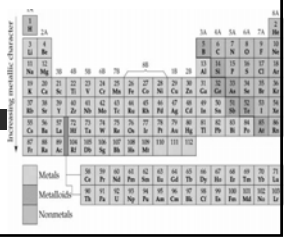
Chapter 6

Ionic Bonds and Periodic Trends

Periodic Table Song



Electron Configurations of Ions

- **Anions**: simply continue to add electrons in the Aufbau order
- Write the electron configuration of:
- F^-
- O^{2-}
- N^{3-}
- $1s^2\ 2s^2\ 2p^6$ for each
- These ions are called ***isoelectronic***.

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Electron Configurations of Ions

- **Cations**: remove electrons in the reverse of the Aufbau order
- Write the electron configuration of:
- Al^{3+}
- Mg^{2+}
- Na^+
- $1s^2\ 2s^2\ 2p^6$
- These ions are also ***isoelectronic***.
- Problems 6.1, 6.2

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Electron Configurations of Ions

- Transition metal cations: remove electrons from the *s* orbital first, then in the reverse Aufbau order
- Write the electron configurations for:
- Cr^{2+}
  - ♦ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^4$
- Ti^{3+}
  - ♦ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^1$
- Fe^{3+}
  - ♦ $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^5$

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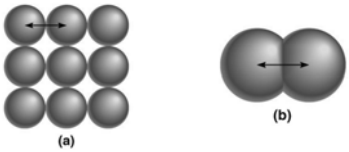
Atomic Radius, Section 5.15

- Atomic radius can be predicted by looking at elements' numbers of electrons
- Definition: one-half the distance between two nuclei in two adjacent atoms



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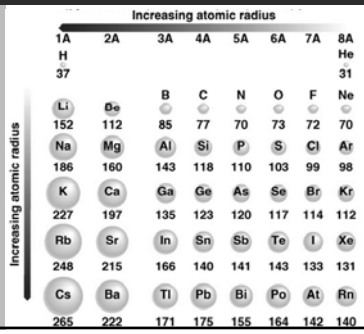
Atomic Radius

Atomic Radii



What trends do you notice (left to right; top to bottom)? Why?

Problem 5.21

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Effective Nuclear Charge

Effective Nuclear Charge

- Electrons are all attracted to the nucleus, but electrons in inner shells shield protons and reduce attractive forces of valence electrons.
- The ***effective nuclear charge*** (Z_{eff}) is the amount of positive charge from the nucleus that is **perceived** by an electron.
  - ♦ In a row (or period) in the Periodic Table, the number of protons increases, but the number of inner e^- (shielding e^-) stays the same. Atoms on the right side of the Table can pull e^- in more tightly.

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Trends in Atomic Radius

- Which element in each pair has a larger atomic radius? Why?
- F or Cl
- C or N
- Rb or Ca
- Na or Mg
- K or Na

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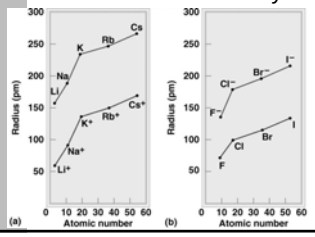
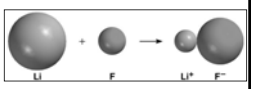
Ionic Radius

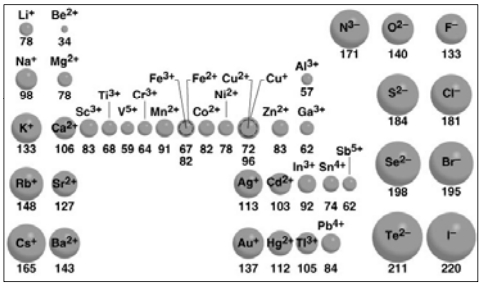
- Radius of cation or anion
- Which gets larger as it goes from atom to ion: cation or anion? Why?

Ionic Radii





Trends in Ionic Radii



Z_{eff} for Isoelectronic Series

- In an isoelectronic series, all ions have the same number of electrons, but the number of **protons** increases from most negative to most positive ion.
- Therefore, the radius of the most positive ion is smallest and the most negative ion is largest.
- Place the following ions in order of increasing ionic radius: O^{2-}, Al^{3+}, Na^+, F^-, Mg^{2+}, N^{3-}

In each pair, which is larger? Why?

- Be^{2+} or B^{3+}
- Al^{3+} or P^{3-}
- Ca^{2+} or Ca
- K or Ca
- O^{2-} or F^-
- Problem 6.3, 6.4

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Ionization Energy

- Equation representing ionization energy: $X(g) \rightarrow X^{+}(g) + e^{-}$
- Define ***ionization energy***

Mg → Mg^{+1}

IE Defined

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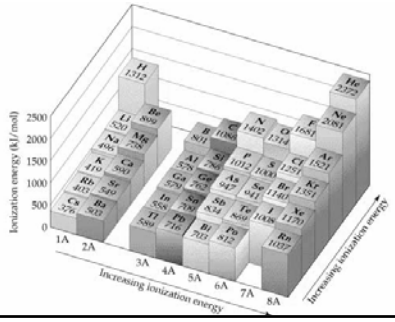
Periodicity of First Ionization Energy

Ionize Trends



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Trends in First Ionization Energy

- Which member of each pair has the larger first ionization energy? Why?
- F or Cl
- N or C
- O or F
- Na or Mg
- K or Na
- Worked Ex. 6.1; Problem 6.5

