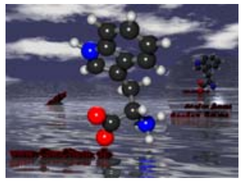
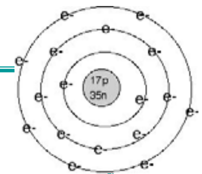
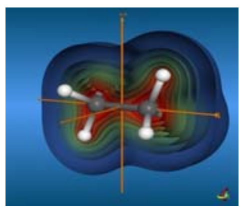




Chapter 7

Electron Configuration and the Periodic Table



Preview

- History of the periodic table.
- Classification of elements in the periodic table.
- Atomic properties from the periodic table (*periodicity*),
  - atomic radius.
  - ionization energy.
  - electron affinity.
  - metallic properties.
- Electron configuration for ions.
- Section 7.7 is a reading assignment and will not be included in the exams.

Development of the Periodic Table

- The main objective from constructing the periodic table is to represent the *patterns* observed in chemical and physical properties for elements.
- Main features of historical development:
  - Elements were generally arranged according to the increase in their atomic masses.
  - In 1864, Newlands showed that chemical properties seemed to repeat for every eight elements (*the law of octaves*).
  - Newlands's work was found to be inadequate for elements beyond calcium.

Development of the Periodic Table

- The basis of today's periodic table was the effort of Mendeleev and Meyer. In 1869, they tabulated the elements based on a phenomenon they called *periodicity*.
- This allowed the scientists to predict the existence of some elements as well as their properties.

Chapter 7 Section 1

Mendeleev's Periodic Table

TABELLE II

REIHEN	GRUPPE I. — R^2O	GRUPPE II. — RO	GRUPPE III. — R^2O^3	GRUPPE IV. RH^4 RO^2	GRUPPE V. RH^3 R^2O^5	GRUPPE VI. RH^2 RO^3	GRUPPE VII. RH R^2O^7	GRUPPE VIII. — RO^4
1	H = 1							
2	Li = 7	Be = 9,4	B = 11	C = 12	N = 14	O = 16	F = 19	
3	Na = 23	Mg = 24	Al = 27,3	Si = 28	P = 31	S = 32	Cl = 35,5	
4	K = 39	Ca = 40	— = 44	Ti = 48	V = 51	Cr = 52	Mn = 55	Fe = 56, Co = 59, Ni = 59, Cu = 63.
5	(Cu = 63)	Zn = 65	— = 68	— = 72	As = 75	Se = 78	Br = 80	
6	Rb = 85	Sr = 87	?Yt = 88	Zr = 90	Nb = 94	Mo = 96	— = 100	Ru = 104, Rh = 104, Pd = 106, Ag = 108.
7	(Ag = 108)	Cd = 112	In = 113	Sn = 118	Sb = 122	Te = 125	J = 127	
8	Cs = 133	Ba = 137	?Di = 138	?Ce = 140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er = 178	?La = 180	Ta = 182	W = 184	—	Os = 195, Ir = 197, Pt = 198, Au = 199.
11	(Au = 199)	Hg = 200	Tl = 204	Pb = 207	Bi = 208	—	—	
12	—	—	—	Th = 231	—	U = 240	—	— — — —

Eka-Aluminum

Note that elements are ordered by their atomic masses

Chapter 7 Section 1

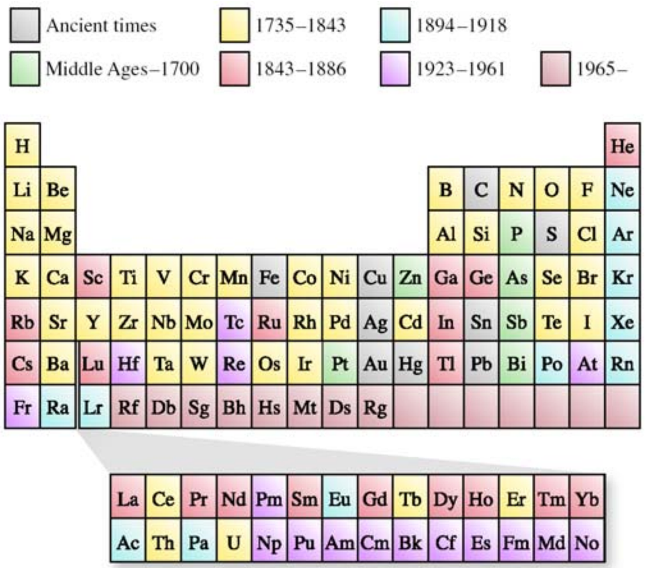
Development of the Periodic Table

- Mendeleev correctly predicted the existence and properties of an element that he called "Eka-aluminum". Four years later, the element Ga was discovered.

