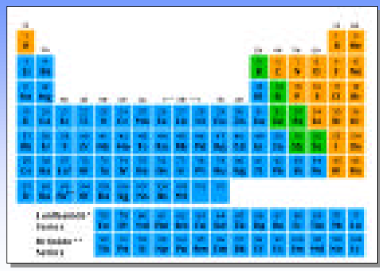
ATOMIC ELECTRON CONFIGURATIONS AND PERIODICITY

1



Arrangement of Electrons in Atoms

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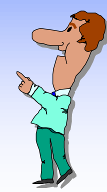
Electrons in atoms are arranged as

SHELLS (n)

SUBSHELLS (l)

ORBITALS (m_l)



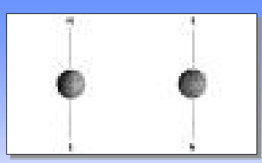
Arrangement of Electrons in Atoms

3

Each orbital can be assigned no more than 2 electrons!

This is tied to the existence of a 4th quantum number, the **electron spin quantum number, m_s.**

4



Electron Spin Quantum Number, m_s

Can be proved experimentally that electron has a spin. Two spin directions are given by m_s where m_s = +1/2 and -1/2.

Electron Spin Quantum Number

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Diamagnetic: NOT attracted to a magnetic field

Paramagnetic: substance is attracted to a magnetic field. Substance has **unpaired electrons**.

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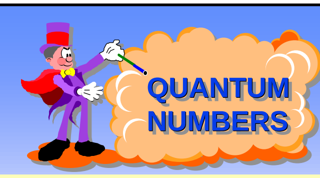

n --->	shell	1, 2, 3, 4, ...
l --->	subshell	0, 1, 2, ... n - 1
m_l --->	orbital	-l ... 0 ... +l
m_s --->	electron spin	+1/2 and -1/2

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Pauli Exclusion Principle

No two electrons in the same atom can have the same set of 4 quantum numbers.

That is, each electron has a unique address.

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Electrons in Atoms

When n = 1, then l = 0

this shell has a single orbital (1s) to which 2e- can be assigned.

When n = 2, then l = 0, 1

2s orbital	2e-
three 2p orbitals	6e-
TOTAL =	8e-

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Electrons in Atoms

When n = 3, then l = 0, 1, 2

3s orbital	2e-
three 3p orbitals	6e-
five 3d orbitals	10e-
TOTAL =	18e-

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Electrons in Atoms

When n = 4, then l = 0, 1, 2, 3

4s orbital	2e-
three 4p orbitals	6e-
five 4d orbitals	10e-
seven 4f orbitals	14e-
TOTAL =	32e-

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Assigning Electrons to Atoms

- Electrons generally assigned to orbitals of successively higher energy.
- For H atoms, E = - C(1/n^2). E depends only on n.
- For many-electron atoms, energy depends on both n and l.
- See Figure 8.6, page 342 and Screen 8. 5.

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Assigning Electrons to Subshells

- In H atom all subshells of same n have same energy.
- In many-electron atom:

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

b) for subshells of same n + l, subshell with lower n is lower in energy.

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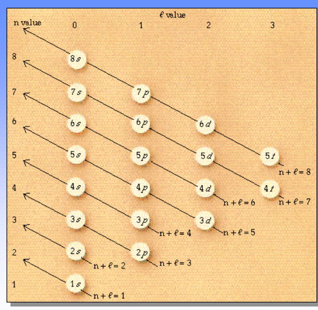

Electron Filling Order

Figure 8.7

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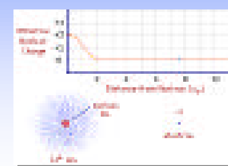
Effective Nuclear Charge

The reason for difference in energy for 2s and 2p subshells, for example, is effective nuclear charge, Z*.



