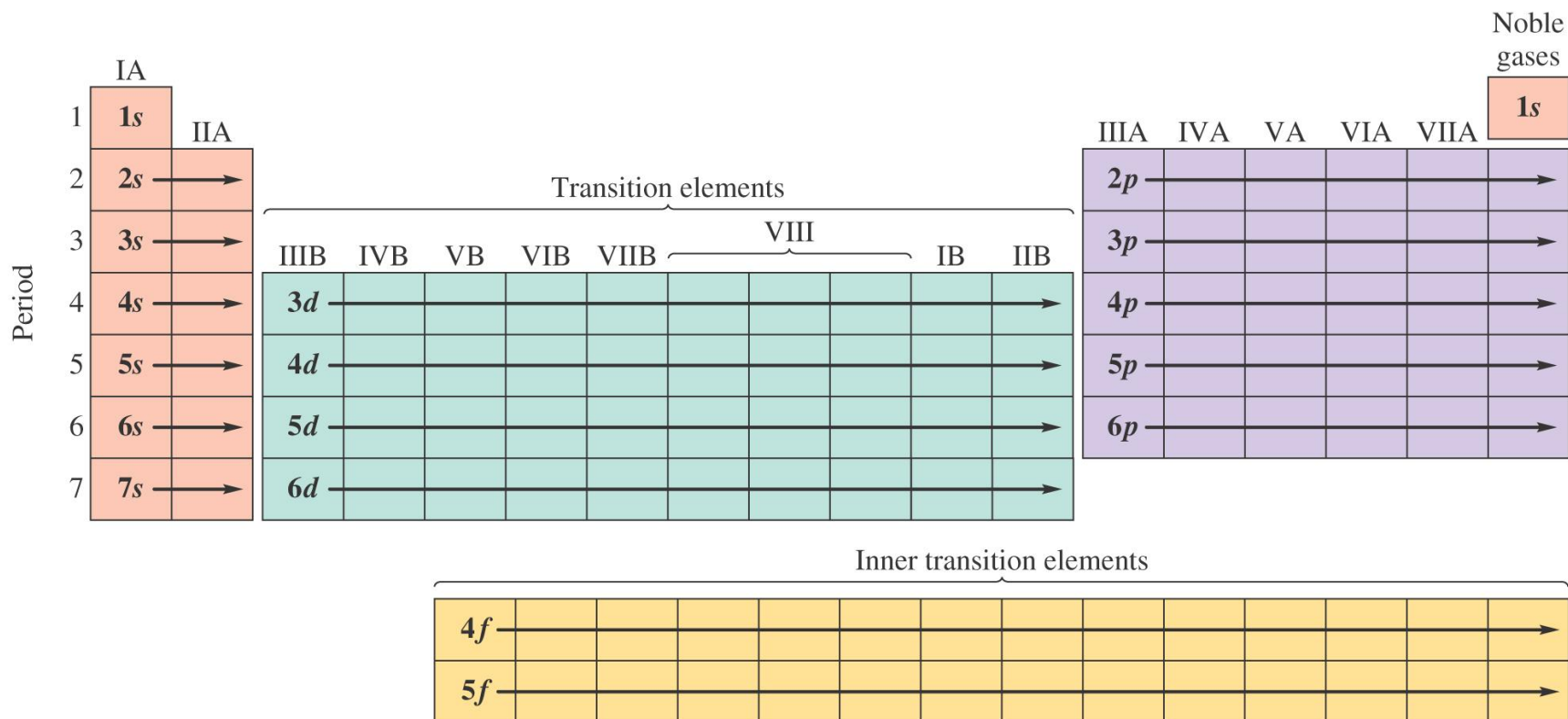


ELECTRON CONFIGURATION AND THE PERIODIC TABLE

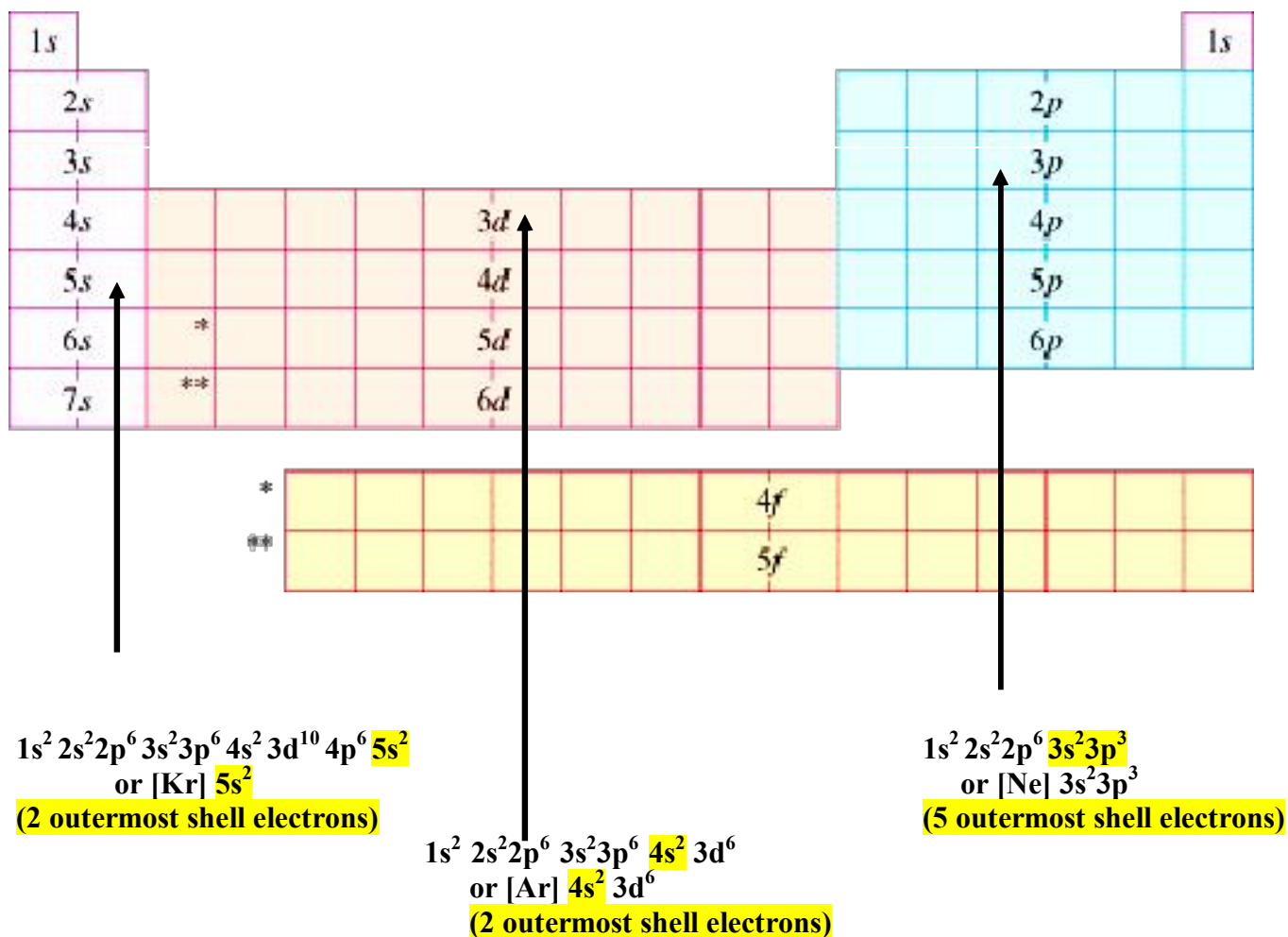
- The electrons in an atom fill from the lowest to the highest orbitals.
- The knowledge of the location of the orbitals on the periodic table can greatly help the writing of electron configurations for large atoms.



