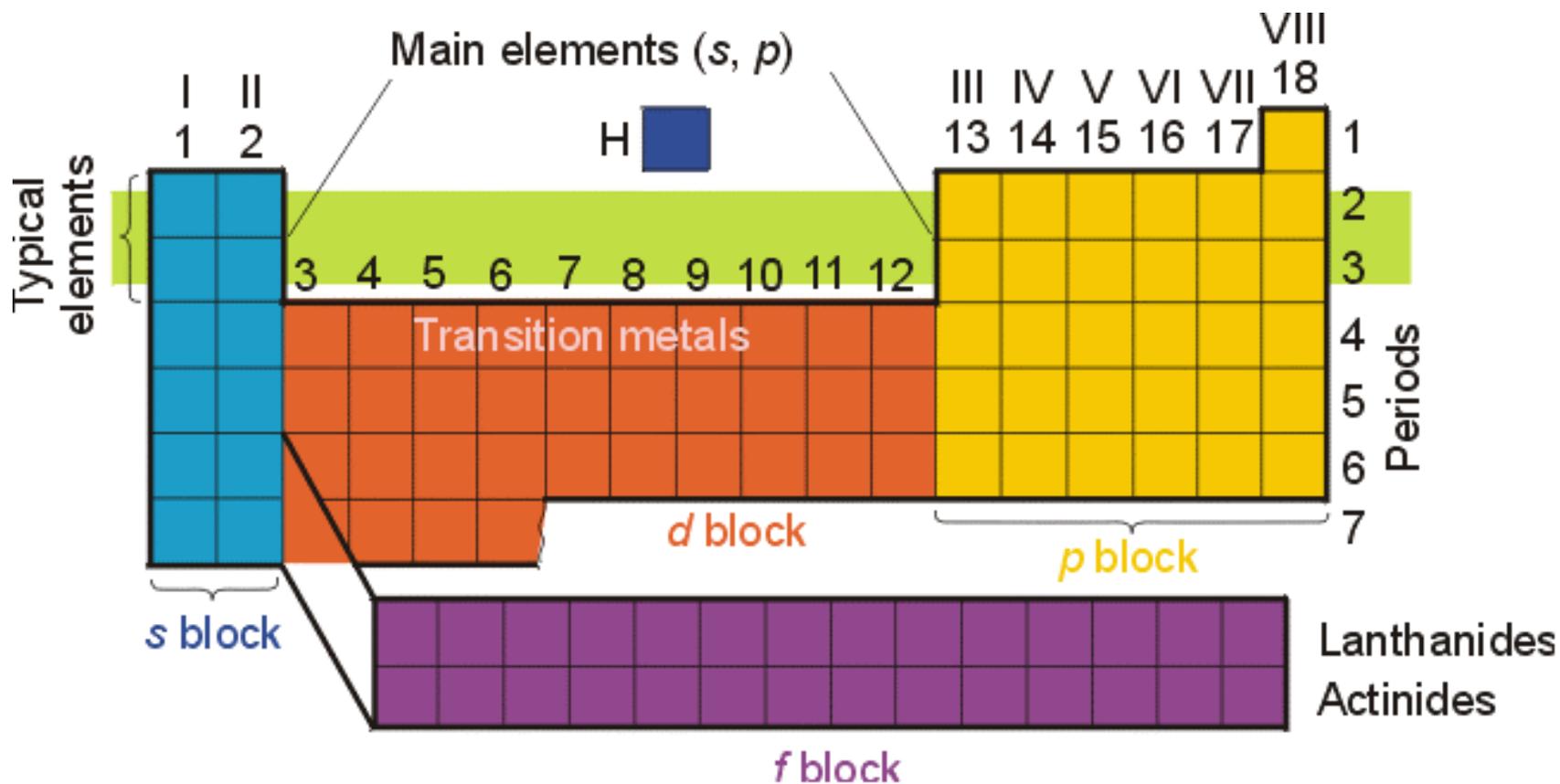


# Modern Main Group Chemistry

Primarily, the chemistry of the p-block elements.



# Chem 59-651

- Group 1 - Alkali metals
- Group 2 - Alkaline earth metals
- Group 13 - Earth metals (Triels)
- Group 14 - no name

- Group 15 - Pnictogens (Pentels)
- Group 16 - Chalcogens
- Group 17 - Halogens
- Group 18 - Noble Gases

*s*-block: 1 New Designation, IA Original Designation

*s*-block: 18, VIIIA

Non-Metals: 13-17, IIIA-VIIA

Atomic #, Symbol, Atomic Mass

*d*-block: Transition Metals

*p*-block

Phases: Solid, Liquid, Gas

(Mass Numbers in Parentheses are from the most stable of common isotopes.)

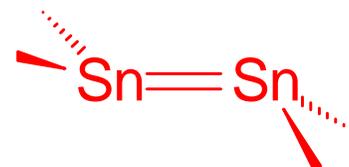
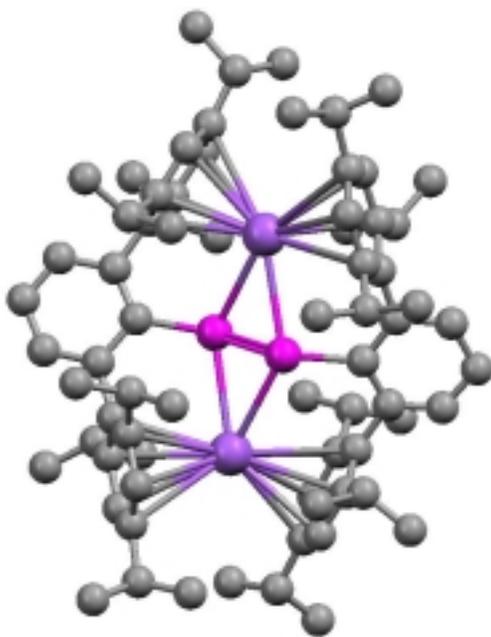
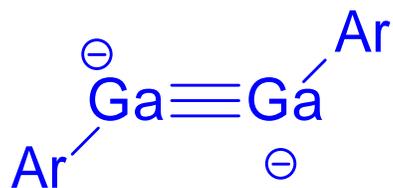
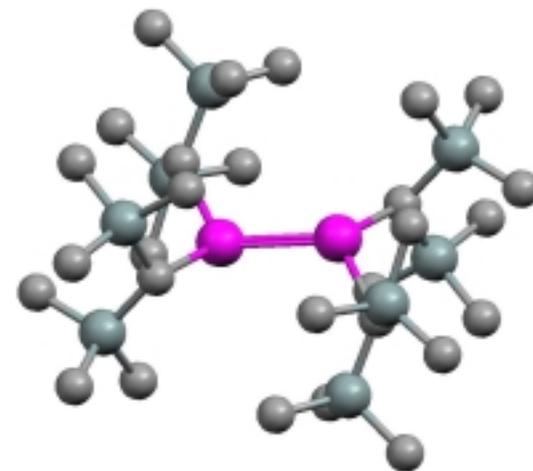
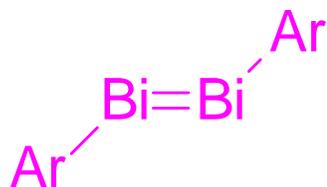
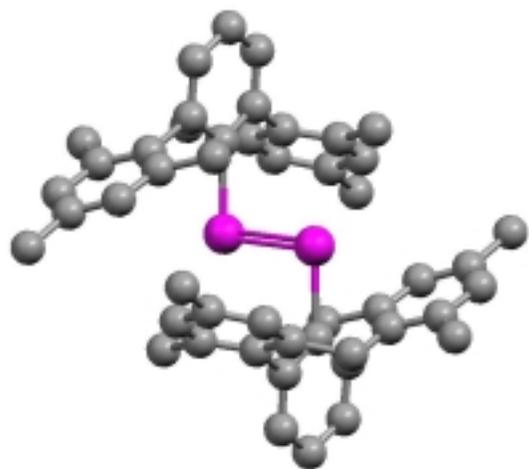
Rare Earth Elements

Lanthanide Series

Actinide Series

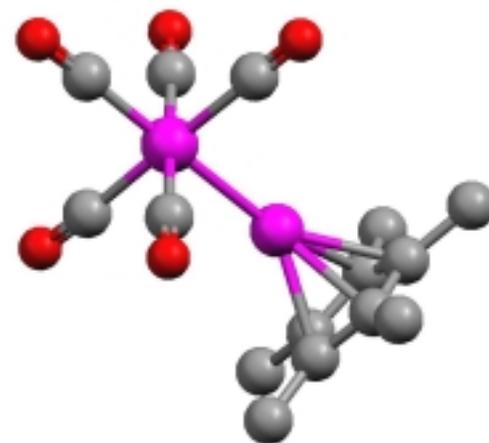
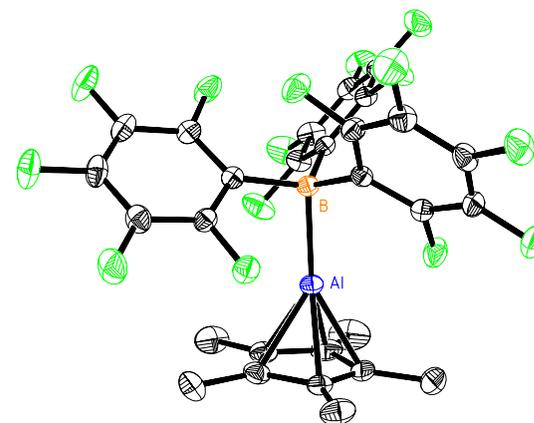
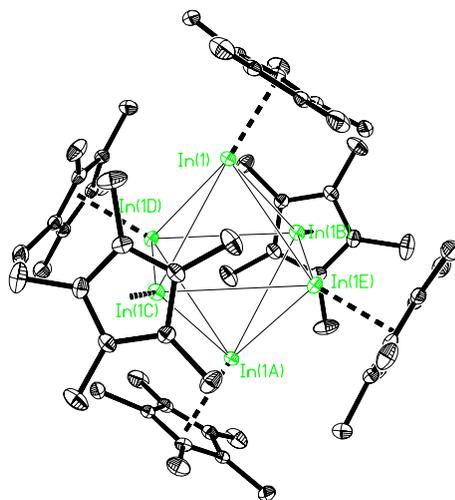
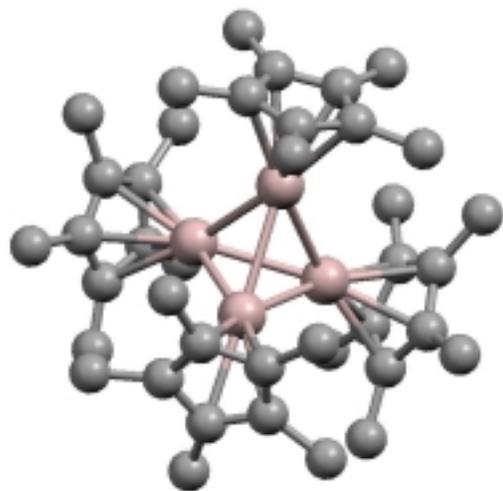
1	1 H 1.0094	2 He 4.00260																	
2	3 Li 6.941	4 Be 9.0122	<i>d</i> -block										5 B 10.81	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.179	
3	11 Na 22.990	12 Mg 24.305	13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.06	17 Cl 35.453	18 Ar 39.948	<i>p</i> -block										
4	19 K 39.098	20 Ca 40.08	21 Sc 44.956	22 Ti 47.88	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.847	27 Co 58.933	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.80	
5	37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.91	54 Xe 131.29	
6	55 Cs 132.91	56 Ba 137.33	57 La to 71	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)	
7	87 Fr (223)	88 Ra 226.03	89 Ac to 103	104 Unq (261)	105 Unp (262)	106 Unh (263)	107 Uns (262)	108 Uno (265)	109 Une (266)	110 Uun (267)									
			<i>d</i> -block										<i>f</i> -block						
			57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97		
			89 Ac 227.03	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np 237.05	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)		

## Multiple bonds between main group elements

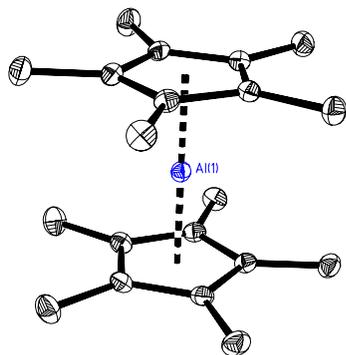


What do such compounds teach us about the rules of bonding ?