Properties	Eka-aluminum (Ea)	Gallium (Ga)
Atomic mass	68 amu	69.9 amu
Melting point	Law	30.15˚C
Density	5.9 g/cm^3	5.94 g/cm^3
Oxide formula	Ea_2O_3	Ga_2O_3

Chapter 7 Section 1

Development of the Periodic Table



Chapter 7 Section 1

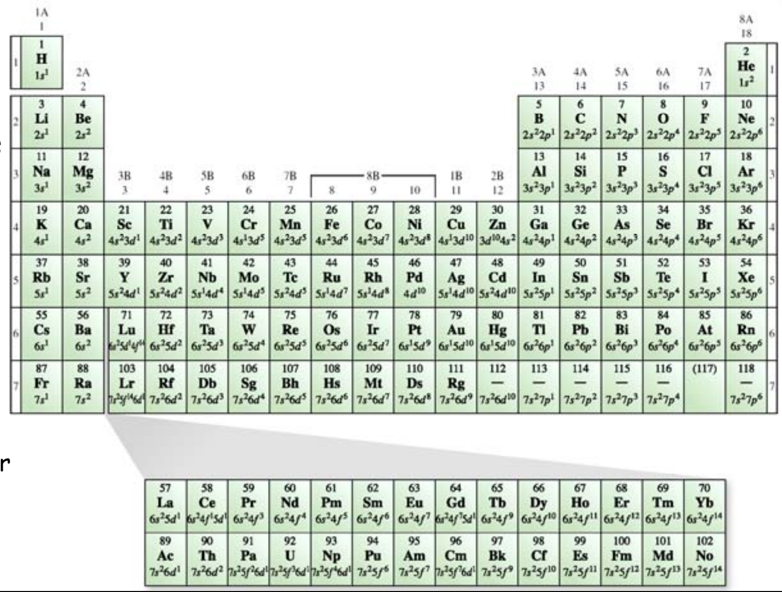
Development of the Periodic Table

- Although Mendeleeve's model was a good one, it could not explain inconsistencies, for instance, all elements were not in order according to atomic mass. (Ar and K, for instance).
- In 1913, Moseley explained the discrepancy. He discovered a correlation between the number of protons (atomic number) and frequency of X rays generated.
- Today, elements are arranged in order of *increasing atomic number*.

Chapter 7 Section 2

The Modern Periodic Table

The configurations shown are those for the ***outermost electrons***, which are the electrons involved in chemical bonding and that are responsible for the chemical properties.



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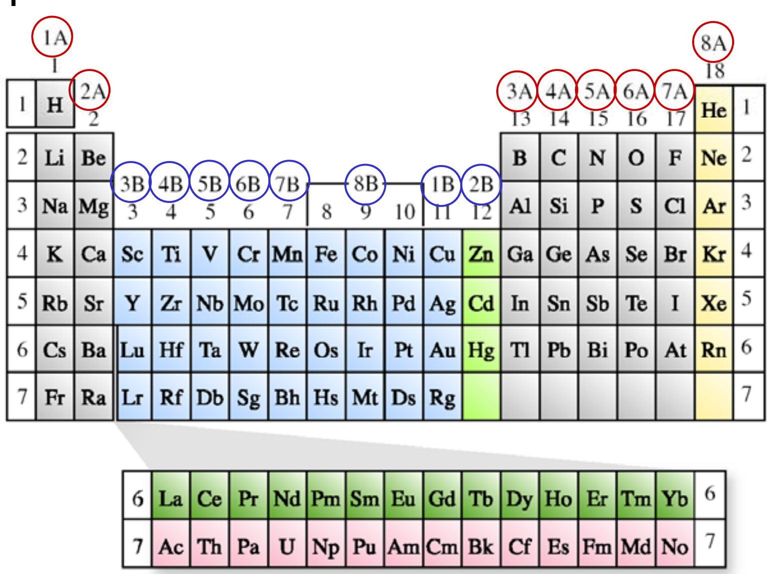
Chapter 7 Section 2