For A Groups (Representative Elements):
 (s and p blocks) $n = \text{period number}$

For B Groups (Transition Elements)
 d blocks: $n = \text{period number} - 1$
 f blocks: $n = \text{period number} - 2$

- The electronic configuration of an element can now be given if the position in the Periodic Table is known:



- Group VIIIA elements (Noble Gases) have a stable outermost shell electron configuration:

Element	Orbital Notation	Electron Configuration	# of Outermost Shell Electrons
He	$\frac{\uparrow\downarrow}{1s}$	$1s^2$	2
Ne	$\frac{\uparrow\downarrow}{2s} \frac{\downarrow\uparrow}{2p} \frac{\uparrow\downarrow}{2p} \frac{\uparrow\downarrow}{2p}$	$2s^2 2p^6$	8
Ar	$\frac{\uparrow\downarrow}{3s} \frac{\downarrow\uparrow}{3p} \frac{\uparrow\downarrow}{3p} \frac{\uparrow\downarrow}{3p}$	$3s^2 3p^6$	8
Kr	$\frac{\uparrow\downarrow}{4s} \frac{\downarrow\uparrow}{4p} \frac{\uparrow\downarrow}{4p} \frac{\uparrow\downarrow}{4p}$	$4s^2 4p^6$	8
Xe	$\frac{\uparrow\downarrow}{5s} \frac{\downarrow\uparrow}{5p} \frac{\uparrow\downarrow}{5p} \frac{\uparrow\downarrow}{5p}$	$5s^2 5p^6$	8
Rn	$\frac{\uparrow\downarrow}{6s} \frac{\downarrow\uparrow}{6p} \frac{\uparrow\downarrow}{6p} \frac{\uparrow\downarrow}{6p}$	$6s^2 6p^6$	8

- The electron configuration of $ns^2 np^6$ (or ns^2) is referred to as a “noble gas core”
- Chemical Properties of elements are determined by the outermost shell electrons (electrons in highest energy level); these electrons are involved in bonding.

VALENCE ELECTRONS:

- Valence electrons are those outside the noble-gas core
- For transition elements, the “s” and “p” electrons of the outermost shell (highest energy level) (“d” electrons are not valence electrons)

NOTE:

- Most transition elements have 2 valence electrons
Reason: They fill “d” or “f” orbitals which are not part of the outer most shell
- Exceptions: Transition elements with irregular electronic configurations: **Cr, Cu, etc.**

MAGNETIC PROPERTIES OF ATOMS

- Every electron acts like a small magnet
- Magnetic attractions from 2 electrons with opposite spin cancel each other

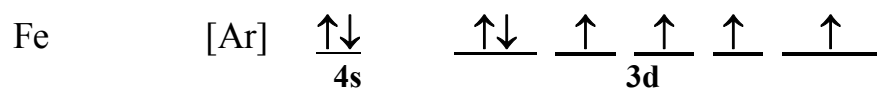
It follows:

- An atom with paired electrons has no magnetism.
- An atom with unpaired electrons (excess of one spin) exhibits magnetism.
- An element with atoms that exhibit magnetism is attracted to a strong magnet

DIAMAGNETIC ELEMENTS	PARAMAGNETIC ELEMENTS
<ul style="list-style-type: none"> • Are not attracted by a magnetic field • Sometimes even repelled by a magnetic field • Atoms contain paired electrons only 	<ul style="list-style-type: none"> • Are weakly attracted by a magnetic field • Atoms contain unpaired electrons
$\text{Hg} \rightarrow [\text{Xe}] 6s^2 4f^{14} 5d^{10}$	$\text{Na} \rightarrow [\text{Ne}] 3s^1$ $\text{C} \rightarrow [\text{He}] 2s^2 2p^2$

FERROMAGNETIC ELEMENT

- Are strong permanent magnetism in iron due to the alignment of many unpaired electrons



Examples:

1. Write complete and condensed electron configuration for bromine ($Z=35$) and determine the number of valence electrons.

2. Write complete and condensed electron configuration for palladium ($Z=46$) and determine the number of valence electrons.

3. Write complete and condensed electron configuration for antimony ($Z=51$) and determine the number of valence electrons.

PERIODIC PROPERTIES

- Periodic properties are properties of the elements that are repeated according to a regular “periodic” trend.

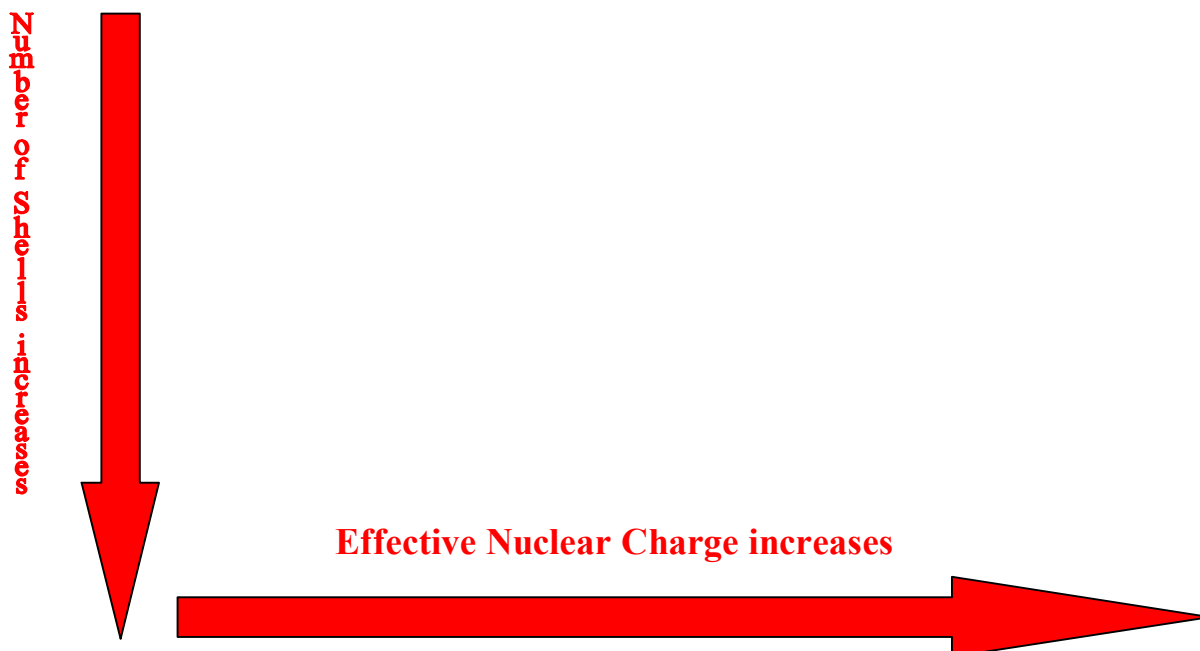
Three important periodic properties will be discussed:

1. **Atomic Radius** (size of the atoms)
2. **Ionization Energy (I.E.)** - energy needed to remove the outermost electron from a neutral atom in the gaseous state to form a positive ion.
3. **Electron Affinity (E.A.)** - energy (absorbed or **released**) during the process of adding an electron to a neutral atom in the gaseous state to form a negative ion

The variation of these 3 Periodic Properties will be discussed :

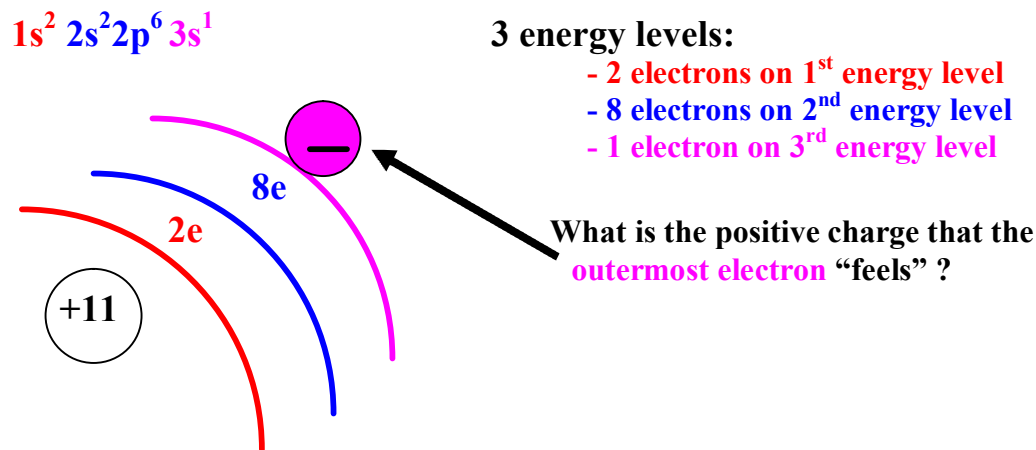
- within a group (vertical trend)
- along a period (horizontal trend)

Two important factors determine these trends:



Effective Nuclear Charge

- Effective nuclear charge (Z_{eff}) is the positive charge that an electron experiences from the nucleus.
- Consider the electron configuration of the Na atom ($Z = 11$)



1. **It is not +11**, since its electrons of the 1st and 2nd energy level (10 electrons carry a charge equal to -10) = cancel out some of the $+11$ charge of the nucleus (‘shielding effect’)
2. **It is not +1**, since the 10 core electrons (charge = -10) cannot completely cancel out 10 positive charges of the nucleus.

