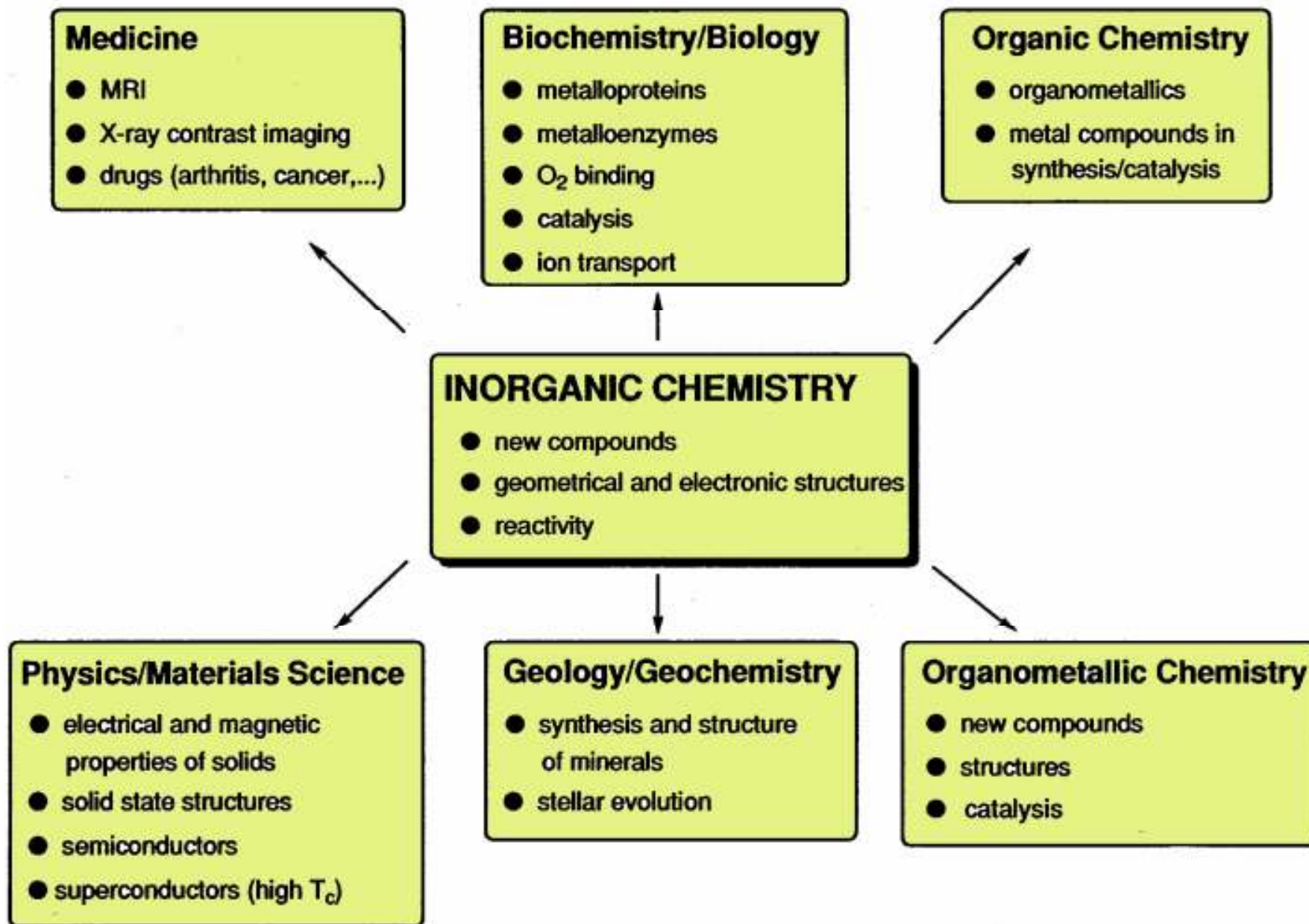


CH105(I)

Part I: Inorganic Chemistry

What is the scope of inorganic chemistry?



Modern Inorganic Chemistry

Inorganic Materials

Inorganic Biological Chemistry

Inorganic Drugs

Organometallic Chemistry

Catalysis

[These Influence life & life style]

These are built based on simple principles of chemistry in general and inorganic chemistry in particular

Modern Inorganic Chemistry: Vast & Vital

Electronic Configuration (s,p,d,f blocks)

**Atoms (in elemental or metallic)
Ions (in compounds or complexes)
Variable positive or negative charges or
oxidation states**

**Size & charge, ionization potential,
electron affinity, oxidation state, redox
potential, polarisability, HSAB,
coordination characteristics, CFSE,...**

Modern Inorganic Chemistry: Vast & Vital

Coordination characteristics

- Coordination number & geometry (CFSE),
- Overall charge & size (Redox)
- Ligating environment (HSAB)
- Electron density & orbital orientations leading to bonding (covalent, non-covalent & ionic)
- Ligand exchange & reactivity
- Physical states (solid, liquid & gas)
- Atoms (metals, nonmetals, metalloids) to molecules to supramolecules to materials**
- Properties at nano (nanosci. & Tech.) & bulk**

Interpretations
Explanations
Substantiations

Are not necessarily reflected in the slides,
but are reflected in the lecture.

Please Do Attend All the Classes.

Please DONOT miss any class

Course Coverage

- 1. Some properties of elements & compounds**
- 2. Basic principles of extraction of metals from ores & purification**
- 3. Transition metal chemistry & some applications**
- 4. Magnetism & some applications**
- 5. Biological inorganic chemistry & Inorganic compounds in medicine**

Recommended Text Books

(1) Concise Inorganic Chemistry

- J.D. Lee

(2) Shriver & Atkins' Inorganic Chemistry

**P. Atkins, T. Overton, J. Rourke,
M. Weller, F. Armstrong**

Topic I

Some properties of elements & compounds

Periodic Table of the Elements

Periodic Table of the Elements																		18	
1 1IA 1A																	VIIIA 8A		
1 H Hydrogen 1.0079	2 He Helium 4.00260																		
3 Li Lithium 6.941	4 Be Beryllium 9.01218											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.00674	8 O Oxygen 15.9994	9 F Fluorine 18.998403	10 Ne Neon 20.1797		
11 Na Sodium 22.989769	12 Mg Magnesium 24.305	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8 VIII 8	9 VIII 8	10 VIII 8	11 IB 1B	12 IIB 2B	13 Al Aluminum 26.981539	14 Si Silicon 28.0855	15 P Phosphorus 30.973762	16 S Sulfur 32.06	17 Cl Chlorine 35.4527	18 Ar Argon 39.948		
19 K Potassium 39.0983	20 Ca Calcium 40.078	21 Sc Scandium 44.95591	22 Ti Titanium 47.88	23 V Vanadium 50.9415	24 Cr Chromium 51.9961	25 Mn Manganese 54.938	26 Fe Iron 55.847	27 Co Cobalt 58.9332	28 Ni Nickel 58.6934	29 Cu Copper 63.546	30 Zn Zinc 65.39	31 Ga Gallium 69.723	32 Ge Germanium 72.64	33 As Arsenic 74.92159	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.80		
37 Rb Rubidium 85.4678	38 Sr Strontium 87.62	39 Y Yttrium 88.90585	40 Zr Zirconium 91.224	41 Nb Niobium 92.90638	42 Mo Molybdenum 95.94	43 Tc Technetium 98.9062	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.9695	46 Pd Palladium 106.42	47 Ag Silver 107.8682	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.71	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.90447	54 Xe Xenon 131.29		
55 Cs Cesium 132.90543	56 Ba Barium 137.327	57-71	72 Hf Hafnium 178.49	73 Ta Tantalum 180.9479	74 W Tungsten 183.85	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.22	78 Pt Platinum 195.08	79 Au Gold 196.9665	80 Hg Mercury 200.59	81 Tl Thallium 204.3833	82 Pb Lead 207.2	83 Bi Bismuth 208.98037	84 Po Polonium [209]	85 At Astatine 209	86 Rn Radon 222		
87 Fr Francium 223	88 Ra Radium 226	89-103	104 Rf Rutherfordium [261]	105 Db Dubnium [262]	106 Sg Seaborgium [266]	107 Bh Bohrium [264]	108 Hs Hassium [269]	109 Mt Meitnerium [268]	110 Ds Darmstadtium [289]	111 Rg Roentgenium [272]	112 Cn Copernicium [285]	113 Uut Ununtrium unknown	114 Fl Flerovium [289]	115 Uup Ununpentium unknown	116 Lv Livermorium [293]	117 Uus Ununseptium unknown	118 Uuo Ununoctium unknown		
Lanthanide Series		57 La Lanthanum 138.9055	58 Ce Cerium 140.115	59 Pr Praseodymium 140.90765	60 Nd Neodymium 144.24	61 Pm Promethium 144.9127	62 Sm Samarium 150.36	63 Eu Europium 151.9655	64 Gd Gadolinium 157.25	65 Tb Terbium 158.92534	66 Dy Dysprosium 162.50	67 Ho Holmium 164.93032	68 Er Erbium 167.26	69 Tm Thulium 168.93421	70 Yb Ytterbium 173.04	71 Lu Lutetium 174.967			
Actinide Series		89 Ac Actinium 227	90 Th Thorium 232.0377	91 Pa Protactinium 231.03688	92 U Uranium 238.02891	93 Np Neptunium 237.04817	94 Pu Plutonium 244.0642	95 Am Americium 243.0613	96 Cm Curium 247.0754	97 Bk Berkelium 247.0754	98 Cf Californium 251.0765	99 Es Einsteinium [254]	100 Fm Fermium 257.10528	101 Md Mendelevium 258.1	102 No Nobelium 259	103 Lr Lawrencium [262]			
Alkali Metal		Alkaline Earth		Transition Metal		Basic Metal		Semimetal		Nonmetal		Halogen		Noble Gas		Lanthanide		Actinide	
©2014 Todd Helmenstine chemistry.about.com scienceready.com																			