Classification of Elements



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Chapter 7 Section 2

The Modern Periodic Table

- Classification of elements based on the outermost electrons:
  - ***Main group elements*** - "representative elements" Group 1A-7A.
  - ***Noble gases*** - Group 8A all have ns^2np^6 configuration (exception-He).
  - ***Transition elements*** - 1B, 3B - 8B "*d*-block". (have incompletely filled *d* subshells or produce ions with incompletely filled *d* subshells)
  - ***Lanthanides/actinides*** - "*f*-block". (incompletely filled *f* subshells)

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Chapter 7 Section 2

Electron Configuration of a Particular Group

- In general, a particular group in the periodic table has a distinct electron configuration.

TABLE 7.1 Electron Configurations of Group 1A and Group 2A Elements

Group 1A		Group 2A	
Li	$[He]2s^1$	Be	$[He]2s^2$
Na	$[Ne]3s^1$	Mg	$[Ne]3s^2$
K	$[Ar]4s^1$	Ca	$[Ar]4s^2$
Rb	$[Kr]5s^1$	Sr	$[Kr]5s^2$
Cs	$[Xe]6s^1$	Ba	$[Xe]6s^2$
Fr	$[Rn]7s^1$	Ra	$[Rn]7s^2$

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Chapter 7 Section 2

Electron Configuration and Predicting Chemical Properties

- ***Valence electrons*** are the outermost electrons and are involved in bonding.
- Similarity of valence electron configurations helps predict chemical properties.
  - Groups **1A**, **2A** and **8A** all have similar properties to other members of their respective groups.
  - Groups **3A** - **7A** show a considerable variation among properties from metallic, metalloid, to nonmetallic.
  - **Transition metals** do not always exhibit regular patterns in their electron configurations but have some similarities as a whole such as colored compounds and multiple oxidation states.

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Chapter 7 Section 3

Effective Nuclear Charge

- ***Z*** (***nuclear charge***) : the number of protons in the nucleus of an atom.
- Z_{eff} (***effective nuclear charge***) : the actual magnitude of the positive charge "experienced" by an electron in the atom.
- $Z = Z_{eff}$ only in the hydrogen atom.

  $Z > Z_{eff}$ in all other atoms where more than one electron are there.

  Why?

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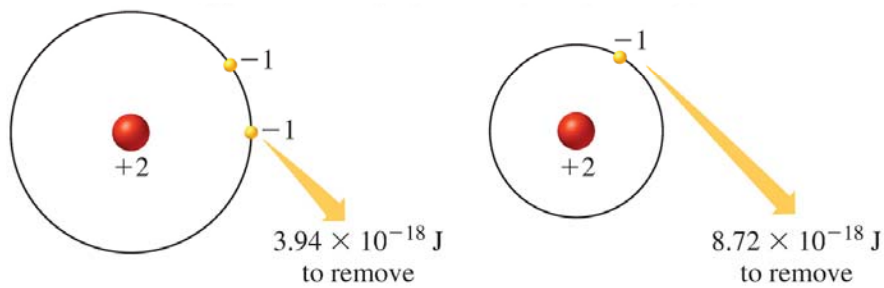
Shielding



- ***Shielding*** occurs when an electron in a many-electron atom is partially shielded from the positive charge of the nucleus by other electrons in the atom.

Trend of Effective Nuclear Charge

- Z_{eff} increases as going across a period of the periodic table.

	Li	Be	B	C	N	O
Z	3	4	5	6	7	8
$\mathbf{Z_{eff}}$ (felt by the valence electrons)	1.28	1.91	2.42	3.14	3.83	4.45

That is because the number of core electrons is the same. Only the value of *Z* and the number of valence electrons increase.

$$Z_{eff} = Z - \sigma$$

σ represents the shielding constant (greater than 0 but less than *Z*)

Chapter 7 Section 3

Trend of Effective Nuclear Charge

- Z_{eff} increases *less significantly* as going from top to bottom in the periodic table.

  That is because there is an additional shell of core electrons that shield the valence electrons from the nucleus.

- Many physical and chemical properties of elements depend on Z_{eff}.