Reason: The 10 core electrons are in the nucleus, but some distance away

- Note: The closer to the nucleus the core electrons are, the more effective they are in canceling out some of the positive charge of the nucleus (shielding effect)

shielding effect
of a 1st shell electron



shielding effect
of a 2nd shell electron

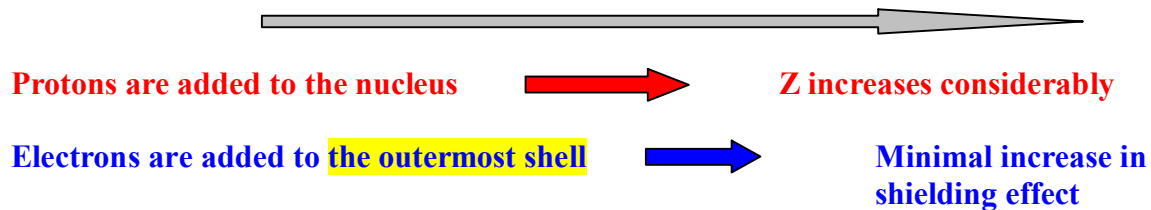
3. The outermost electron is attracted to the nucleus by a positive charge which is less than the actual nuclear charge ($Z = +11$) because of the shielding effect of the core electrons, but more than $+1$

Effective Nuclear Charge = Z_{eff} = $+2.8$ (in this particular case)

- **The Nuclear Charge an outermost electron experiences is reduced by the shielding effect of other electrons ($Z_{\text{eff}} = Z - \text{Shielding Effect}$)**

$$\begin{array}{c}
 \mathbf{Z_{eff}} \\
 \nearrow \\
 \mathbf{Effective\ Nuclear\ Charge}
 \end{array}
 =
 \begin{array}{c}
 \mathbf{Z} \\
 \nwarrow \\
 \mathbf{Actual\ Nuclear\ Charge}
 \end{array}
 -
 \mathbf{Shielding\ effect}$$

Along a period, moving from left to right:



$$Z_{\text{eff}} = Z - \text{Shielding effect}$$

Effective Nuclear Charge \rightarrow Z_{eff}

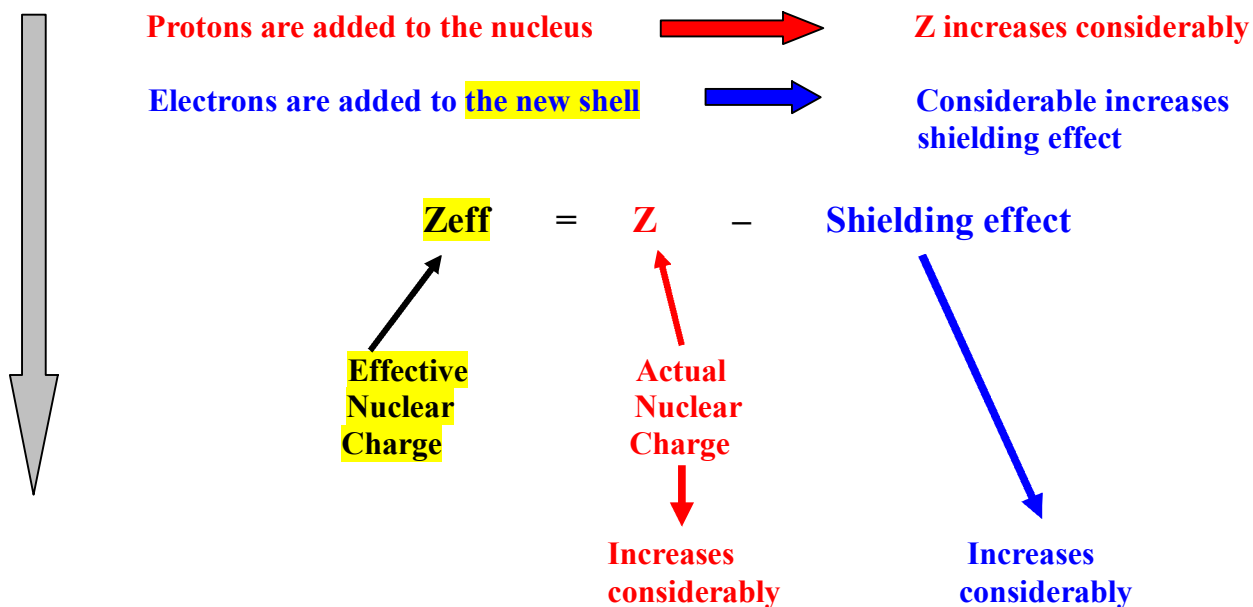
Actual Nuclear Charge \rightarrow Z

Increases considerably \rightarrow Z

Minimal increase \rightarrow Shielding effect

Moving across a period Z_{eff} increases (Z increases a lot, Shielding Effect increases little)

Along a group, moving from up to down:

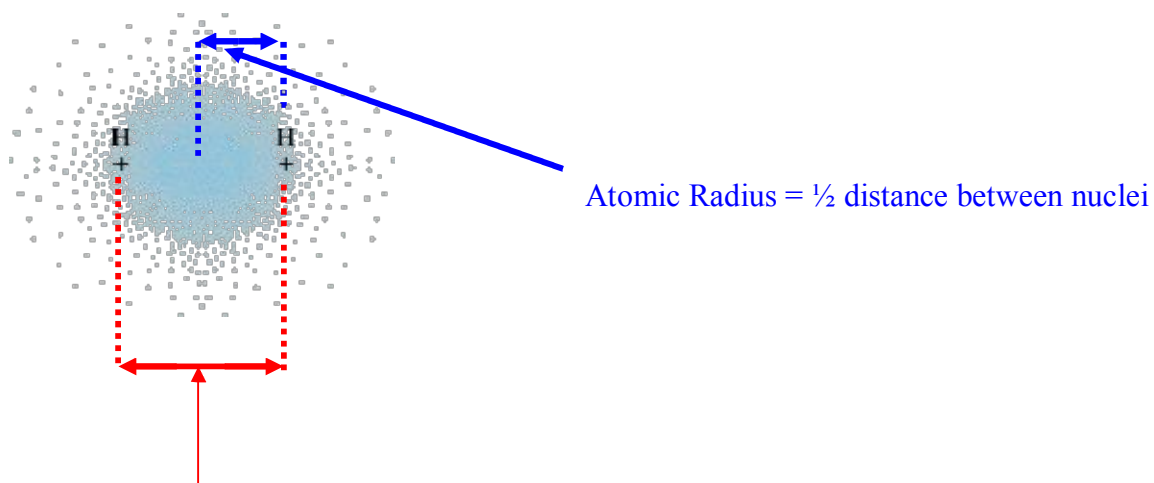


Moving down a group Z_{eff} does NOT increase (Z increases a lot, Shielding Effect increases a lot)

ATOMIC RADIUS

- Values of Atomic radii are obtained from measurements of distances between the nuclei of atoms in the chemical bonds of molecular substances.

