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Unit II - Lecture 9

Chemistry

The Molecular Nature of Matter and Change

Fifth Edition

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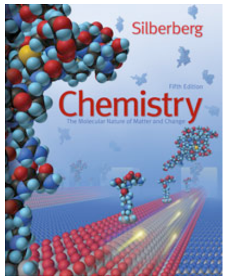


Electron Configuration and Chemical Periodicity

8.4 Trends in Three Key Atomic Properties

8.5 Atomic Structure and Chemical Reactivity

Figure 8.14 Defining metallic and covalent radii.

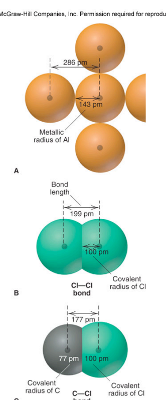
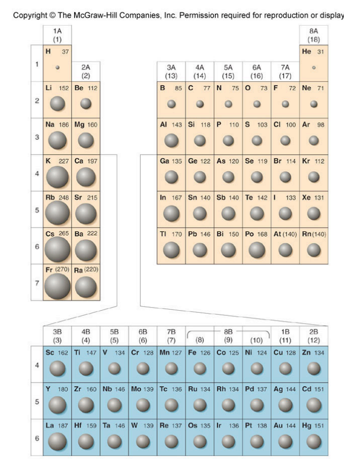


Figure 8.15

Atomic radii of the main-group and transition elements.



Sample Problem 8.3 Ranking Elements by Atomic Size

PROBLEM: Using only the periodic table (not Figure 8.15), rank each set of main-group elements in order of *decreasing* atomic size:

(a) Ca, Mg, Sr (b) K, Ga, Ca (c) Br, Rb, Kr (d) Sr, Ca, Rb

PLAN: Elements in the same group increase in size and you go down a group; elements decrease in size as you go across a period.

SOLUTION:

(a) Sr > Ca > Mg — These elements are in Group 2A(2).

(b) K > Ca > Ga — These elements are in Period 4.

(c) Rb > Br > Kr — Rb has a higher energy level and is far to the left. Br is to the left of Kr.

(d) Rb > Sr > Ca — Ca is one energy level smaller than Rb and Sr. Rb is to the left of Sr.

Figure 8.16 Periodicity of atomic radius.

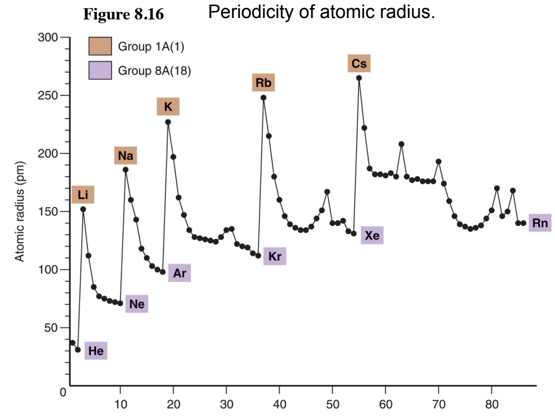


Figure 8.17 Periodicity of first ionization energy (IE_1).

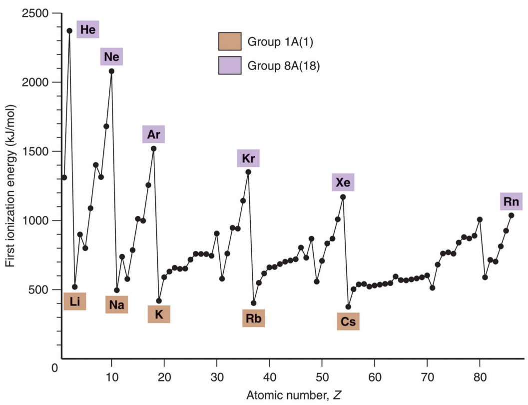


Figure 8.18 First ionization energies of the main-group elements.

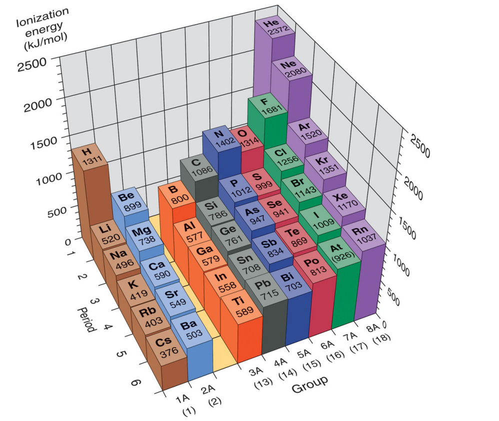
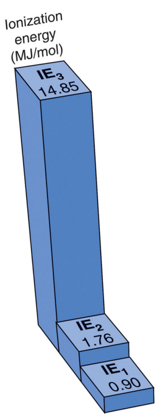


Figure 8.19 The first three ionization energies of beryllium (in MJ/mol).