Periodic Table of the Elements

Periodic Table of the Elements

1 H Hydrogen 1.008																	2 He Helium 4.003
3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.811	6 C Carbon 12.011	7 N Nitrogen 14.007	8 O Oxygen 15.999	9 F Fluorine 18.998	10 Ne Neon 20.180
11 Na Sodium 22.990	12 Mg Magnesium 24.305											13 Al Aluminum 26.982	14 Si Silicon 28.086	15 P Phosphorus 30.974	16 S Sulfur 32.06	17 Cl Chlorine 35.453	18 Ar Argon 39.948
19 K Potassium 39.098	20 Ca Calcium 40.078	21 Sc Scandium 44.956	22 Ti Titanium 47.867	23 V Vanadium 50.942	24 Cr Chromium 51.996	25 Mn Manganese 54.938	26 Fe Iron 55.845	27 Co Cobalt 58.933	28 Ni Nickel 58.693	29 Cu Copper 63.546	30 Zn Zinc 65.38	31 Ga Gallium 69.723	32 Ge Germanium 72.631	33 As Arsenic 74.922	34 Se Selenium 78.96	35 Br Bromine 79.904	36 Kr Krypton 83.798
37 Rb Rubidium 85.468	38 Sr Strontium 87.62	39 Y Yttrium 88.906	40 Zr Zirconium 91.224	41 Nb Niobium 92.906	42 Mo Molybdenum 95.95	43 Tc Technetium 98.907	44 Ru Ruthenium 101.07	45 Rh Rhodium 102.906	46 Pd Palladium 106.42	47 Ag Silver 107.868	48 Cd Cadmium 112.411	49 In Indium 114.818	50 Sn Tin 118.710	51 Sb Antimony 121.760	52 Te Tellurium 127.6	53 I Iodine 126.905	54 Xe Xenon 131.29
55 Cs Cesium 132.905	56 Ba Barium 137.327	57-71	72 Hf Hafnium 178.49	73 Ta Tantalum 180.948	74 W Tungsten 183.84	75 Re Rhenium 186.207	76 Os Osmium 190.23	77 Ir Iridium 192.225	78 Pt Platinum 195.085	79 Au Gold 196.967	80 Hg Mercury 200.592	81 Tl Thallium 204.384	82 Pb Lead 207.2	83 Bi Bismuth 208.980	84 Po Polonium 209	85 At Astatine 209	86 Rn Radon 222
87 Fr Francium 223	88 Ra Radium 226	89-103	104 Rf Rutherfordium 261	105 Db Dubnium 262	106 Sg Seaborgium 266	107 Bh Bohrium 269	108 Hs Hassium 269	109 Mt Moscovium 288	110 Ds Darmstadtium 289	111 Rg Roentgenium 289	112 Cn Copernicium 277	113 Nh Nihonium 286	114 Fl Flerovium 289	115 Mc Moscovium 289	116 Lv Livermorium 293	117 Ts Tennessine unknown	118 Og Oganesson unknown
57 La Lanthanum 138.905	58 Ce Cerium 140.116	59 Pr Praseodymium 140.908	60 Nd Neodymium 144.242	61 Pm Promethium 144.913	62 Sm Samarium 150.36	63 Eu Europium 151.964	64 Gd Gadolinium 157.25	65 Tb Terbium 158.925	66 Dy Dysprosium 162.500	67 Ho Holmium 164.930	68 Er Erbium 167.259	69 Tm Thulium 168.930	70 Yb Ytterbium 173.055	71 Lu Lutetium 174.967			
89 Ac Actinium 227	90 Th Thorium 232	91 Pa Protactinium 231	92 U Uranium 238	93 Np Neptunium 237	94 Pu Plutonium 244	95 Am Americium 243	96 Cm Curium 247	97 Bk Berkelium 247	98 Cf Californium 251	99 Es Einsteinium 252	100 Fm Fermium 257	101 Md Mendelevium 258	102 No Nobelium 259	103 Lr Lawrencium 262			
Alkali Metal	Alkaline Earth	Transition Metal	Basic Metal	Semimetal	Nonmetal	Halogen	Noble Gas	Lanthanide	Actinide								

**Few concepts which are
important in explaining the
trends in the properties of atoms
(you have already studied)**

Effective nuclear charge

Penetration of orbitals

Nuclear influence on electrons

Shielding/Screening

The energy order of orbitals for a given quantum number depends on shielding effects (σ), effective nuclear charge (Z^* or Z_{eff}) & penetration of orbitals

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

How to estimate the Z^* ? (tutorial)

{If the electron resides in s or p orbital}

1. Electrons in principal shell higher than the e^- in question: contribute 0 to σ
2. Each e^- in the same principal shell: contribute 0.35 to σ
3. Electrons in $(n-1)$ shell: each contribute 0.85 to σ
4. Electrons in deeper shell: each contribute 1.00 to σ

P.S.: There may be other ways of calculating these as given in the literature. Please stick to this procedure as far as this course is concerned.

How to estimate the Z^* ? (tutorial)

Example: Calculate the Z^* for the 2p electron

Fluorine ($Z = 9$) $1s^2 2s^2 2p^5$

Screening constant for one of the outer electron (2p):

$$6 \text{ (six) (two } 2s \text{ e- and four } 2p \text{ e-)} = 6 \times 0.35 = 2.10$$

$$2 \text{ (two) } 1s \text{ e-} = 2 \times 0.85 = 1.70$$

$$\sigma = 1.70 + 2.10 = 3.80 \quad \& \quad Z^* = 9 - 3.80 = 5.20$$

P.S.: There may be other ways of calculating these as given in the literature. Please stick to this procedure as far as this course is concerned.

How to estimate the Z^* ? (tutorial)

{If the e^- resides in a 'd' or 'f' orbital}

- 1. All e^- 's in higher principal shell contribute 0**
- 2. Each e^- in the same shell contributes 0.35**
- 3. All inner shells in (n-1) and lower contributes 1.00**

P.S.: There may be other ways of calculating these as given in the literature. Please stick to this procedure as far as this course is concerned.

Trends in Z^* (Z_{eff}): Down a group

	n	Z	Z^*
H	1	1	1.0
Li	2	3	1.3
Na	3	11	2.5
K	4	19	2.2
Rb	5	37	2.2
Cs	6	55	2.2

Effective nuclear charge Z^* increases very slowly down a group for the “valence electron”.
Example of Valence configuration as ‘ ns^1 ’

Consequence of this?

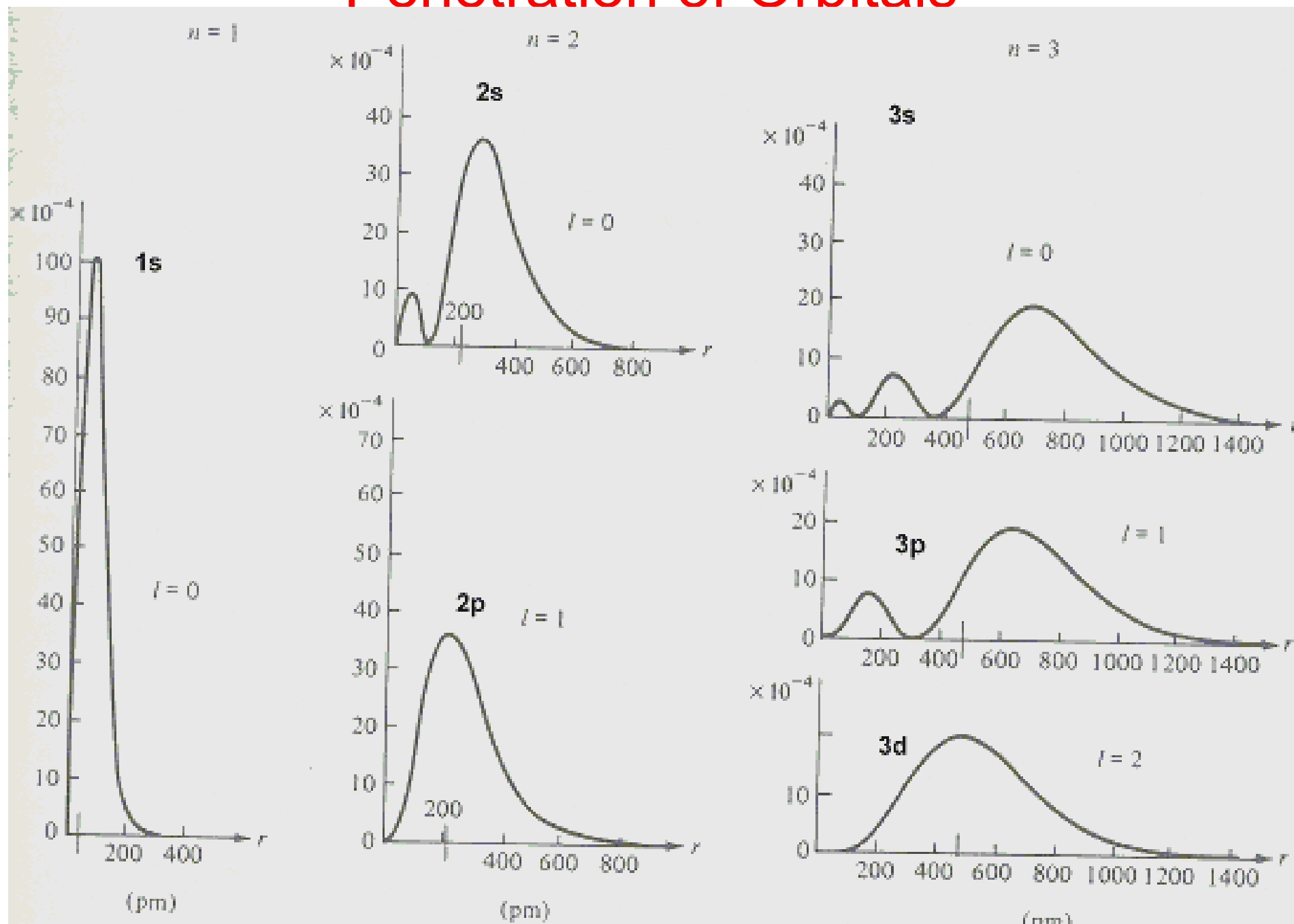
Trends in Z^* (Z_{eff}): Across a period