Example: experimental Determination of the radius of H atom:



distance between nuclei is determined experimentally (by X-ray crystallography)

- Atomic Radius determined in this manner are referred to as Covalent Radii.
- Atoms are very small; consequently, atomic radii have very small values
- Values of Atomic Radii are listed on the Periodic Table (back side) and are commonly expressed in:

Angstroms ($1 \text{ \AA} = 10^{-10} \text{ m}$)

or

Nanometers ($1 \text{ nm} = 10^{-9} \text{ m}$)

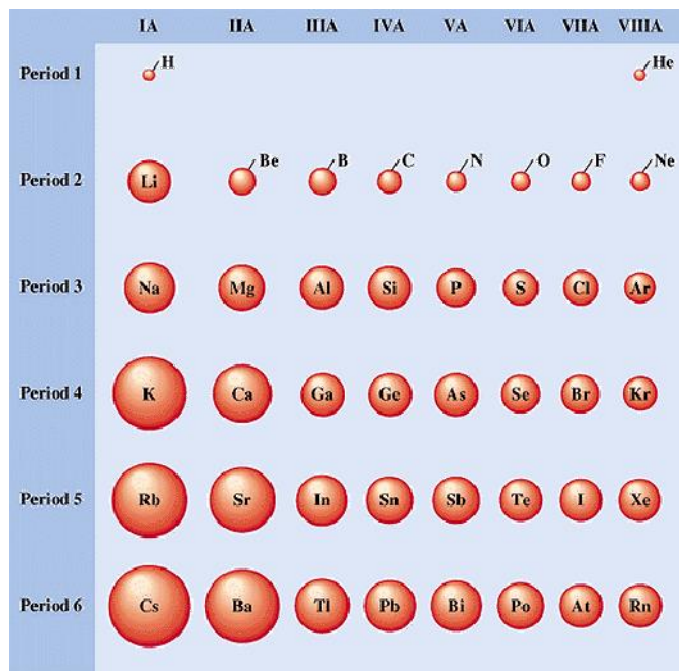
or

Picometers ($1 \text{ pm} = 10^{-12} \text{ m}$)

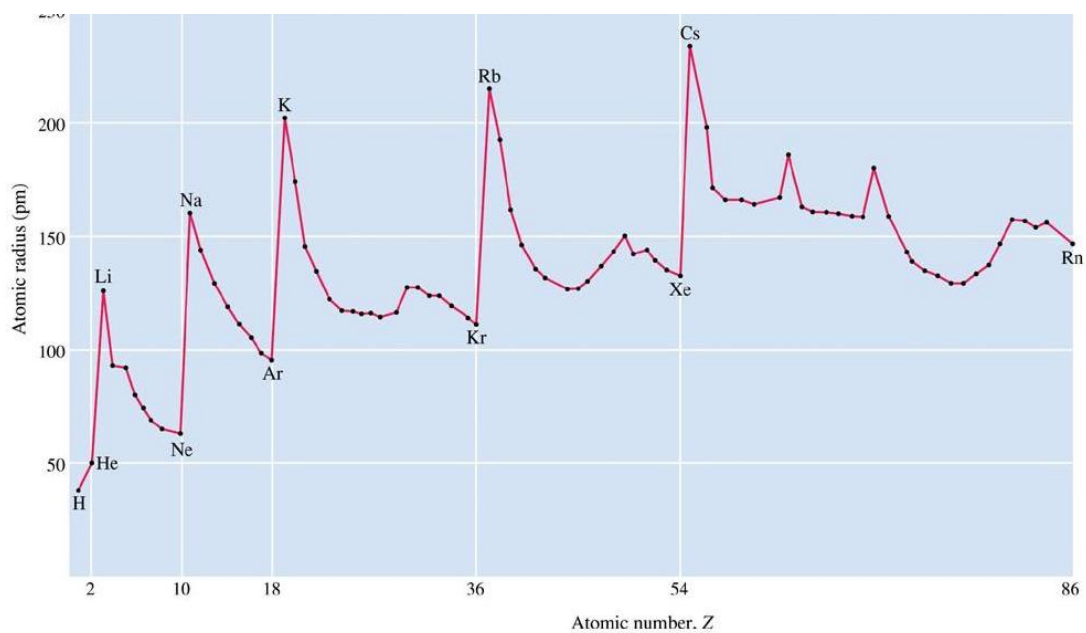
- H is the smallest atom: Covalent radius of H atom = 0.32 Å

ATOMIC RADII

Atomic Size Increases



Atomic Size Increases

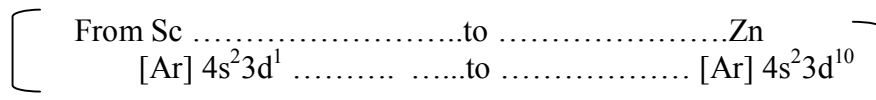


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ATOMIC RADII

Main-Group Elements												Main-Group Elements										
1 IA												13 IIIA		14 IVA		15 VA		16 VIA		17 VIIA		18 VIIIA
1	1 H 1.00794											5 B 10.811	6 C 12.011	7 N 14.00674	8 O 15.9994	9 F 18.9984032	10 Ne 20.1797					
2	3 Li 6.941	4 Be 9.012182	Transition Metals								11 Na 22.989768	12 Mg 24.3050	13 Al 26.981538	14 Si 28.0855	15 P 30.973762	16 S 32.066	17 Cl 35.453	18 Ar 39.948				
3	11 Na 22.989768	12 Mg 24.3050	3 Al 26.981538	4 Si 28.0855	5 P 30.973762	6 S 32.066	7 Cl 35.453	8 Ar 39.948	9 K 39.0983	10 Ca 40.078	21 Sc 44.955910	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.93805	26 Fe 55.845	27 Co 58.93320	28 Ni 58.6934	29 Cu 63.546	30 Zn 65.39		
4	19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.88	23 V 50.9415	24 Cr 51.9961	25 Mn 54.93805	26 Fe 55.845	27 Co 58.93320	28 Ni 58.6934	29 Cu 63.546	30 Zn 65.39	31 Ga 69.723	32 Ge 72.61	33 As 74.92159	34 Se 78.95	35 Br 79.904	36 Kr 83.80				
5	37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.29				
6	55 Cs 132.90543	56 Ba 137.327	57 La* 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.85	75 Re 186.207	76 Os 190.23	77 Ir 192.2217	78 Pt 195.08	79 Au 196.96654	80 Hg 200.59	81 Tl 204.387	82 Pb 207.2	83 Bi 208.98037	84 Po (209)	85 At (210)	86 Rn (222)				
7	87 Fr (223)	88 Ra (226)	89 Ac** (227)	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (264)	108 Hs (265)	109 Mt (266)	110 Ds (269)	111 Rg (272)	112 Cn (277)										

Note: There is little variation in atomic Size throughout a row of Transition Elements:
Reason: Consider the elements completing the 3d subshell:



Moving along period 4 from left to right:

- **1 proton is added when moving from one transition element to the next**
ACTUAL NUCLEAR CHARGE (Z) INCREASES CONSIDERABLY

Sc

- 1 electron is added to the 3d subshell of the 3rd shell (an inner shell)
SHIELDING EFFECT CAUSED BY ADDED ELECTRON INCREASES CONSIDERABLY

Recall: **Z_{eff} = Z - Shielding Effect**
increases increases
considerably considerably

Result: **Z_{eff} does not change**

- **Insignificant variation in Atomic Size for 3d block elements**
This trend holds true for all transition elements (d block elements)

IONIZATION ENERGY

- Ionization Energy (IE) is the energy needed to remove the outermost electron from a neutral atom in the gaseous state to form a positive ion.
- Ionization Energy :
 - is commonly measured in kJ/mol
 - is also referred to as Ionization Potential (electrical energy) and can also be measured in Volts
 - is listed on the Periodic Table (back side)

Consider the removal of the outermost shell electron from sodium:



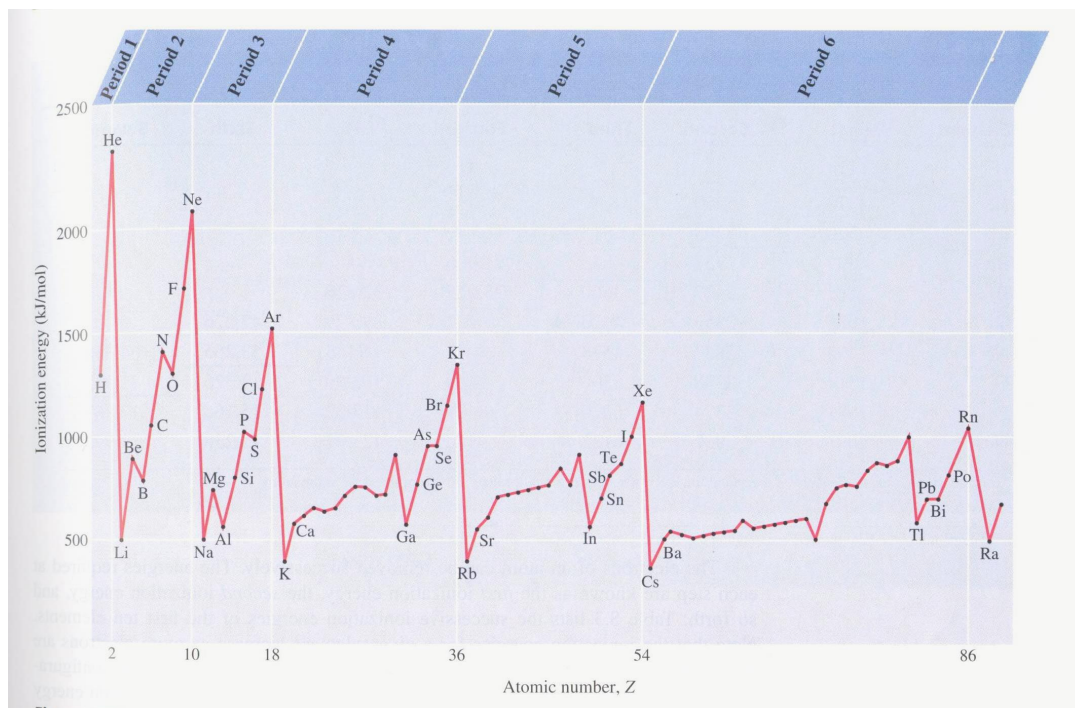
Consider the removal of a second electron:



NOTES:

1. Removal of an electron is an endothermic process (it requires energy to remove an electron)
Atoms do not “lose” electrons; Energy is required to remove the electrons
2. All the electrons can be removed successively from an atom.
The energies required at each step are known as:
First Ionization Energy (IE₁), Second Ionization Energy (IE₂), Third Ionization Energy (IE₃)
3. In general:
The Ionization Energies for a given element increase as more electrons are removed:
 $\text{IE}_1 < \text{IE}_2 < \text{IE}_3 < \text{IE}_4 < \text{IE}_5 < \text{IE}_6 < \text{IE}_7 < \text{IE}_8$ and so on.
Reason: The electron is being removed from ions with increasingly larger positive charges
(The larger the positive charge of the ion, the stronger the electron is attracted)

IONIZATION ENERGY



- Down a group IE decreases, due to increased atomic size (the larger the atom, the less strongly it holds its electrons in the outer shell)
- Across a period, IE generally increases, due to increased effective nuclear charge (the greater the nuclear charge, the stronger the atom hold its electrons in the outer shell)
- Some anomalies in IE trend occur in each period (e.g. in period 2, B is lower than Be and O is lower than N; similarly in period 3, Al is lower than Mg and S is lower than P).
- These anomalies are due to the increased stabilities of the electron shell arrangements for each atom.

- The increase in the values of successive Ionization Energies **is not gradual**:

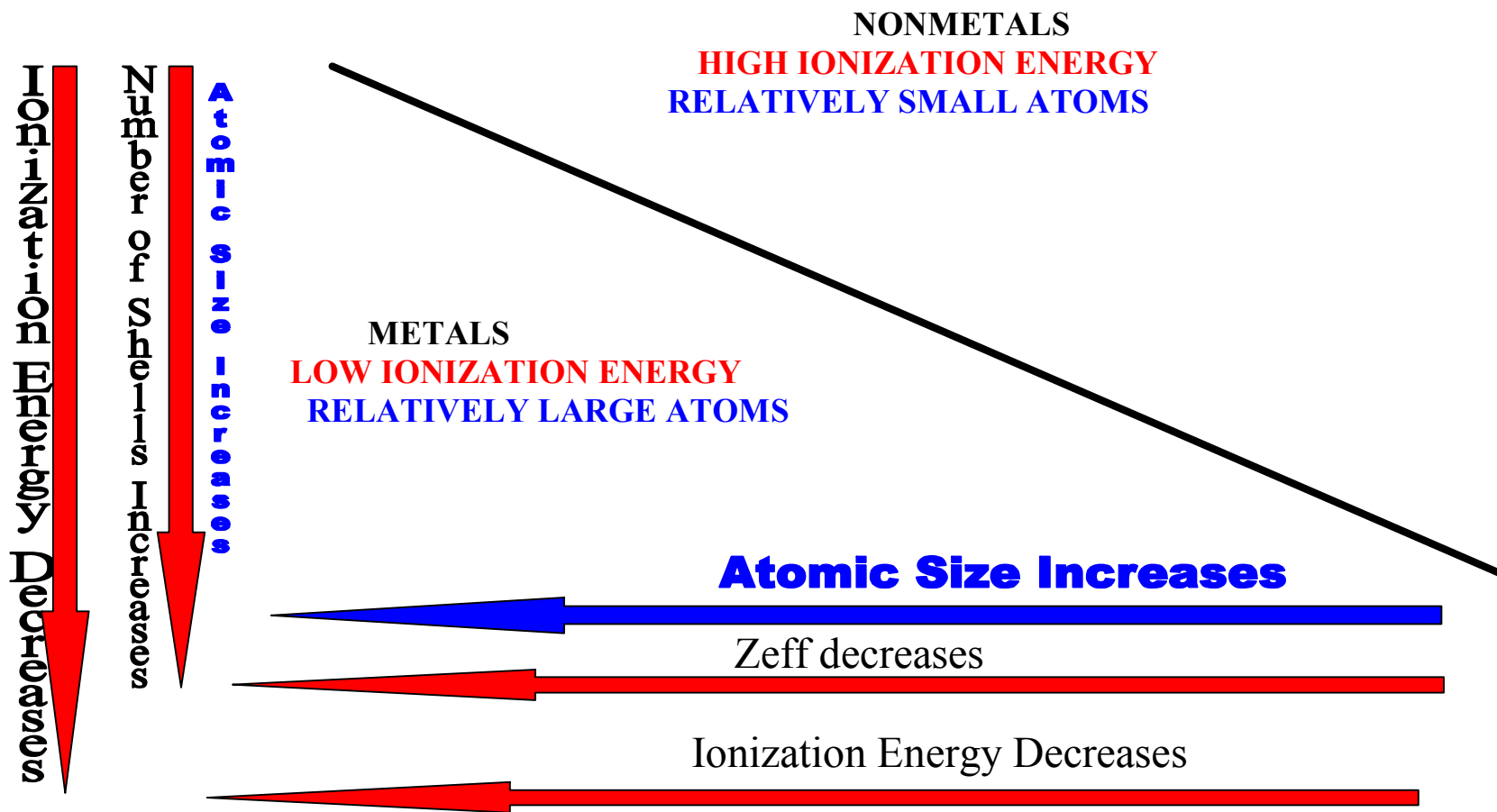
Element	Successive Ionization Energies (kJ/mol)						
	IE ₁	IE ₂	IE ₃	IE ₄	IE ₅	IE ₆	IE ₇
H	1312						
He	2372	5250					
Li	520	7298	11,815				
Be	899	1757	14,848	21,006			
B	801	2427	3660	25,025	32,826		
C	1086	2353	4620	6222	37,829	47,276	
N	1402	2857	4578	7475	9445	53,265	
O	1314	3388	5300	7469	10,989	13,326	71,333
F	1681	3374	6020	8407	11,022	15,164	17,867

- NOTE: A very large jump in the I E value occurs after all the valence electrons have been removed

Element	Successive Ionization Energies (kJ/mol)						
	IE ₁	IE ₂	IE ₃	IE ₄	IE ₅	IE ₆	IE ₇
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	
O	1314	3388	5300	7469	10,989	13,326	71,333
F	1681	3374	6020	8407	11,022	15,164	17,867

- Once the valence electrons are removed, a stable noble-gas configuration is obtained. Further removal of electrons will involve electrons from a complete inner shell.
- For purposes of comparing the chemical behavior of different elements, usually only the First Ionization Energy (IE₁) is compared.

