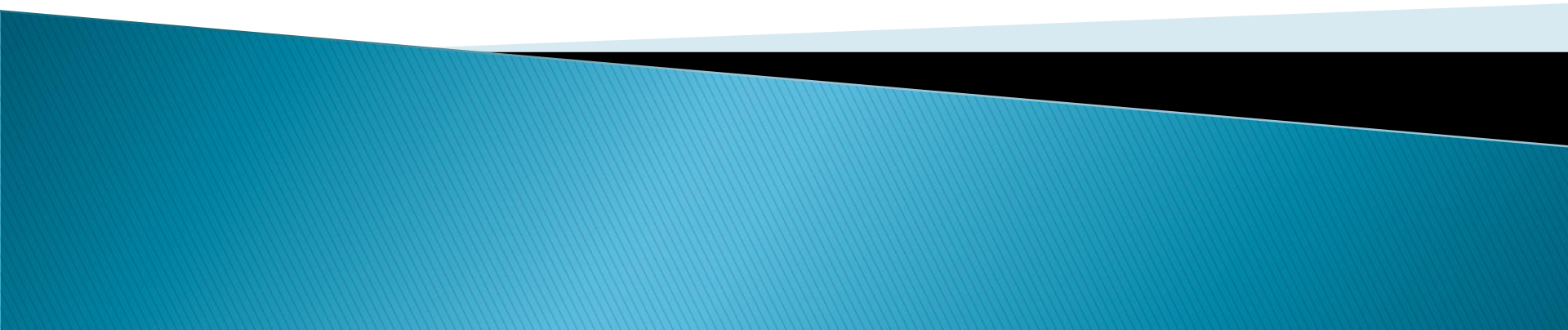# Electron Configurations
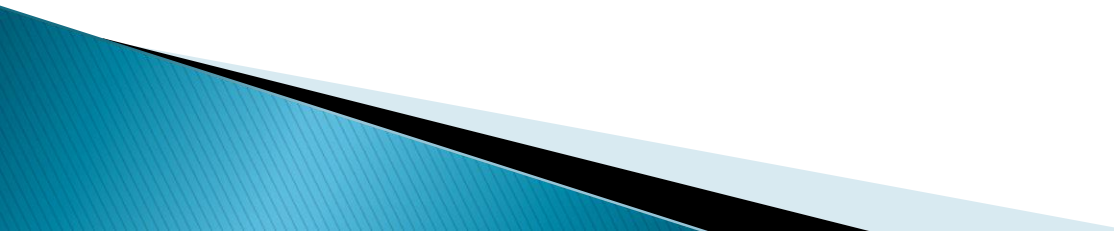
Using Subshell Notation and Orbital Diagrams

How are they placed?

- All electrons in an atom are placed into shells according to electron energies.

- These shells are labeled with the integers, starting from one up to infinity, with one being the lowest energy.

The Ground State

▸ While there are an infinite number of electron configurations for any atom, a method of describing an atom or an ion is based on the lowest energy state called the ground state (any other configuration with electron(s) in higher energy states are called *excited states*).