Sample Problem 8.4 Ranking Elements by First Ionization Energy

PROBLEM: Using the periodic table only, rank the elements in each of the following sets in order of *decreasing* IE_1:

(a) Kr, He, Ar (b) Sb, Te, Sn (c) K, Ca, Rb (d) I, Xe, Cs

PLAN: IE decreases as you proceed down in a group; IE increases as you go across a period.

SOLUTION:

(a) He > Ar > Kr Group 8A(18) - IE_1 decreases down a group.

(b) Te > Sb > Sn Period 5 elements – IE_1 increases across a period.

(c) Ca > K > Rb Ca is to the right of K; Rb is below K.

(d) Xe > I > Cs I is to the left of Xe; Cs is farther to the left and down one period.

Table 8.5 Successive Ionization Energies of the Elements Lithium Through Sodium

Z	Element	Number of Valence Electrons	Ionization Energy (MJ/mol)* IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
3	Li	1	0.52	7.30	11.81							
4	Be	2	0.90	1.76	14.85	21.01			CORE ELECTRONS			
5	B	3	0.80	2.43	3.66	25.02	32.82					
6	C	4	1.09	2.35	4.62	6.22	37.83	47.28				
7	N	5	1.40	2.86	4.58	7.48	9.44	53.27	64.36			
8	O	6	1.31	3.39	5.30	7.47	10.98	13.33	71.33	84.08		
9	F	7	1.68	3.37	6.05	8.41	11.02	15.16	17.87	92.04	106.43	
10	Ne	8	2.08	3.95	6.12	9.37	12.18	15.24	20.00	23.07	115.38	131.43
11	Na	1	0.50	4.56	6.91	9.54	13.35	16.61	20.11	25.49	28.93	141.37

*MJ/mol, or megajoules per mole = 10^3 kJ/mol.

Sample Problem 8.5 Identifying an Element from Successive Ionization Energies

PROBLEM: Name the Period 3 element with the following ionization energies (in kJ/mol) and write its electron configuration:

IE_1	IE_2	IE_3	IE_4	IE_5	IE_6
1012	1903	2910	4956	6278	22,230

PLAN: Look for a large increase in energy which indicates that all of the valence electrons have been removed.

SOLUTION:

The largest increase occurs after IE_5, that is, after the 5th valence electron has been removed. Five electrons would mean that the valence configuration is $3s^2 3p^3$ and the element must be phosphorous, P (Z = 15).

The complete electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^3$.

Figure 8.20 Electron affinities of the main-group elements.

1A (1)	2A (2)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
H −72.8							He (0.0)
Li −59.6	Be ≤0	B −26.7	C −122	N +7	O −141	F −328	Ne (+29)
Na −52.9	Mg ≤0	Al −42.5	Si −134	P −72.0	S −200	Cl −349	Ar (+35)
K −48.4	Ca −2.37	Ga −28.9	Ge −119	As −78.2	Se −195	Br −325	Kr (+39)
Rb −46.9	Sr −5.03	In −28.9	Sn −107	Sb −103	Te −190	I −295	Xe (+41)
Cs −45.5	Ba −13.95	Tl −19.3	Pb −35.1	Bi −91.3	Po −183	At −270	Rn (+41)

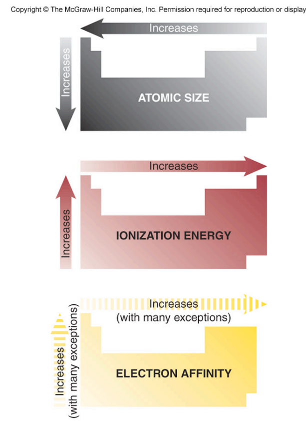


Figure 8.21 Trends in three atomic properties.

Figure 8.22 Trends in metallic behavior.

