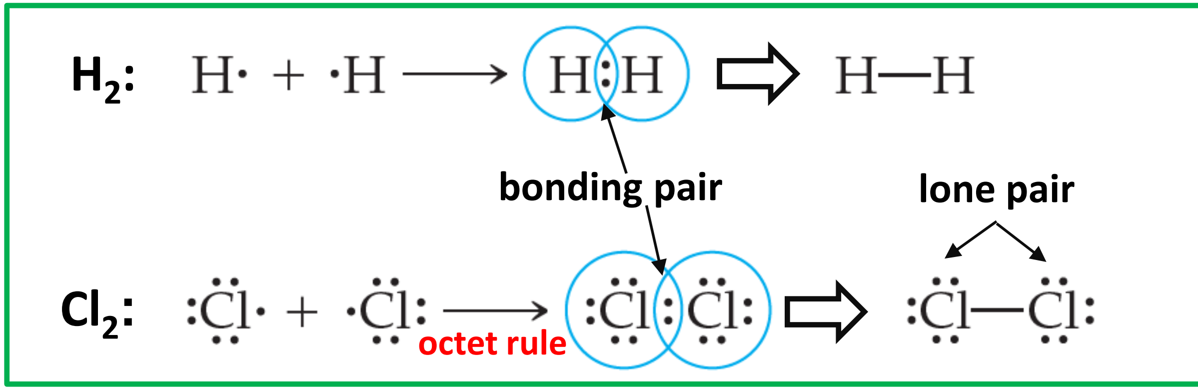


Covalent Bonding

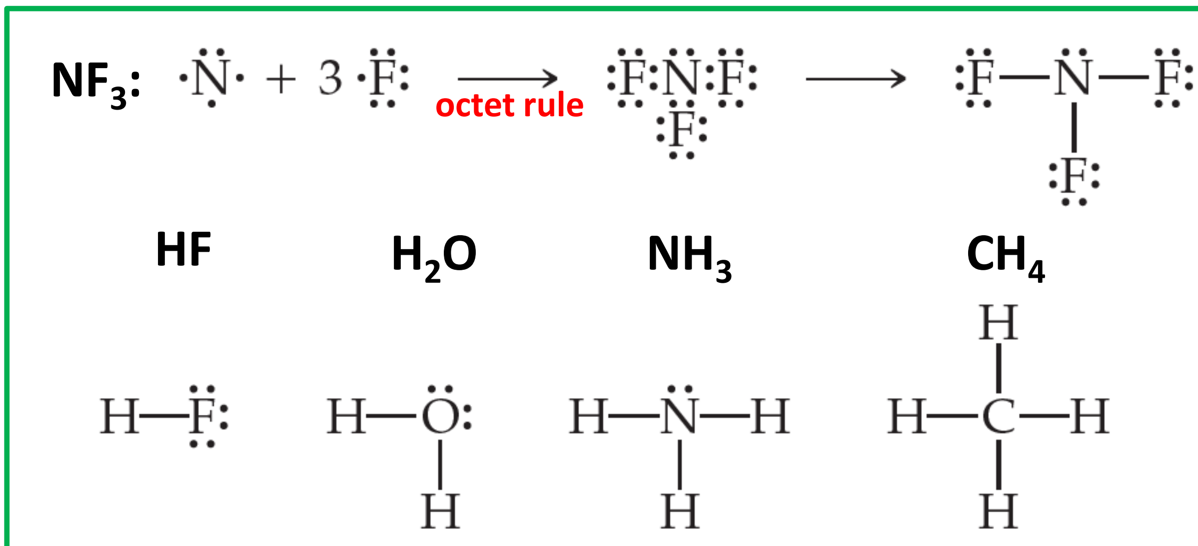
The shared electron is called ***bonding pair*** can be presented using a line between two atoms, while for the rest electrons are what we called ***lone pair*** electrons.