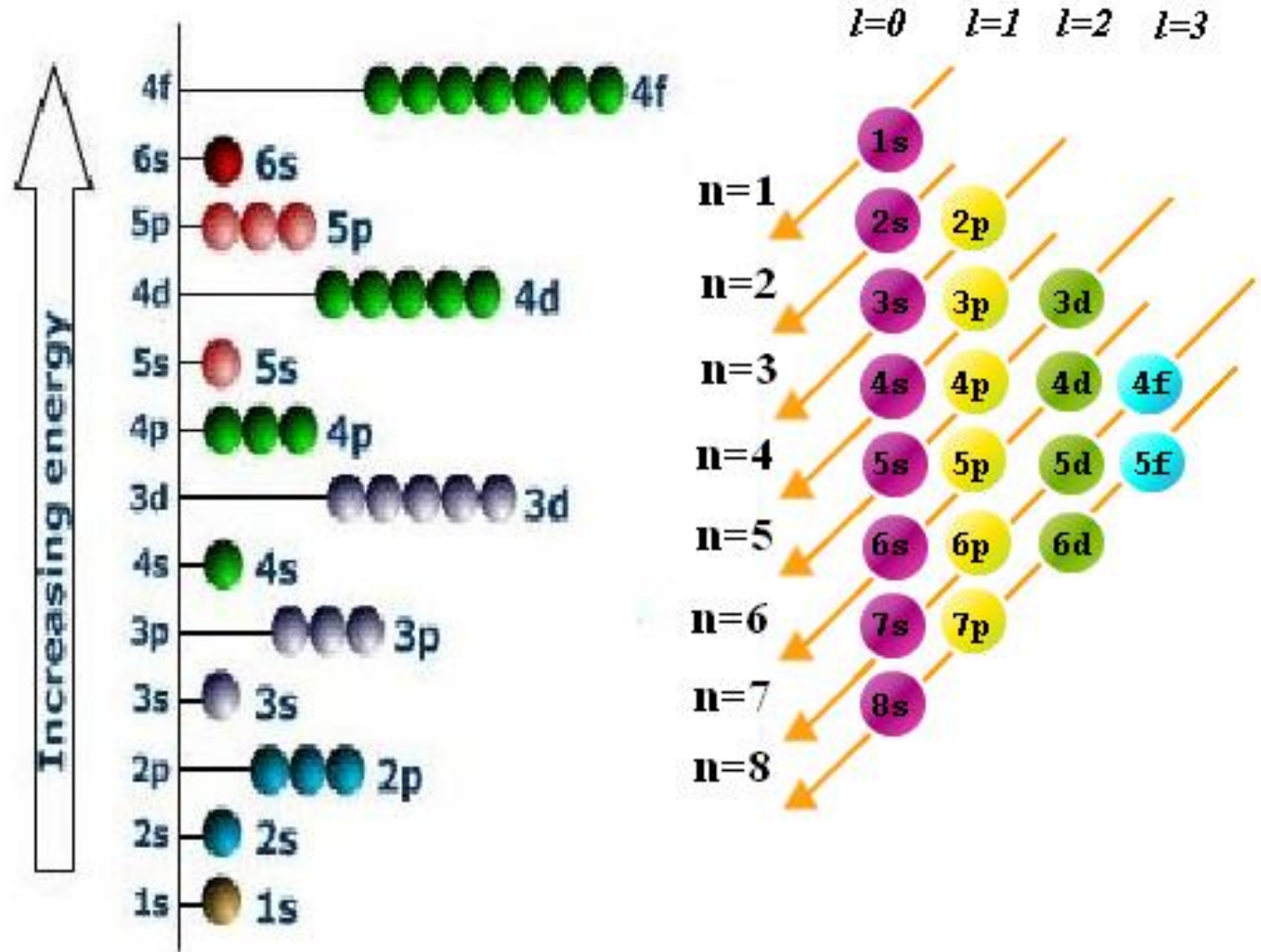▸ The ground state is determined by always placing each electron into the lowest energy subshell.

The order

▸ The order for filling subshells is (sometimes called the Aufbau filling order):

1s 2s 2p 3s 3p 4s 3d 4p 5s 4d 5p 6s 4f 5d 6p 7s 5f 6d 7p

Aufbau Principle–Energy Levels

Electrons are placed in the lowest energy level first.

How Many Can Each Shell Hold?

- Each subshell has a maximum number of electrons that it can hold.
- The s-type subshells can hold a maximum of two electrons
- the p-type subshells can hold a maximum of six electrons
- the d-type subshells can hold a maximum of ten electrons
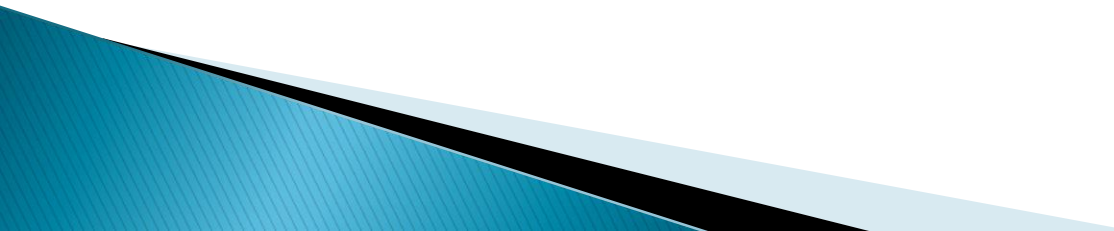- the f-type subshells can hold a maximum of fourteen electrons.

