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Energy level diagrams

These diagrams show the relative energies of electrons in various orbitals under normal conditions.

Three rules must be used when placing electrons in various sublevels.

1) **Aufbau Principle**: Electrons are always placed in the lowest energy level available

2) **Hund's Rule**: When filing orbitals within any given sublevel, each orbital within the sublevel must be half-filled before it can pair up.

3) **Pauli Exclusion Principle**: Orbitals can be empty, contain 1 electron, but are filled when 2 electrons are present. If two electrons occupy the same orbit they mist have opposite spins. *This means that no two electrons have the same quantum numbers.

Guidelines for filling Orbitals:

1) Place electrons into the orbitals in order of increasing energy levels.
2) Completely fill orbitals of the same energy level before preceding to the next orbital.
3) When electrons are added to the same sublevel, each orbital receives an electron before pairing.
4) When electrons are added individually to different orbitals of the same energy, the must have the same spin.

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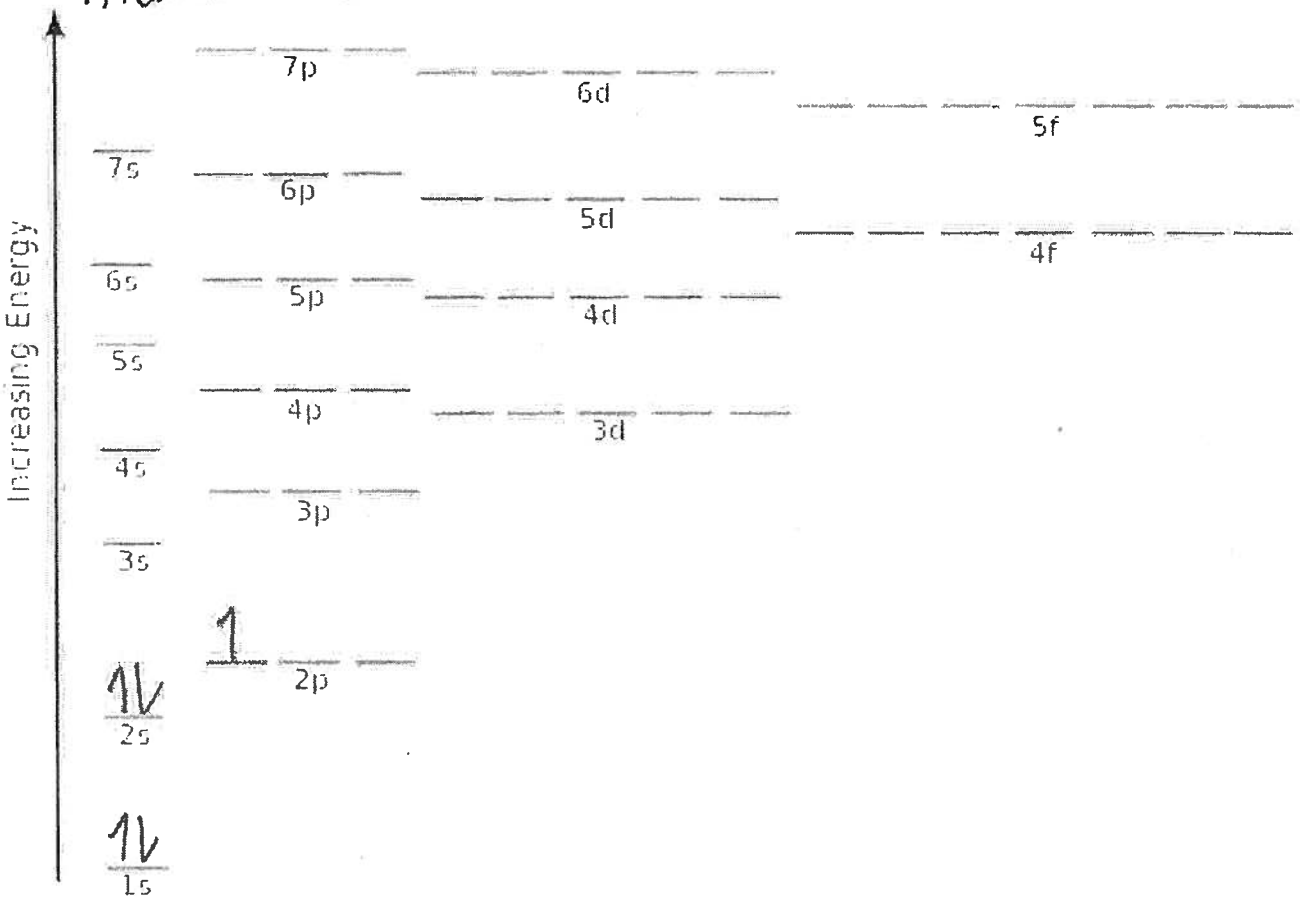
Example: Fill in the energy level diagram for Boron