SUMMARY OF ATOMIC RADIUS & IONIZATION ENERGY



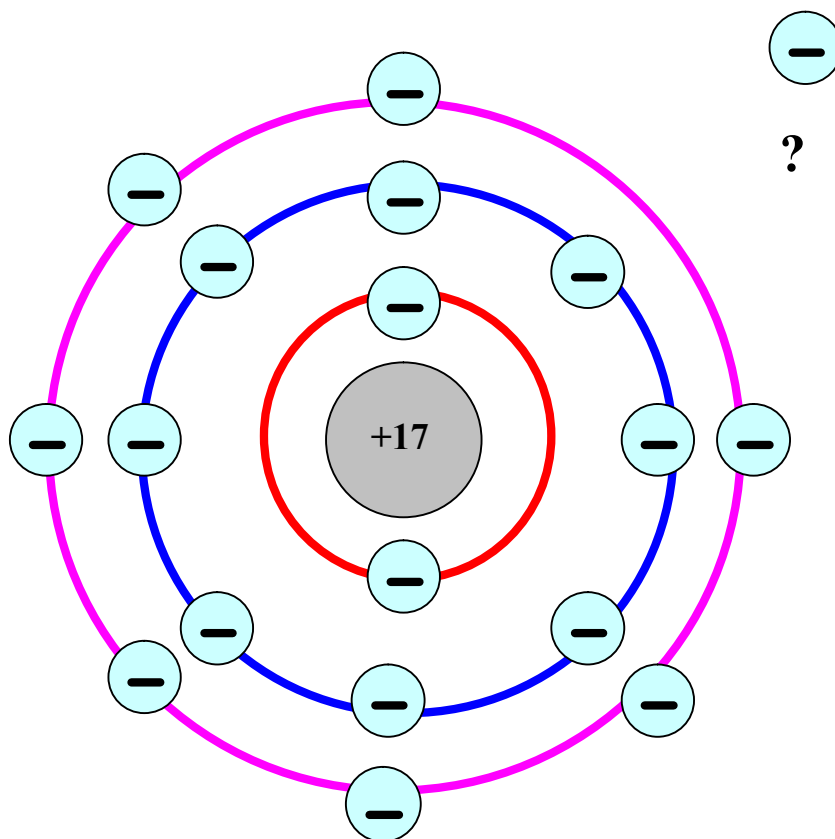
ELECTRON AFFINITY (EA)

- Consider the electron configuration of the Cl atom ($Z = 17$)



3 energy levels:

- 2 electrons on the 1st energy level
- 8 electrons on the 2nd energy level
- 7 electrons on the 3rd energy level

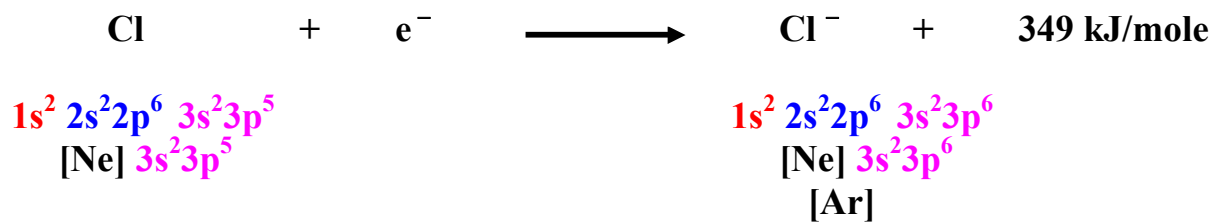
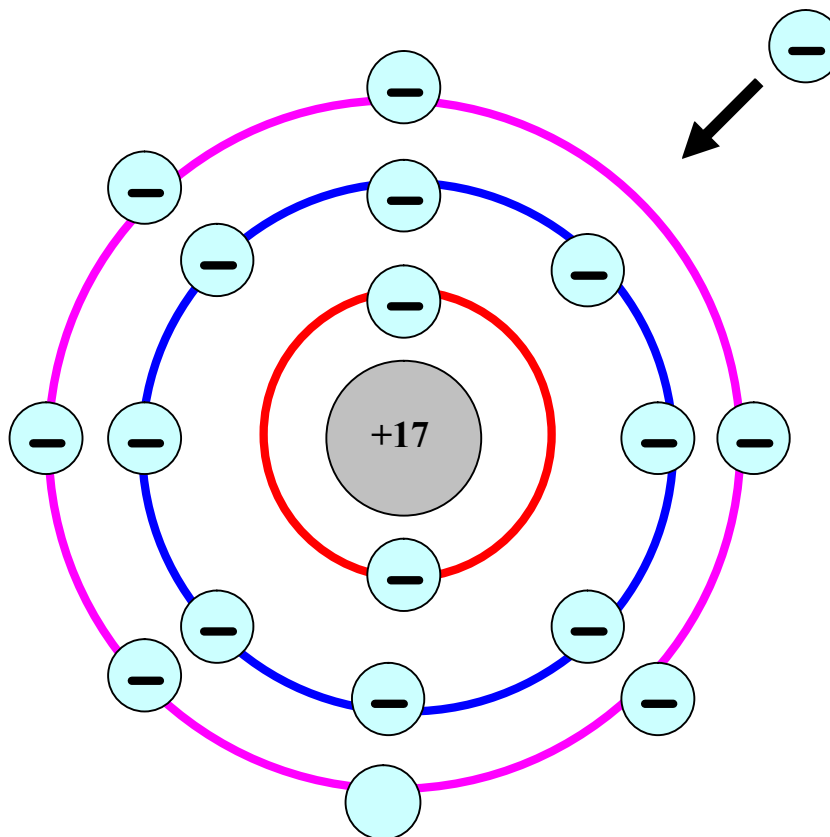


- What is the effect (if any) of the Cl atom on an outside electron ?

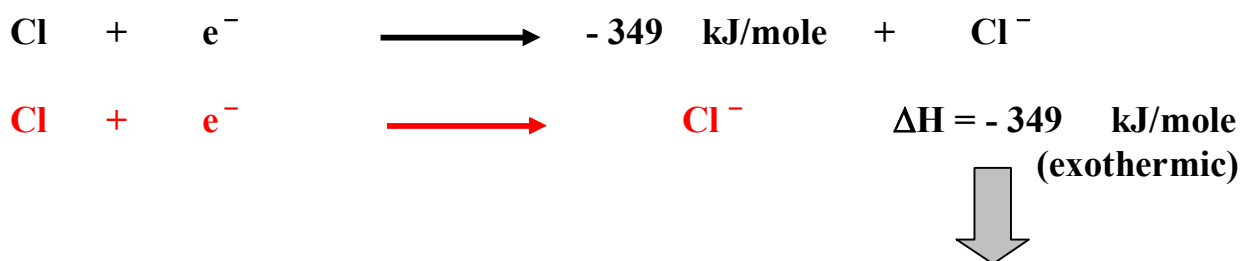
Hypothetically, there are 3 possibilities:

- The atom is neutral (+ 17 nuclear charge is canceled out by the 17 negative electrons)
As such, **the outside electron should be unaffected by the neighboring Cl atom**
- The 17 electrons of the Cl atom (charge = -17) **should repel the outside electron**
- The 17 electrons of the Cl atom cancel out the positive nuclear charge of +17 only partially (shielding effect). The positively charged nucleus creates an electric field which **should attract the outside electron**