Lewis structures (or Lewis dot structures) of molecules

Homoatomic Molecules
(single bond case)



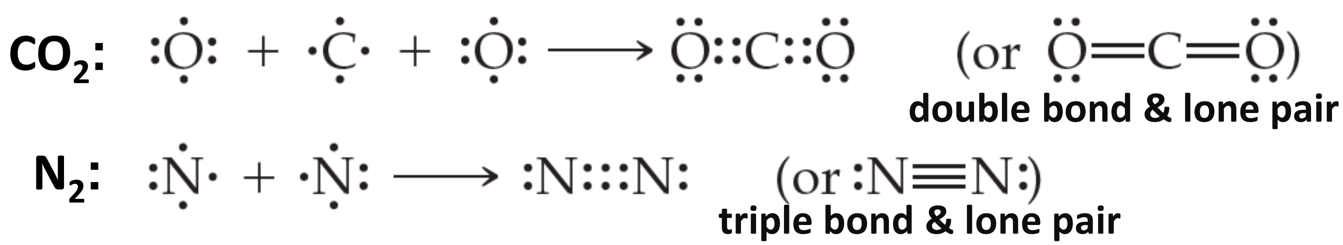
Heteroatomic Molecules
(single bond case)

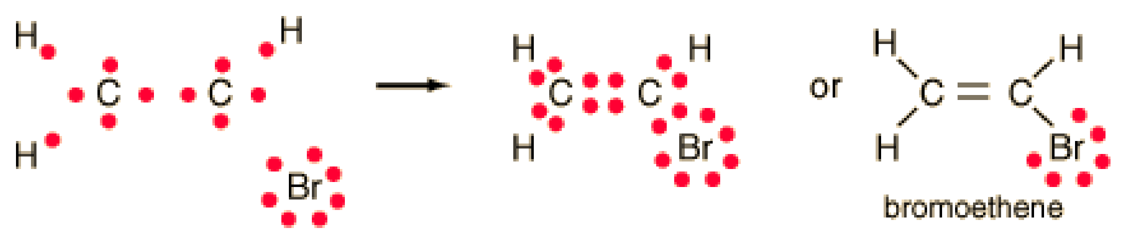


Covalent Bonding

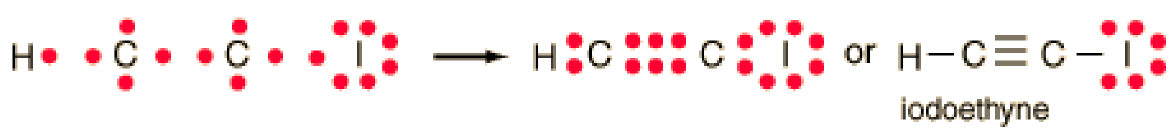
Instead of ***single bond***… ***double & triple bond*** may also appear to fulfill the octet rule!