How to write…

▸ To write an electron configuration using the subshell notation, a combination of the subshell followed by a superscript indicating the number of electrons in that subshell is used.

▸ Thus, for the first two elements, we would write their electron configurations as:

   H: $1s^1$ and He: $1s^2$

How to write...

▸ For the next element, Li, we can't put a third electron into the 1s subshell because it is full. Thus, we would need to got to the next available subshell – the 2s.

▸ Li: $1s^2 2s^1$

How to write…

- The filling of subshells would continue to build upon the previous element and fill subshells completely before going on to the next subshell. You can see that, however, this would get to be quite a chore when we reach larger elements like lead (Pb).

- Thus, it is preferential to use a shorthand method that utilizes the configuration of the noble gases (group 18).

Electron Configuration

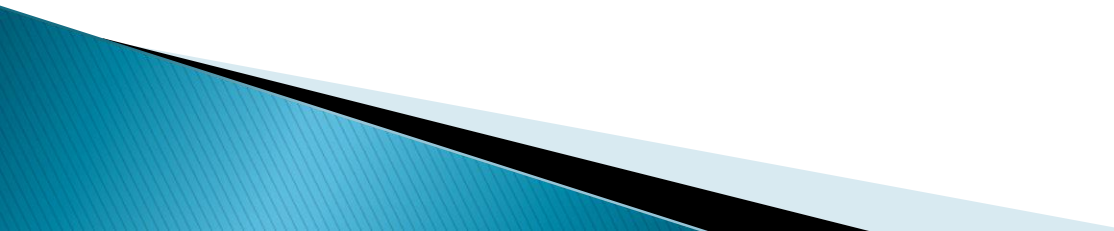Shows the arrangement of electrons in an atom by energy level.

Number of
electrons
in the orbital

$$1s^1$$

Value of n
(principal
energy level)

Type (shape) of
orbital
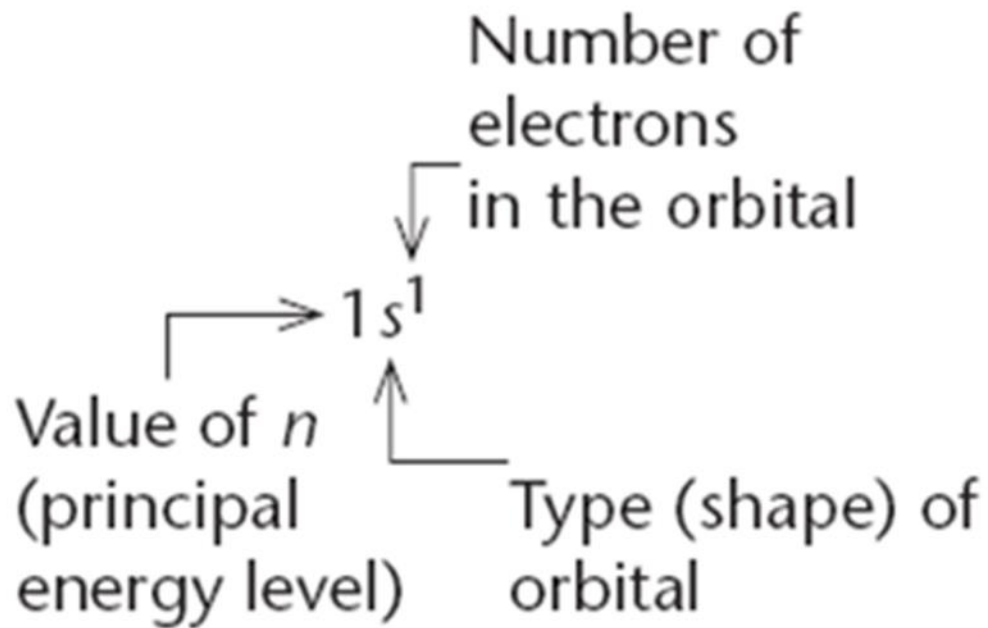
Shorthand…

▸ To do a shorthand configuration for any element, count backwards from that element until you reach a noble gas.