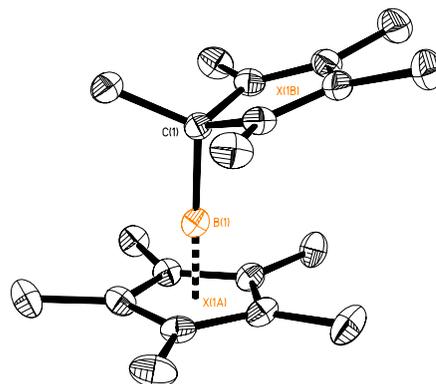
# Low oxidation state and cluster compounds



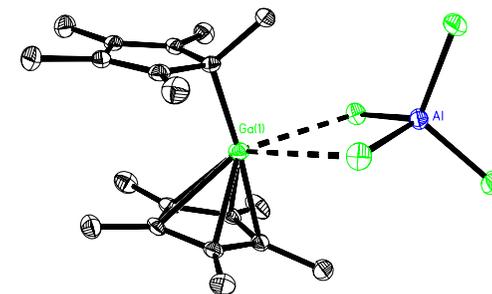
# Main Group Metalloenes



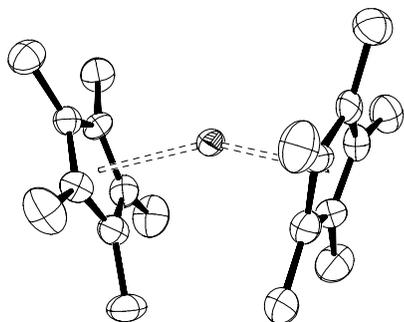
$\text{Cp}^*_2\text{M}$  (M = Be, Al, Mg)



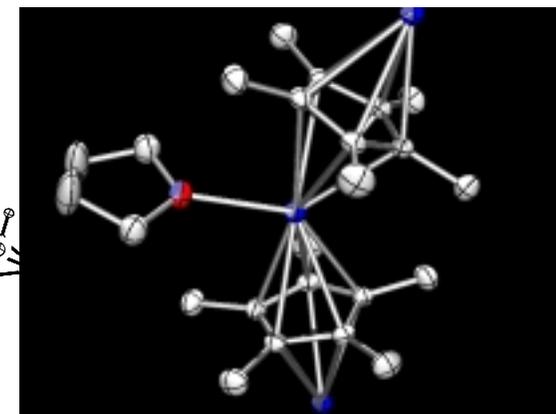
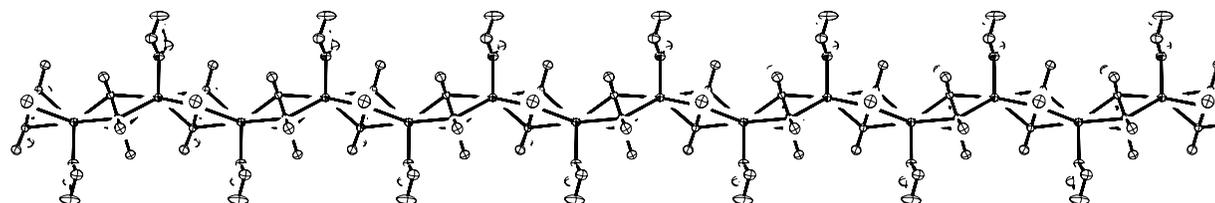
$\text{Cp}^*_2\text{B}^+$



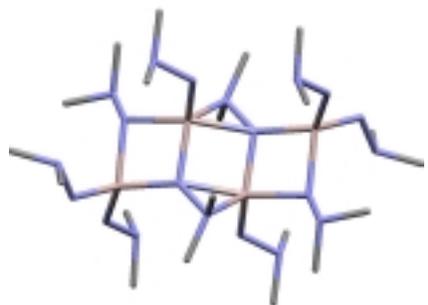
$[\text{Cp}^*_2\text{Ga}^+][\text{AlCl}_4^-]$



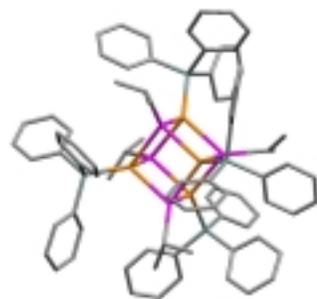
$\text{Cp}^*_2\text{M}$  (M = Si, Ge, Sn)



# Single-source Precursors for Materials



AlN



InP

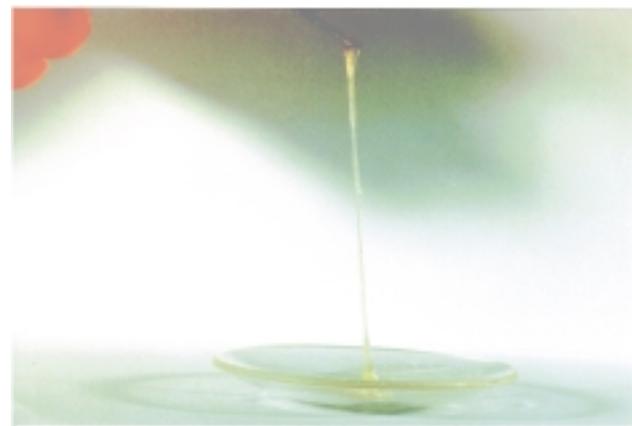
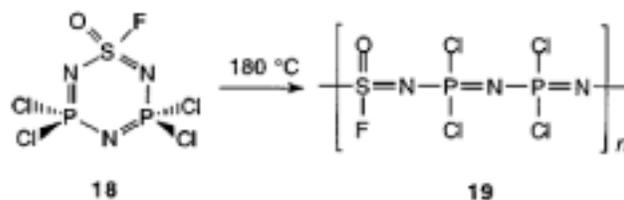
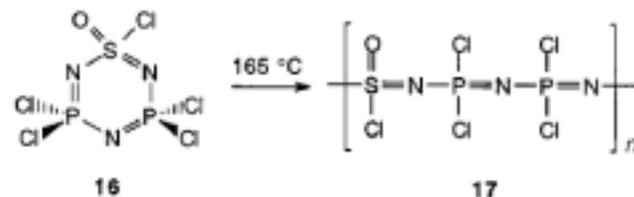
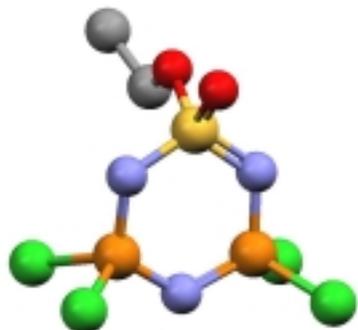
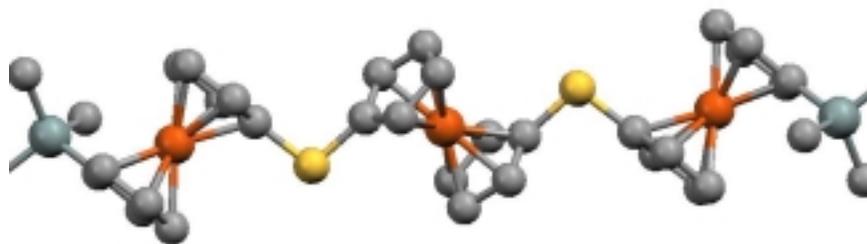
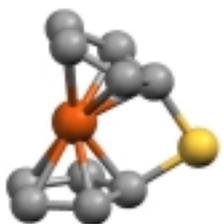
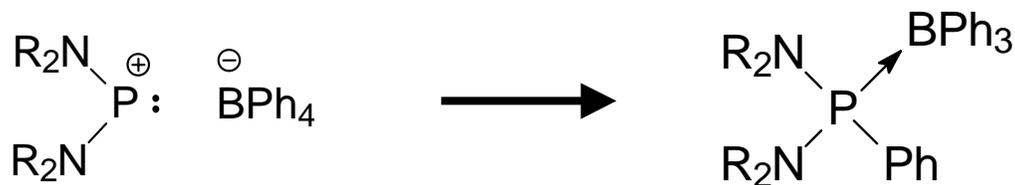
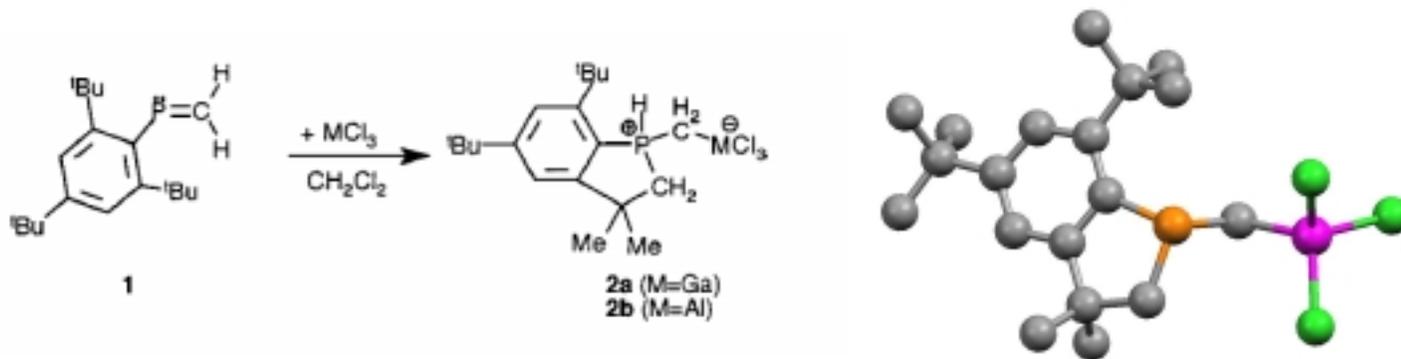
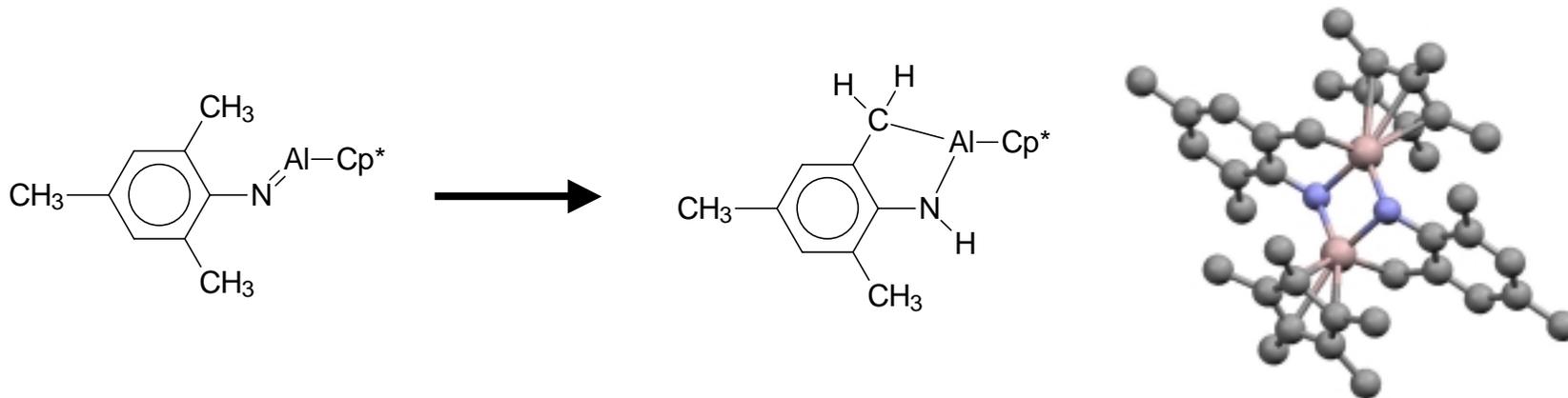


