Chem 103, Section F0F
Unit II - Quantum Theory and Atomic Structure
Lecture 9

- The periodic trends for three key properties of the elements
- How the electronic structure of the elements affects their chemical reactivity.

Lecture 9 - Electron Configuration

- Reading in Silberberg
  - Chapter 8, Section 4 Trends in Three Key atomic Properties
  - Chapter 8, Section 5 Atomic Structure and Chemical Reactivity

Lecture 9 - Introduction

In lab we compared the chemical reactivity for elements at different locations on the periodic table.
- We defined chemical reactivity for metals as a willingness to give up electrons.

Lecture 9 - Introduction

In lab we compared the chemical reactivity for elements at different locations on the periodic table.
- We defined chemical reactivity for nonmetals as a willingness to accept electrons.

Lecture 9 - Introduction

Today we will discuss the question: What does the electron configuration of an element tell us about its physical and chemical properties?
- We will focus on trends for three properties:
  - Trends in atomic size
  - Trends in ionization energy
    - The energy required to remove an electron from an atom in the gaseous state.
  - Trends in electron affinity
    - The energy change when an electron is added to an atom in the gaseous state.

Lecture 9 - Trends in Atomic Size

There are different ways to measure size, we will use the following definitions:
- For metals
  - 1/2 the distance between one atom and its neighbor in a metallic crystal
Lecture 9 - Trends in Atomic Size

There are different ways to measure size, we will use the following definitions:
- For nonmetals, which typically form covalent bonds with other nonmetals - 1/2 the distance between one atom and the neighbor it is bonded to.

Trends among the main-group elements:
- As \( n \) (the period) increases, the radius increases.
  - Each shell places the electrons further from the nucleus.
  - There are also more inner electrons, which shield the outer electrons, lowering the \( Z_{	ext{eff}} \).
- As you move across a period, the size decreases.
  - The added electrons are in the same shell, but the \( Z_{	ext{eff}} \) increases, pulling these electrons closer to the nucleus.

The effect of electron repulsions and shielding on orbital energy:
- The electrons feel not only the attraction of the nucleus, but also the repulsion of the other electrons.
- Shielding by inner electrons greatly lowers the effective nuclear charge (\( Z_{	ext{eff}} \)).

The effect of nuclear charge (\( Z \))
- Higher nuclear charge lowers the energy of an energy level.

Comparing the 1s orbital of H and He*
Lecture 9 - Trends in Atomic Size

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• As you move across a period, the size decreases.
  - The added electrons are in the same shell, but the \( Z_{\text{eff}} \) increases, pulling these electrons closer to the nucleus.
• We will focus on the main group elements and not worry about the subtleties of the transition metals (pp318-320).

Lecture 9 - Trends in Atomic Size

Moving through the periodic table, we see the atomic size is periodic.

Lecture 9 - Trends in Ionization Energy

The ionization energy (IE) is the energy required to completely remove 1 electron from an atom in its gaseous state.
• Because energy is entering the atom (system), it is positive.

\[
\text{Atom}_{(g)} \rightarrow \text{Ion}^+_{(g)} + e^- \quad \Delta E = \text{IE}_1 > 0
\]

• The energy for removing a second electron (IE\(_2\)) is always greater than the energy for removing the first electron (IE\(_1\))

\[
\text{Ion}^+_{(g)} \rightarrow \text{Ion}^{2+}_{(g)} + e^- \quad \Delta E = \text{IE}_2 > \text{IE}_1
\]

Lecture 9 - Trends in Ionization Energy

• For the hydrogen atom, the ionization energy (IE) can be calculated by combining Rydberg’s equation with Planck’s equation and finding the energy it takes to move the electron from \( n = 1 \) to \( n = \infty \).

\[
\frac{1}{\lambda} = R_c \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right) \quad \text{(Rydberg’s Equation)}
\]

\[
E = \frac{hc}{\lambda} \quad \text{(Planck’s Equation)}
\]

\[
= \frac{hc}{\lambda} \left( \frac{1}{n_1^2} - \frac{1}{n_2^2} \right)
\]

For 1 mole of atoms

\[
= [2.179 \times 10^{-18} \text{ J} \times 6.022 \times 10^{23} \text{ mol}] = 1.312,000 \text{ J/mol}
\]

\[
E = 1,312 \text{ kJ/mol}
\]
Atoms with low IE’s tend to form cations (lose electrons),
Atoms with high IE’s tend to form anions (gain electrons)
  • Like the atomic sizes, the IE’s vary periodically:

Trends
  • As n increases ionization energies decrease
    - As the atoms get larger, the electrons removed are located further from the nucleus.