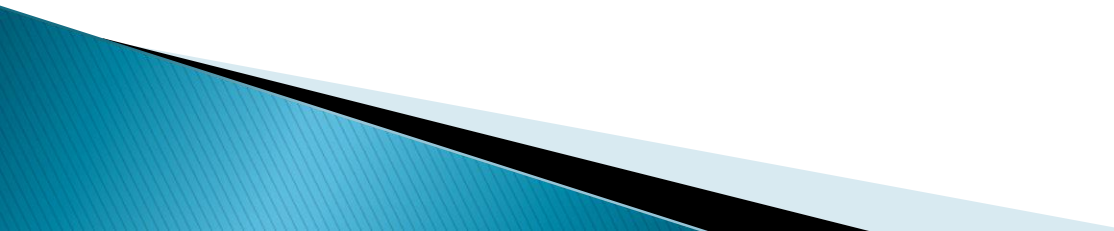▸ Write that element in brackets. Then, continue forward with next subshell(s)

Shorthand…

▸ For example, if we wanted to do the shorthand configuration for sodium (Na), you would count back one element to neon (Ne).

▸ Put this element symbol in brackets and then, noting that the next correct subshell is 3s, include the rest of the electrons as we did with the smaller elements.

▸ Na: [Ne] $3s^1$

Why do we do this?
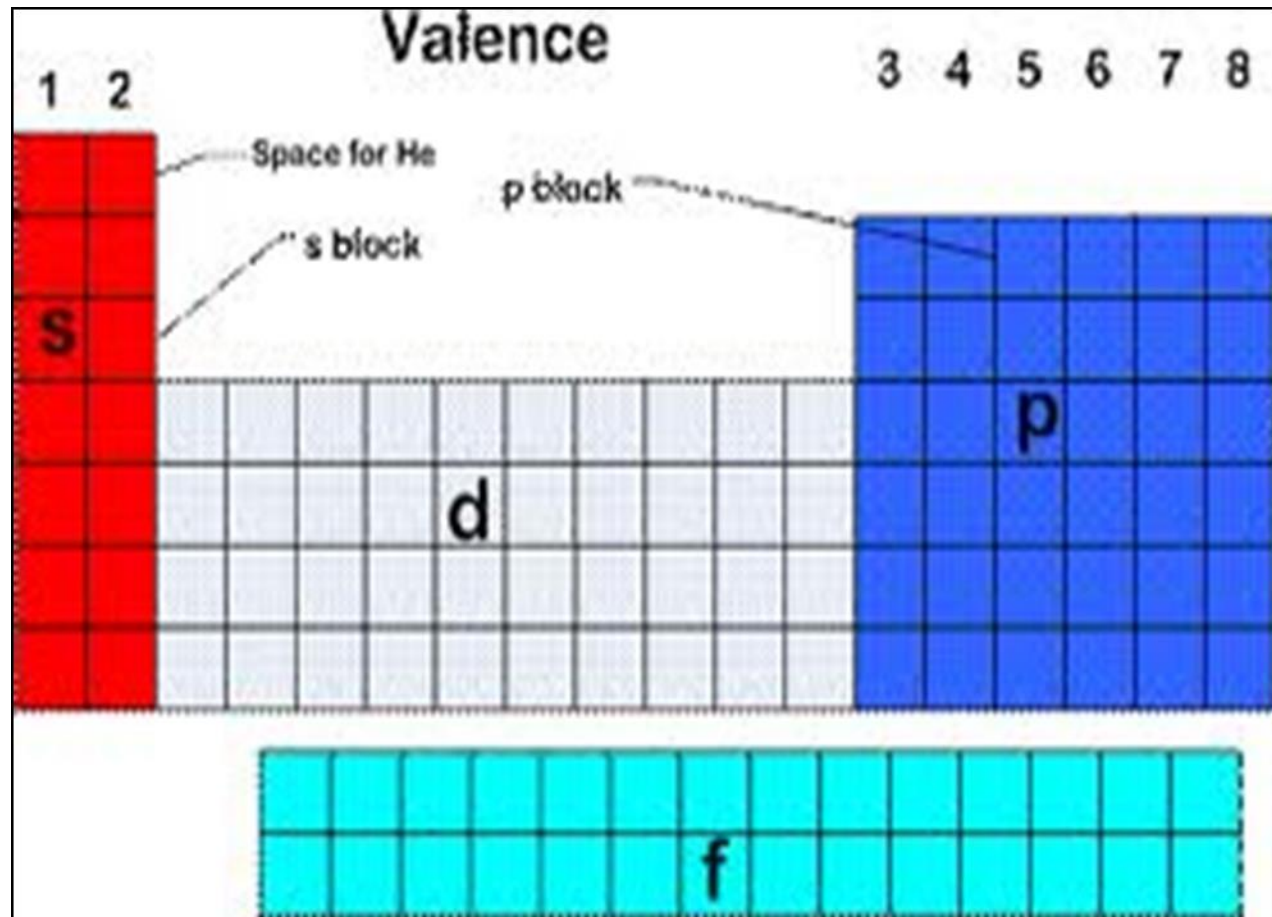
- One of the reasons for doing this is also to distinguish between core and valence electrons.