Fig. 5 A typical polythiophosphazene elastomer

# Highly-reactive species: C-H and other bond activation



## Concepts/Approaches that we must review

- Some aspects of periodicity (effective nuclear charge, electronegativity, the sizes of atoms and orbitals, contractions, relativistic effects)
- Valence Shell Electron Repulsion Theory (shapes)
- Valence Bond Theory (hybridization of orbitals, the use of d-orbitals)
- Molecular Orbital Theory and diatomic MO diagrams
- Symmetry and Point Groups
- Group Theory, Character Tables and polyatomic MO diagrams

**You can find most of these in Inorganic text books or on my web site.**

# Trends for Atomic Properties in the Periodic Table

Understanding how and why properties change from element to element requires us to consider:

1. The electron configuration of the atom or ion (the filling order)
2. The type of valence orbitals involved (size, shape, shielding and penetration)
3. The effective nuclear charge felt by electrons in valence orbitals
4. "Oddities"



**PERIODIC CHART OF THE ELEMENTS**

IA	IIA	IIIB	IVB	VB	VIB	VIIIB	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	VIIIA	INERT GASES	
1 H 1.00797															1 H 1.00797	2 He 4.0026	
3 Li 6.939	4 Be 9.0122									5 B 10.811	6 C 12.0112	7 N 14.0067	8 O 15.9994	9 F 18.9984	10 Ne 20.183		
11 Na 22.9898	12 Mg 24.312									13 Al 26.9815	14 Si 28.086	15 P 30.9738	16 S 32.064	17 Cl 35.453	18 Ar 39.948		
19 K 39.102	20 Ca 40.08	21 Sc 44.956	22 Ti 47.90	23 V 50.942	24 Cr 51.996	25 Mn 54.9380	26 Fe 55.847	27 Co 58.9332	28 Ni 58.71	29 Cu 63.54	30 Zn 65.37	31 Ga 69.72	32 Ge 72.59	33 As 74.9216	34 Se 78.96	35 Br 79.909	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.905	40 Zr 91.22	41 Nb 92.906	42 Mo 95.94	43 Tc (99)	44 Ru 101.07	45 Rh 102.905	46 Pd 106.4	47 Ag 107.870	48 Cd 112.40	49 In 114.82	50 Sn 118.69	51 Sb 121.75	52 Te 127.60	53 I 126.904	54 Xe 131.30
55 Cs 132.905	56 Ba 137.34	*57 La 138.91	72 Hf 178.49	73 Ta 180.948	74 W 183.85	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.09	79 Au 196.967	80 Hg 200.59	81 Tl 204.37	82 Pb 207.19	83 Bi 208.980	84 Po (210)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	†89 Ac (227)	104 Rf (261)	105 Db (262)	106 Sg (262)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 ? (271)	111 ? (272)	112 ? (277)						

Numbers in parenthesis are mass numbers of most stable or most common isotope.

Atomic weights corrected to conform to the 1963 values of the Commission on Atomic Weights.

The group designations used here are the former Chemical Abstract Service numbers.

\* Lanthanide Series

58 Ce 140.12	59 Pr 140.907	60 Nd 144.24	61 Pm (147)	62 Sm 150.35	63 Eu 151.96	64 Gd 157.25	65 Tb 158.924	66 Dy 162.50	67 Ho 164.930	68 Er 167.26	69 Tm 168.934	70 Yb 173.04	71 Lu 174.97
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† Actinide Series

90 Th 232.038	91 Pa (231)	92 U 238.03	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (249)	99 Es (254)	100 Fm (253)	101 Md (256)	102 No (256)	103 Lr (257)
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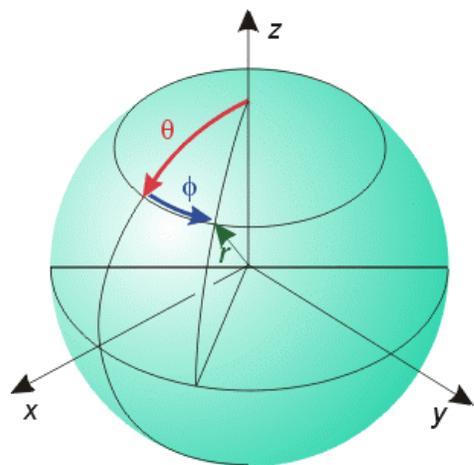


# Wave functions and Orbitals

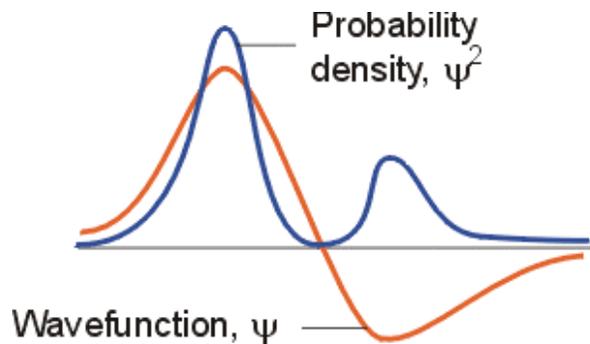
$\Psi$  = wave function

$\Psi^2$  = probability density

$4\pi r^2 \Psi^2$  = radial probability function



polar coordinates

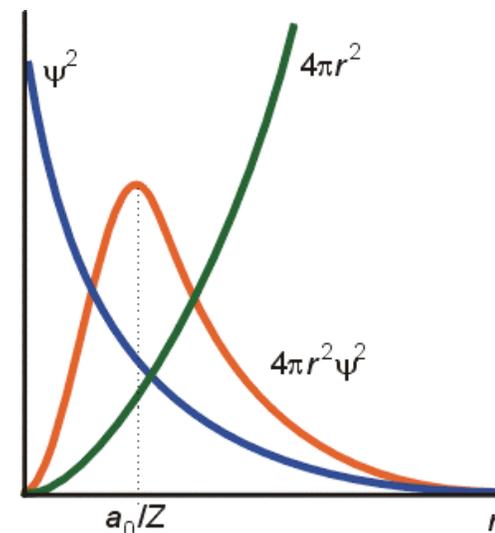


$$\hat{H}\Psi = E\Psi$$

$$\Psi = R_{nl}(r) Y_{lm_l}(\theta, \phi)$$

$R_{nl}(r)$  – radial function

$Y_{lm_l}(\theta, \phi)$  – angular function



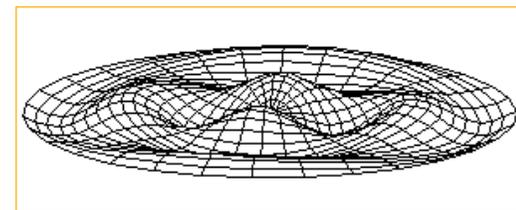
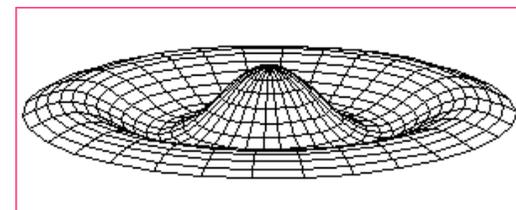
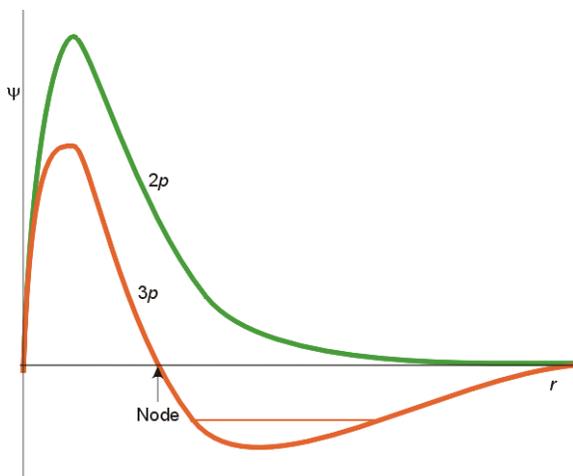
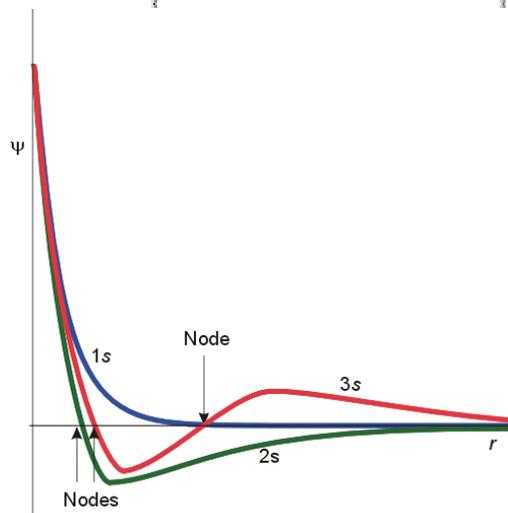
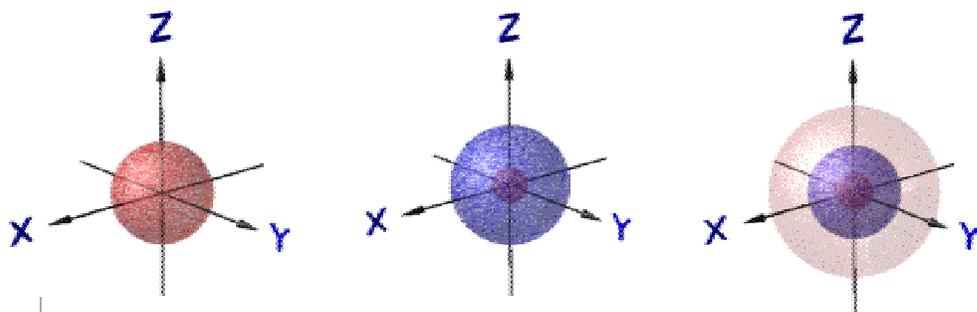
Nodes: surfaces where there is 0 probability of finding an electron