Effective nuclear charge Z^* increases rapidly along a period. For example, take period two

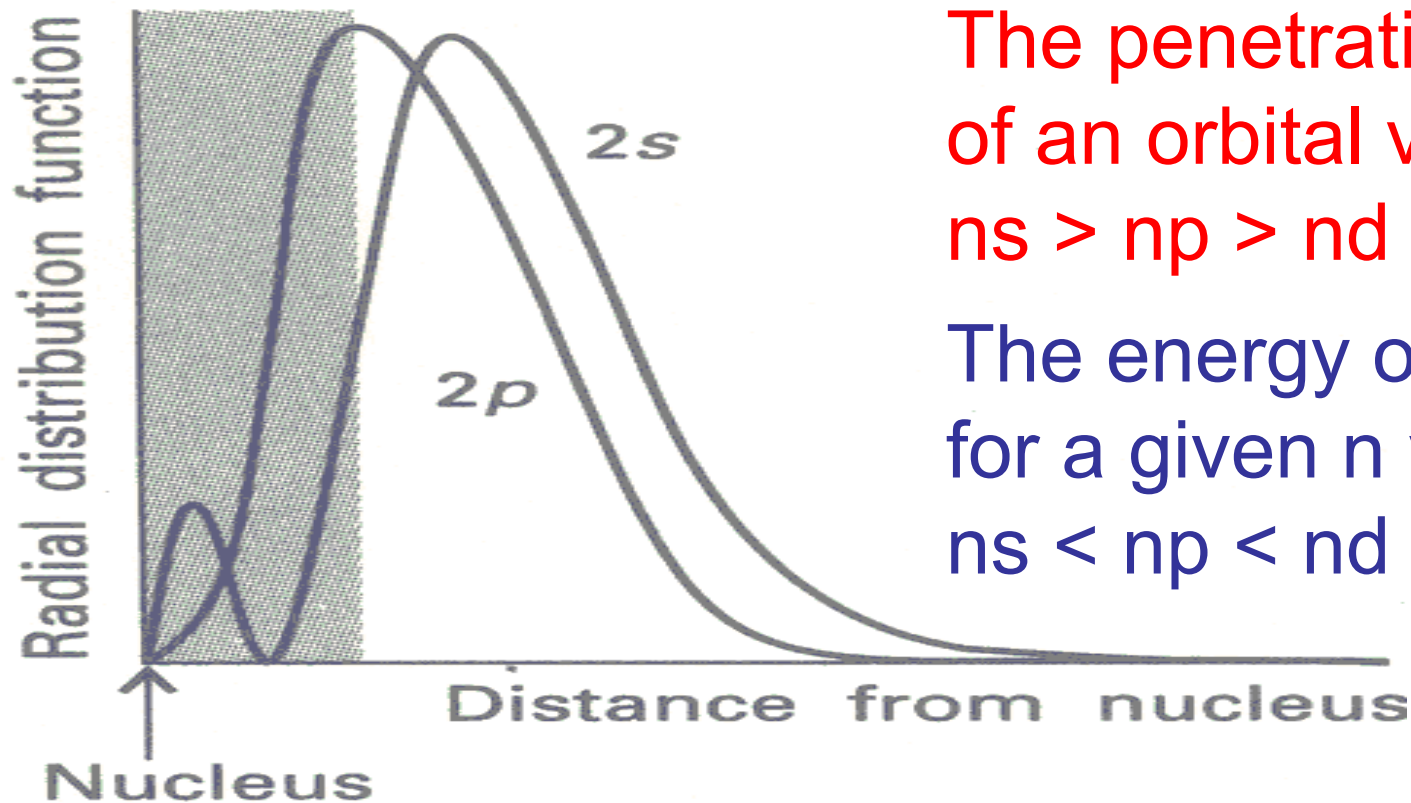
Li	Be	B	C	N	O	F	Ne
3	4	5	6	7	8	9	10
1.3	1.95	2.4	3.1	3.8	4.5	5.1	5.8
$2s^1$	$2s^2$	$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$

Consequence of this?

Penetration of Orbitals



Penetration of Orbitals



The penetration potential of an orbital varies as:

$$ns > np > nd > nf$$

The energy of the orbitals for a given n varies as:

$$ns < np < nd < nf$$

The penetration of 2s electron through the inner core is greater than that of a 2p electron because the latter vanishes at the nucleus. Therefore, the 2s electrons are less shielded than the 2p electrons.

Consequence: Influence of nucleus on e's

Two electrons present in the same d-orbital repel each other more strongly than do two electron in the same s-orbital.

The electrons present in f are much less influenced by the nucleus as compared to d, those present in d much less influenced as compared to p, than s, etc.

It is essential to consider all contributions to the energy of a configuration, and just not one-electron orbital energies.

On going from

**Atom → Molecule → Compound
or Material**

The following properties are important in order to address the compounds formed

- Atomic size (radius)
- Ionic size (radius)
- Ionization energy
- Electron affinity
- Electronegativity
- Hard soft acid base (HSAB)
- Polarizability
- Oxidation states
- Redox properties

Size (Radius)

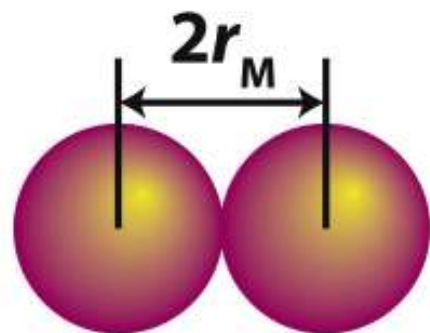


Figure 1.10a
Brown & LeMay: General Chemistry, Seventh Edition
© 2004 H. C. Brown, P. M. LeMay, Jr., Chemistry, 7th Edition, H. C. Brown and P. M. LeMay, Jr.

The METALLIC RADIUS is half of the experimentally determined distance between the nuclei of nearest neighbors in the solid

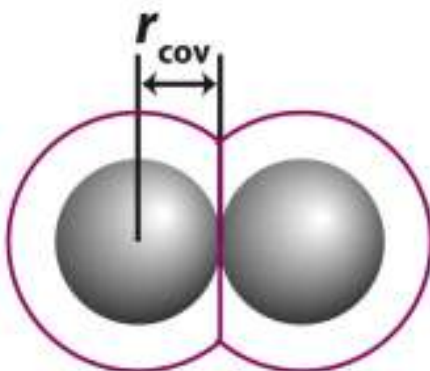


Figure 1.10b
Brown & LeMay: General Chemistry, Seventh Edition
© 2004 H. C. Brown, P. M. LeMay, Jr., Chemistry, 7th Edition, H. C. Brown and P. M. LeMay, Jr.

The COVALENT RADIUS of a non-metallic element is half of the experimentally determined distance between the nuclei of nearest neighbors in the solid

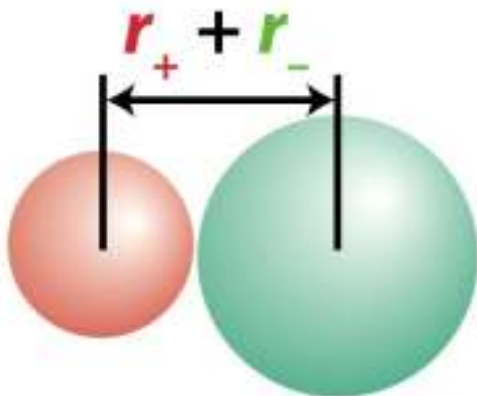


Figure 1.10c
Brown & LeMay: General Chemistry, Seventh Edition
© 2004 H. C. Brown, P. M. LeMay, Jr., Chemistry, 7th Edition, H. C. Brown and P. M. LeMay, Jr.