Atomic # = 5



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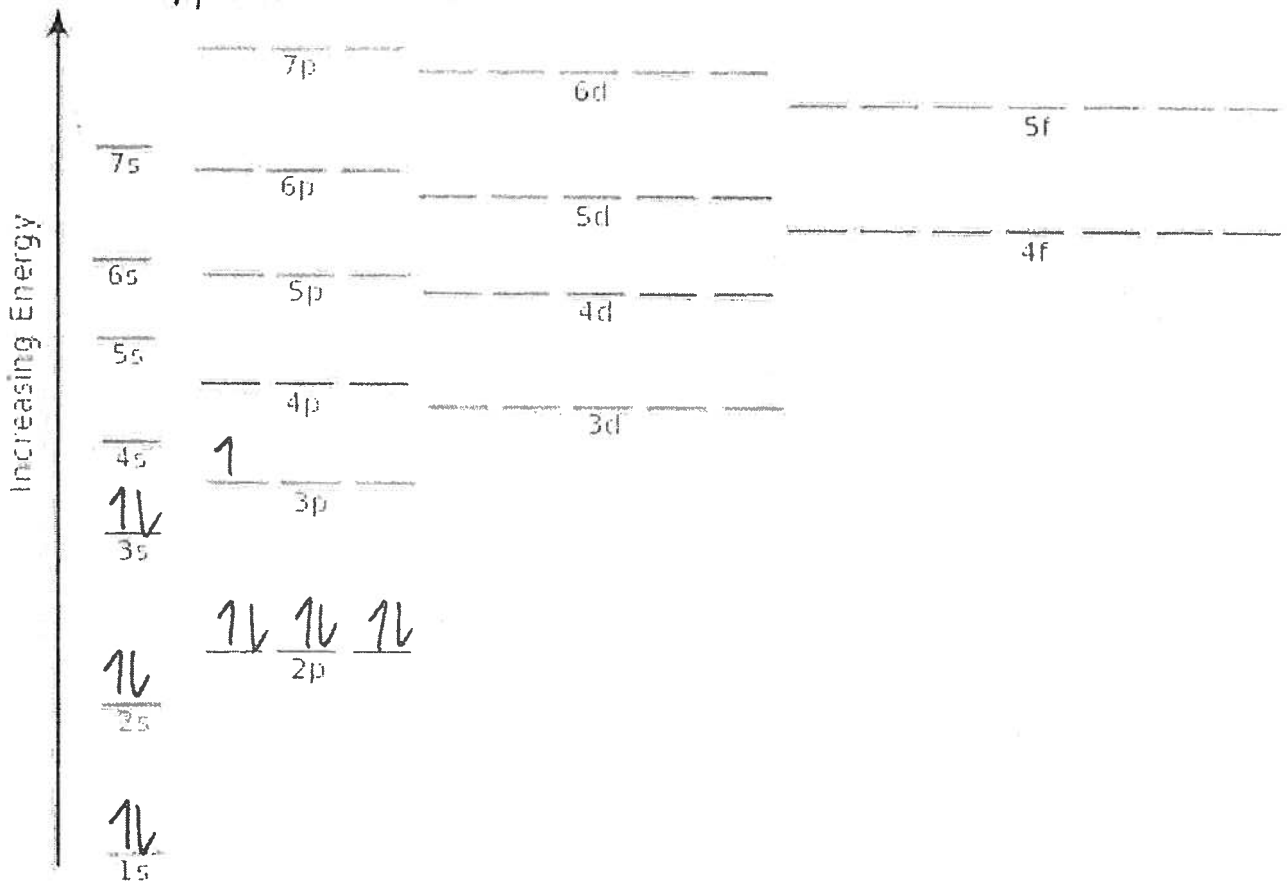
Example: Fill in the energy level diagram for Aluminum

