Lecture 7: Periodic properties

Read: BLB 2.5; 7.1–7.6
HW: BLB 2.37; 7.11, 23, 25, 27, 31, 45a–c, 47a, e, f, 53, 61, 94
Sup 7:1–12

Know:
• screening effects
• periodic properties
  atomic and ion sizes
  isoelectronic series
  electron configurations of ions
  ionization energies
  electron affinities

Exam 1: Monday, Feb 9 @ 6:30!!! start preparing now!! only non-
text programmable calculators are allowed—no PDAs, blackberries,
cell phones, etc. will be permitted. Bring: pencils, student ID and a
calculator—Absolutely NO text-programmable calculators or wireless devices
(will be checked)

Form a study group, use the CRC, take advantage of SI (info on web), use the online
resources, and work those problems—practice, practice, practice

Bonus deadline for Skill check tests 3 & 4 is tomorrow Thursday,
1/29 @ midnight; Bonus deadline for Skill check test 5 is
Thursday, 2/5 & Skill check test 6 is SUNDAY, 2/8

Sheets's office hours: Mondays 12:30-2; Tuesdays 10:30-12 in
324 Chem (or 326 Chem).
Key to understanding periodic trends

As \( n \uparrow \), atomic orbitals become & less stable

As \( Z (\# \text{ protons}) \uparrow \), any given orbital becomes & more stable; \( Z_{\text{eff}} \uparrow \)

Trade-off: \( Z \cdots e^- \) attraction
\( e^- \cdots e^- \) repulsion

recall Coulomb’s Law: \( E \propto \frac{Q_1 Q_2}{d} \)

column = “group” or “family”; row = “period”
Electron configuration & the periodic table

• electron configuration determines

1. periodic table organization (Chap 6)

2. properties of the elements
   atomic size
   ionization energy
   electron affinities
   reactivity

• properties of elements determined by $(n)$ & $(\ell)$ of orbitals and
  atomic z (#) (nuclear charge)

• valence shell configuration is key to element’s properties
  ⇒ valence $e^-$ determine the chemistry of an element!

• core electrons: smaller radius, lower E, mostly not participating in chemistry of reactions
Screening, again (see p 3 ff of Lecture 5 notes)

• $Z_{\text{eff}} = Z - S$

• weak screening by $e^-$ with same $n$

$S$ is *average* number of $e^-$ *between* nucleus & $e^-$ of interest; often (but not always) use number of core $e^-$
Ion sizes

What is happening when we make a cation or anion??

• *cations* are **less** than parent atoms; why??

• *anions* are **greater** than parent atoms; why??

• atom sizes **go down** going down family; ion sizes also
Isoelectronic series

- *isoelectronic:* same number of e\(^-\), same e\(^-\) configuration

- example

10 e\(^-\) each: [He]2s\(^2\)2p\(^6\)

\[
\begin{array}{cccccc}
Z \\
& 8 & 9 & 10 & 11 & 12 & 13 \\
O^{2-} & F^- & Ne & Na^+ & Mg^{2+} & Al^{3+} \\
\end{array}
\]

\(Z_{\text{eff}}\) size

For this example, \(Z_{\text{eff}} = Z - 2\) from [He]
Makin’ ions

• to form a **cation**: remove e\(^{-}\) first from orbitals with **highest** n; these are the **valence** e\(^{-}\)

• in **transition metals (TM)**: s electrons are part of valence e\(^{-}\)

\[
\text{Ag} \quad [\text{Kr}] \ 5s^1 4d^{10}
\]

\[
\text{Ag}^+ \quad [\text{Kr}] \ 4d^{10}
\]

• **when forming TM ions:**
  remove e\(^{-}\) first
  then maybe remove e\(^{-}\), if needed

ions with different charges may be formed

\[
\text{Fe} \quad [\text{Ar}] \ 4s^2 3d^6
\]

\[
\text{Fe}^{2+} \quad [\text{Ar}] \ 3d^6
\]

\[
\text{Fe}^{3+} \quad [\text{Ar}] \ 3d^5
\]

• to form an **anion**: add e\(^{-}\) to empty or partially filled orbital with **available** n
Ionization energy (IE)

- energy to remove 1 e\textsuperscript- in the gas phase

\[ \Delta E = I_1 \]

- \( I_1 \): first ionization energy (IE)

\[ M(g) \rightarrow M^+(g) + e^- \quad \Delta E = I_1 \]

- \( I_2 \): second ionization energy

\[ M^+(g) \rightarrow M^{2+}(g) + e^- \quad \Delta E = I_2 \]

- \( I_n > 0 \); therefore energy is **endothermic**
Ionization energy (IE)