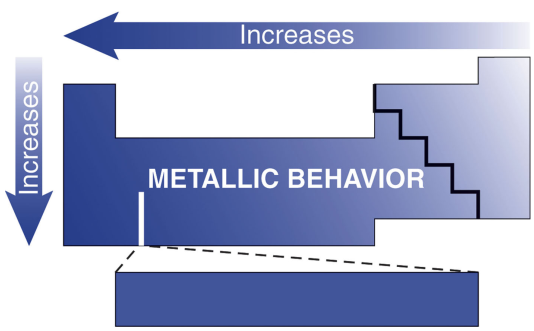


Figure 8.23 The change in metallic behavior of Group 5A(15) and Period 3.

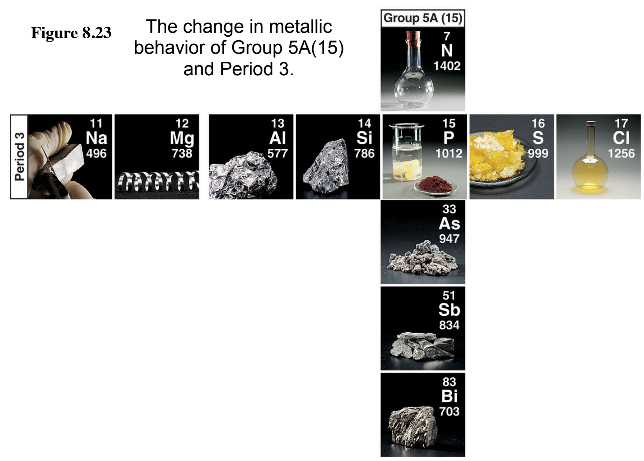


Figure 8.24 The trend in acid-base behavior of element oxides.

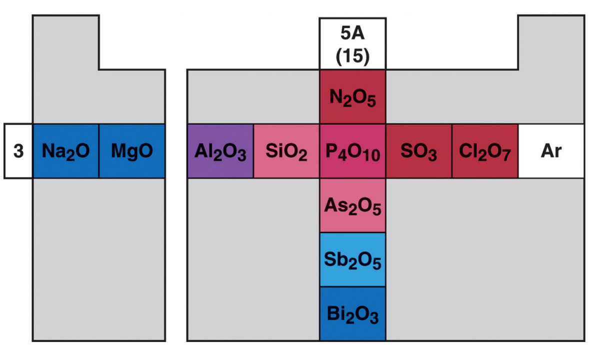
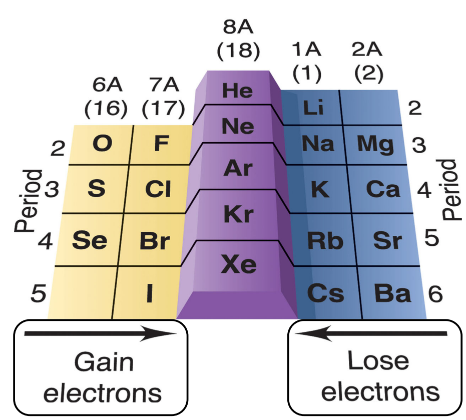


Figure 8.25 Main-group ions and the noble gas electron configurations.



Sample Problem 8.6 Writing Electron Configurations of Main-Group Ions

PROBLEM: Using condensed electron configurations, write reactions for the formation of the common ions of the following elements:

(a) Iodine (Z = 53) (b) Potassium (Z = 19) (c) Indium (Z = 49)

PLAN: Ions of elements in Groups 1A(1), 2A(2), 6A(16), and 7A(17) are usually isoelectronic with the nearest noble gas.

Metals in Groups 3A(13) to 5A(15) can lose their *np* or *ns* and *np* electrons.

SOLUTION:

(a) Iodine (Z = 53) is in Group 7A(17) and will gain one electron to be isoelectronic with Xe: I ([Kr] $5s^24d^{10}5p^5$) + e^- $\rightarrow$ I^- ([Kr] $5s^24d^{10}5p^6$)

(b) Potassium (Z = 19) is in Group 1A(1) and will lose one electron to be isoelectronic with Ar: K ([Ar] $4s^1$) $\rightarrow$ K^+ ([Ar]) + e^-

(c) Indium (Z = 49) is in Group 3A(13) and can lose either one electron or three electrons: In ([Kr] $5s^24d^{10}5p^1$) $\rightarrow$ In^+ ([Kr] $5s^24d^{10}$) + e^+

In ([Kr] $5s^24d^{10}5p^1$) $\rightarrow$ In^{3+}([Kr] $4d^{10}$) + $3e^-$

Figure 8.26 The Period 4 crossover in sublevel energies.

