Ch 7: Atomic Electron Configurations & Chemical Periodicity

1. Electron Spin & Pauli Exclusion Principle
2. Polyelectronic Atoms
3. Electron Configurations
4. Ionic Electron Configurations
5. Periodic Trends & Chemical Properties

7.1 Pauli Exclusion Principle
Two Electrons Can Occupy an Orbital According to the **Pauli Exclusion Principle**

Note: Electrons Have Intrinsic Spin
-A Spinning Charged Particle Creates a Magnetic Field, and Just Like Two Magnets, Will Repel Unless North-South Pole Coupled, **Two Electrons Must Have Opposite Spins to Be in the Same Orbit**

$m_s$ – fourth quantum number

Summary of Quantum numbers
-Electrons can be described by orbitals which can in turn be described by 4 quantum numbers
-No two electrons can have the same set of quantum numbers

$n$ – Principle Q# - describes “shell”
$l$ – angular momentum Q# - describes shape
$m_l$ – magnetic Q# - describes orientation
$m_s$ – spin Q# - describes intrinsic spin

Periodic Table & Quantum Numbers

Orbital Energies for Multiple Electron Atoms

- The 4s orbital “penetrates” the 3d orbital and feels a greater effective nuclear charge ($Z^*$)
7.3 Electron Configurations

General Rules

1. Electrons Fill Lowest Energy Orbitals First (Aufbau Principle)
2. For Orbitals of Equal Energy, Fill Each Orbital With One $e^-$ Before Placing 2 in an Orbital (Hunds Rule)
3. Each Orbital Can Only Have Two Electrons of Opposite Spin (Pauli Exclusion Principle)
4. Determine # of Electrons in Atom (or Ion), and Fill the Orbitals According to the Above Rules

Electron Configuration of Hydrogen

Number of electrons = 1

Electron Configuration of Lithium

Number of electrons = 3

Electron Configuration of Carbon

Number of electrons = 6

Electron Configuration of Sodium (11 electrons)

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

Inner Shells can be abbreviated by Nobel gas they represent (Ne = 1s$^2$2s$^2$2p$^6$)

Alternative (preferred) Notation use Nobel gas symbol to represent Core Electrons

Na: [Ne]3s$^1$

Write the Electron Configuration of Vanadium and Chromium
Order of Orbital Energies

Cr: [Ar]4s\(^1\)3d\(^5\)

Filled and 1/2 Filled Sublevels Are Exceptionally Stable, This Is the Reason Behind the Stability of Noble Gasses and the Previous Electron Configuration of Chromium.

7.4 Electron Configuration of Ions

1. Anions – Add electrons to Lowest Energy Unoccupied Orbital
   - Tend to form ions which are isoelectronic to nearest noble gas
   
   S\(^2\) : [Ne]3s\(^2\)3p\(^6\)
   P\(^3\) : [Ne]3s\(^2\)3p\(^6\)

Electron Configuration of Ions

1. Cations – Remove electrons the orbital which is furthest from the nucleus (note how the 4s is both closer and further from the nucleus than the 3d

V: [Ar]4s\(^2\)3d\(^3\)

What is V\(^+2\)?

Ionic Radii

Anions Are Larger Than Parent Atom

Ionic Radii

Cations Are Smaller Than Parent Atom

7.5 Periodic Trends in Atomic Properties

1. Ionization Energy - ease of loosing an electron
2. Electron Affinity – ease of gaining an electron
3. Atomic radii – “size of atom”
Atomic Radii

How are atomic radii determined?

1/2 the Internuclear Bond Length for a Homonuclear Bond

\[
\text{Atomic Radii of Carbon} = .77 \ \text{Å}
\]

Atomic Size

What 2 Factors Influence Atomic Size?

1. **Principle quantum #**: as \( n \) increases, the orbitals increase in size

2. **Effective Nuclear Charge**: Core electrons shield valence electrons from the nucleus thus reducing the positive charge the valence electron feels

Why Do Atoms Get Larger Going Down a Group?

Why do Atoms Get Smaller Going Across a Period?
First Ionization Energy
- the amount of energy required to remove an electron from a gas phase atom (forming a cation)

\[ \text{M(g)} \rightarrow \text{M}^+(\text{g}) + e^- \]

First Ionization Energy
Why does the ionization energy tend to increase going across a period?
The effective nuclear charge increases, making it harder to remove an electron

First Ionization Energy
Why does the ionization energy decrease going Be to B and Mg to Al
The 2s or 3s orbitals are respectively filled and have a partial shielding effect on the 2p\(^1\) & 3p\(^1\) electrons of B & Al
(Decreasing the effective nuclear charge)

First Ionization Energy
Why Does the Ionization Energy Tend to Decrease in Going From N to O & P to S?
The spin pairing energy associated with 2 electrons in a single orbital of the ns\(^2\)np\(^4\) electron configuration is repulsive and weakens the nuclear attraction

First Ionization Energy
Why Does the Ionization Energy Decrease Going Down a Group?
The valence electron tends to be further from the nucleus, making it easier to be removed

Higher Ionization Energies
Note: You can move more than one electron from an atom

First Ionization Energy
\[ \text{M(g)} \rightarrow \text{M}^+(\text{g}) + e^- \]
Second Ionization Energy
\[ \text{M}^+(\text{g}) \rightarrow \text{M}^{2+}(\text{g}) + e^- \]
Which is Larger?
Higher Ionization Energies

First Ionization Energy

\[ \text{Li}(g) \rightarrow \text{Li}^+(g) + e^- \quad I_1 = 513.3 \text{ kJ/mol} \]

Second Ionization Energy

\[ \text{Li}^+(g) \rightarrow \text{Li}^{+2}(g) + e^- \quad I_2 = 7298 \text{ kJ/mol} \]

Electron Affinities

Energy must be added for the reaction to proceed

\[ \text{Ar}([\text{Ne}]:3s^23p^6) + e^- \rightarrow \text{Ar}([\text{Ne}]:3s^23p^64s^1) \]

What Does a Negative Electron Affinity Mean?

Periodic Trends: Electron Affinity

Electron Affinity Tends to Decrease Going Across a Period With Some Notable Exceptions (becoming more exothermic)

Is It Easier to Ionize an Element With a High or a Low Ionization Energy?

Note, Electron Affinities Can Be Both Endothermic and Exothermic.

What Does a Positive Electron Affinity Mean?
Periodic Trends: Electron Affinity

**Why is group IIA > IA?**

Group IIA has a filled s orbital

**Periodic Trends: Electron Affinity**

**Why is group VA > IVA?**

Group VA has a 1/2 filled p shell

**Periodic Trends: Electron Affinity**

**Why is group VIIIA Positive?**

The e⁻ is being added to the next shell which is shielded from the nucleus

**Noble Gases**

Which Group Tends to Have the Highest Ionization Energy?

Which Group Tends to Have the Highest Electron Affinity?

**Metals**

Tend to Have Low Ionization Energies
- Easily Oxidized (lose electrons)
- Form Cations

Charge of cations can be determined by successive ionization energies (table 7.2)

**Nonmetals**

- Tend to have low (negative) electron affinities
- Easily Reduced (gain electrons)
- Tend to Form anions

Note: Noble gases are an exception