Number of radial nodes =  $n - l - 1$

Number of angular nodes: 0 for s orbitals

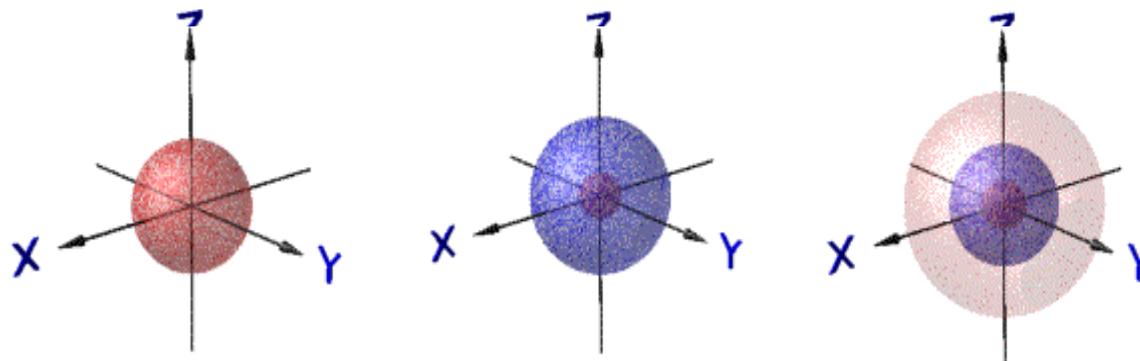
1 for p orbitals

2 for d orbitals (except  $d_{z^2}$ )

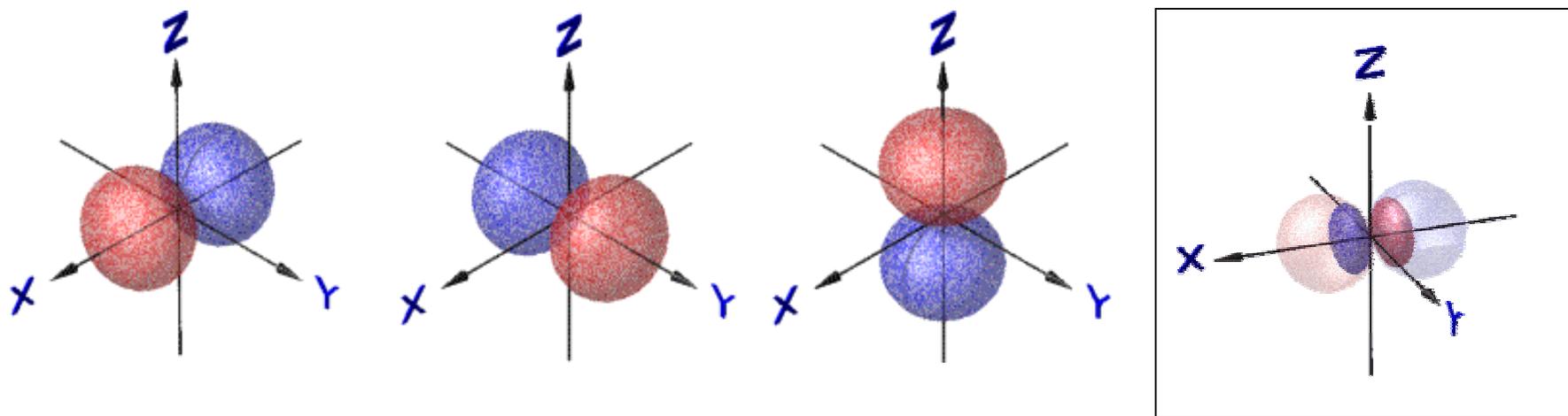


An applet demonstrating nodes for 2D wavefunctions can be found at:

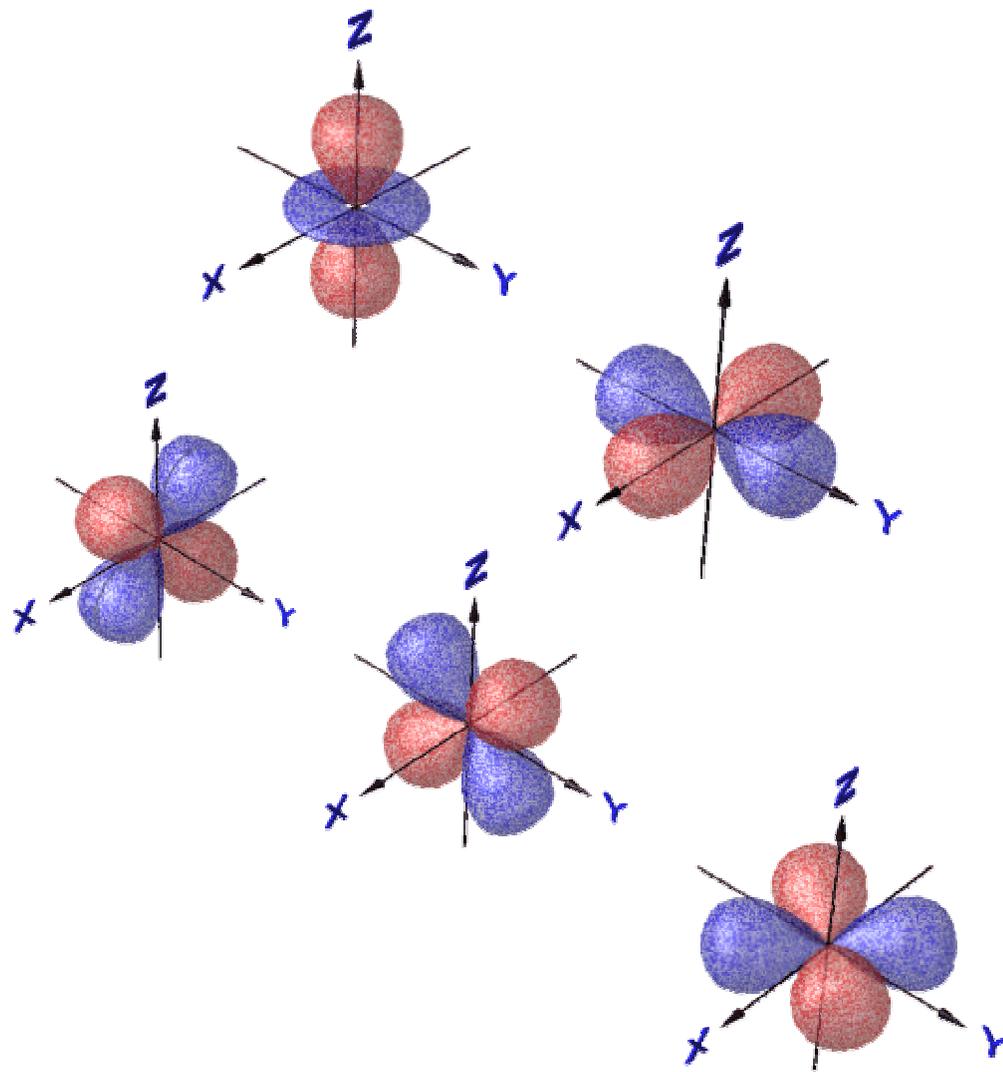
<http://artsci-ccwin.concordia.ca/facstaff/a-c/bird/c241/java/drums/drums.html>



s orbitals



p orbitals



d orbitals

## Effective Nuclear Charge, $Z^*$

The presence of other electrons around a nucleus “screens” an electron from the full charge of the nucleus.

We can approximate the energy of the electrons by modifying the Bohr equation to account for the lower “effective” nuclear charge:

$$E_n = -R \left( \frac{Z^{*2}}{n^2} \right)$$

$$Z^* = Z - \sigma$$

$Z^*$  is the effective nuclear charge

$Z$  is the atomic number

$\sigma$  is the shielding or screening constant

Slater's rules for the prediction of  $\sigma$  for an electron:

1. Group electron configuration as follows:  
(1s)(2s,2p)(3s,3p)(3d)(4s,4p)(4d)(4f)(5s,5p) etc.
2. Electrons to the right (in higher subshells and shells) of an electron do not shield it.
3. For ns or np valence electrons:
  - a) each other electron in the same group contributes 0.35 (0.30 for 1s)
  - b) each electron in an n-1 group contributes 0.85
  - c) each electron in an n-2 or lower group contributes 1.00
4. For nd or nf valence electrons:
  - a) each other electron in the same group contributes 0.35
  - b) each electron in a lower group (to the left) contributes 1.00

Example with a valence electron on oxygen: O,  $Z = 8$

Electron configuration:  $1s^2 2s^2 2p^4$

a)  $(1s^2) (2s^2 2p^4)$

$$Z^* = Z - \sigma$$

b)  $\sigma = (2 * 0.85) + (5 * 0.35) = 3.45$

$$Z^* = Z - \sigma$$

$$Z^* = 8 - 3.45 = 4.55$$

This electron is actually held with about 57% of the force that one would expect for a +8 nucleus.

Example with two electrons for nickel: Ni,  $Z = 28$

Electron configuration:  $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

$(1s^2) (2s^2 2p^6) (3s^2 3p^6) (3d^8) (4s^2)$

$$Z^* = Z - \sigma$$

For a 3d electron:

$$\sigma = (18 * 1.00) + (7 * 0.35) = 20.45$$

1s,2s,2p,3s,3p                      3d

$$Z^* = Z - \sigma \quad Z^* = 28 - 20.45 = 7.55$$

For a 4s electron:

$$\sigma = (10 * 1.00) + (16 * 0.85) + (1 * 0.35) = 23.95$$

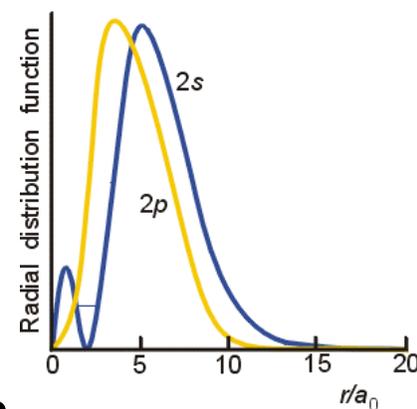
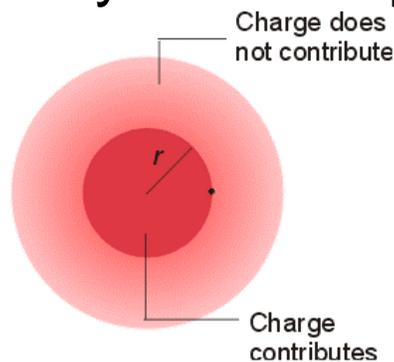
1s,2s,2p                      3s,3p,3d                      4s

$$Z^* = Z - \sigma \quad Z^* = 28 - 23.95 = 4.05$$

The basis of Slater's rules for  $\sigma$

s and p orbitals have better "penetration" to the nucleus than d (or f) orbitals for any given value of n

i.e. there is a greater probability of s and p electrons being near the nucleus



This means:

1. ns and np orbitals completely shield nd orbitals
2. (n-1) s and p orbitals don't completely shield n s and p orbitals



## Shielding and Effective Nuclear Charge

**The energy of valence electrons in an atom/ion changes with the loss of addition of an electron.**

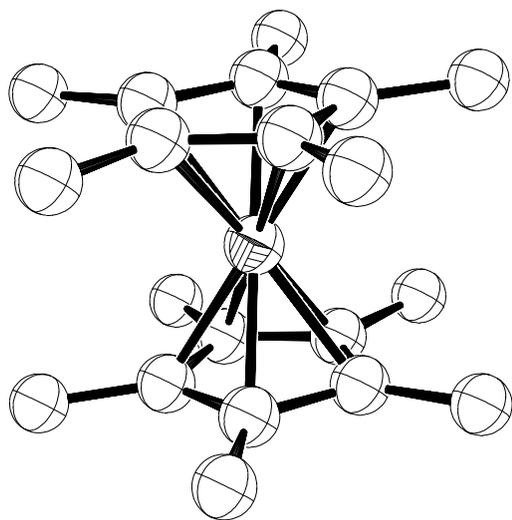
Slater's rules are only approximate and can give poor predictions. For example:

They ignore the differences in penetration between s and p orbitals. Real s and p orbitals do not have the same energy.