- $I_1$ (kJ/mol)

<table>
<thead>
<tr>
<th>Na</th>
<th>Mg</th>
<th>Al</th>
<th>Si</th>
<th>P</th>
<th>S</th>
<th>Cl</th>
<th>Ar</th>
</tr>
</thead>
<tbody>
<tr>
<td>495</td>
<td>738</td>
<td>578</td>
<td>786</td>
<td>1012</td>
<td>1000</td>
<td>1251</td>
<td>1521</td>
</tr>
</tbody>
</table>

- $e^-$ closer to nucleus, more difficult to remove

- as atomic size increases, IE ↑

- therefore, IE ↑ left to right

- exceptions: extra energy to remove $e^-$ from filled subshells (Mg, Ar) or from half-filled subshell (P)
Electron affinity (EA)

• energy to add $1 \ e^-$ in the gas phase

$$M(g) + e^- \rightarrow M^-(g) \quad \Delta E = EA$$

• EA: either endothermic ($\Delta E +$, energetically unfavorable) or exothermic ($\Delta E -$, energetically favorable)

• EA of cation is always exothermic ($\Delta E -$); because opposite process of IE

$$M^+(g) + e^- \rightarrow M(g) \quad EA(M^+) = - I_1(M)$$

• complex property due to trade-off between

$Z^+ \cdot \cdot \cdot e^- \quad \text{attraction}$

$e^- \cdot \cdot \cdot e^- \quad \text{repulsion}$
Electron affinity trends

• most negative (energetically favorable) for halogens (group 7A)  why??

• group 2A (Be, Mg, Ca)
  do not want to fill a new subshell
  have positive values (energetically unfavorable) for EA (unstable negative ions)

• group 1A (Li, Na, K)
  negative ions are not stable but have ns\(^2\) configuration

• noble gases (group 8A)
  have positive values for EA  why??
Review of periodic trends

- **atomic size**

- **ionization energy**

- **electron affinity**: both + and −; group 7A most negative

- **metallic character:**
  - metals lose e\(^{-}\) (oxidation)
  - nonmetals gain e\(^{-}\) (reduction)
The modern periodic table

• elements listed in order of increasing atomic number (Z)

• **elements** in same column (group or family) have similar chemical properties

• **metals** toward left side of PT reactivity ↑ going down group

  • most metallic character, lower left of PT

• **nonmetals** toward right side of PT reactivity ↑ upward

  • most nonmetallic character, upper right of PT

• **noble gases** at far right of PT: quite inert
Metals

• chemistry of metals: form

  group 1A (alkali): 1+ cations
  on atom, valence e\(^-\): \( \text{ns}^1 \)
  most active (Cs, Fr)

  group 2A (alkaline earth): 2+ cations
  on atom, valence e\(^-\): \( \text{ns}^2 \)
  less reactive than group 1A

  transition: 1+, 2+, 3+ cations

• reactivity \( \uparrow \) as IE \( \downarrow \); easier to lose that e\(^-\) as \( n \uparrow \)
Nonmetals

• **group 8A (noble gases)***
on *atom*, valence e\(^{-}\): \(\text{ns}^2\text{np}^6\)
  almost completely inert (except \(\text{XeF}_n\))
  all gases

• **chemistry of nonmetals: form***
  
  **group 7A (halogens):** 1– anions
  on *atom*, valence e\(^{-}\): \(\text{ns}^2\text{np}^5\)
  most reactive nonmetals, particularly \(\text{F}_2\)
  reactions dominated by:
  \[
  \text{X}_2 + 2\text{e}^- \leftarrow 2\text{X}^- 
  \]

  • **reactivity** of halogens ↑ as EA ↑

  **group 6A (oxygen family):** 2– anions
  on *atom*, valence e\(^{-}\): \(\text{ns}^2\text{np}^4\)
Nonmetals (cont.)

• compounds which are entirely nonmetals are (usually) **compounds** (CO$_2$, hydrocarbons)

• seven diatomics (H$_2$, N$_2$, O$_2$, F$_2$, Cl$_2$, Br$_2$, I$_2$) *why??* will see in Chap 8

• compounds: metal + nonmetal are **compounds**

\[
2 \text{Al(s)} + 3 \text{Br}_2(\text{l}) \rightarrow 2 \text{AlBr}_3(\text{s})
\]
Before next class:

Read: BLB 2.6–2.8; 8.1–8.4
HW: BLB 2:51,53,57,59,65,67; 8:16,22,25,29,37,39
    Sup 2:4

Know:
• chemical bonding
  ionic bonding
  covalent bonding
  metallic bonding
• Lewis symbols
• lattice energy