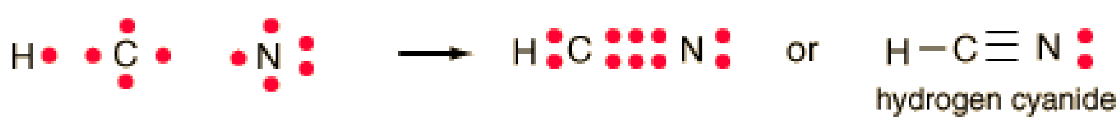
multiple bonds case





single/double/triple bond and lone pair co-exist in a molecule





Basic Concept of Chemical Bonding

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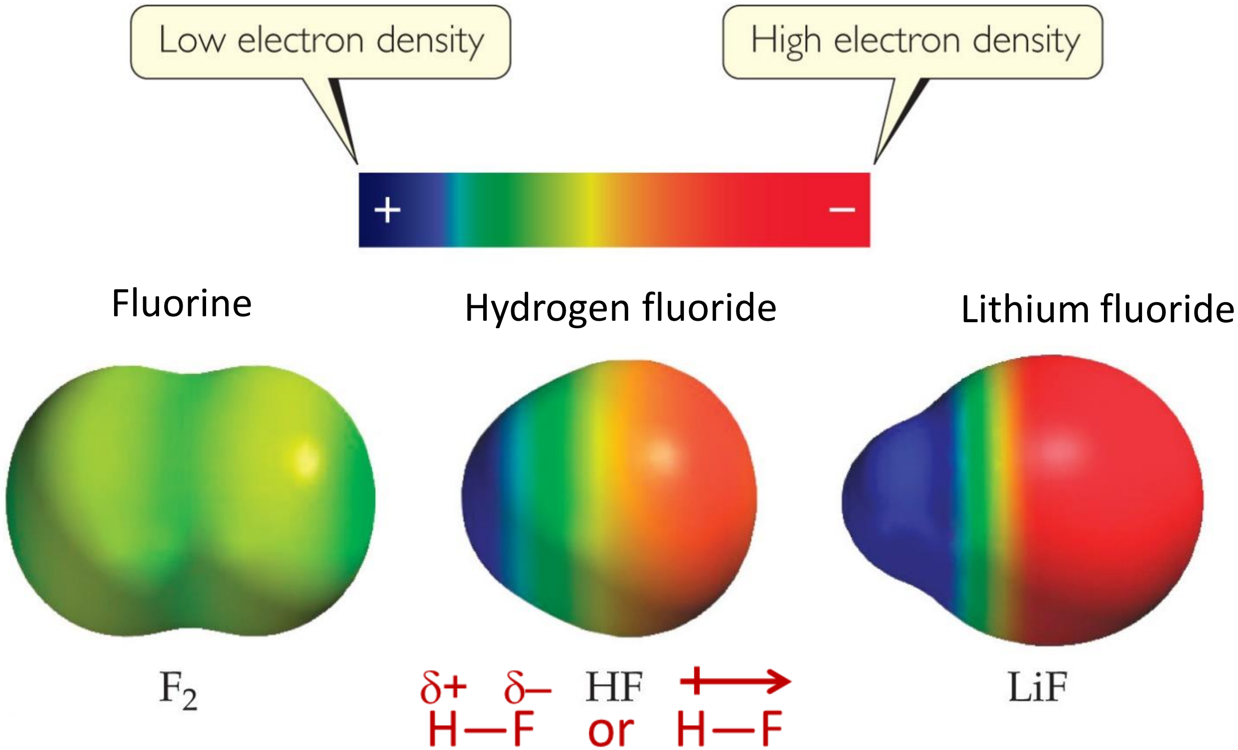
Bond Polarity and Electronegativity

Bonding electron pairs between two atoms are “not” always shared equally.

Asymmetrical sharing of electrons results in ***bond polarity***.



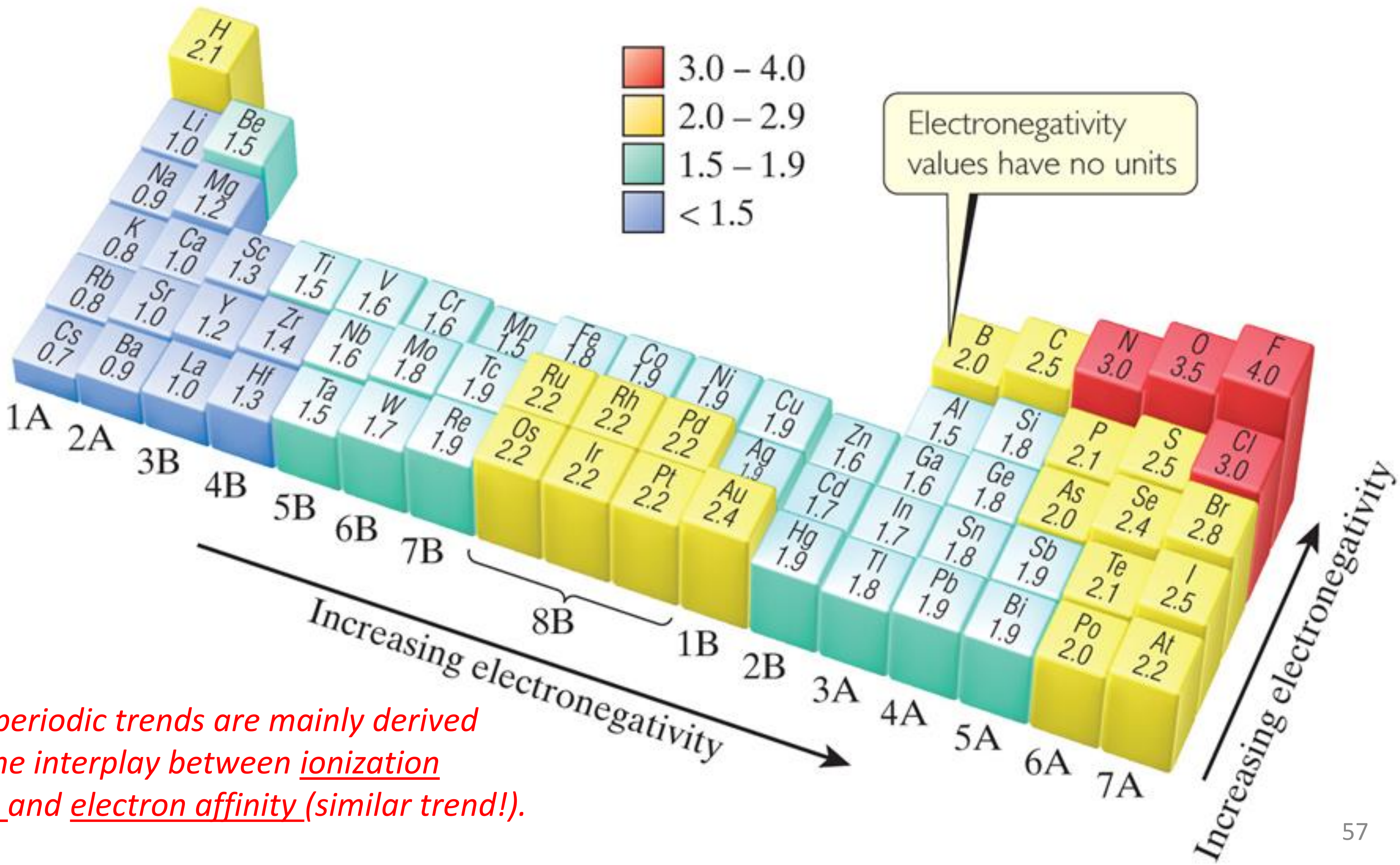
- The **F—F** bond is ***nonpolar*** as the electrons are ***symmetrically*** shared.
- The **H—F** bond is ***polar*** as the electrons are ***asymmetrically*** shared.