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Effective Nuclear Charge, Z*

- **Z* is the nuclear charge experienced by the outermost electrons.** See p. 344 and Screen 8.6.
- Explains why E(2s) < E(2p)
- Z* increases across a period owing to incomplete shielding by inner electrons.
- Estimate Z* by --> [Z - (no. inner electrons)]
- Charge felt by 2s e- in Li Z* = 3 - 2 = 1
- Be Z* = 4 - 2 = 2
- B Z* = 5 - 2 = 3 and so on!

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Writing Atomic Electron Configurations

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

SPECTROSCOPIC NOTATION for H, atomic number = 1

$1s^1$

1 — value of n; s — value of l; superscript 1 — no. of electrons

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Writing Atomic Electron Configurations

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

ORBITAL BOX NOTATION for He, atomic number = 2

$1s^2$ [↓↑] 1s — Arrows depict electron spin

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

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

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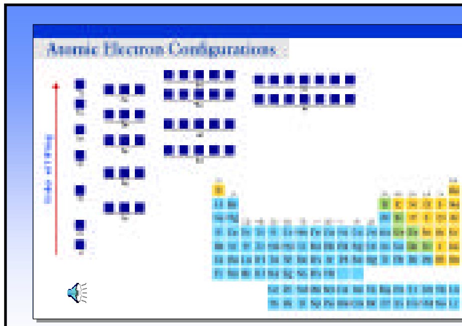
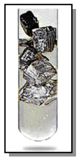


See "Toolbox" for Electron Configuration tool.

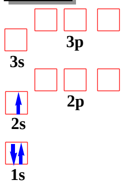
Lithium



Group 1A

Atomic number = 3

$1s^2 2s^1$ ---> 3 total electrons

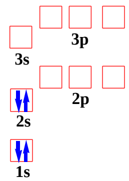


Beryllium

Group 2A

Atomic number = 4

$1s^2 2s^2$ ---> 4 total electrons

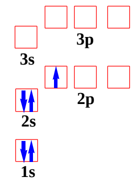


Boron

Group 3A

Atomic number = 5

$1s^2 2s^2 2p^1$ ---> 5 total electrons

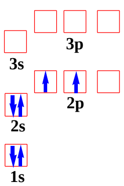


Carbon

Group 4A

Atomic number = 6

$1s^2 2s^2 2p^2$ ---> 6 total electrons



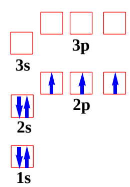
Here we see for the first time **HUND'S RULE**. When placing electrons in a set of orbitals having the same energy, we place them singly as long as possible.

Nitrogen

Group 5A

Atomic number = 7

$1s^2 2s^2 2p^3$ ---> 7 total electrons

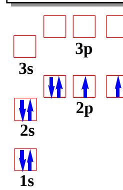


Oxygen

Group 6A

Atomic number = 8

$1s^2 2s^2 2p^4$ ---> 8 total electrons



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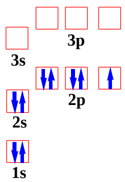
Fluorine

Group 7A

Atomic number = 9

$1s^2 2s^2 2p^5$ --->

9 total electrons



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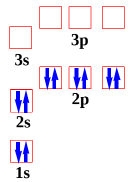
Neon

Group 8A

Atomic number = 10

$1s^2 2s^2 2p^6$ --->

10 total electrons



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

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Electron Configurations of p-Block Elements

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Sodium

Group 1A

Atomic number = 11

$1s^2 2s^2 2p^6 3s^1$ or

"neon core" + $3s^1$

[Ne] $3s^1$ (uses rare gas notation)

Note that we have begun a new period.

All Group 1A elements have [core]ns^1 configurations.

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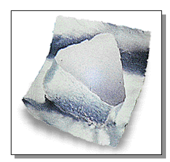
Aluminum

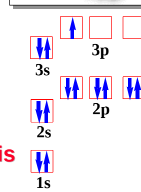


Group 3A

Atomic number = 13

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

[Ne] $3s^2 3p^1$



All Group 3A elements have

[core] $ns^2 np^1$ configurations where n is the period number.