- The outside electron is attracted to the Cl atom, and is accepted by the Cl atom.
- The addition of the electron to the Cl atom does not require energy (not endothermic)
- The electron is “naturally” attracted and releases energy (exothermic process)



Recall that the heat term is always written on the reactant side (Thermochemistry)



ELECTRON AFFINITY

- ELECTRON AFFINITY (EA)** is the energy change (commonly released) for the process of adding an electron to a neutral atom in the gaseous state to form a negative ion. (it is a measure of how easily an atom accepts an electron)

NOTE:

- The majority of atoms have a high tendency to gain electrons (ex: Cl)

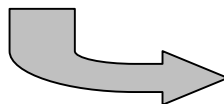
The atoms have **HIGH EA's**

HIGH EA

means



LARGE NEGATIVE EA



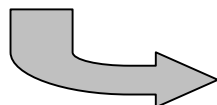
Large Amount of Energy given off when an electron is gained "willingly" by the neutral atom (stable ion formed) (addition of electron is exothermic)

- Relatively few atoms resist the addition of an electron(ex:Noble Gases and some metals)

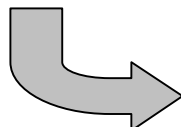
They have **LOW EA's**

LOW EA

means

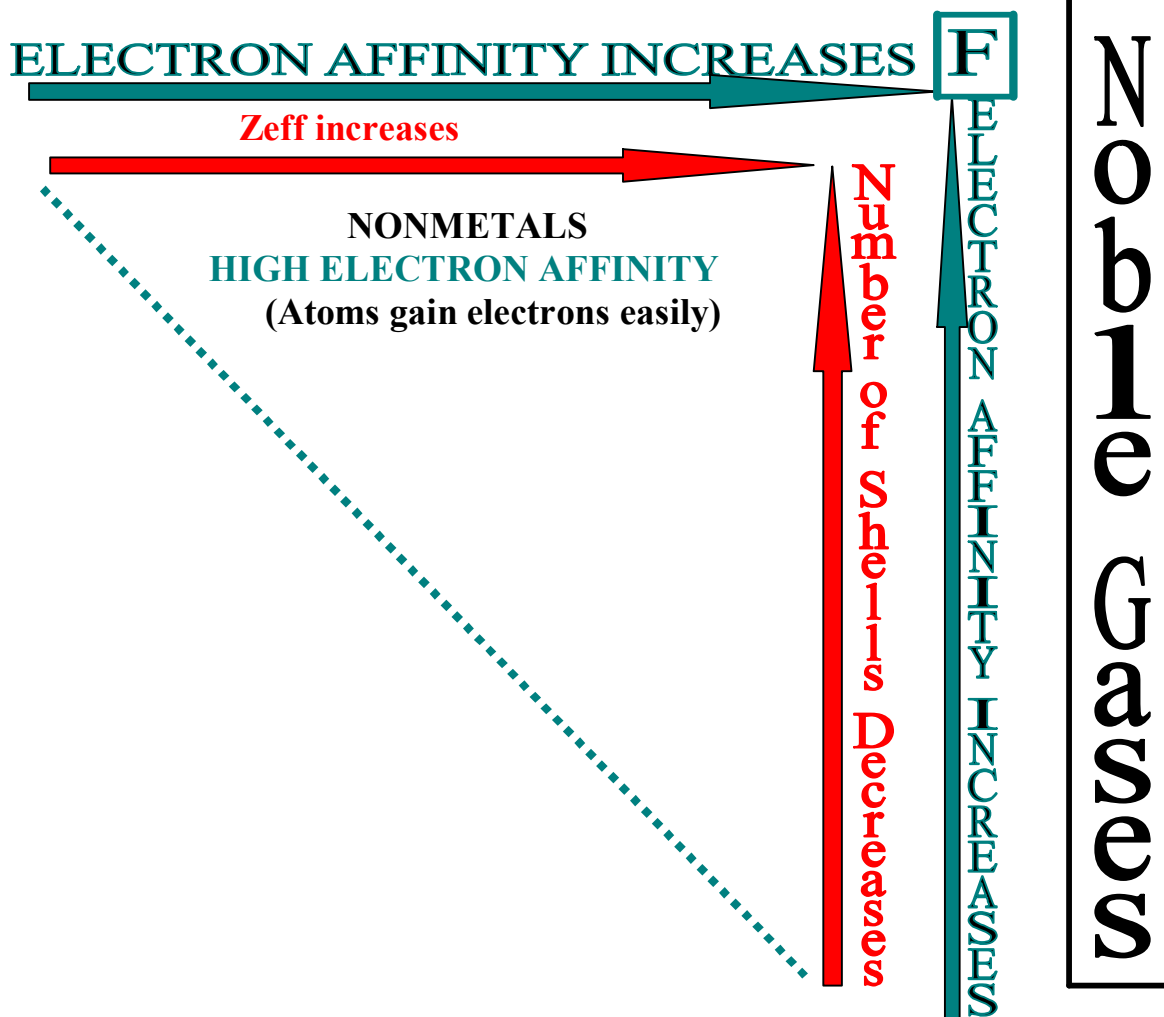


Zero or Positive EA



Energy is required when an electron is forced in the neutral atom (unstable ion formed) (addition of the electron is endothermic)

ELECTRON AFFINITY TRENDS



VARIATIONS IN ELECTRON AFFINITIES

I. Effect caused by Interelectron repulsions between outer shell electrons

	IVA	VA	VIA	VIIA
Period 2	C	N	O	F
Period 3	Si	P	S	Cl

Expected Trend of EA



EA
Should
increase

	IVA	VA	VIA	VIIA
Period 2	C	N	O	F
EA (kJ/mol)	-122	0	-141	-328
Period 3	Si	P	S	Cl
EA (kJ/mol)	-134	-72	-200	-349

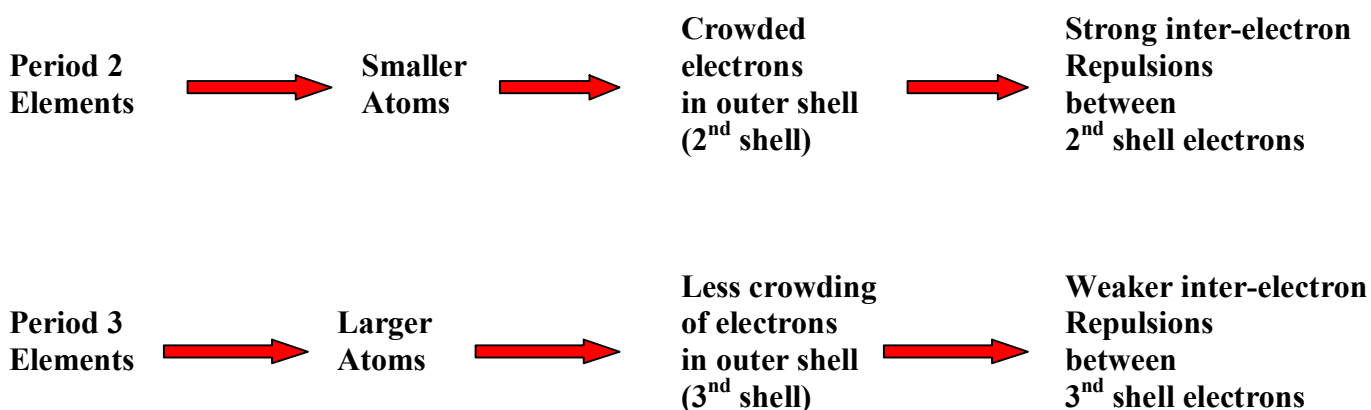
Actual Trend of EA



EA
increases

Reason:

- Atoms of Period 2 elements are much, much smaller than the atoms of Period 3 elements (large energy gap between the 2nd and the 3rd energy levels; this energy gap diminishes as one moves to 4th, 5th, 6th, 7th energy levels)
- In small atoms there is very strong inter-electron electron repulsion (electrons are crowded). The mutual repulsions between 2nd shell electrons act against adding an additional electron to the atom.