- Core electrons are held very tightly by the atom and do not interact when bonds are formed to make compounds. These include the noble gas configuration plus any completely filled d-subshells.

- The valence, or outer shell, electrons are the ones that interact with each other when bonds formed are therefore very important.
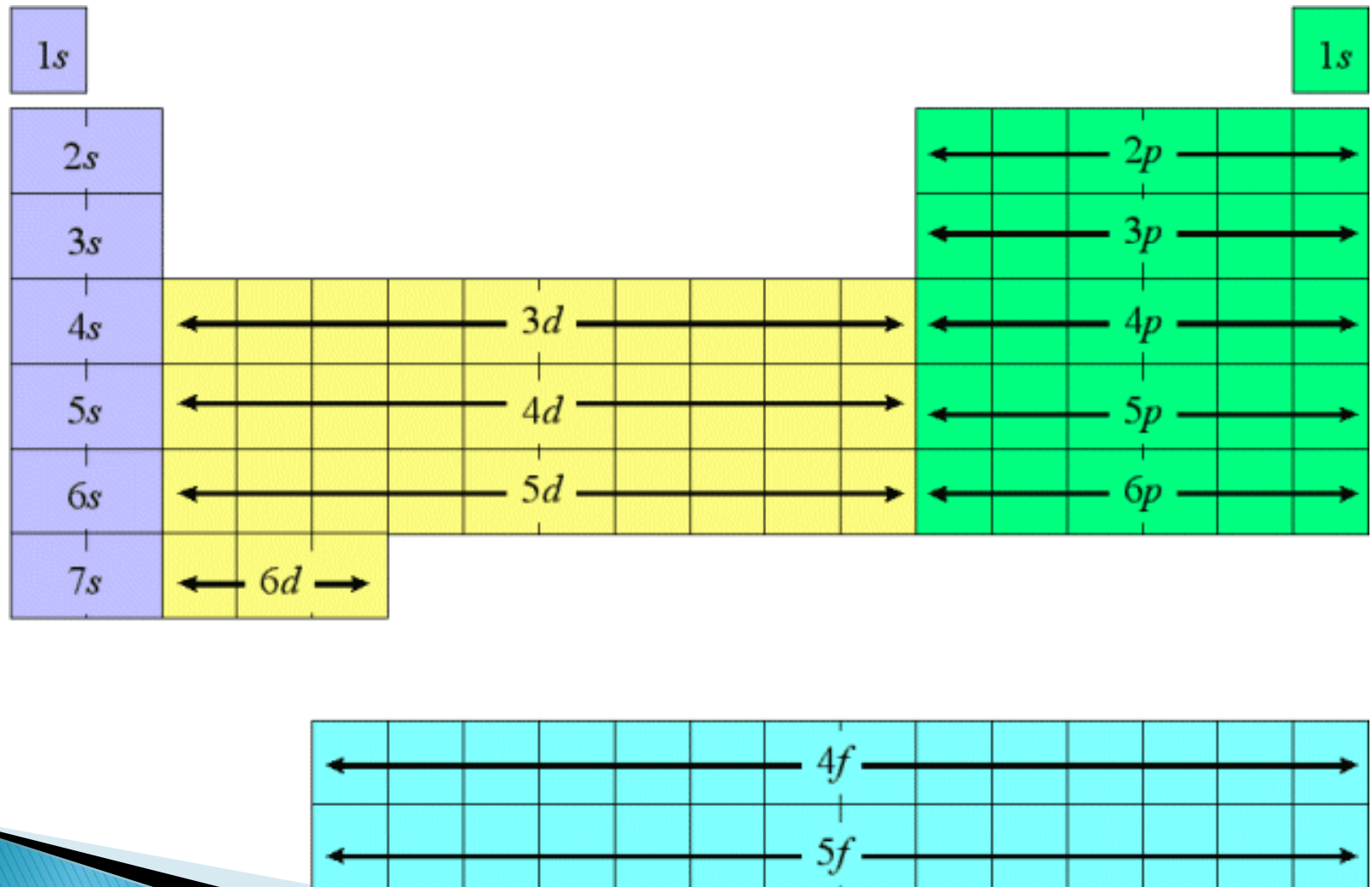
Why do we do this?