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Electron Affinity

- Equation representing electron affinity: $X(g) + e^{-} \rightarrow X^{-}(g)$
- Define ***electron affinity***
- Ionization energies are positive values (require input of energy). Electron affinities are negative for most atoms and for all cations.
- Greater attraction between atom and electron results in more negative EA (e.g., halogens).
- Worked Ex. 6.3; Problem 6.10

E- Affinity

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Electron Affinity (kJ/mol)

EA Trends

TABLE 8.3 Electron Affinities (kJ/mol) of Some Representative Elements and the Noble Gases*

1A	2A	3A	4A	5A	6A	7A	8A
H							He
73							< 0
Li	Be	B	C	N	O	F	Ne
60	≤ 0	27	122	0	141	328	< 0
Na	Mg	Al	Si	P	S	Cl	Ar
53	≤ 0	44	134	72	200	349	< 0
K	Ca	Ga	Ge	As	Se	Br	Kr
48	2.4	29	118	77	195	325	< 0
Rb	Sr	In	Sn	Sb	Te	I	Xe
47	4.7	29	121	101	190	295	< 0
Cs	Ba	Tl	Pb	Bi	Po	At	Rn
45	14	30	110	110	?	?	< 0

* The electron affinities of the noble gases, Be, and Mg have not been determined experimentally, but are believed to be close to zero or negative.

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Octet Rule

- Elements react in order to obtain 8 valence (outermost) electrons. They do this by **transferring** electrons or **sharing** electrons.
  - ♦ Metals tend to lose electrons (low I.E.), nonmetals tend to gain electrons (high E.A.). Nonmetal atoms donate electrons to metals to make ionic bonds. Both elements achieve an octet.
  - ♦ Covalent bonds form when elements have to share electrons in order to get 8 valence electrons (an octet) (nonmetals + nonmetals).

The Ionic Bond

- What do the following have in common?
  - ♦ LiF
  - ♦ CaO
  - ♦ Mg_3N_2

- Look at Ionization Energies and Electron Affinities
  - ♦ Low ionization energy → cations
  - ♦ High electron affinity → anions

Lattice Energy: $F = k^*(z_1 \cdot z_2 / d^2)$

- ***Lattice energy***: energy required to completely separate one mole of a solid ionic compound into gaseous ions

TABLE 6.3 Lattice Energies of Some Ionic Solids (kJ/mol)

	Anion				
Cation	F^-	Cl^-	Br^-	I^-	O^{2-}
Li^+	1036	853	807	757	2925
Na^+	923	787	747	704	2695
K^+	821	715	682	649	2360
Be^{2+}	3505	3020	2914	2800	4443
Mg^{2+}	2957	2524	2440	2327	3791
Ca^{2+}	2630	2258	2176	2074	3401
Al^{3+}	5215	5492	5361	5218	15,916

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Relative Lattice Energies

- Charges and sizes of ions determine the value of the lattice energy
- Larger lattice energy → more stable crystal → stronger ionic bond
  - ♦ Charges are greater, or
  - ♦ Sizes are smaller (lattice energy is inversely related to size)
    - Smaller radii → nuclei more strongly attracted to opposite electrons

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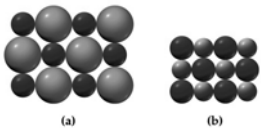
Relative Lattice Energies

- Which compound in each pair is more stable? Is this determined by a high or low lattice energy?
  - ♦ NaCl or $MgCl_2$
  - ♦ MgO or Na_2O
  - ♦ NaCl or KCl
  - ♦ NaBr or NaCl
- Worked Ex. 6.5 →
- Problem 6.12, 6.13



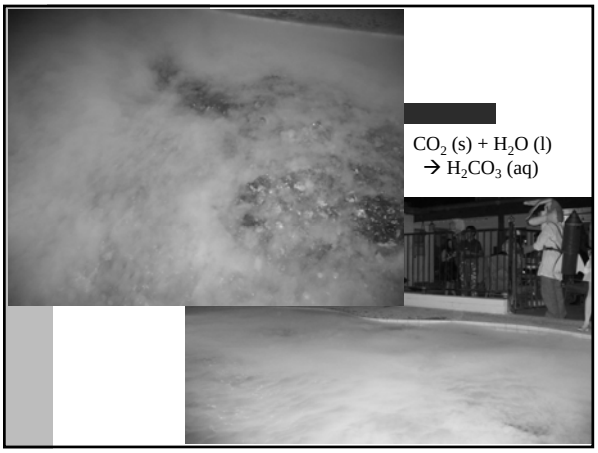
(a) (b)

Metals versus Nonmetals

- Metals tend to react with water to form bases:
  - ♦ $2Na\ (s) + 2H_2O \rightarrow 2NaOH + H_2$
  - ♦ $MgO\ (s) + H_2O \rightarrow Mg(OH)_2$

Alkali metals + water

- Nonmetals tend to react with water to form acids:
  - ♦ $2F_2\ (g) + 2H_2O \rightarrow 4HF + O_2$
  - ♦ $CO_2\ (g) + H_2O \rightarrow H_2CO_3$



$CO_2\ (s) + H_2O\ (l) \rightarrow H_2CO_3\ (aq)$

Sections 7 – 11

- These cover general trends in reactivities of groups of elements. We won't cover these in detail.