II. Effect caused by the stability of half-filled subshells

	IIIA	IVA	VA	VIA	VIIA
Period 2	B	C	N	O	F
Period 3	Al	Si	P	S	Cl

Expected trend of EA  EA should increase

	IIIA	IVA	VA	VIA	VIIA
Period 2 E.A. (kJ/mol)	B -27	C -122	N 0	O -141	F -328
Period 3 E.A. (kJ/mol)	Al -44	Si -134	P -72	S -200	Cl -349

E.A. increases   E.A. increases 

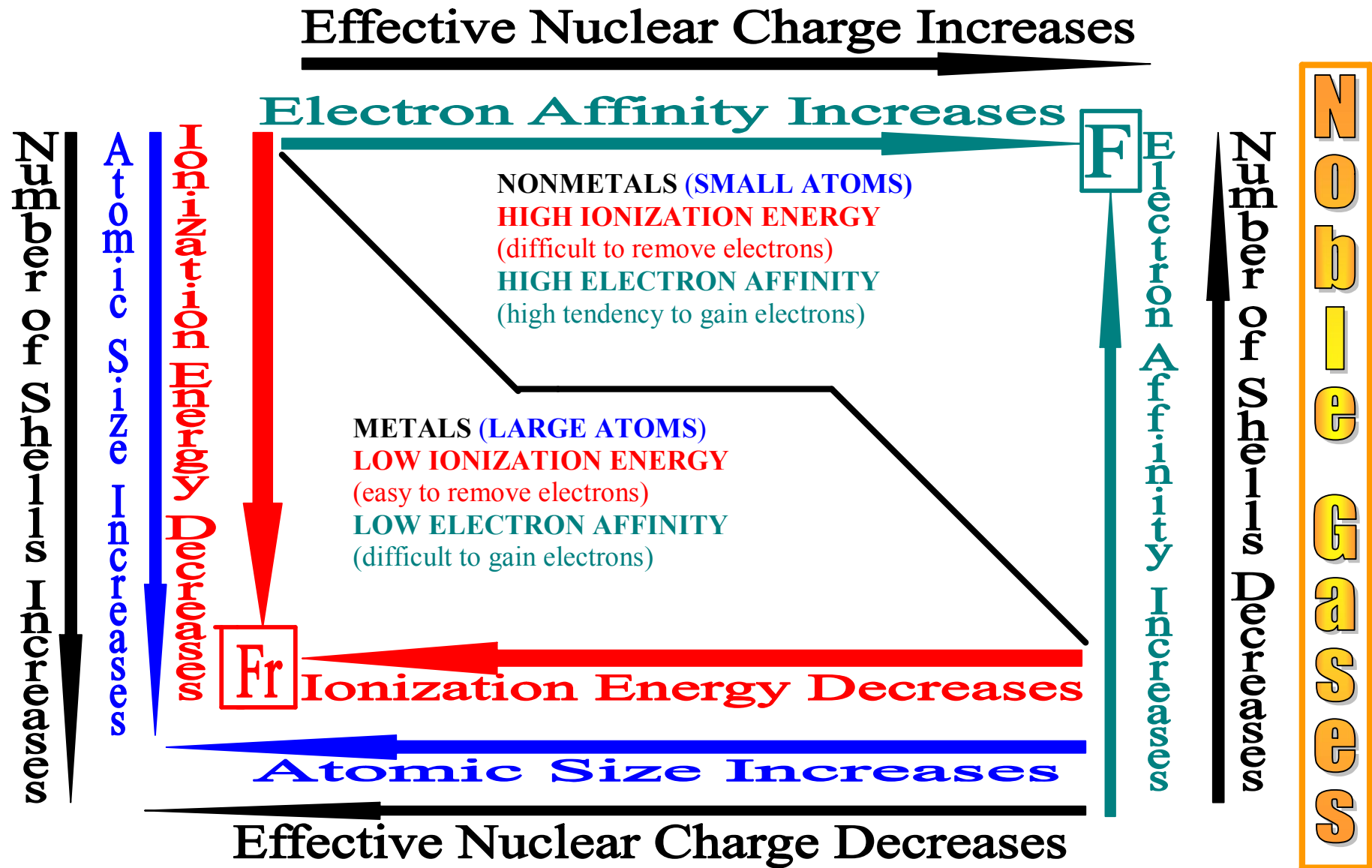
Break in trend !
E.A. is less than expected !

Reason:

- The more stable half-filled shell of group VA elements rejects electrons, leading to a lower EA.

	IVA	VA
Period 2	C [He] $\uparrow\downarrow_{2s}$ \uparrow_{2p} \uparrow_{2p} $\underline{\hspace{0.5cm}}$	N [He] $\uparrow\downarrow_{2s}$ \uparrow_{2p} \uparrow_{2p} \uparrow_{2p} $\underline{\hspace{0.5cm}}$
Period 3	Si [He] $\uparrow\downarrow_{3s}$ \uparrow_{3p} \uparrow_{3p} $\underline{\hspace{0.5cm}}$	P [He] $\uparrow\downarrow_{3s}$ \uparrow_{3p} \uparrow_{3p} $\underline{\hspace{0.5cm}}$

SUMMARY OF TRENDS IN PERIODIC PROPERTIES



METALLIC CHARACTER

- The physical and chemical properties of the main-group elements clearly display periodic character.
- The classification of elements in the periodic table into metals, non-metals and metalloids is based on fundamental periodic trends.
- The more an element exhibits the physical and chemical properties of metals, the greater its **metallic character**.

Metallic character increases

1 IA												18 VIIIA					
1 H 1.00794	2 He 4.002602											13 Al 26.981538	14 Si 28.0855	15 P 30.973761	16 S 32.065	17 Cl 35.453	18 Ar 39.948
3 Li 6.941	4 Be 9.012182											5 B 10.811	6 C 12.0107	7 N 14.0067	8 O 15.9994	9 F 18.9984032	10 Ne 20.1797
11 Na 22.989770	12 Mg 24.3050	3 III B	4 IV B	5 V B	6 VI B	7 VII B	8 VIII B			10 IB	12 IIB	13 Al 26.981538	14 Si 28.0855	15 P 30.973761	16 S 32.065	17 Cl 35.453	18 Ar 39.948
19 K 39.0983	20 Ca 40.078	21 Sc 44.955910	22 Ti 47.867	23 V 50.9415	24 Cr 51.9961	25 Mn 54.938049	26 Fe 55.845	27 Co 58.933200	28 Ni 58.6934	29 Cu 63.546	30 Zn 65.409	31 Ga 69.723	32 Ge 72.64	33 As 74.92160	34 Se 78.96	35 Br 79.904	36 Kr 83.798
37 Rb 85.4678	38 Sr 87.62	39 Y 88.90585	40 Zr 91.224	41 Nb 92.90638	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.90550	46 Pd 106.42	47 Ag 107.8682	48 Cd 112.411	49 In 114.818	50 Sn 118.710	51 Sb 121.760	52 Te 127.60	53 I 126.90447	54 Xe 131.293
55 Cs 132.90545	56 Ba 137.327	57 La* 138.9055	72 Hf 178.49	73 Ta 180.9479	74 W 183.84	75 Re 186.207	76 Os 190.23	77 Ir 192.217	78 Pt 195.078	79 Au 196.96655	80 Hg 200.59	81 Tl 204.3833	82 Pb 207.2	83 Bi 208.98038	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89 Ac** (227)	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (277)	109 Mt (268)	110 Uun (281)	111 Uuu (272)	112 Uub (285)	114 Uuq (289)		116 Uuh (292)			

- The basic-acidic behavior of the oxides of elements is a good indicator of their metallic character.
- Oxides are classified as basic or acidic based on their reactions with acids and bases.
- A basic oxide is an oxide that reacts with acids. Most metal oxides are basic.
- An acidic oxide is an oxide that reacts with bases. Most non-metal oxides are acidic oxides.
- An amphoteric oxide is an oxide that has both basic and acidic properties.