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Chapter 7 Section 4

Periodic Trends in Properties of Elements

- We are going to predict the following important atomic properties from the periodic table:
  - ***Atomic Radius***: obtained from the distances between atoms in chemical compounds.
  - ***Ionization Energy***: minimum energy required to remove electrons from a gaseous atom or ion.
  - ***Electron Affinity***: change in energy associated with the addition of an electron to a gaseous atom.
  - ***Metallic Properties***: metallic characters.

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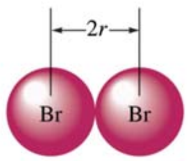
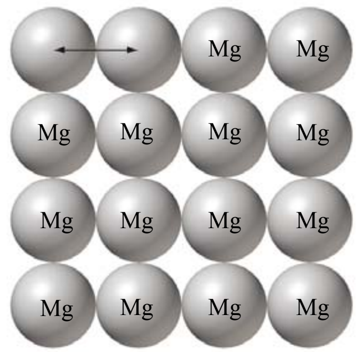
Chapter 7 Section 4

Atomic Radius

- It is half the distance between the nuclei of two adjacent identical atoms.



Covalent radius



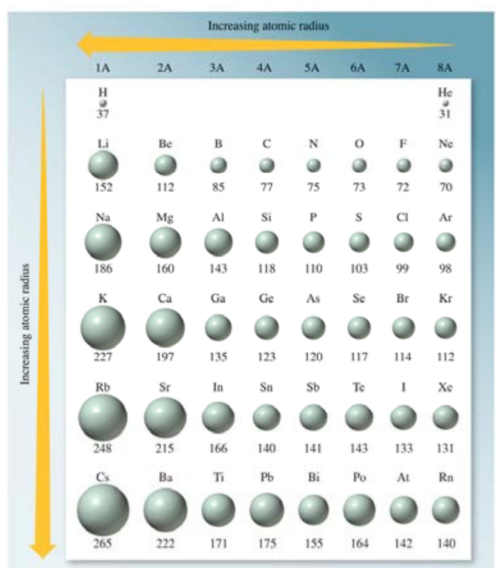
Metallic radius

Chapter 7 Section 4

Atomic Radius

- Atomic radii *increase* in going from top to bottom because the size of the orbitals increases.

- Atomic radii *decrease* in going form left to right because both the effective nuclear charge (Z_{eff}) increases.

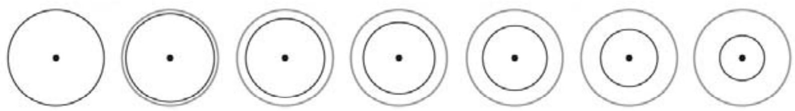


Chapter 7 Section 4

Atomic Radius

- Atomic radii *decrease* in going form left to right because both the effective nuclear charge and the charge of the valence shell increase.

  This results in an increase of the electrostatic attraction force between the nucleus and the electrons.



	Li	Be	B	C	N	O	F
Z_{eff}	1.28	1.91	2.42	3.14	3.83	4.45	5.10
Charge on valence shell	−1	−2	−3	−4	−5	−6	−7
$F \propto$	$\frac{(+1.28)(-1)}{d^2}$	$\frac{(+1.91)(-2)}{d^2}$	$\frac{(+2.42)(-3)}{d^2}$	$\frac{(+3.14)(-4)}{d^2}$	$\frac{(+3.83)(-5)}{d^2}$	$\frac{(+4.45)(-6)}{d^2}$	$\frac{(+5.10)(-7)}{d^2}$

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Chapter 7 Section 4

Atomic Radius

Arrange the following groups of atoms in order of increasing size.

(a) Rb, Na, Be. Be < Na < Rb

(b) Sr, Ne, Se. Ne < Se < Sr

(c) P, Fe, O. O < P < Fe

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Chapter 7 Section 4

Ionization Energy (*IE*)

- It is the minimum energy required to remove an electron from an atom in the gaseous phase.

$$X(g) \rightarrow X^{+}(g) + e^{-}$$

- For example,

$$Na(g) \rightarrow Na^{+}(g) + e^{-} \qquad IE_1(Na) = 495.8 \text{ kJ/mol}$$

  - 495.8 kJ of energy is required to remove 1 mole of electrons from 1 mole of gaseous sodium atoms.
  - IE_1 refers to the minimum energy required to remove the most loosely held electron, which is the outermost electron.