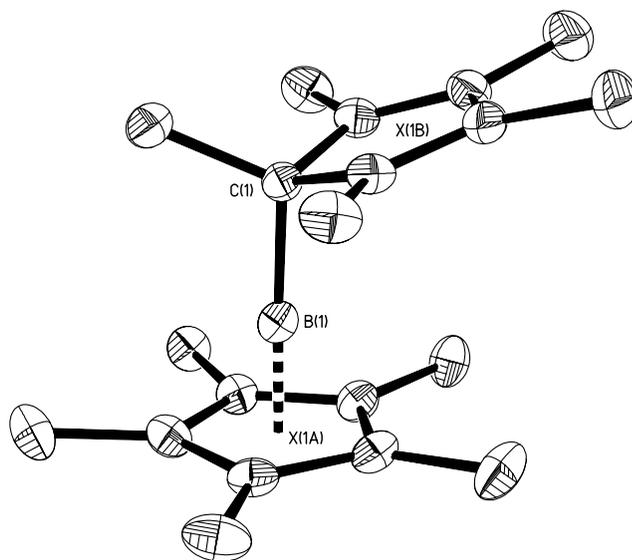
They assume that all electrons in lower shells shield outer electrons equally effectively.

## Shielding and Effective Nuclear Charge

Effective nuclear charge can be used to rationalize properties such as the size of atoms and ions.



$\text{Cp}^*_2\text{Be}$



$\text{Cp}^*_2\text{B}^+$

Be and  $\text{B}^+$  are *isoelectronic* ( $1s^2 2s^2$ ) but very different because of effective nuclear charge.

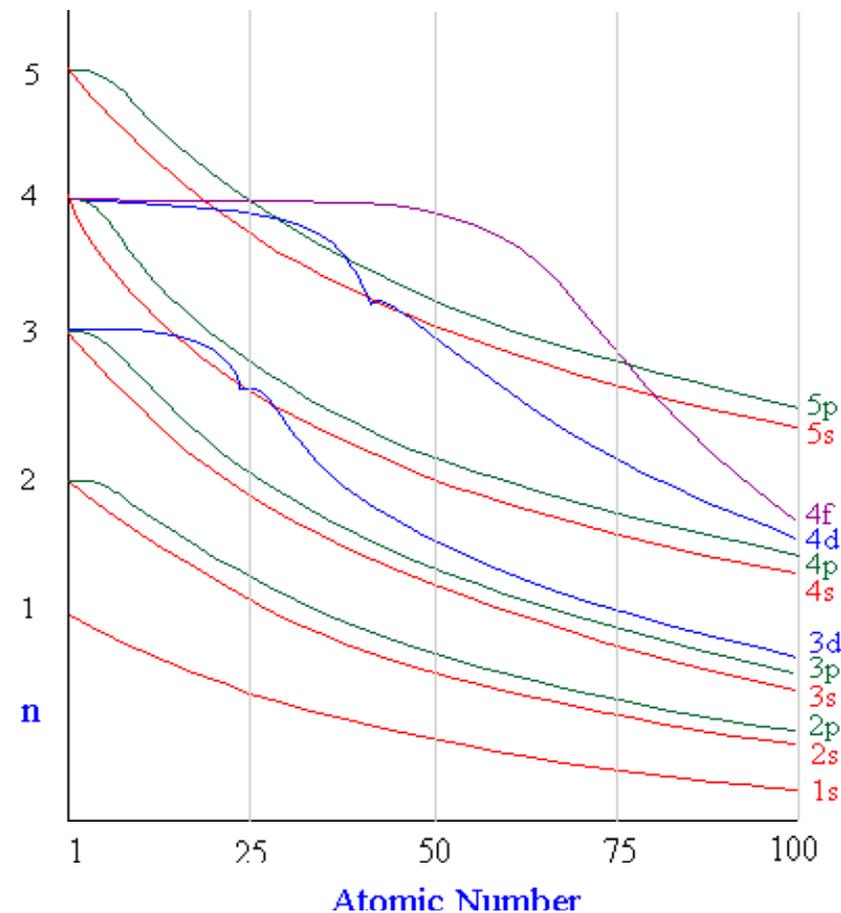
## Orbital energy levels and atomic number

For atoms other than hydrogen:  
Orbital energy depends on  $n$  and  $l$

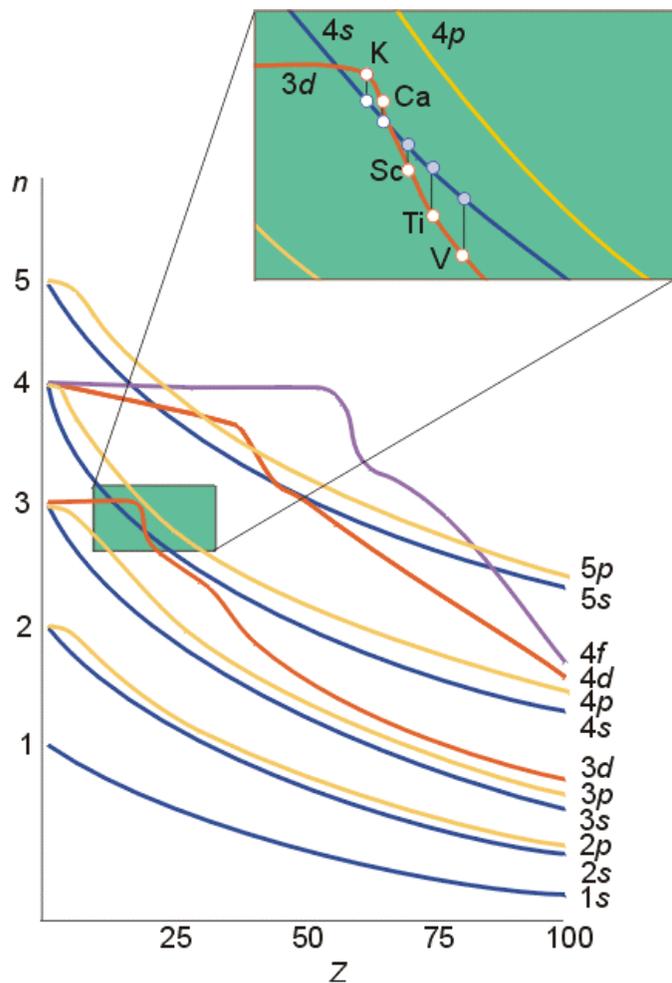
Ordering of orbital energies:  
 $ns < np < nd < nf$

Remember:

This ordering is due to the different penetrating ability of the different types of orbitals and the different effective nuclear charges felt by the electrons in those orbitals.



# Chem 59-651 An anomaly of the periodic table



The 4s orbitals are lower in energy than the 3d orbitals for K and Ca.

**This is only for the free atoms! In molecules 3d are lower in energy than 4s!**

This is assumed to be an accident of nature but it is consistent throughout the table.

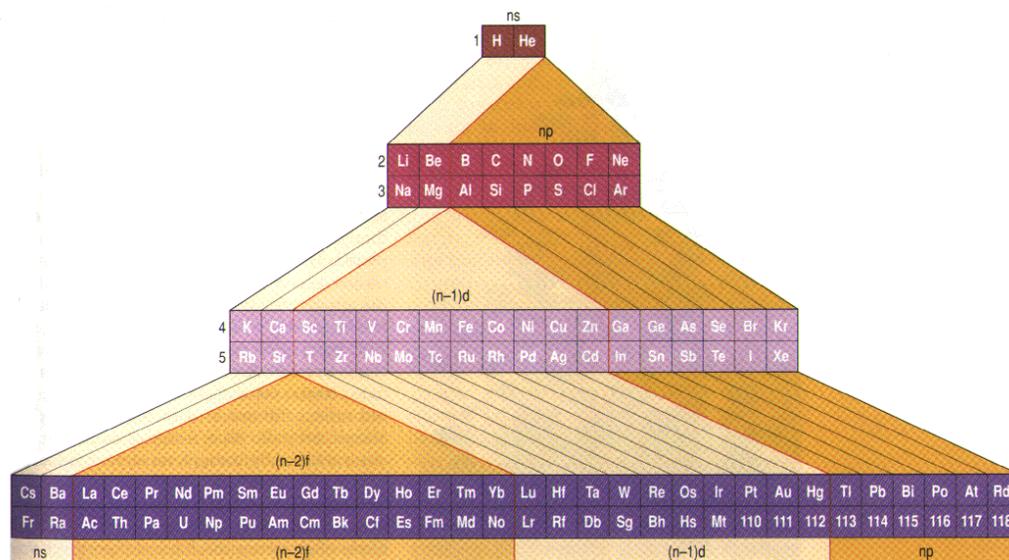


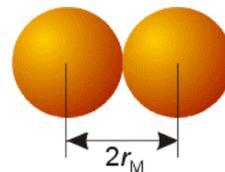
Figure 8. Alternative and updated periodic table, adapted from tables developed by Thomas Bayley, Jørgen Thomsen and Niels Bohr. These tables all depict the symmetrical nature of the periodic law regarding the increase in intervals before periodicity occurs in every other period with the exception of the first one. The lines denote chemical analogies.



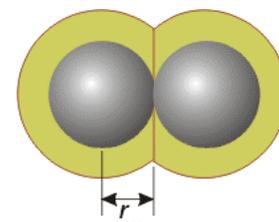
# The Size of Atoms and Ions

## Radii of neutral atoms

The atomic radius of an atom is defined as half the distance between the nuclei in a homonuclear bond.