Trends
  • From left to right in a period, ionization energies increase.
    - The Z_eff is increasing while the atoms are getting smaller, so the electrons are held more strongly.

Rank the elements Potassium (K), Argon (Ar) and Neon (Ne) in order of increasing IE, based on their position in the periodic table.
A) Ar < Ne < K
B) Ar < K < Ne
C) K < Ar < Ne
D) Ne < Ar < K

Successive ionizations
  • Successive ionizations show increasing ionization energies
    - This is because the resulting ion becomes increasingly positively charged.
  • There is also a substantial jump in IE when dropping down to a lower energy level to remove one of the inner (core) electrons.

Successive ionizations
  • There is also a substantial jump in IE when dropping down to a lower energy level to remove one of the inner (core) electrons.

<table>
<thead>
<tr>
<th>Period</th>
<th>Element</th>
<th>Number of Valence Electrons</th>
<th>Ionization Energy (kcal/mol)*</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Li</td>
<td>1</td>
<td>5.35</td>
</tr>
<tr>
<td></td>
<td>Be</td>
<td>2</td>
<td>8.61</td>
</tr>
<tr>
<td></td>
<td>B</td>
<td>3</td>
<td>9.03</td>
</tr>
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</tr>
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<td></td>
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<td>5</td>
<td>15.72</td>
</tr>
<tr>
<td></td>
<td>O</td>
<td>6</td>
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</tr>
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<td></td>
<td>F</td>
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<td></td>
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<td>18</td>
<td>28.13</td>
</tr>
<tr>
<td></td>
<td>Xe</td>
<td>18</td>
<td>30.39</td>
</tr>
</tbody>
</table>

Table 6.5 Successive Ionization Energies of the Elements Utilized Through Solution

<table>
<thead>
<tr>
<th>Z</th>
<th>Element</th>
<th>Number of Valence Electrons</th>
<th>Ionization Energy (kcal/mol)*</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td>Li</td>
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</tr>
</tbody>
</table>

*Values in parentheses are ions. - IE=Ionization Energy.
Lecture 9 - Clicker Question 2

Element Q is in Period 3 and has the following ionization energies (in kJ/mol):

<table>
<thead>
<tr>
<th>IE1</th>
<th>IE2</th>
<th>IE3</th>
<th>IE4</th>
<th>IE5</th>
<th>IE6</th>
</tr>
</thead>
<tbody>
<tr>
<td>577</td>
<td>1816</td>
<td>2744</td>
<td>11,576</td>
<td>14,829</td>
<td>18,375</td>
</tr>
</tbody>
</table>

Element Q is
A) Al  
B) Be  
C) B  
D) C  
E) N

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Lecture 9 - Electron Affinity

Electron affinity (EA) is the energy that is released or gained when an atom adds an electron.

\[ \text{Atom}_{(g)} + e^- \rightarrow \text{ion}_{(g)} \quad \Delta E = \text{EA} \]

- Usually energy is released when an electron is gained (\(\Delta E < 0\)), but there are some exceptions.
  - Unlike ionization energy, EA can be either positive or negative.

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Lecture 9 - Electron Affinity

The electron affinity for adding a second electron (EA\(_2\)) is always positive because energy must be added to bind a negatively charged electron to an already negatively charged ion.

\[ \text{ion}_{(g)} + e^- \rightarrow \text{ion}^2_{(g)} \quad \Delta E = \text{EA}_2 > 0 \]

The trends for electron affinities are not as easy to predict as those for the ionization energies.

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Lecture 9 - Summarizing Trends

Trends in Ionization Energies (IE) and Electron Affinities (EA):

- **Reactive nonmetals**, in particular the halogens (7A)
  - Have high (positive) ionization energies (IE)
    - This makes it unfavorable for them to lose an electron
  - At the same time, they have highly negative electron affinities (EA)
    - This makes them more likely to gain an electron
  - Together, these properties make it more favorable for these elements to become negatively charged ions (anions).