A bond is ionic/polar/nonpolar can be estimated by a quantity called **electronegativity**.

Bond Polarity and Electronegativity

Electronegativity is a measure of the ability of *an atom in a molecule* to attract electrons to itself (no unit!!!!).

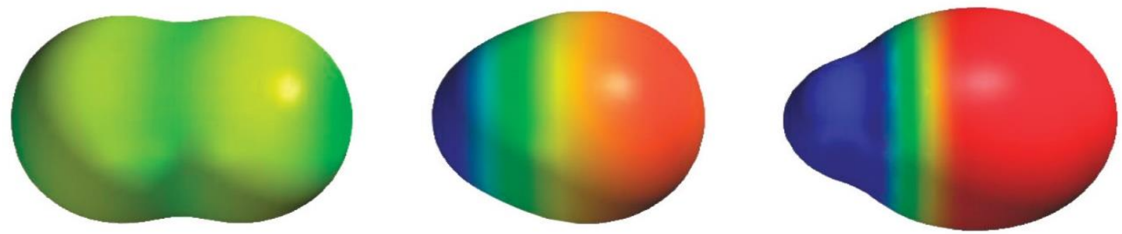
It increases from left to right across a period and from bottom to top of a group.



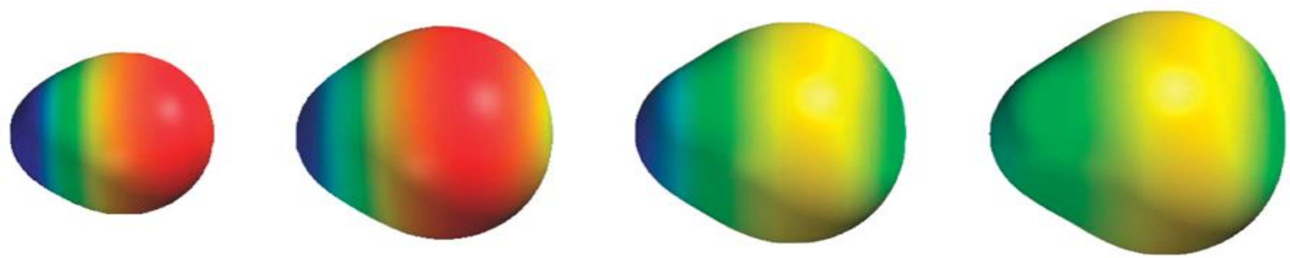
These periodic trends are mainly derived from the interplay between ionization energy and electron affinity (similar trend!).

Bond Polarity and Electronegativity

We can estimate the polarity of a bond! The greater the difference in electronegativity between two atoms, the covalent bond formed by these two atoms will be more polar!