1 Metallic radius

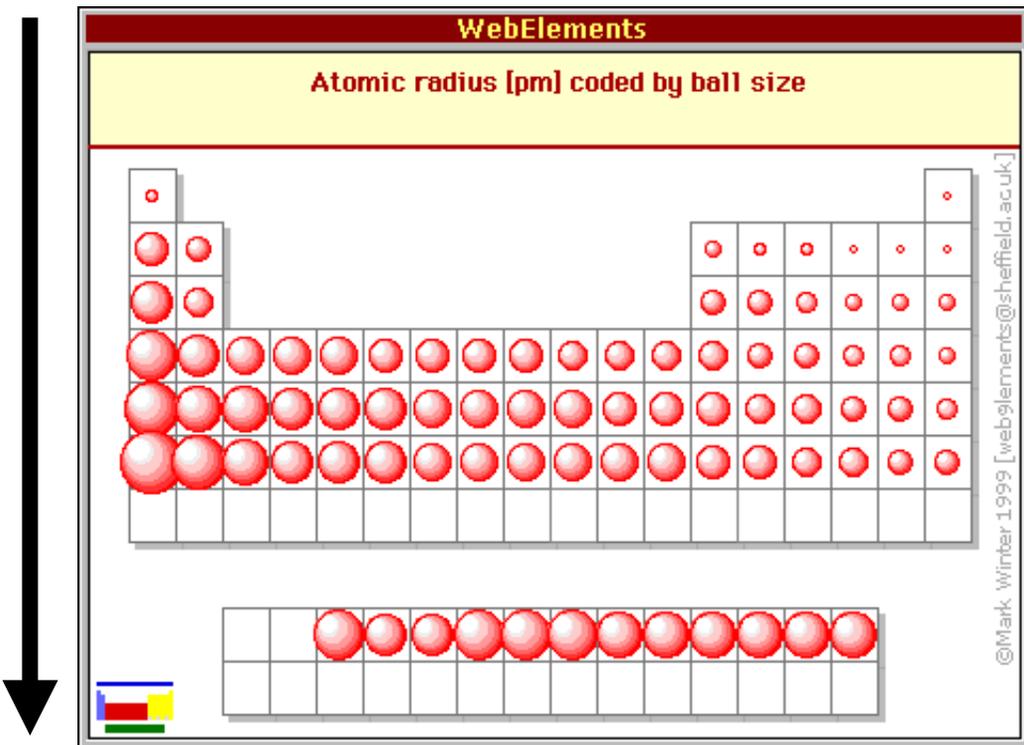


2 Covalent radius

r decreases

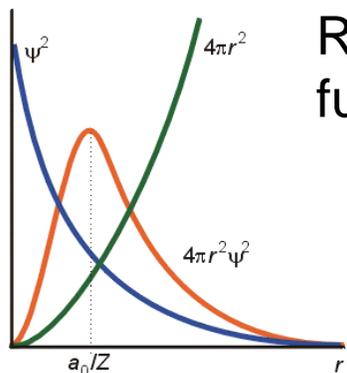


r increases



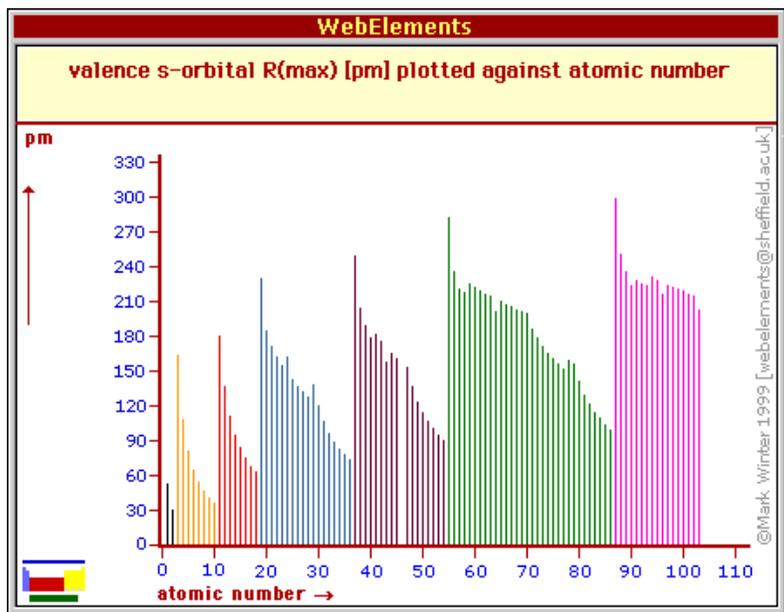
In general:

- radii decrease across a period because of increasing  $Z^*$ .
- radii increase down a group because of the increasing distance of the electrons from the nucleus.
- increasing distance from the nucleus outweighs effective nuclear charge for atomic radii down a group.



Remember that the maximum probability for an orbital moves further away from the nucleus with increasing  $n$ .

$0.1 \text{ nm} = 1 \text{ \AA} = 100 \text{ pm}$



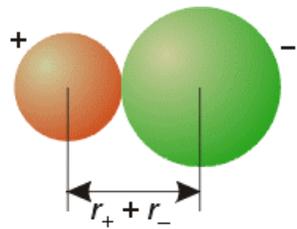
Atomic Radii (pm)

1A	2A	3A	4A	5A	6A	7A	8A
Li 152	Be 112	B 85	C 77	N 75	O 73	F 72	Ne 71
Na 186	Mg 160	Al 143	Si 118	P 110	S 103	Cl 100	Ar 98
K 227	Ca 197	Ga 135	Ge 122	As 120	Se 119	Br 114	Kr 112
Rb 248	Sr 215	In 167	Sn 140	Sb 140	Te 142	I 133	Xe 131
Cs 265	Ba 222	Tl 170	Pb 146	Bi 150	Po 168	At (140)	Rn (141)

The “d-block contraction” causes Ga to be about the same size as Al. This is caused by the introduction of the 3d elements which cause a vastly larger  $Z^*$  for Ga.

Bohr model for H;  $\text{radius}(n) = n^2 a_0$

# Radii of ions



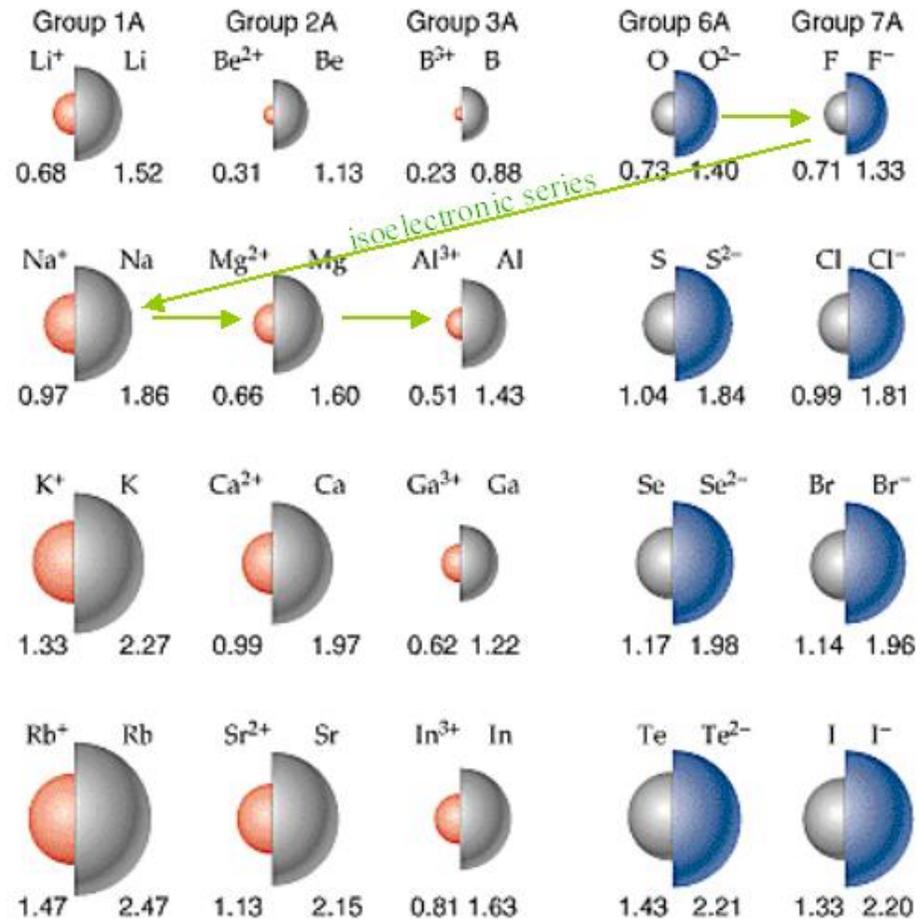
3 Ionic radius

This is a “self-consistent” scale based on  $O^{2-} = 1.40$  (or  $1.38$ ) Å.

Ionic radii depend on the magnitude of the charge of the ion and its environment.

Positively charged ions are smaller than their neutral analogues because of increased  $Z^*$ .

Negatively charged ions are larger than their neutral analogues because of decreased  $Z^*$ .



**Same periodic trends as atomic radii for a given charge**

## The utility of covalent and ionic radii

The radii tabulated in various books (e.g. Huheey, Keiter and Keiter) allow us to predict the bond length we would expect to see for a new bond.

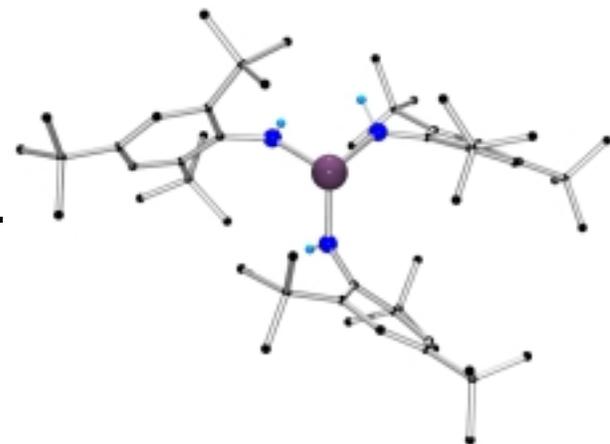
Example:

What is the expected bond length for a single Sb-N bond ?

For N,  $r_{\text{cov}} = 0.70 \text{ \AA}$  and for Sb,  $r_{\text{cov}} = 1.41 \text{ \AA}$

Using these values, an Sb-N bond should be  $2.11 \text{ \AA}$ .

The experimental distance is  $2.05 \text{ \AA}$ .



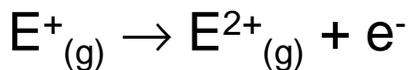
For covalent radii, the predictions will be the best for atoms that have similar *electronegativities*. If the electronegativities are very different, the predicted distance will be too long because it ignores the coulombic attraction.

# Ionization Enthalpy, $\Delta H^\circ_{ie}$ (ionization potential)

*The enthalpy change for ionization by loss of electron(s)*



$\Delta H^\circ_{ie}$  "First ionization potential"



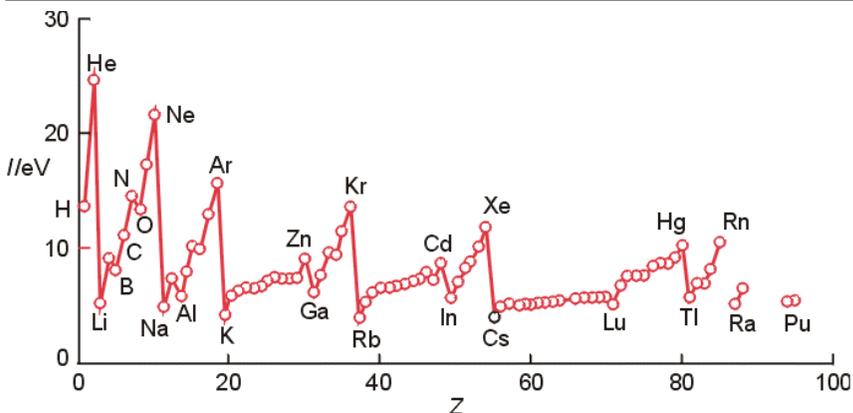
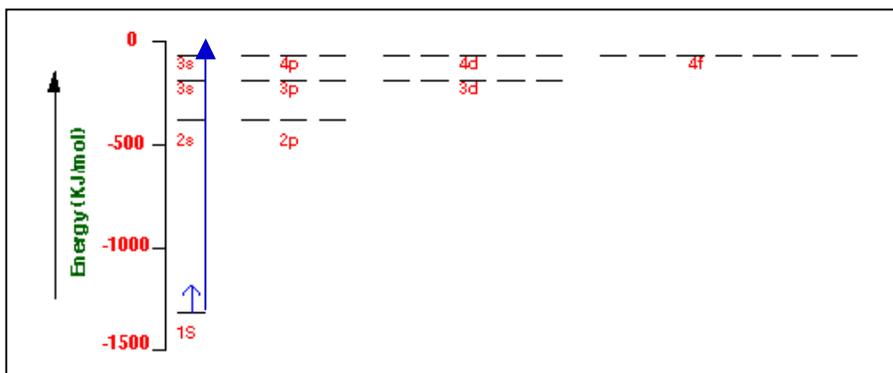
$\Delta H^\circ_{ie}$  "Second ionization potential" > first



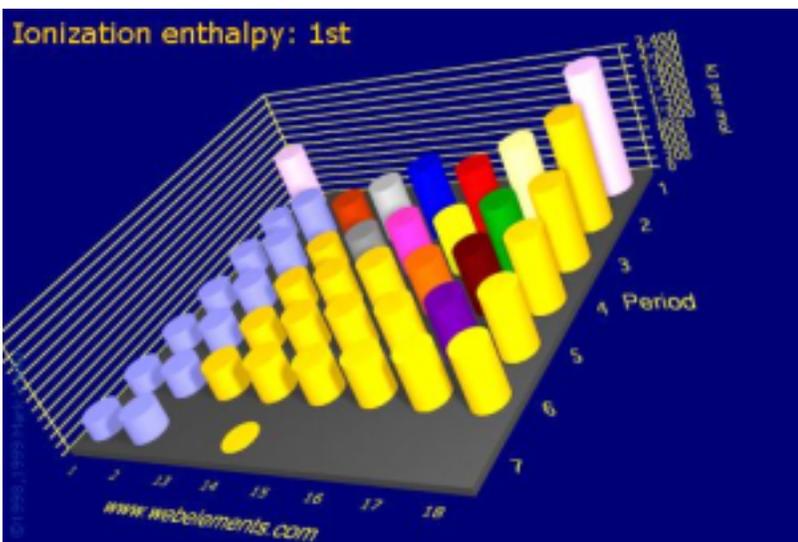
$\Delta H^\circ_{ie}$  "Third ionization potential" > second

...