The IONIC RADIUS of an element is related to the distance between the nuclei of neighboring cations and anions
Ionic radius of O^{2-} is 1.40 \AA ;
What is the ionic radius for Mg^{2+} ?
Measure the Mg-O distance in MgO and subtract 1.40 \AA

Trend in Size (Radius)

In a period, left to right: atomic radius decreases

1. n (number of shells) remain constant.
2. Z increases (by one unit)
3. Z^* increases (by 0.65 unit)
4. Electrons are pulled close to the nucleus by the increased Z^*

In a group, top to bottom: atomic radius increases

1. n increases
2. Z increases
3. No dramatic increase in Z^* - almost remains constant

Trends in atomic radii

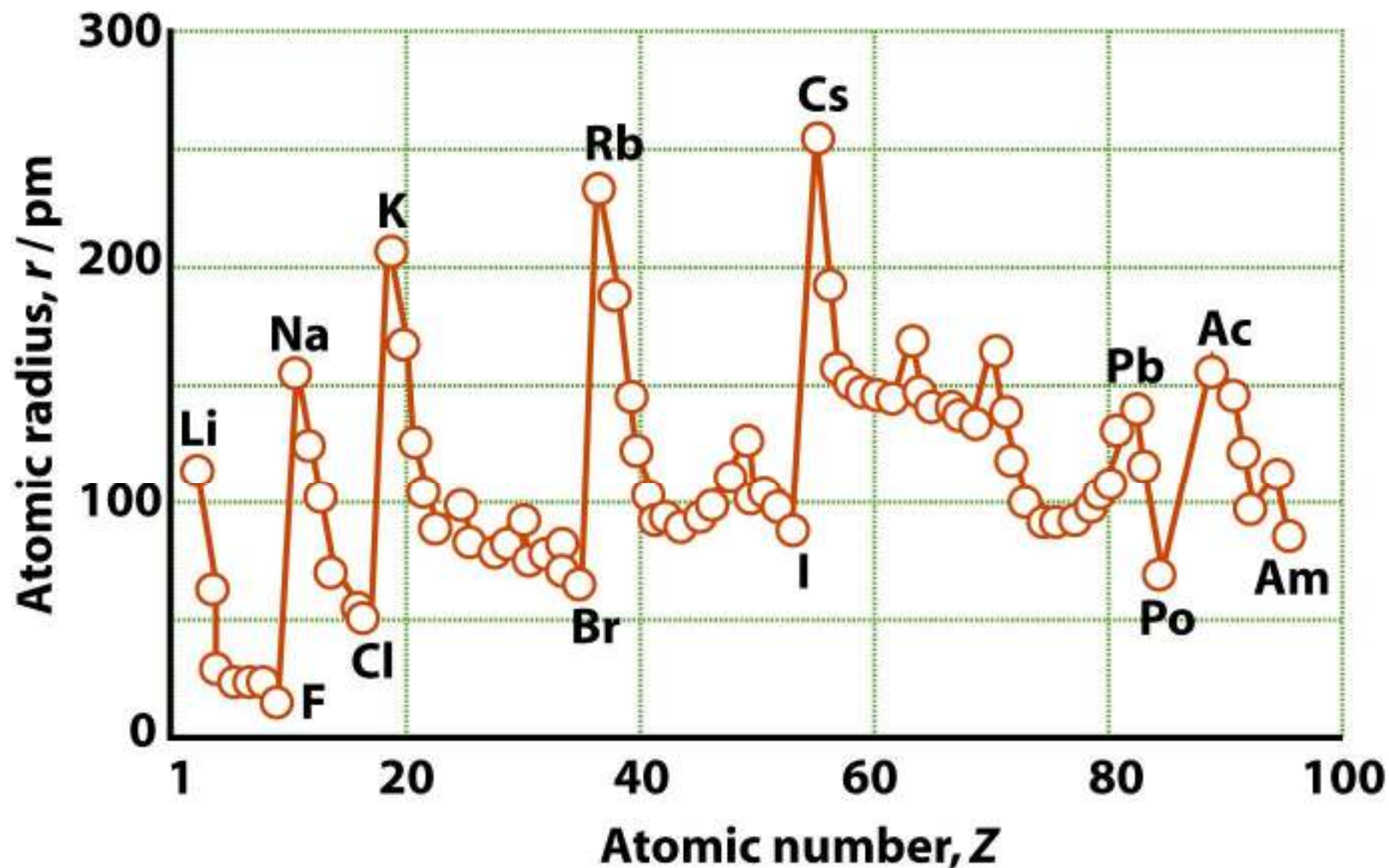


Figure 1-24

Shriver & Atkins Inorganic Chemistry, Fourth Edition

© 2006 by D. F. Shriver, P. W. Atkins, T. L. Overton, J. P. Rourke, M. T. Weller, and F. A. Armstrong

Table of atomic radii

Table 1.4 Atomic radii, r/pm^*

Li	Be											B	C	N	O	F
157	112											88	77	74	66	64
Na	Mg											Al	Si	P	S	Cl
191	160											143	118	110	104	99
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br
235	197	164	147	135	129	137	126	125	125	128	137	153	122	121	117	114
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I
250	215	182	160	147	140	135	134	134	137	144	152	167	158	141	137	133
Cs	Ba	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi		
272	224	172	159	147	141	137	135	136	139	144	155	171	175	182		

*The values refer to coordination number 12 (see Section 3.2).

Decreases with increase in Z in a period
Increases moving down a group

Metallic Radius in 5d- elements

Metallic radii of 5d- block elements are expected to be larger than that of the 4d- elements, but found that these are not larger. Of course these are larger than 3d- block elements.

Lanthanide Contraction

f-orbitals have poor shielding properties;
low penetrating power.

Table of ionic radii

All anions are larger than their parent atoms

The cations are smaller than their parent atoms

Table 1.5 Ionic radii, r/pm^*						
Li⁺	Be²⁺	B³⁺		N³⁻	O²⁻	F⁻
59(4)	27(4)	11(4)		146	135(2)	128(2)
76(6)					138(4)	131(4)
					140(6)	133(6)
					142(8)	
Na⁺	Mg²⁺	Al³⁺		P³⁻	S²⁻	Cl⁻
99(4)	49(4)	39(4)		212	184(6)	181(6)
102(6)	72(6)	53(6)				
118(8)	89(8)					
K⁺	Ca²⁺	Ga³⁺		As³⁻	Se²⁻	Br⁻
138(6)	100(6)	62(6)		222	198(6)	196(6)
151(8)	112(8)					
159(10)	123(10)					
160(12)	134(12)					
Rb⁺	Sr²⁺	In³⁺	Sn²⁺	Sn⁴⁺	Te²⁻	I⁻
152(6)	118(6)	79(6)	83(6)	69(6)	221(6)	220(6)
160(8)	125(8)	92(8)	93(8)			
173(12)	144(12)					
Cs⁺	Ba²⁺	Tl³⁺				
167(6)	149(6)	88(6)				
174(8)	156(8)	Tl⁺				
188(12)	175(12)	164(6)				

* Numbers in parentheses are the coordination number of the ion. For more values, see *Resource section 1*.

Ionisation Energy (IE)

The minimum energy needed to remove an electron from a gas phase atom

Depends on: (a) Size of the atom - IE decreases as the size of the atom increases; (b) Nuclear Charge - IE increases with increase in nuclear charge; (c) The type of electron - Shielding effect

1st IE: H = 1312 KJ mol⁻¹ Li = 520 KJ mol⁻¹

Reasons: (1) Average distance of 2s electron is greater than that of 1s; (2) Penetration effect; 3. Electronic configuration

Trends in Ionisation Energy (IE)

On moving down the group: IE decreases

- 1. nuclear charge increases**
- 2. Z^* due to screening is (very small) & almost constant down the group**
- 3. number of shells increases, hence atomic size increases.**
- 4. there is an increase in the number of e's which shield the valence electrons from the nucleus.**

On moving across the period: IE increases

- 1. the atomic size decreases**
- 2. nuclear charge increases (does not decrease as much as it happens down gr)**

Trends in Ionisation Energy (IE)