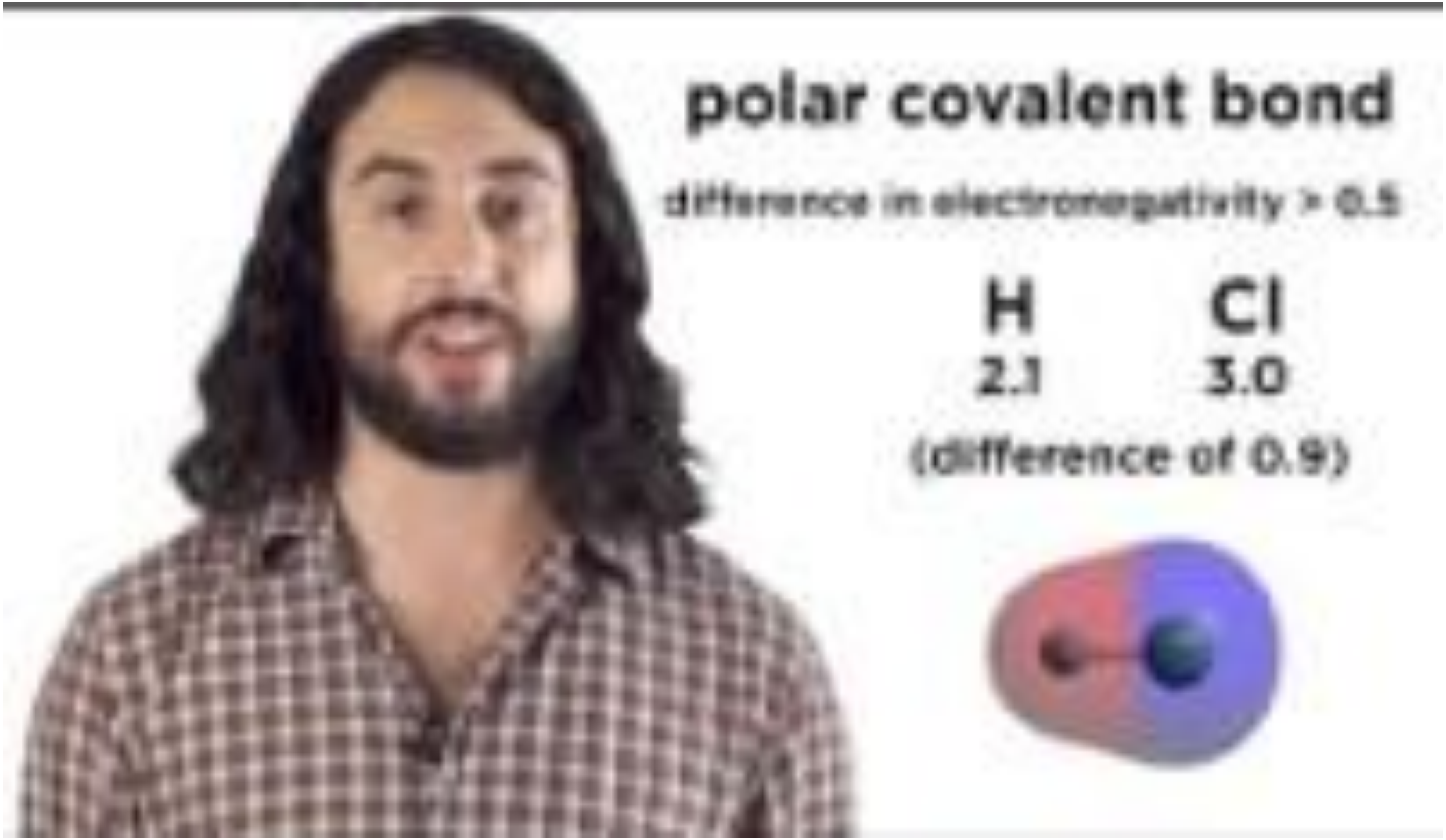


	F_2	HF	LiF
Electronegativity difference	$4.0 - 4.0 = 0$	$4.0 - 2.1 = 1.9$	$4.0 - 1.0 = 3.0$
Type of bond	Nonpolar covalent	Polar covalent	Ionic



	HF	HCl	HBr	HI
Electronegativity difference	1.9	0.9	0.7	0.4
Polarity of covalent bond	High ←			Low

Summary Video



https://www.youtube.com/watch?v=PoQjsnQmxok&ab_channel=ProfessorDaveExplains

The End

&

Thanks For Your Attention !