- These are always the outermost shell of electrons plus any unfilled d-subshell.
- Note that for any main group element, the number of valence electrons always equals the group number!

- For sodium, it has ten core electrons (neon has ten electrons) and one valence electron (it is in group 1A).

Valence Electrons

- The periodic table organizes valence electrons.
- The number of valence electrons are written above each column in the diagram.

Configuration Chart

Detailed Configuration Chart

$1s^1$																	$1s^2$
$2s^1$	$2s^2$											$2p^1$	$2p^2$	$2p^3$	$2p^4$	$2p^5$	$2p^6$
$3s^1$	$3s^2$											$3p^1$	$3p^2$	$3p^3$	$3p^4$	$3p^5$	$3p^6$
$4s^1$	$4s^2$	$3d^1$	$3d^2$	$3d^3$	$3d^5$	$3d^5$	$3d^6$	$3d^7$	$3d^8$	$3d^{10}$	$3d^{10}$	$4p^1$	$4p^2$	$4p^3$	$4p^4$	$4p^5$	$4p^6$
$5s^1$	$5s^2$	$4d^1$	$4d^2$	$4d^4$	$4d^5$	$4d^5$	$4d^7$	$4d^8$	$4d^{10}$	$4d^{10}$	$4d^{10}$	$5p^1$	$5p^2$	$5p^3$	$5p^4$	$5p^5$	$5p^6$
$6s^1$	$6s^2$		$5d^2$	$5d^3$	$5d^4$	$5d^5$	$5d^6$	$5d^7$	$5d^9$	$5d^{10}$	$5d^{10}$	$6p^1$	$6p^2$	$6p^3$	$6p^4$	$6p^5$	$6p^6$
$7s^1$	$7s^2$		$6d^2$	$6d^3$	$6d^4$	$6d^5$	$6d^6$	$6d^7$	$6d^8$	$6d^{10}$	$6d^{10}$	$7p^1$	$7p^2$	$7p^3$	$7p^4$	$7p^5$	$7p^6$

$5d^1$	$4f^1$	$4f^3$	$4f^4$	$4f^5$	$4f^6$	$4f^7$	$4f^7$	$4f^9$	$4f^{10}$	$4f^{11}$	$4f^{12}$	$4f^{13}$	$4f^{14}$	$4f^{14}$
$6d^1$	$6d^2$	$5f^2$	$5f^3$	$5f^4$	$5f^6$	$5f^7$	$5f^7$	$5f^9$	$5f^{10}$	$5f^{11}$	$5f^{12}$	$5f^{13}$	$5f^{14}$	$5f^{14}$

Orbital Diagram

▸ Another way of writing the electron configuration.

▸ An orbital is a potential space for an electron.

▸ Atoms can have many potential orbitals.