Atomic # = 13



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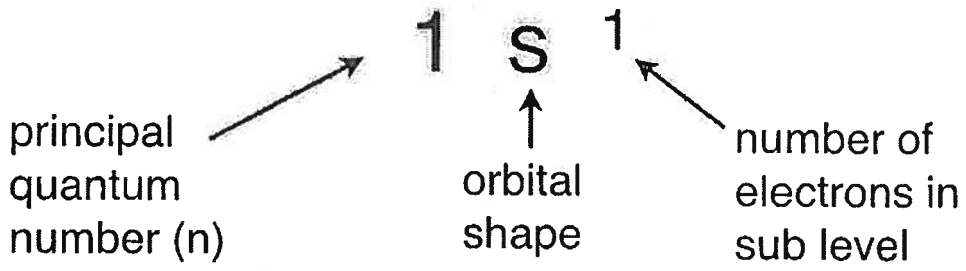
Electron Configuration

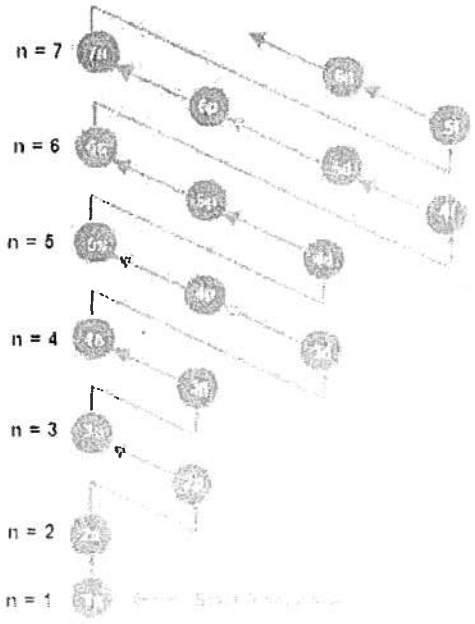
Electron configurations provide the same information as energy level diagrams, but in a more condensed format.

Guidelines writing electron configurations:



The order of filling puts the electrons as close to the nucleus as possible.

To remember the correct order of filling you can use:



Or you can use this friendly shortcut:

$7s^2$	$7p^6$		
$6s^2$	$6p^6$	$6d^{10}$	
$5s^2$	$5p^6$	$5d^{10}$	$5f^{14}$
$4s^2$	$4p^6$	$4d^{10}$	$4f^{14}$
$3s^2$	$3p^6$	$3d^{10}$	
$2s^2$	$2p^6$		
$1s^2$			

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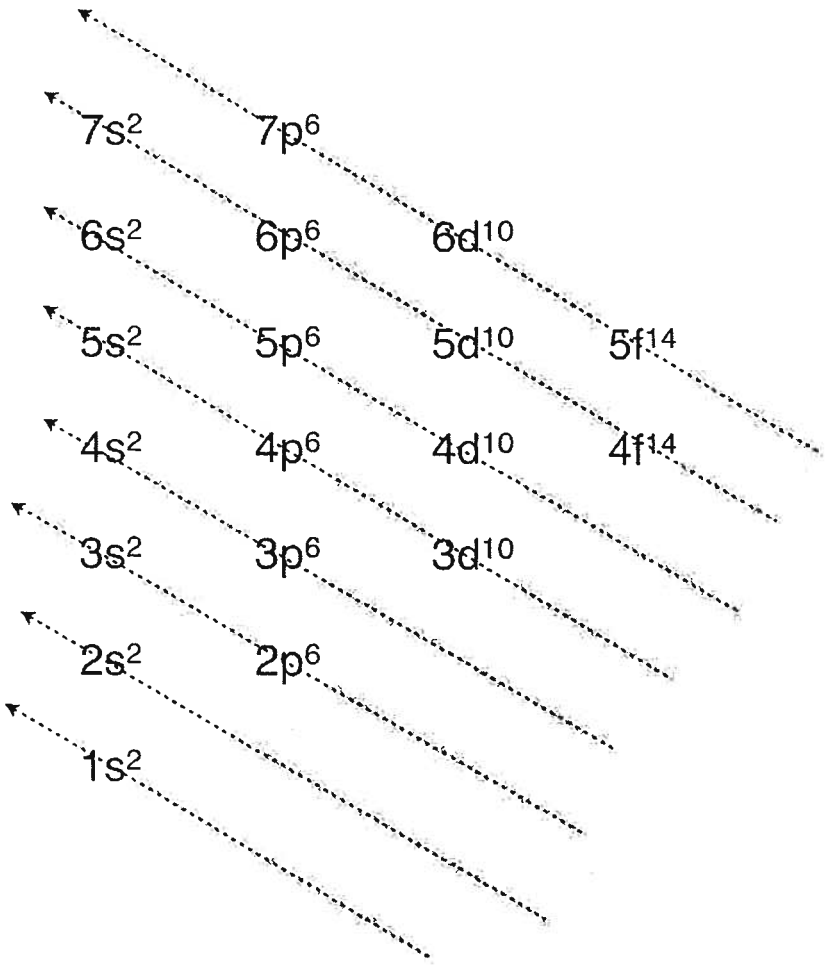
Or you can use this friendly shortcut:



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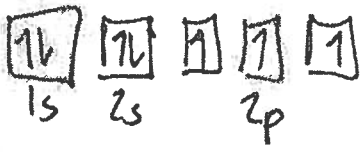
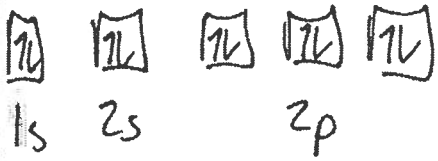
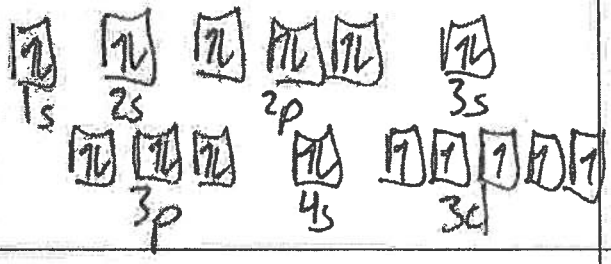
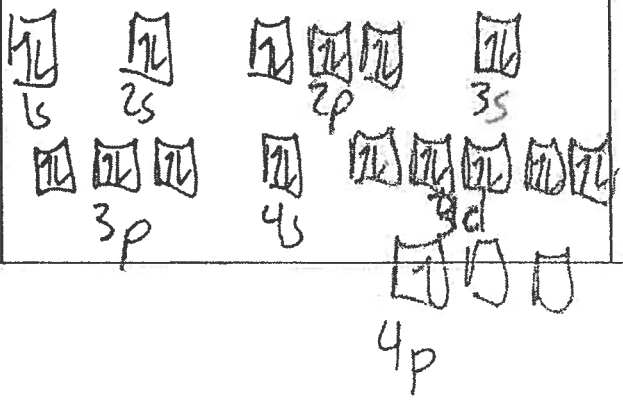
Examples

Element	Atomic Number	Orbital Diagram	Electron Configuration
N	7		$1s^2 2s^2 2p^3$
Ne	10		$1s^2 2s^2 2p^6$
Mn	25		$1s^2 2s^2 2p^6 3s^2 3p^6$ $4s^2 3d^5$
Ga	31		$1s^2 2s^2 2p^6 3s^2 3p^6$ $4s^2 3d^{10} 4p^1$

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Condensed Format

Electron configurations can also be written using a "Condensed Format". The non-valence electrons (electrons in full shells) are represented by the symbol of the nearest preceding Noble Gas, enclosed in square brackets, followed by the electron configurations for the valence electrons in the outer energy level.

	Standard Format	Condensed Format
H	$1s^1$	$1s^1$
N	$1s^2\ 2s^2\ 2p^3$	[He] $2s^2\ 2p^3$
S	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	[Ne] $3s^2\ 3p^4$
Ba	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2$	[Xe] $6s^2$

Exceptions:

A slight revision of the general rules for writing the elements chromium and copper are needed.

Experimental evidence suggests half filled and filled sublevels are more stable then empty sublevels. This is more important for d orbitals compared to s orbitals.

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Chromium

Predicted	Experimentally Determined
[Ar] $4s^2 3d^4$	[Ar] $4s^1 3d^5$
	This electron arrangement allows all sublevels to be half filled providing greater stability

Copper

Predicted	Experimentally Determined
[Ar] $4s^2 3d^9$	[Ar] $4s^1 3d^{10}$
	All d orbitals are filled, providing greater stability.