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Chapter 7 Section 4

Ionization Energy (*IE*)

- Generally, ionization energy increases as Z_{eff} increases.

➲ On going down a group (from top to bottom), the value of IE_1 *decreases*.

➲ On going across a period (left to right), the value of IE_1 *increases*.

1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H 1312							He 2372
Li 520	Be 899	B 800	C 1086	N 1402	O 1314	F 1681	Ne 2080
Na 496	Mg 738	Al 577	Si 786	P 1012	S 999	Cl 1256	Ar 1520
K 419	Ca 590	Ga 579	Ge 761	As 947	Se 941	Br 1143	Kr 1351
Rb 403	Sr 549	In 558	Sn 708	Sb 834	Te 869	I 1009	Xe 1170
Cs 376	Ba 503	Tl 589	Pb 715	Bi 703	Po 813	At (926)	Rn 1037

IE_1 values for main group elements. (kJ/mol)

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Chapter 7 Section 4

Interruption in the Periodic Trend of the Ionization Energies

1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H 1312							He 2372
Li 520	Be 899	B 800	C 1086	N 1402	O 1314	F 1681	Ne 2080
Na 496	Mg 738	Al 577	Si 786	P 1012	S 999	Cl 1256	Ar 1520
K 419	Ca 590	Ga 579	Ge 761	As 947	Se 941	Br 1143	Kr 1351
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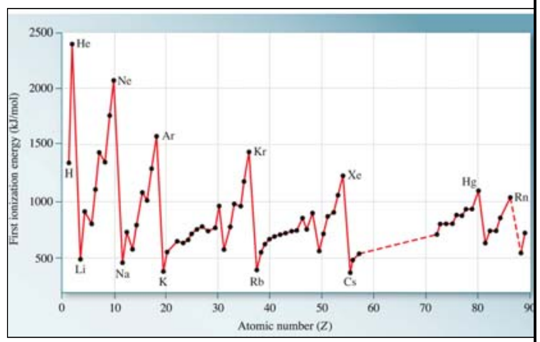
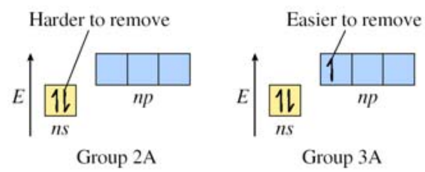
Cs 376	Ba 503	Tl 589	Pb 715	Bi 703	Po 813	At (926)	Rn 1037

IE_1 values for main group elements. (kJ/mol)



IE_1(Be) > IE_1(B) because the 2*s* electrons provide some shielding for the 2*p* electron from the nuclear charge.



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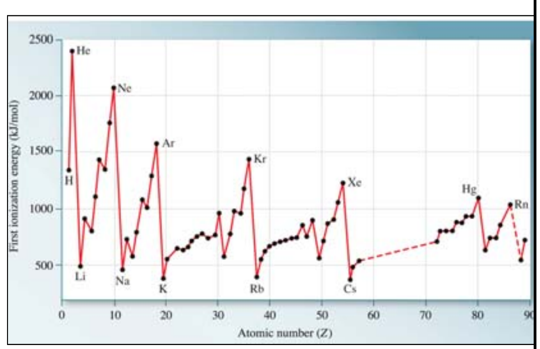
Chapter 7 Section 4

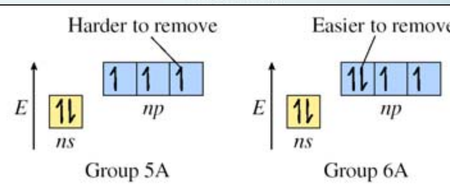
Interruption in the Periodic Trend of the Ionization Energies

1A 1	2A 2	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
H 1312							He 2372
Li 520	Be 899	B 800	C 1086	N 1402	O 1314	F 1681	Ne 2080
Na 496	Mg 738	Al 577	Si 786	P 1012	S 999	Cl 1256	Ar 1520
K 419	Ca 590	Ga 579	Ge 761	As 947	Se 941	Br 1143	Kr 1351
Rb 403	Sr 549	In 558	Sn 708	Sb 834	Te 869	I 1009	Xe 1170
Cs 376	Ba 503	Tl 589	Pb 715	Bi 703	Po 813	At (926)	Rn 1037

IE_1 values for main group elements. (kJ/mol)



IE_1(N) > IE_1 (O) because of the electron repulsion in the doubly occupied 2*p* orbital.