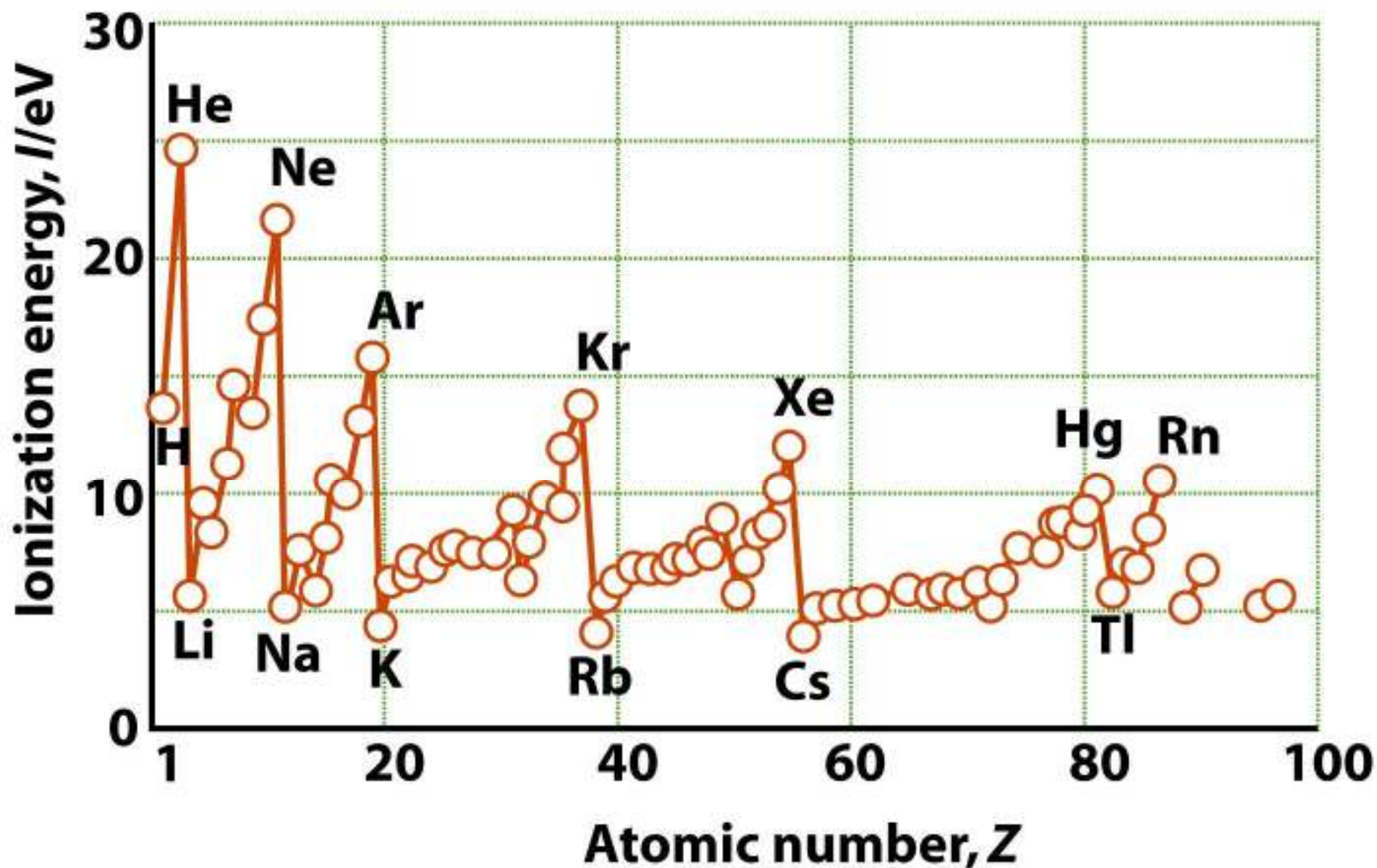


Figure 1-25

Shriver & Atkins Inorganic Chemistry, Fourth Edition

© 2006 by D. F. Shriver, P. W. Atkins, T. L. Overton, J. P. Rourke, M. T. Weller, and F. A. Armstrong

Table 1.6 First and second (and some higher) ionization energies of the elements, $I/(kJ\ mol^{-1})$

H							He
1312							2373
							5259
Li	Be	B	C	N	O	F	Ne
513	899	801	1086	1402	1314	1681	2080
7297	1757	2426	2352	2855	3386	3375	3952
11809	14844	3660	4619	4577	5300	6050	6122
		25018					
Na	Mg	Al	Si	P	S	Cl	Ar
495	737	577	786	1011	1000	1251	1520
4562	1476	1816	1577	1903	2251	2296	2665
6911	7732	2744	3231	2911	3361	3826	3928
		11574					
K	Ca	Ga	Ge	As	Se	Br	Kr
419	589	579	762	947	941	1139	1351
3051	1145	1979	1537	1798	2044	2103	3314
4410	4910	2963	3302	2734	2974	3500	3565
Rb	Sr	In	Sn	Sb	Te	I	Xe
403	549	558	708	834	869	1008	1170
2632	1064	1821	1412	1794	1795	1846	2045
3900	4210	2704	2943	2443	2698	3197	3097
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
375	502	590	716	704	812	926	1036
2420	965	1971	1450	1610	1800	1600	
3400	3619	2878	3080	2466	2700	2900	

Table of ionization energies

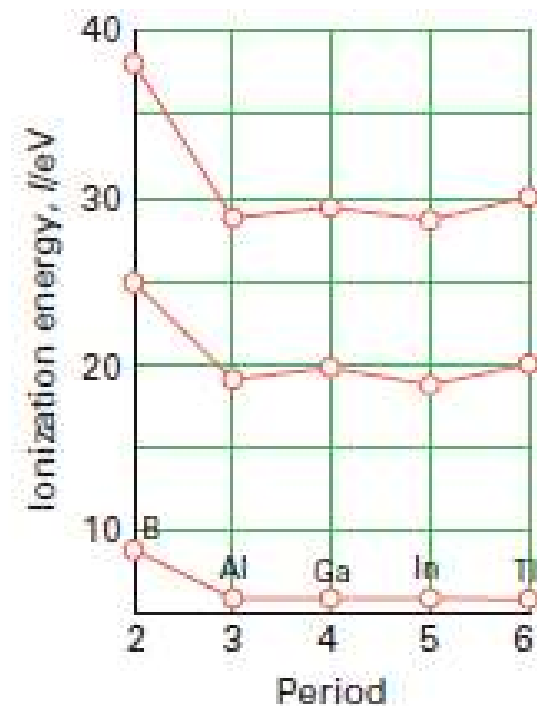


Figure 1.26 The first, second, and third ionization energies of the elements of Group 13. Successive ionization energies increase, but there is no clear pattern of ionization energies down the group.

Electron affinity (EA)

- the amount of energy associated with the gain of electrons**

The greater the energy released in the process of taking up the extra electron, greater is the EA

The EA of an atom is a measure of the tightness with which it binds an additional electron to itself.

Trends in Electron affinity (EA)

On moving across a period,

-As the size decreases, the force of attraction by the nucleus increases. Consequently, the atom has a greater tendency to attract added electron, i.e., EA electron affinity increases

-Generally the EA's of metals are low while those of non-metals are high

-Halogens have high EA. This is due to their strong tendency to change their configuration to ns^2np^6

On moving down a group,

- the atomic size increases and therefore, the effective nuclear attraction decreases and thus electron affinity decreases

Thus, electron affinity follows the left-right trend of electronegativity but not the up-down trend.

The following data are quoted in **kJ/mol**. Elements marked with an asterisk are expected to have negative or near-zero electron affinities on quantum mechanical grounds.

V • T • E																		
Electron affinities in the periodic table																		
Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	H 73																	He -1894
2	Li 60	Be -231											B 27	C 122	N -134	O 141	F 328	Ne *
3	Na 53	Mg *											Al 42	Si 134	P 72	S 200	Cl 349	Ar -1106
4	K 48	Ca 2	Sc 18	Ti 8	V 51	Cr 65	Mn (-96)	Fe 15	Co 64	Ni 112	Cu 119	Zn *	Ga 41	Ge 119	As 79	Se 195	Br 324	Kr *
5	Rb 47	Sr 5	Y 30	Zr 41	Nb 86	Mo 72	Tc *	Ru 101	Rh 110	Pd 54	Ag 126	Cd *	In 39	Sn 107	Sb 101	Te 190	I 295	Xe (-5)
6	Cs 46	Ba 14	*	Hf (11)	Ta 31	W 79	Re 14	Os 104	Ir 150	Pt 205	Au 223	Hg *	Tl 36	Pb 35	Bi 91	Po (127)	At (269)	Rn *
7	Fr (47)	Ra (10)	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Nh	Fl	Mc	Lv	Ts (173)	Og (5)
			*	La 45	Ce 55	Pr 93	Nd (16)	Pm (12)	Sm (16)	Eu 11	Gd (13)	Tb (42)	Dy (34)	Ho (33)	Er (30)	Tm 99	Yb 0	Lu 33
			**	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

Legend
 The number mentioned is **Electron affinity** in **kJ/mol** (rounded)
 For the equivalent value in eV, see: [Electron affinity \(data page\)](#)
 * Denotes elements that are expected to have negative or near-zero electron affinities on quantum mechanical grounds
 Parentheses denote predictions

Electronegativity

measure of the tendency of an element to attract electrons to itself (from its neighbour)

On moving down the group: EN decreases

- Z increases but Z^* almost remains constant
- number of shells (n) increases
- atomic radius increases
- force of attraction between added electron and nucleus decreases

On moving across a period: EN increases

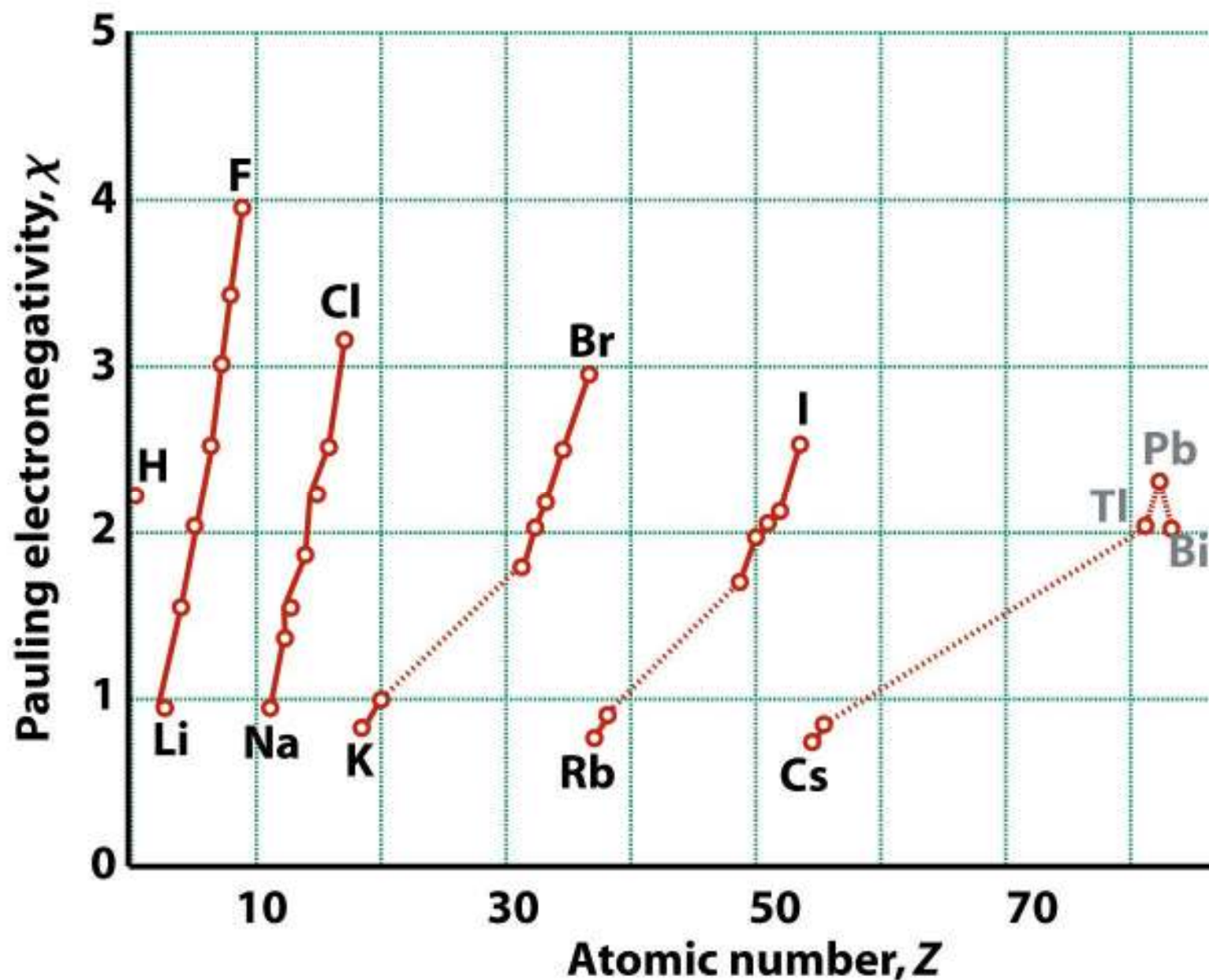
- Z and Z^* increases
- number of shells remains constant
- atomic radius decreases
- force of attraction between added electron and nucleus increases