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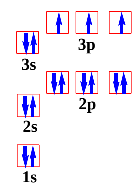
Phosphorus

Group 5A

Atomic number = 15

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

[Ne] $3s^2 3p^3$



All Group 5A elements have

[core] $ns^2 np^3$ configurations where n is the period number.

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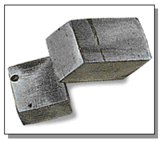
Calcium



Group 2A

Atomic number = 20

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

$[Ar] 4s^2$

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

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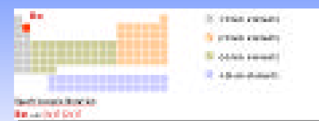
Relationship of Electron Configuration and Region of the Periodic Table



- Gray = s block
- Orange = p block
- Green = d block
- Violet = f block

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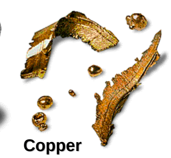
Transition Metals Table 8.4

All 4th period elements have the configuration [argon] $ns^x (n - 1)d^y$ and so are "d-block" elements.


Chromium


Iron


Copper

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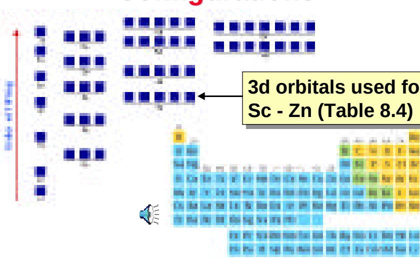
Transition Element Configurations



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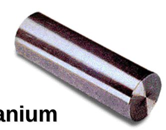
Lanthanides and Actinides

All these elements have the configuration [core] $ns^x (n - 1)d^y (n - 2)f^z$ and so are "f-block" elements.


Cerium
$[Xe] 6s^2 5d^1 4f^1$


Uranium
$[Rn] 7s^2 6d^1 5f^3$

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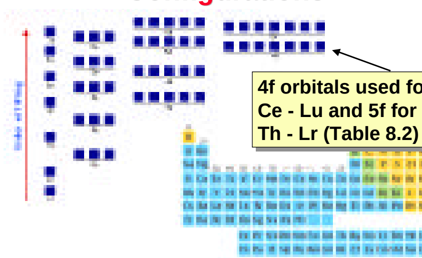
Lanthanide Element Configurations



Ion Configurations

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

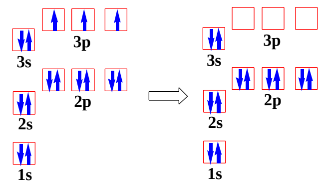
P [Ne] $3s^2$ $3p^3$ - 3e- ---> P^{3+} [Ne] $3s^2$ $3p^0$

Ion Configurations

38

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

P [Ne] $3s^2$ $3p^3$ - 3e- ---> P^{3+} [Ne] $3s^2$ $3p^0$



Ion Configurations

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For transition metals, remove ns electrons and then (n - 1) electrons.

Fe [Ar] $4s^2$ $3d^6$

loses 2 electrons ---> Fe^{2+} [Ar] $4s^0$ $3d^6$

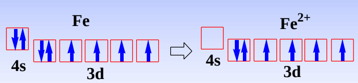
Ion Configurations

40

For transition metals, remove ns electrons and then (n - 1) electrons.

Fe [Ar] $4s^2$ $3d^6$

loses 2 electrons ---> Fe^{2+} [Ar] $4s^0$ $3d^6$



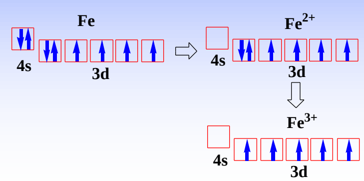
Ion Configurations

41

For transition metals, remove ns electrons and then (n - 1) electrons.

Fe [Ar] $4s^2$ $3d^6$

loses 2 electrons ---> Fe^{2+} [Ar] $4s^0$ $3d^6$



Ion Configurations

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How do we know the configurations of ions?

Determine the magnetic properties of ions.

Ions with **UNPAIRED ELECTRONS** are **PARAMAGNETIC**.

Without unpaired electrons **DIAMAGNETIC**.