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Chapter 7 Section 4

Multiple Ionizations

- Can we remove another electron from the cation generated from the IE_1 step?

- $B(g) \rightarrow B^{+}(g) + e^{-}$ $IE_1 = 800$ kJ/mol
  $B^{+}(g) \rightarrow B^{2+}(g) + e^{-}$ $IE_2 = 2427$ kJ/mol
  $B^{2+}(g) \rightarrow B^{3+}(g) + e^{-}$ $IE_3 = 3660$ kJ/mol
  $B^{3+}(g) \rightarrow B^{4+}(g) + e^{-}$ $IE_4 = 25026$ kJ/mol

  The increase in *IE*'s is because of the increase in the effective nuclear charge

- $IE_1 < IE_2 < IE_3 << IE_4$

  [He]$2s^2 2p^1$ [He]$2s^2$ [He]$2s^1$ $1s^2$

The highest-energy electron (the one that is most loosely bound to the nucleus) will be removed first

Removing a core electron which is closer to the nucleus

Chapter 7 Section 4

Ionization Energy

TABLE 7.3 Ionization Energies (in kJ/mol) for Elements 3 through 11*

	Z	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
Li	3	520	7,298	11,815							
Be	4	899	1,757	14,848	21,007	21,007					
B	5	800	2,427	3,660	25,026	32,827					
C	6	1,086	2,353	4,621	6,223	37,831	47,277				
N	7	1,402	2,856	4,578	7,475	9,445	53,267	64,360			
O	8	1,314	3,388	5,301	7,469	10,990	13,327	71,330	84,078		
F	9	1,681	3,374	6,050	8,408	11,023	15,164	17,868	92,038	106,434	
Ne	10	2,080	3,952	6,122	9,371	12,177	15,238	19,999	23,069	115,380	131,432
Na	11	496	4,562	6,910	9,543	13,354	16,613	20,117	25,496	28,932	141,362

*Cells shaded with blue represent the removal of core electrons.

- I.E. *increases largely* in going from valence-electron removal to core-electron removal.

Chapter 7 Section 4

Exercise

In each of the following which atom (or ion) has the smallest first ionization energy (IE_1)?

(a) Ca, Sr, Ba. Ba

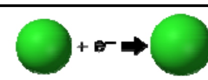
(b) K, Mn, Ga. K

(c) N, O, F. O

(d) S^{2-}, S, S^{2+}. S^{2-}

(e) Cs, Ge, Ar. Cs

Chapter 7 Section 4



Electron Affinity (*EA*)

- The change in energy *released* when an atom in the gaseous phase accepts an electron.

$$X(g) + e^- \rightarrow X^-(g)$$

- For example,

$$Cl(g) + e^- \rightarrow Cl^-(g) \qquad \Delta H = -349.0 \text{ kJ/mol}$$

exothermic process

  - 349 kJ of energy is released when 1 mole of gaseous chlorine atoms accepts 1 mole of electrons.
  - *EA* value of chlorine is +349.0 kJ/mol.
  - The more +ve the *EA* value, the more favorable the process.

Chapter 7 Section 4

Trend of Electron Affinity (*EA*)

- Generally, *EA* *increases* as going across a period from left to right and *decreases* as going down a group from top to bottom.

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)

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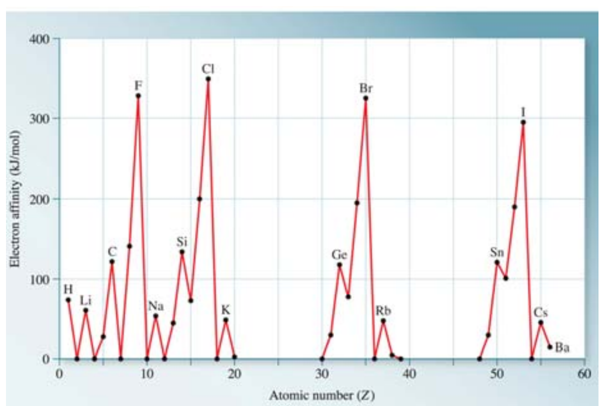
Chapter 7 Section 4

Trend of Electron Affinity (*EA*)

- Generally, *EA* *increases* as going across a period from left to right and *decreases* as going down a group from top to bottom.
- Periodic interruptions still exist in the trend of *EA* values.