Periodic table of electronegativity using the Pauling scale

→ Atomic radius decreases → Ionization energy increases → Electronegativity increases →

Group →	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓ Period																		
1	H 2.20																	He
2	Li 0.98	Be 1.57											B 2.04	C 2.55	N 3.04	O 3.44	F 3.98	Ne
3	Na 0.93	Mg 1.31											Al 1.61	Si 1.90	P 2.19	S 2.58	Cl 3.16	Ar
4	K 0.82	Ca 1.00	Sc 1.36	Ti 1.54	V 1.63	Cr 1.66	Mn 1.55	Fe 1.83	Co 1.88	Ni 1.91	Cu 1.90	Zn 1.65	Ga 1.81	Ge 2.01	As 2.18	Se 2.55	Br 2.96	Kr 3.00
5	Rb 0.82	Sr 0.95	Y 1.22	Zr 1.33	Nb 1.6	Mo 2.16	Tc 1.9	Ru 2.2	Rh 2.28	Pd 2.20	Ag 1.93	Cd 1.69	In 1.78	Sn 1.96	Sb 2.05	Te 2.1	I 2.66	Xe 2.60
6	Cs 0.79	Ba 0.89	*	Hf 1.3	Ta 1.5	W 2.36	Re 1.9	Os 2.2	Ir 2.20	Pt 2.28	Au 2.54	Hg 2.00	Tl 1.62	Pb 1.87	Bi 2.02	Po 2.0	At 2.2	Rn 2.2
7	Fr 0.7	Ra 0.9	**	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uus	Uuo
* Lanthanoids			La 1.1	Ce 1.12	Pr 1.13	Nd 1.14	Pm 1.13	Sm 1.17	Eu 1.2	Gd 1.2	Tb 1.1	Dy 1.22	Ho 1.23	Er 1.24	Tm 1.25	Yb 1.1	Lu 1.27	
** Actinoids			Ac 1.1	Th 1.3	Pa 1.5	U 1.38	Np 1.36	Pu 1.28	Am 1.13	Cm 1.28	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr 1.3	

Trends in electronegativity (Pauling scale)



Polarizability

The polarizability of an atom is its willingness to be distorted by the electric field generated by the neighbouring atom or ion.

The polarizability is greater when the electron cloud can be easily distorted, which is common with the atoms having unfilled atomic orbitals which lie close to the highest energy filled orbitals. (for example, heavy atoms & ions)

Small & highly charged cations have greater polarizing ability

Large & highly charged anions are easily polarized

Cations not possessing noble-gas e- configuration are easily polarized

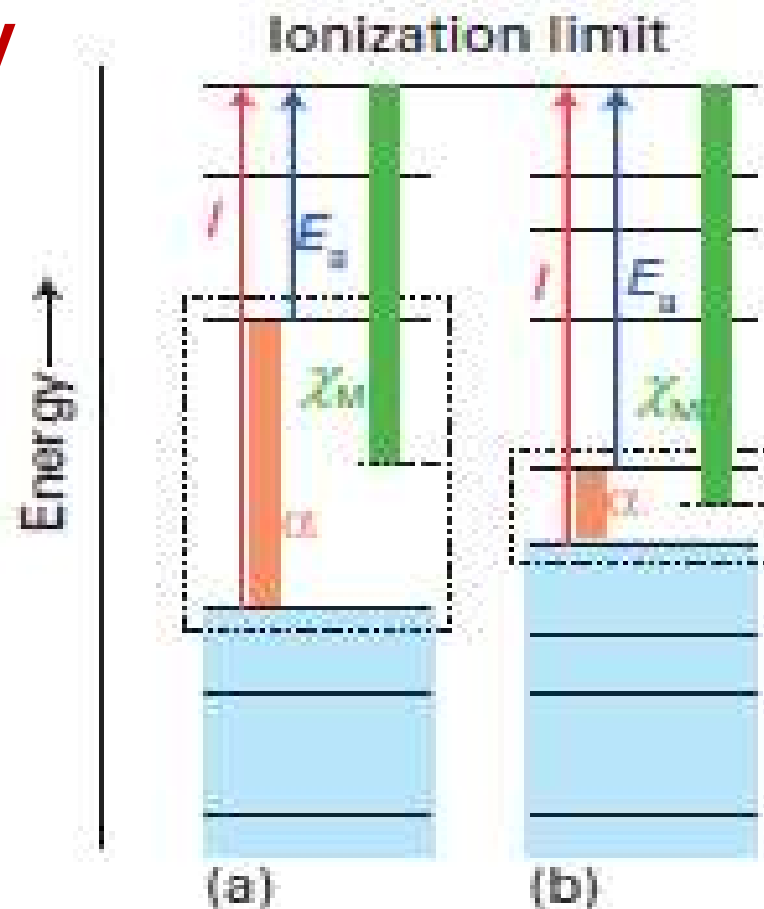


Figure 1.28 The interpretation of the electronegativity and polarizability of an element in terms of the energies of the frontier orbitals (the highest filled and lowest unfilled atomic orbitals). (a) Low electronegativity and polarizability; (b) high electronegativity and polarizability.

- E denotes the strength of the electric field and
- α is the polarizability constant with units of $\text{C m}^2\text{V}^{-1}$.

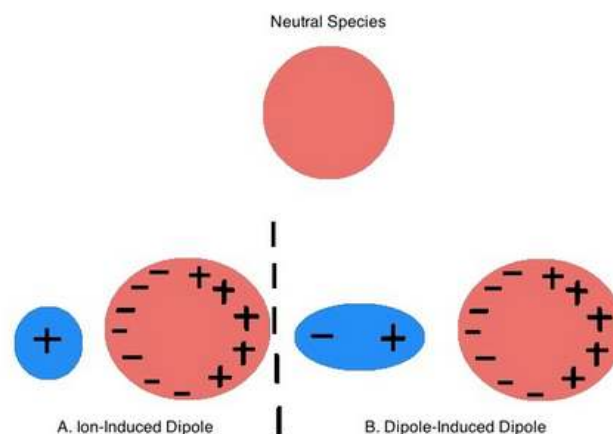


Figure 1.1: A neutral nonpolar species's electron cloud is distorted by A.) an Ion and B.) a polar molecule to induce a dipole moment.

In general, polarizability correlates with the interaction between electrons and the nucleus. The amount of electrons in a molecule affects how tight the nuclear charge can control the overall charge distribution. Atoms with less electrons will have smaller, denser electron clouds, as there is a strong interaction between the few electrons in the atoms' orbitals and the positively charged nucleus. There is also less shielding in atoms with less electrons contributing to the stronger interaction of the outer electrons and the nucleus. With the electrons held tightly in place in these smaller atoms, these atoms are typically not easily polarized by external electric fields. In contrast, large atoms with many electrons, such as negative ions with excess electrons, are easily polarized. These atoms typically have very diffuse electron clouds and large atomic radii that limit the interaction of their external electrons and the nucleus.

Hardness and Softness

[Chemical but not mechanical]

An important concept (HSAB) of compounds formed

Chemical Hardness or Softness of an atom can be correlated with ionization energy (IE), electron affinity (EA), size and polarizability. If the $IE > EA$, the EA can be ignored.