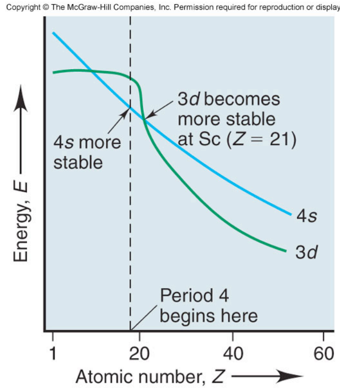
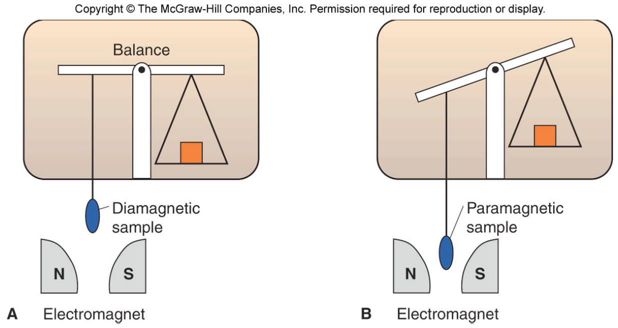


Figure 8.27 Apparatus for measuring the magnetic behavior of a sample.



Sample Problem 8.7 Writing Electron Configurations and Predicting Magnetic Behavior of Transition Metal Ions

PROBLEM: Use condensed electron configurations to write the reaction for the formation of each transition metal ion, and predict whether the ion is paramagnetic.

(a) Mn^{2+}(Z = 25) (b) Cr^{3+}(Z = 24) (c) Hg^{2+}(Z = 80)

PLAN: Write the electron configuration and remove electrons starting with ns to match the charge on the ion. If the remaining configuration has unpaired electrons, it is paramagnetic.

SOLUTION:

(a) Mn^{2+}(Z = 25) Mn ([Ar] $4s^23d^5$) $\rightarrow$ Mn^{2+} ([Ar] $3d^5$) + $2e^-$ paramagnetic

(b) Cr^{3+}(Z = 24) Cr ([Ar] $4s^13d^5$) $\rightarrow$ Cr^{3+} ([Ar] $3d^3$) + $3e^-$ paramagnetic

(c) Hg^{2+}(Z = 80) Hg ([Xe] $6s^24f^{14}5d^{10}$) $\rightarrow$ Hg^{2+} ([Xe] $4f^{14}5d^{10}$) + $2e^-$

not paramagnetic (is diamagnetic)

Figure 8.28 Depicting ionic radius.

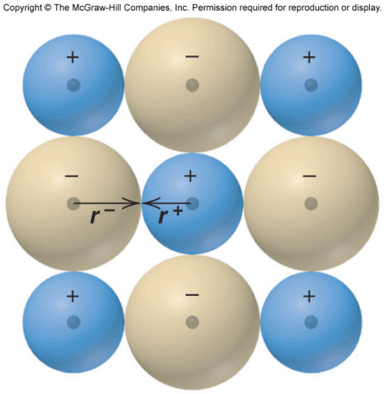
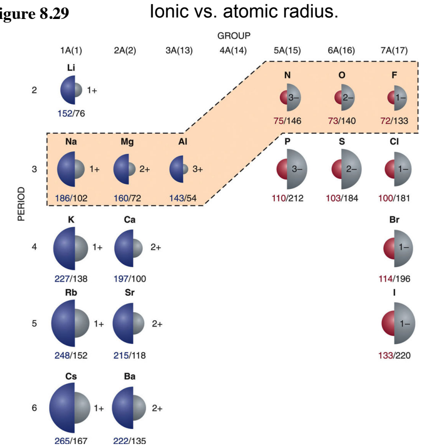


Figure 8.29 Ionic vs. atomic radius.



Sample Problem 8.8 Ranking Ions by Size

PROBLEM: Rank each set of ions in order of *decreasing* size, and explain your ranking:

(a) Ca^{2+}, Sr^{2+}, Mg^{2+} (b) K^+, S^{2-}, Cl^- (c) Au^+, Au^{3+}

PLAN: Compare positions in the periodic table, formation of positive and negative ions and changes in size due to gain or loss of electrons.

SOLUTION:

(a) $Sr^{2+} > Ca^{2+} > Mg^{2+}$ These are members of the same Group 2A(2) and therefore decrease in size going up the group.

(b) $S^{2-} > Cl^- > K^+$ The ions are isoelectronic; S^{2-} has the smallest Z_{eff} and therefore is the largest while K^+ is a cation with a large Z_{eff} and is the smallest.

(c) $Au^+ > Au^{3+}$ The higher the + charge, the smaller the ion.