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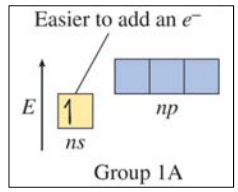
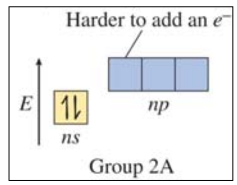
Chapter 7 Section 4

Interruptions in the Periodic Trend of the Electron Affinity

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)

- It is *easier* to add an electron to a group 1A element than to a group 2A element.





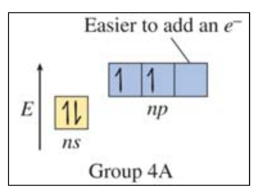
The p orbital is of a higher energy than the s orbital

Chapter 7 Section 4

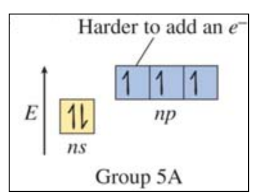
Interruptions in the Periodic Trend of the Electron Affinity

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)

- It is *easier* to add an electron to a group 4A element than to a group 5A element.



"Little repulsion"



"Extra repulsion"

Chapter 7 Section 4

Metallic Character

- ***Metals***
  - Shiny, lustrous, malleable.
  - Good conductors.
  - Low *IE* (easily form cations) .
  - Form ionic compounds with chlorine.
  - Form basic, ionic compounds with oxygen.
- Metallic character *increases* top to bottom in group and *decreases* left to right across a period

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Chapter 7 Section 4

Metallic Character

- ***Nonmetals*** (show trend opposite to metals)
  - Vary in color, not shiny.
  - Brittle.
  - Poor conductors.
  - Form acidic, molecular compounds with oxygen.
  - High *EA* (easily form anions).
- ***Metalloids***
  - Properties between the metals and nonmetals.

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Chapter 7 Section 5

Electron Configuration of Ions

- Electron configuration of ions
  - follows the same rules we studied for neutral atoms.
  - helps explain charges memorized earlier.
- Noble gases (Group 8A) almost completely unreactive due to their electron configuration.

  ns^2np^6 (except He: $1s^2$)
  - No tendency to accept electrons (EA = –ve) or to lose electrons (IE = highly +ve)

Main group elements tend to gain or lose electrons to become ***isoelectronic*** with a noble gas element (same valence electron configuration as the nearest noble gas)

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Chapter 7 Section 5

Ions of Main Group Elements

Main group elements tend to gain or lose electrons to become ***isoelectronic*** with a noble gas element (same valence electron configuration as nearest noble gas)

- Na: $1s^22s^22p^63s^1 \rightarrow Na^+ : 1s^22s^22p^6$

  Na: $[Ne]3s^1 \rightarrow Na^+$: [Ne]

  (Na^+ : 10 electrons - ***isoelectronic*** with Ne)
- Cl: $1s^22s^22p^63s^23p^5 \rightarrow Cl^-\ 1s^22s^22p^63s^23p^6$

  Cl: $[Ne]3s^23p^5 \rightarrow Cl^-$ [Ar]

  (Cl^- : 18 electrons - ***isoelectronic*** with Ar)

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Chapter 7 Section 5

Exercise

- Write the electron configuration for Li^+ and Ba^{2+}.

- List all species that are likely to have the following electron configuration: $1s^22s^22p^6$.

Chapter 7 Section 5

Ions of *d*-Block Elements

- Although the *ns* orbital fills before the *(n -1)d* orbital in transition metals, when a *d*-block element becomes a cation, it loses electrons based on the following:
  - first from the *ns* subshell,
  - then from the *(n -1)d* subshell.

- This explain why many of the transition metals can form ions with a +2 charge.

