Focus

- Electronic Configuration:
  - Describe electron spin quantum number and Pauli’s exclusion principle
  - Rules of electronic configuration
  - Characteristic electronic configuration of groups and its impact on group properties

Arrangement of Electrons in Atoms

Electrons in atoms are arranged as

SHELLS (n) → SUBSHELLS (l) → ORBITALS (m)
Arrangement of Electrons in Atoms

Each orbital can be assigned only ______ !

Because of the magnetic properties of the atom, a fourth quantum number, the ____________ is introduced.

Electron Spin Quantum Number, \( m_s \)

Two spin directions are given by \( m_s \) where \( m_s = + \frac{1}{2} \) and \( -\frac{1}{2} \).

Electron Spin Quantum Number

Definition of some magnetic terminologies:

Substances which are NOT attracted to a magnetic field are called __________________ substances
substances have completely paired electrons

Substances which are attracted to a magnetic field are called ____________: Substance substances.

__________ substances have unpaired electrons.

Substances which are strongly attracted to magnetic field are said to be ________________________
Electron Spin Quantum Number

Many alloys (e.g. Alnico, Nd-Fe-B) exhibit greater ferromagnetism than pure metals.

Electron Spin Quantum Numbers

- \( n \rightarrow \) shell \( 1, 2, 3, 4, ... \)
- \( l \rightarrow \) subshell \( 0, 1, 2, ... n - 1 \)
- \( m_l \rightarrow \) orbital \( -l, 0, +l \)
- \( m_s \rightarrow \) electron spin \( + \frac{1}{2}, - \frac{1}{2} \)

Pauli Exclusion Principle

i.e. even if the 2 electrons have the same \( n, l, \) and \( m_s \), they will have different \( m_s \) (\( - \frac{1}{2} \) or \( + \frac{1}{2} \) ) values.

That is, each electron in an atom has a unique address.
Electrons in Atoms

How many electrons can each orbital, $m_l$ (e.g. s, or $2p_x$ or $3d_{xy}$) hold?

Electrons in Atoms

We know that each orbital in a subshell can hold a maximum of 2 electrons, how many electrons can each subshell, $l$, hold?

# of e⁻ in a subshell, $l$ =

Subshell, $l$ | Orbital | # of orbitals | # of e⁻ in subshell
--- | --- | --- | ---
0 | s | 1 | 2
1 | p | 3 | 6
2 | d | 5 | 10
3 | f | 7 | 14
4 | g | 9 | 18

Electrons in Atoms

What maximum number of electrons can a shell accommodate?

Shell | Subshell | # of orbitals | # of e⁻ |
--- | --- | --- | ---
$n$ | $l$ | $m_l$ | $\sum$ |
1 | | | Total $\#$ e =
### Electrons in Atoms

How many electrons can a shell accommodate? What about each subshell?

<table>
<thead>
<tr>
<th>Shell</th>
<th>Subshell</th>
<th>Orbitals</th>
<th># of e</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[
\begin{array}{c|c|c|c}
 n & l & m_l & \text{# of } \text{e} \\
\hline
2 & & & \\
\end{array}
\]

Total # e =

---

<table>
<thead>
<tr>
<th>Shell</th>
<th>Subshell</th>
<th>Orbitals</th>
<th># of e</th>
</tr>
</thead>
<tbody>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[
\begin{array}{c|c|c|c}
 n & l & m_l & \text{# of } \text{e} \\
\hline
3 & & & \\
\end{array}
\]

Total # e =

---

<table>
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<tr>
<th>Shell</th>
<th>Subshell</th>
<th>Orbitals</th>
<th># of e</th>
</tr>
</thead>
<tbody>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

\[
\begin{array}{c|c|c|c}
 n & l & m_l & \text{# of } \text{e} \\
\hline
4 & & & \\
\end{array}
\]

Total # e =
Assigning Electrons to Atoms

The arrangement of electrons in an atom is called _______________________________

In an atom, electrons are generally assigned to orbitals of successively higher energy.

Assigning Electrons to Atoms

For hydrogen atom, the energy of the electron depends on n only
In H atom all subshells of same n have same energy, i.e.

$2s = 2p$, $3s = 3p = 3d$ and $4s = 4p = 4d = 4f$

Order of increasing energy levels in H atom is:

$n = 1 < n = 2 < n = 3 < n = 4$……………

So

$1s < 2s = 2p < 3s = 3p = 3d < 4s = 4p = 4d = 4f$

Assigning Electrons to Subshells

For atoms that have multiple electrons, energy depends in both n and l

subshells increase in energy as value of $n + l$ increases.

for subshells of same $n + l$, subshell with lower n is lower in energy.
Assigning Electrons to Atoms

Since orbitals are filled in order of increasing \( n + l \) value, 4s orbital should be filled before 3d.

Electron Filling Order

Effective Nuclear Charge

The reason for difference in energy for 2s and 2p subshells, for example, is effective nuclear charge, \( Z^* \).
Effective Nuclear Charge, Z* 

Z* is the nuclear charge experienced by the outermost electrons.