High IE, smaller size, low polarizability -- makes Harder

Low IE, larger size, high polarizability -- makes softer

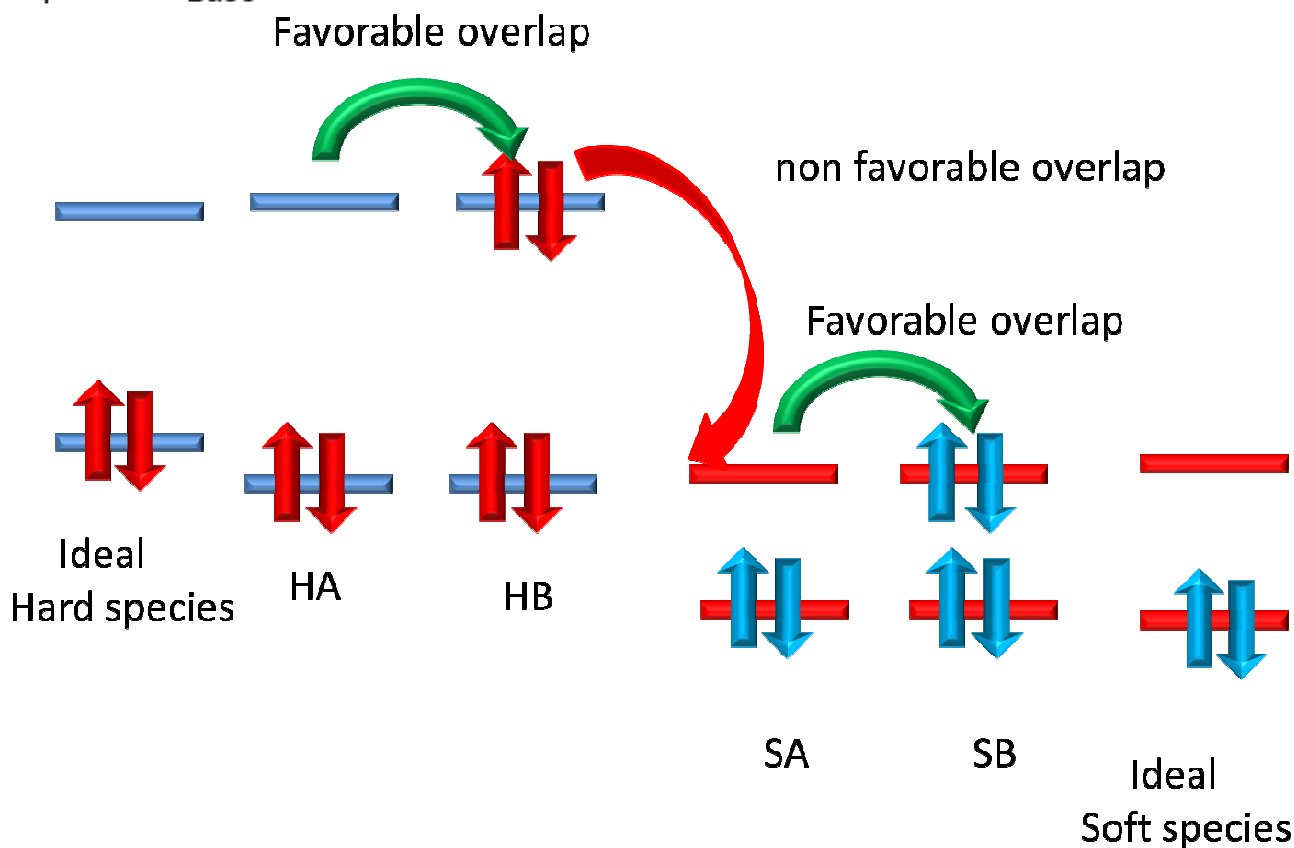
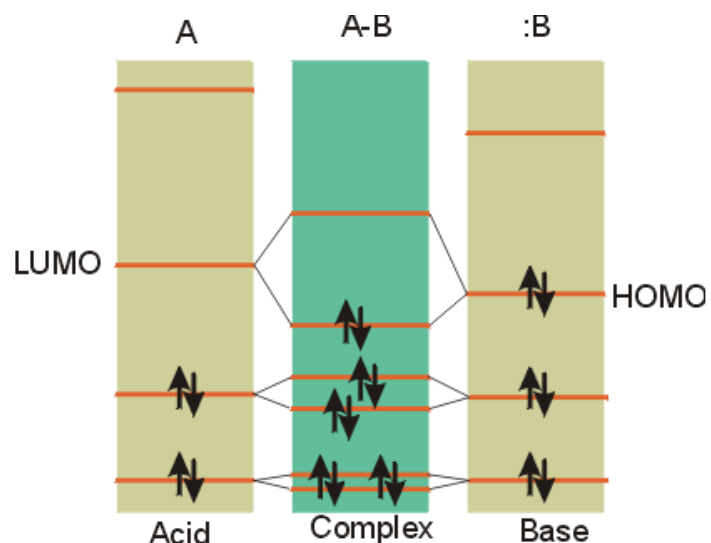
The lighter atoms of a group are chemically harder

The heavier atoms of a group are chemically softer

HSAB: Conceptual

The LUMO and HOMO are called *frontier orbitals*.

If there is a net lowering of energy, the adduct is stable.



Hardness and Softness

The difference between the IE of a neutral atom and its anion: If it is more, the hardness is more & if it is less, the hardness is less or it will be softer.

Pearson's Absolute Hardness $\eta = \frac{I - A}{2}$

Same as saying that the separation between the two frontier orbitals is large it is hard and if small it is soft. .ie. I= HOMO energy and A=LUMO energy

Therefore,

Hard acids tend to bind to hard bases preferentially

Soft acids tend to bind to soft bases preferentially

Example for HSAB

SCN⁻ can bind through either S or N depending upon the HSAB nature of the metal ion.

For eg., Si or Pt

N will prefer Si due to Hard ... Hard type interactions, since 'N' is hard Lewis base, & 'Si' is hard Lewis Acid.

S will prefer Pt due to Soft ... Soft type interactions, since 'S' is soft Lewis base & 'Pt' is soft Lewis Acid

Trends are exhibited,

By keeping the metal same and changing the anion/ligand

By keeping the anion/ligand same and changing the metal

**Bonding/Interaction types: Covalent,
Non-covalent, Ionic**

**Non-covalent interactions are
WEAK Interactions
(between atoms & between molecules)**

**Atoms → Molecules → Supramolecules →
Materials or Solids**

**Hydrogen bonding interactions
Ion – molecular interactions
Vander Waal's interactions**

Van der Waal's Interactions

Three types of Van der Waal's interactions:

(a) Dipole – Dipole Interactions

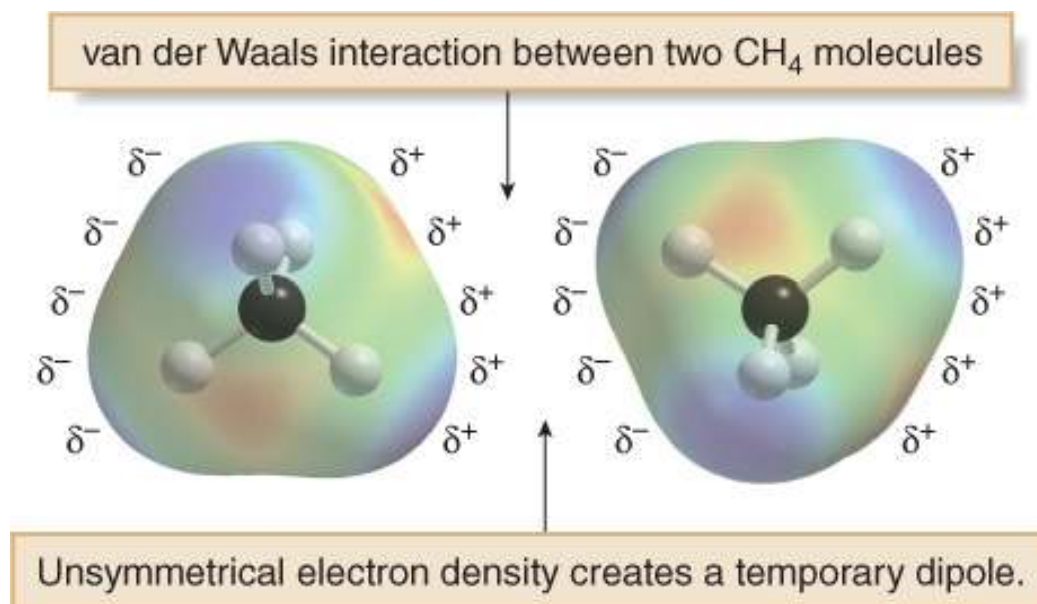
(b) Dipole – Induced Dipole Interactions

**(c) Induced Dipole – Induced Dipole
Transient Dipole – Transient Dipole
(London Dispersion Forces)**

van der Waals Forces (self reading)

- van der Waals forces are also known as London forces.
- They are weak interactions caused by momentary changes in electron density in a molecule.
- They are the only attractive forces present in nonpolar compounds.

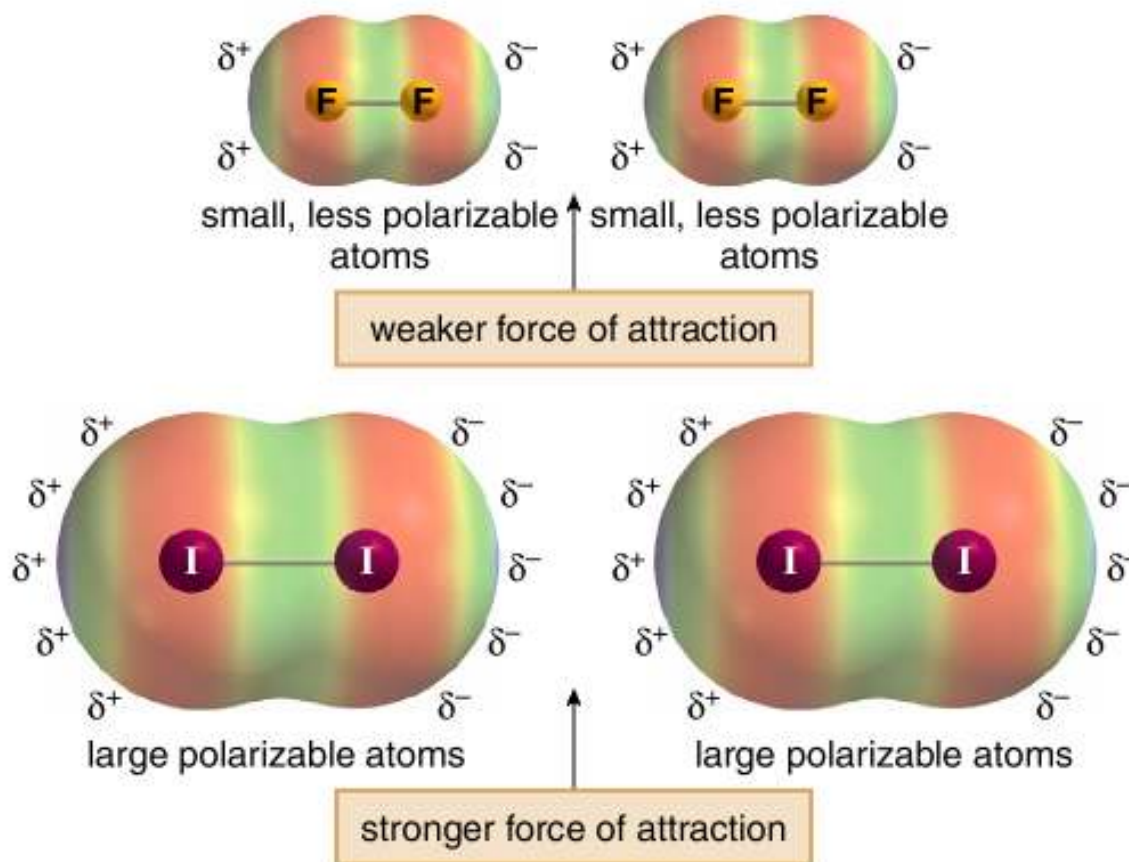
Even though CH_4 has no net dipole, at any one instant its electron density may not be completely symmetrical, resulting in a temporary dipole. This can induce a temporary dipole in another molecule. The weak interaction of these temporary dipoles constitutes van der Waals forces.