Fe: $[Ar]4s^23d^6 \rightarrow Fe^{2+}: [Ar]3d^5$

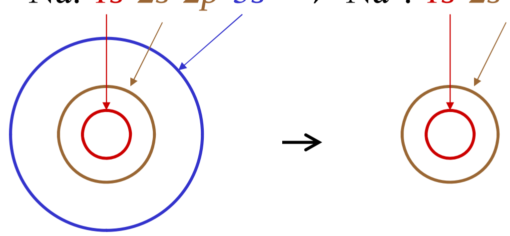
Fe: $[Ar]4s^23d^6 \rightarrow Fe^{3+}: [Ar]3d^5$

Chapter 7 Section 6

Ionic Radius

- When an atom gains or loses electrons, the radius changes.
- **Cations** are always *smaller* than their parent atoms. (often losing an energy level)

Na: $1s^2 2s^2 2p^6 3s^1$ → Na^+: $1s^2 2s^2 2p^6$



Na^+ ion is smaller in size than Na atom

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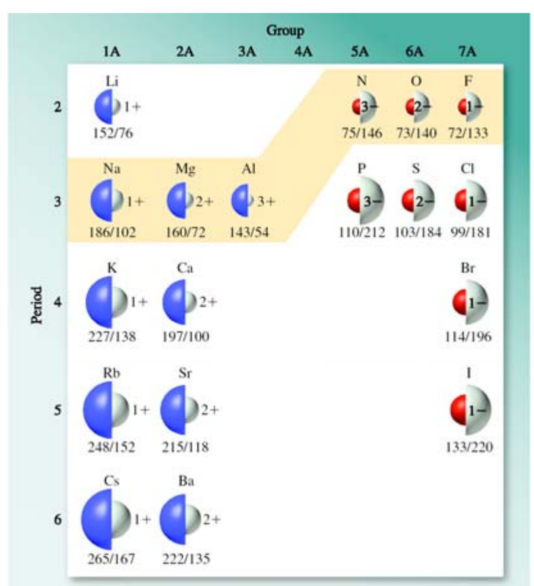
Chapter 7 Section 6

Ionic Radius

- **Cations** are always *smaller* than their parent atoms. (often losing an energy level)
- **Anions** are always *larger* than their parent atoms. (increased e^- repulsions)



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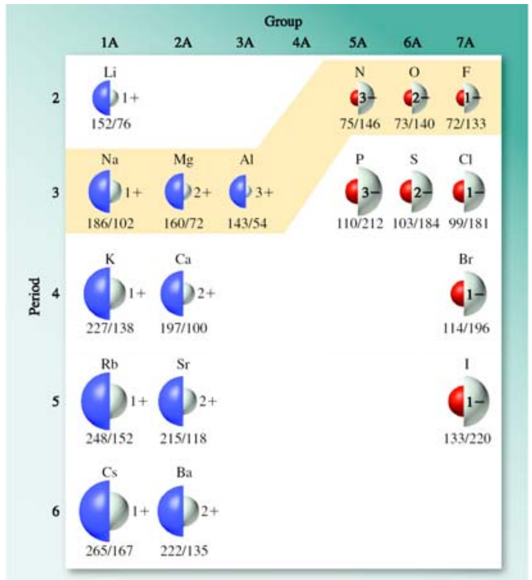
Chapter 7 Section 6

Isoelectronic Series

- ***Isoelectronic atoms***: two or more species having the same electron configuration (same number of electrons) but different nuclear charges.

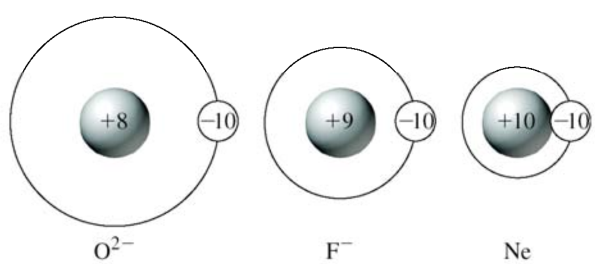
  In this case, size varies significantly.



Chapter 7 Section 6

Isoelectronic Series

- In isoelectronic series
  - the species with the smallest nuclear charge will have the largest radius.
  - the species with the largest nuclear charge will have the smallest radius.



Chapter 7 Section 6

Exercise

Arrange the following ions in order of decreasing size:

Ba^{2+}, Cs^{+}, I^{-}, Sr^{2+} and Te^{2-}

$Te^{2-} > I^{-} > Cs^{+} > Ba^{2+} > Sr^{2+}$