Z* increases across a period owing to incomplete shielding by inner electrons.

To estimate Z* use the expression:

\[ Z* = (Z - \text{no. inner electrons}) \]

• The effective nuclear charge, Z*, felt by the 2s e- in Li:

\[ Z* = 3 - 2 = 1 \]

Effective Nuclear Charge, Z*

Estimate the effective nuclear charge felt by the 2s electrons in Be

Be \[ Z* = 4 - 2 = 2 \]

Z* increases across a period owing to shielding by inner electrons.

Writing Atomic Electron Configurations

Two ways of writing configs. One is called the spectroscopic notation.

SPECTROSCOPIC NOTATION for H, atomic number = 1

1s

no. of electrons

value of n

value of l
Writing Atomic Electron Configurations

Two ways of writing configs. Other is called the orbital box notation.

<table>
<thead>
<tr>
<th>ORBITAL BOX NOTATION for He, atomic number = 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>1s²</td>
</tr>
</tbody>
</table>

One electron has n = 1, l = 0, m_l = 0, m_s = + 1/2
Other electron has n = 1, l = 0, m_l = 0, m_s = - 1/2

Lithium

Group 1A
Atomic number = 3
Total # of e⁻ = 3

Beryllium

Group 2A
Atomic number = 4
4 total electrons
Boron
Group 3A
Atomic number = 5
Total # of e⁻ = 5

Carbon
Group 4A
Atomic number = 6
Total # of e⁻ = 6

Here we see for the first time HUND’S RULE. When filling orbitals of similar energies (e.g. 2pₓ, 2pᵧ, and 2pᵦ), fill them singly first, before pairing up.

Nitrogen
Group 5A
Atomic number = 7
Total # of e⁻ = 7
Oxygen
Group 6A
Atomic number = 8
Total # of e⁻ = 8

Fluorine
Group 7A
Atomic number = 9
Total # of e⁻ = 8

Neon
Group 8A
Atomic number = 10
Total # of e⁻ = 10

Note that we have reached the end of the 2nd period, and the 2nd shell is full!
**Sodium**

Group 1A  
Atomic number = 11  
Total # of e⁻ = 11  
1s² 2s² 2p⁶ 3s¹ or  
“neon core” + 3s¹  
[Ne] 3s¹ (uses noble gas notation)

Note that we have begun a new period.

All Group 1A elements have [core]ns¹ configurations where n is the period number.

---

**Magnesium**

Group 2A  
Atomic Number = 12  
1s² 2s² 2p⁶ 3s²

Noble gas notation: ______________

All Group 2A elements have __________ configurations where n is the period number.

---

**Aluminum**

Group 3A  
Atomic number = 13  
1s² 2s² 2p⁶ 3s² 3p¹

Noble gas notation: ______________

All Group 3A elements have __________ configurations where n is the period number.
Phosphorus

Group 5A
Atomic number = 15
1s² 2s² 2p⁶ 3s² 3p³

Noble gas notation: __________________

All Group 5A elements have ______________ configurations where n is the period number.

Calcium

Group 2A
Atomic number = 20
1s² 2s² 2p⁶ 3s² 3p⁶ 4s²

Noble gas notation: __________________

All Group 2A elements have [core]ns² configurations where n is the period number.

Relationship of Electron Configuration and Region of the Periodic Table

- Yellow = s block  ns¹ & ns²
- Green = p block  ns² npₓ  (x = 1-6 or grp# - 2)
- Violet = d block  ns¹(n-1)d’y  (y = 1-10)
- Purple = f block  ns¹(n-1)f’z  (z = 1-14)

See page 299, fig 8.7
**Transition Metals**

*Table 8.4*

All 4th period elements have the configuration \([\text{argon}] \, ns^x \, (n - 1)d^y\) and so are “d-block” elements.

- Chromium
- Iron
- Copper

**Transition Element Configurations**

- 3d orbitals used for Sc - Zn (Table 8.4)

**Lanthanides and Actinides**

All these elements have the configuration \([\text{core}] \, ns^x \, (n - 1)d^y \, (n - 2)f^z\) and so are “f-block” elements.

- Cerium
  - \([\text{Xe}] \, 6s^2 \, 5d^1 \, 4f^1\)
- Uranium
  - \([\text{Rn}] \, 7s^2 \, 6d^1 \, 5f^3\)

---
To form cations from elements remove 1 or more e- from subshell of highest n [or highest (n + l)].

\[ \text{P} [\text{Ne}] 3s^2 \ 3p^3 \rightarrow \text{P}^{3+} [\text{Ne}] 3s^2 \ 3p^0 \]
For transition metals, remove ns electrons and then (n - 1)d electrons.

Fe [Ar] 4s² 3d⁶
loses 2 electrons \( \rightarrow \) Fe²⁺ [Ar] 4s⁰ 3d⁶

Atoms and ions with unpaired electrons are paramagnetic.

Atoms or ions with paired electrons are diamagnetic.