van der Waals Forces (self reading)

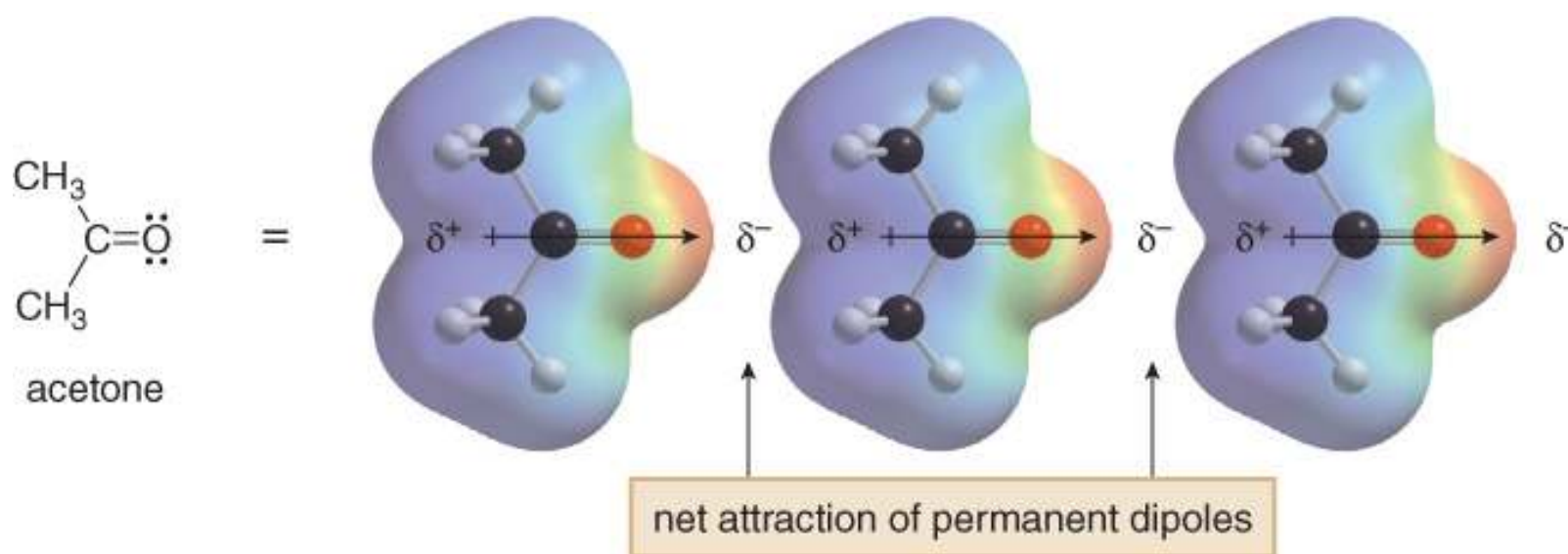
- van der Waals forces are also affected by polarizability.
- **Polarizability** is a measure of how the electron cloud around an atom responds to changes in its electronic environment.

Larger atoms, like iodine, which have more loosely held valence electrons, are more polarizable than smaller atoms like fluorine, which have more tightly held electrons. Thus, two F_2 molecules have little attractive force between them since the electrons are tightly held and temporary dipoles are difficult to induce.



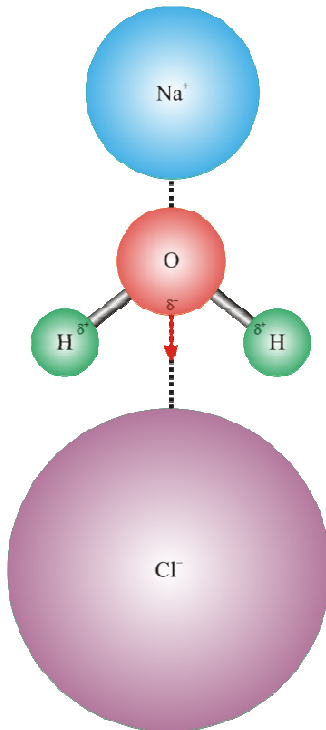
Dipole-Dipole Interactions (self reading)

- Dipole—dipole interactions are the attractive forces between the permanent dipoles of two polar molecules.
- Consider acetone (below). The dipoles in adjacent molecules align so that the partial positive and partial negative charges are in close proximity. These attractive forces caused by permanent dipoles are much stronger than weak van der Waals forces.



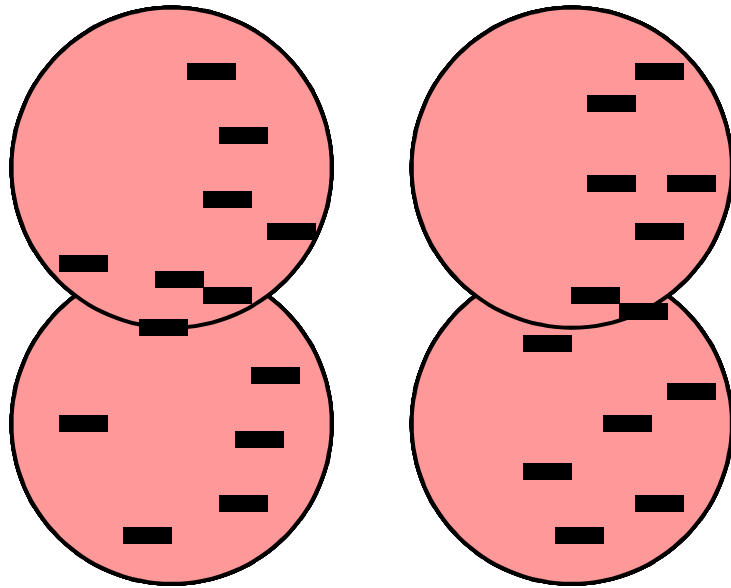
Ion-Dipole (self reading)

- Permanent dipole interacts with an ion.
- This explains for example the solubility of NaCl in water.
- The figure below shows the interaction of Na^+ and Cl^- ions interacting with the permanent dipoles in a water molecule.



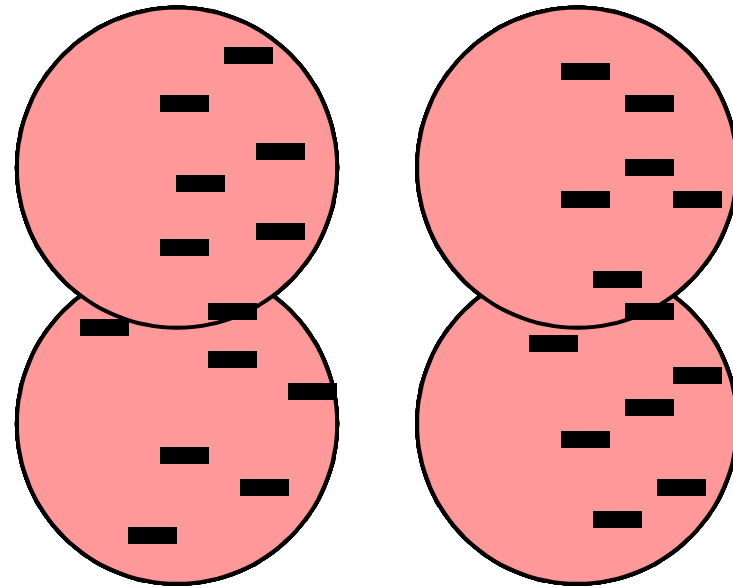
London forces (self reading)

Instantaneous dipole:



Eventually electrons are situated so that tiny dipoles form

Induced dipole:



A dipole forms in one atom or molecule, inducing a dipole in the other

IUPAC Nomenclature of elements

With atomic number above 100 (tutorial)

• Digit	Name	Abbreviation	
• 0	nil	n	
• 1	un	u	E. g.,
• 2	bi	b	114
• 3	tri	t	Ununquadium
• 4	quad	q	Uuq
• 5	pent	p	
• 6	hex	h	118
• 7	sept	s	Ununoctium
• 8	oct	o	
• 9	enn	e	Uuo