...



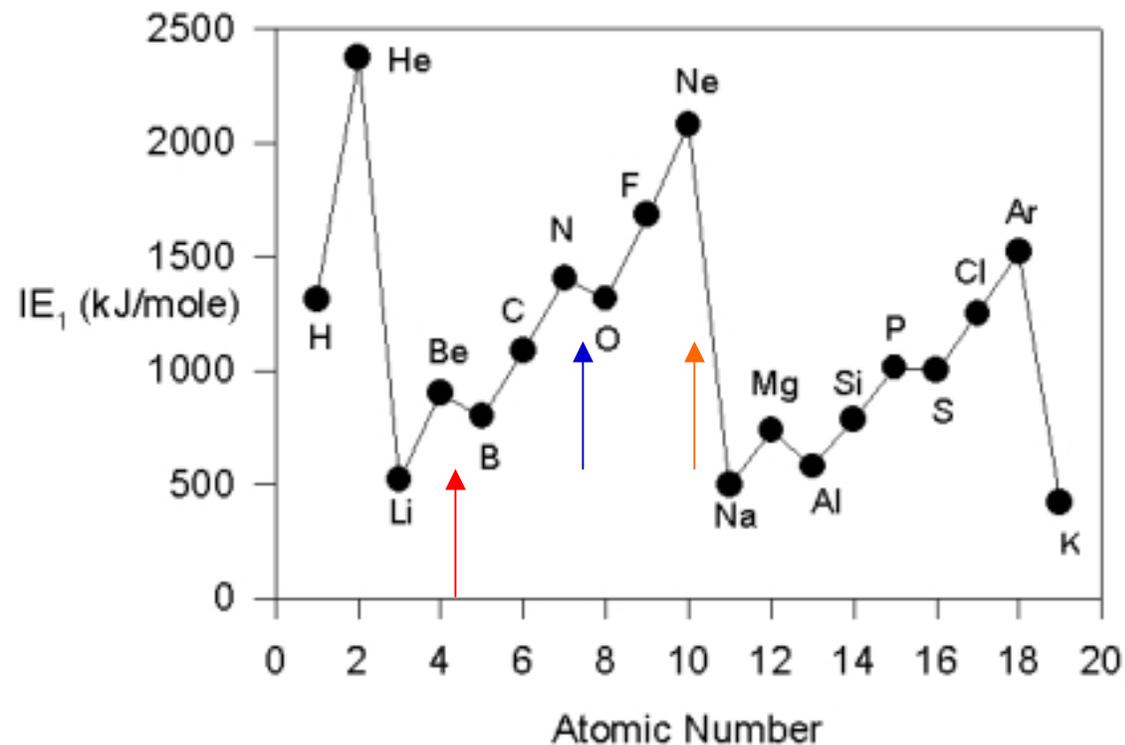
$\Delta H^\circ_{ie}$  increases



$\Delta H^\circ_{ie}$  decreases

# Features and anomalies in the trend of first ionization energies

## First Ionization Energy versus Atomic Number

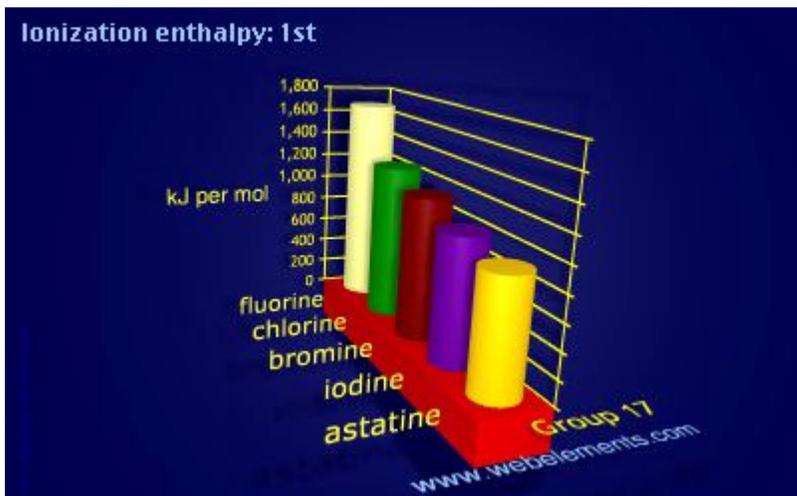


→ B:  $[\text{He}]2s^2 2p^1$  – p orbitals are more effectively shielded than s orbitals so smaller  $Z^*$  and lower  $\Delta H^\circ_{\text{ie}}$ . (also  $[\text{He}]2s^2$  is a full subshell)

→ O:  $[\text{He}]2s^2 2p^4$  – first pairing of electrons causes repulsion so loss of one electron is more favourable.

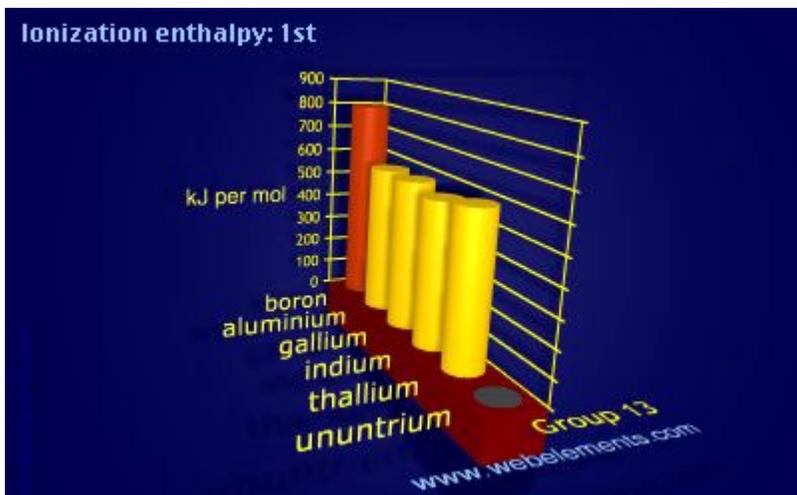
→ Na:  $[\text{Ne}]3s^1$  – expected from lower  $Z^*$  and greater distance of the electron from the nucleus at the start of a new shell.

## Features and anomalies in the trend of first ionization enthalpies



Group 17 is “normal” –  $\Delta H_{ie}^{\circ}$  decreases down the group as one would expect based on the increasing distance of the electrons from the nucleus.

Group 13 has unusual features –  $\Delta H_{ie}^{\circ}$  does not decrease down the group (and is higher for Tl than for Al, Ga or In).



$\Delta H_{ie}^{\circ}$  is greater than expected for Ga because of the greater  $Z^*$  caused by the presence of the 3d elements.

$\Delta H_{ie}^{\circ}$  is greater than expected for Tl because of the “greater  $Z^*$ ” caused by relativistic effects.

# Electronegativity, $\chi$

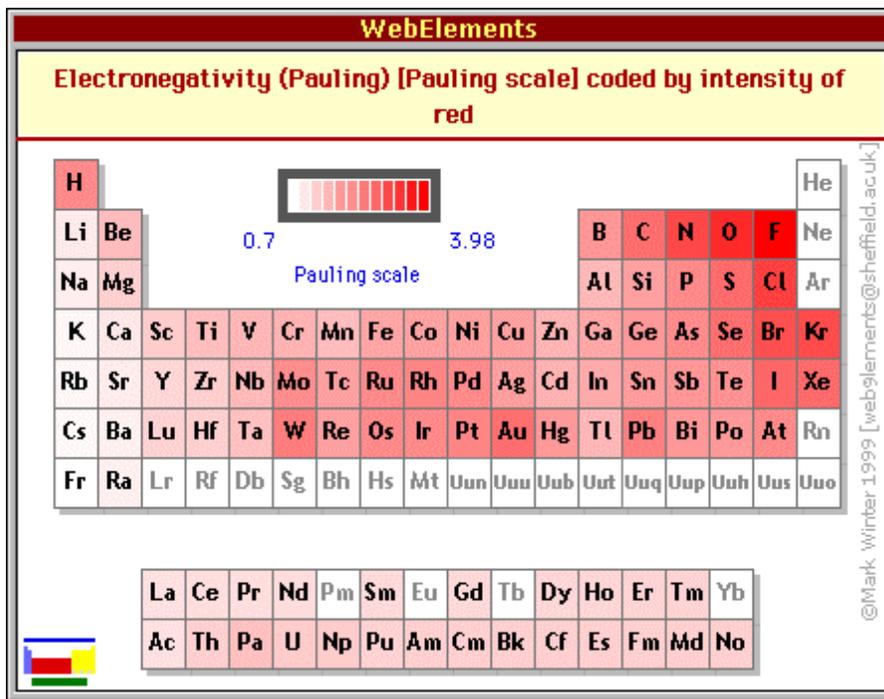
The ability of an atom *in a molecule* to attract electrons in a bond to itself.



Linus Pauling

$\chi$  decreases

$\chi$  increases

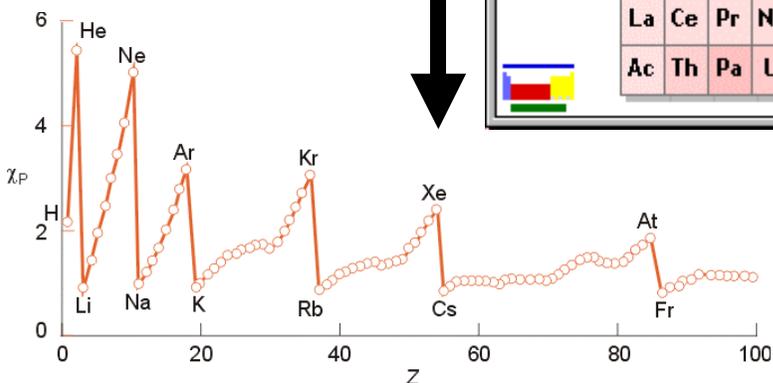


First Year rule

$\Delta X > 2$  : ionic

$2 > \Delta X > 0.5$  : polar

$\Delta X < 0.5$  : covalent



Traditional scale goes from 0 to 4 with  $\chi$  of F set to 4.