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Lecture 9 - Electron Affinity

Trends in Ionization Energies (IE) and Electron Affinities (EA):

- **Reactive metals**, in particular alkali and alkaline earth metals (1A and 2A)
  - Have low ionization energies (IE)
    - This makes it easy for them to lose an electron.
  - At the same time they have low electron affinities (EA)
    - This makes it less likely for them to gain an electron.
  - Together, these properties make it more favorable for these elements to become positively charged ions (cations).
Lecture 9 - Electron Affinity

Trends in Ionization Energies (IE) and Electron Affinities (EA):

- **Noble gases**, the elements in group 8A:
  - High positive ionization energies (IE).
  - Have positive electron affinities (EA).
  - Together, these properties make it unfavorable for these elements to either gain or lose electrons.

Lecture 9 - Summarizing Trends

Atomic Structure and Chemical Reactivity

- **Trends in Metallic vs Nonmetallic Behaviors**
  - The further down and to the left on the periodic table, the more metallic an element's properties become.

- You can skip the discussion of the acid-base behavior of the element oxides (pp 326-327)

Lecture 9 - Properties of Monoatomic Ions

Main group elements (1A, 2A, 6A, 7A) lose (1A & 2A) or gain (6A & 7A) electrons to become isoelectronic (same number of electrons) with the nearest noble gas.

<table>
<thead>
<tr>
<th>Metal</th>
<th>Nonmetal</th>
</tr>
</thead>
<tbody>
<tr>
<td>Shiny</td>
<td>Dull</td>
</tr>
<tr>
<td>High melting points</td>
<td>Low melting points</td>
</tr>
<tr>
<td>Electrical conductors</td>
<td>Electrical insulators</td>
</tr>
<tr>
<td>Thermal conductors</td>
<td>Thermal insulators</td>
</tr>
<tr>
<td>Malleable solids</td>
<td>Brittle solids</td>
</tr>
<tr>
<td>Lose electrons in chemical reactions</td>
<td>Gain electrons in chemical reactions</td>
</tr>
</tbody>
</table>
Lecture 9 - Properties of Monoatomic Ions

For larger metals, with d orbitals, only the s and p electrons are lost.
- Giving them pseudo-noble gas configurations.

Lecture 9 - Clicker Question 3

Write the electron configuration for Aluminum (Al). Write the electron configuration for an aluminum ion (Al$^{3+}$).

Which noble gas is Al$^{3+}$ isoelectronic with?
A) He  
B) Ne  
C) Ar  
D) Kr  
E) Xe

Lecture 9 - Properties of Monoatomic Ions

Transition metals have too many electrons to lose become isoelectronic with a noble gas.
- They tend to lose their ns electrons, plus some of their (n-1)d electrons.
  - Typically the ns electrons are lost before the (n-1) electrons.
- They often form ions with different charge states one charge.

Lecture 9 - Properties of Monoatomic Ions

General rules for predicting the electronic configurations of ions:
- For main group, s-block metals (1A & 2A)
  - Remove all electrons with the highest n value.
- For main-group, p-block metals (3A & 4A)
  - Remove np electrons before ns electrons.
- For transition, d-block metals
  - Remove ns electrons before (n-1)d electrons.
- For nonmetals (5A, 6A & 7A)
  - Add electrons to the p orbitals of the highest n value.
Lecture 9 - Properties of Monoatomic Ions

Magnetic properties of ions
• If an ion leaves some of the electrons unpaired, the ion will be attracted by a magnetic field
  - The ion is said to be paramagnetic.
• If an ion has all its electrons paired, the ion is weakly repelled by a magnetic field
  - The ion is said to be diamagnetic.

Lecture 9 - Clicker Question 4
Write a condensed electronic configuration for Cr\(^{3+}\).
Is Cr\(^{3+}\) paramagnetic?
A) Yes
B) No

Lecture 9 - Clicker Question 5
Write a condensed electronic configuration for Cu\(^+\).
Is Cu\(^+\) paramagnetic?
A) Yes
B) No

Lecture 9 - Properties of Monoatomic Ions

Ionic Size vs Atomic Size
• Cations are smaller than their parent atoms.
  - The outer electrons have been removed leaving only the inner (core) electrons.
    ▶ The resulting net positive charge, along with reduced electron repulsion, leads to a contraction in the size.
• Anions are larger than their parent atoms.
  - Electrons are added to the outer level
    ▶ The increased electron repulsion leads to a swelling in the size.

Unit II - Up Next
Unit III - Combining Atoms to Make Compounds, Part I: Ionic Compounds
• Atomic properties and chemical bonds
• The ionic bonding model.
The End