Pauling's definition:

Pauling reasoned that the dissociation energy of a purely covalent bond A-B should be the mean of the dissociation energies for the homonuclear bonds A-A and B-B. Any additional energy must be caused by electrostatic attraction between A and B (attributed to ionic character in a bond). The ionic character must be related to the difference in the electronegativities of A and B. He calculated this difference as follows:

$$D_{(A-B),\text{theory}} = \frac{1}{2} (D_{(A-A)} + D_{(B-B)})$$

$$\Delta'_{(A-B)} = D_{(A-B),\text{experimental}} - D_{(A-B),\text{theory}}$$

$$X_A - X_B = 0.102 (\Delta'_{(A-B)})^{1/2}$$

$\Delta'_{(A-B)}$  is the *ionic resonance energy* in kJ/mol ( $\Delta_{(A-B)}$  is in eV)

A-B  $\leftrightarrow$  A<sup>+</sup> B<sup>-</sup>  
0.102 is a conversion from kJ/mol to eV

Mulliken's definition:

Mulliken figured that the electronegativity of an element must be related to the energies of gaining and losing electrons. Specifically an atom that binds its electrons strongly (large  $\Delta H_{ie}^\circ$ ) and gains other electrons readily (very positive EA or very negative  $\Delta H_{ea}^\circ$ ) should do the same in molecules. Thus Mulliken calculated the electronegativity of an atom as the mean of the ionization potential and the electron affinity.



Robert  
Mulliken

For A-B, the electronegativity difference between  $A^+B^-$  and  $A^-B^+$  is given by:

$$X_A - X_B = \frac{1}{2} ([IP_A + EA_A] - [IP_B + EA_B])$$

$$X_A = \frac{1}{2} ([IP_A + EA_A]) \text{ (these are then scaled to fit the } \approx 0\text{-}4 \text{ scale)}$$

This method makes a lot of sense, but is not used because values of  $\Delta H_{ea}^\circ$  have not been accurately determined for many elements.

The Allred-Rochow definition:

The assumption is that the force that will draw an electron toward an atom is proportional to the effective nuclear charge of that atom and related to the distance of the electron from the nucleus.

$$\text{Force} = \frac{Z^* e^2}{4\pi\epsilon_0 r^2}$$

$Z^*$  = effective nuclear charge

$e$  = charge of electron

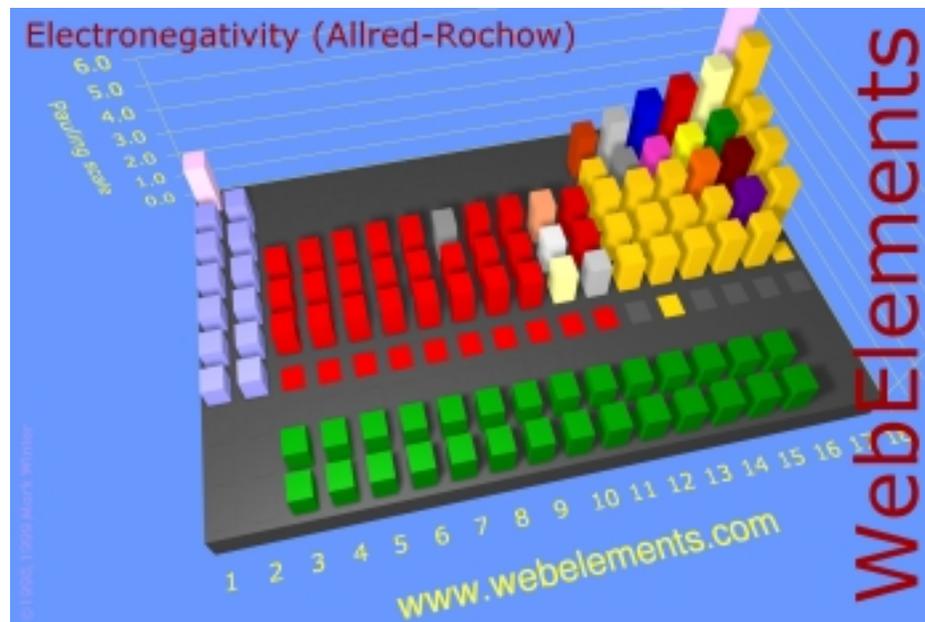
$\epsilon_0$  = permittivity of a vacuum

$r$  = atomic radius

The equation:

$$X = 0.359 (Z^*/r^2) + 0.744$$

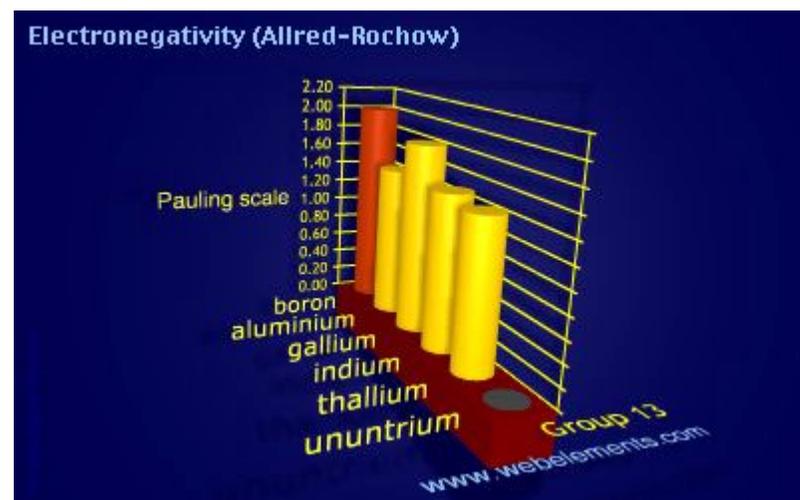
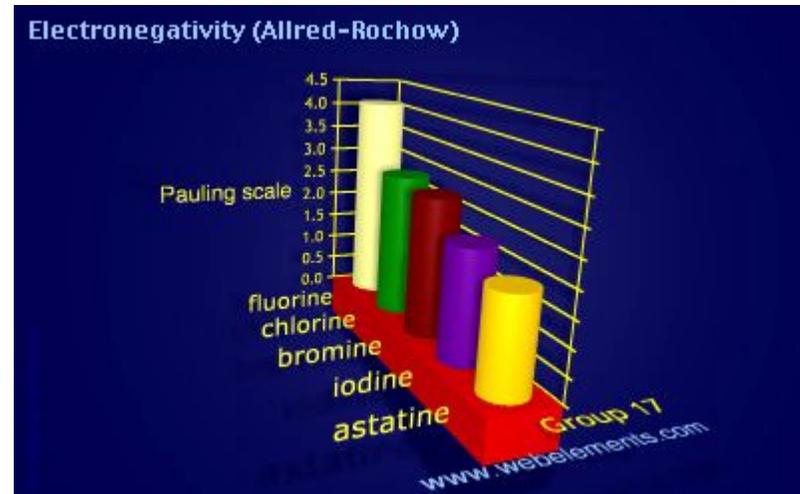
puts the calculated values on the Pauling scale. This definition is useful because it can be applied to many more atoms and is one of the most used scales.



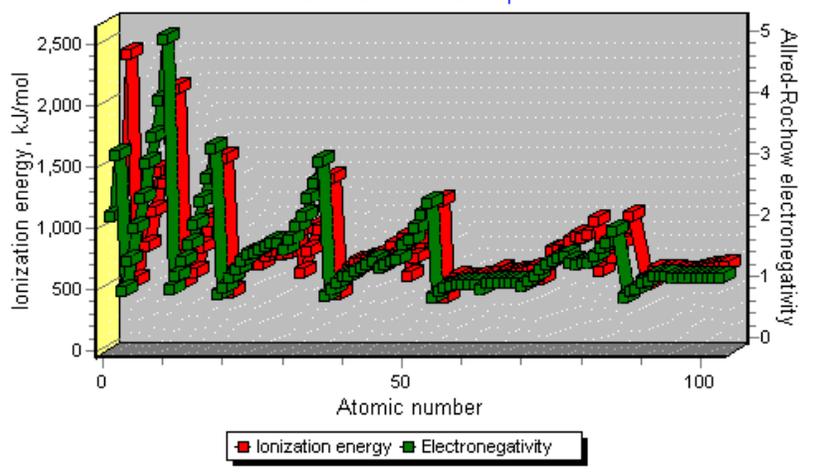
# Electronegativity, $\chi$

There are several other definitions based on different assumptions and methods, such as quantum mechanical calculations (Boyd) or spectroscopic measurements (Allen), but the values for elements usually end up around the same.

Trends in electronegativities are similar to those found for ionization enthalpies.



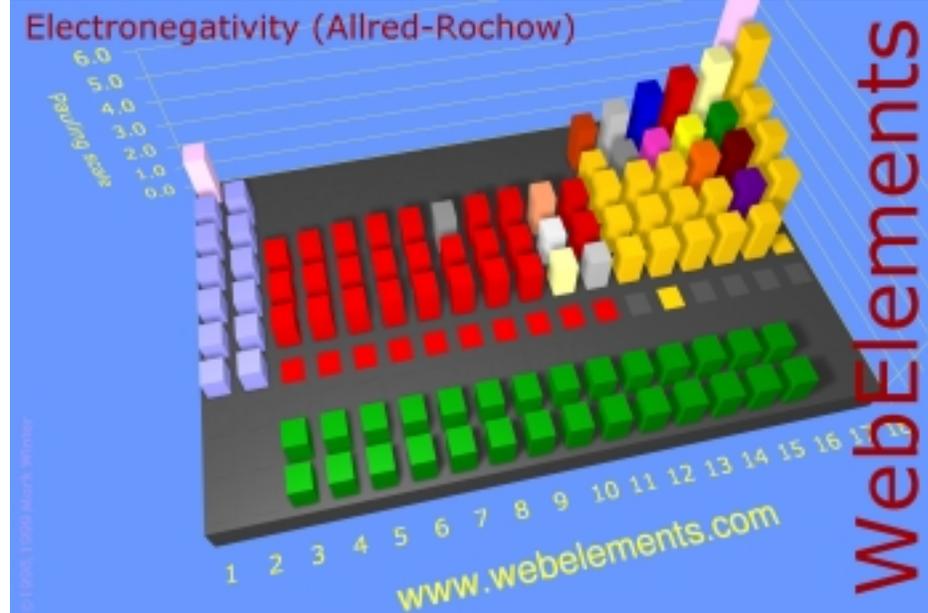
Ionization Energy and Electronegativity  
Periodic Variation Compared



# Chem 59-651

The trends in electronegativities and ionization enthalpies explain many features of chemistry such as the “diagonal relationship” (X) and the position of the metallic and non-metallic elements ( $\Delta H^\circ_{ie}$ ).

Electronegativity also lets us predict the polarity of bonds and chemical reactivity.



I A																	III A	IV A	V A	VIA	VII A	VIII A	
1 H 1.01																	5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18	
3 Li 6.94	4 Be 9.01																	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
11 Na 22.99	12 Mg 24.30	III B	IV B	V B	VI B	VII B	VIII B	VIII B	VIII B	I B	II B	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3						
19 K 39.1	20 Ca 40.08	39 Y 44.96	40 Zr 47.88	41 Nb 50.94	42 Mo 52.00	43 Tc 54.94	44 Ru 55.85	45 Rh 58.93	46 Pd 58.69	47 Ag 63.55	48 Cd 65.38	69 Tl 69.72	70 Pb 207.2	71 Bi 209.0	72 Po (209)	73 At (210)	74 Rn (222)						
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3						
55 Cs 132.9	56 Ba 137.3	57 La 138.9	72 Hf 178.5	73 Ta 181.0	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (209)	85 At (210)	86 Rn (222)						
87 Fr (223)	88 Ra 226.0	89 Ac 227.0	104 Rf (261)	105 Db (262)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)															
* Lanthanide Series			58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (145)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.2	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0							
** Actinide Series			90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np 237.0	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (260)							