▸ Orbitals are represented by boxes grouped by sublevel with small arrows indicating the electrons.

1s

↑

Orbital diagram

1s

↑↓
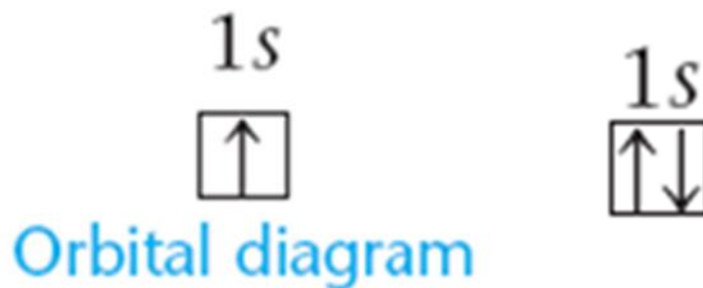
Orbital Diagram Notation

- When writing longhand or shorthand notation each p orbital, d orbital and f orbital gets and individual variable

- So, for 2p or instance the first 2p orbital would be 2px, the second 2py and the last 2pz. A full 2p would look like this:

     $2px^2$ $2py^2$ $2px^2$

This is used instead of the boxes and arrows to show the correct distribution of the Aufbau Principle and Hund's Rule.

Pauli Exclusion Principle

▸ An atomic orbital can hold a maximum of 2 electrons and those 2 electrons must have opposite spins.

▸ An electron is represented by an arrow.

▸ Spin is represented by the arrow facing up or down.

1s

Hydrogen and Helium

- **Electron Configuration:**
  - Hydrogen can only fill the first Principal Energy Level labeled "1".
  - Hydrogen can only fill the first orbital labeled "s".
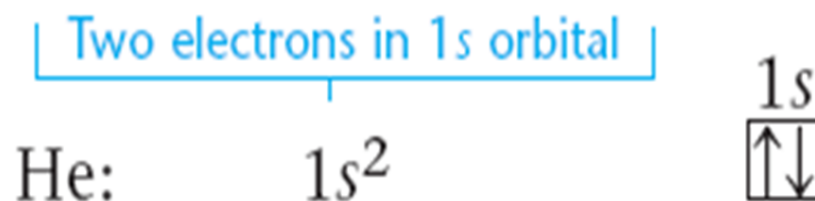- **Orbital Diagram**
  - Hydrogen has an atomic number of 1, so it has 1 electron available to place in the orbital diagram.

- **Electron Configuration:**
  - Helium can only fill the first Principal Energy Level labeled "1".
  - Helium can only fill the first sublevel labeled "s".
- **Orbital Diagram**
  - Helium has an atomic number of 2, so it has 2 electrons available to place in the orbital diagram.

$$1s$$
$$\uparrow$$

H: $1s^1$

Configuration   Orbital diagram

Two electrons in 1s orbital
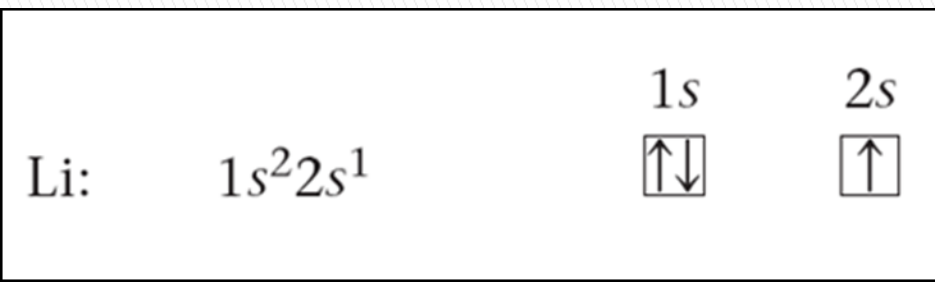
He: $1s^2$

$$1s$$
$$\uparrow\downarrow$$

Lithium and Beryllium

▶ Electron Configuration:
- Lithium can fill the first Principal Energy Level labeled "1" and the second PEL labeled "2".
- Lithium can only fill the first sublevel labeled "s".

▶ Orbital Diagram
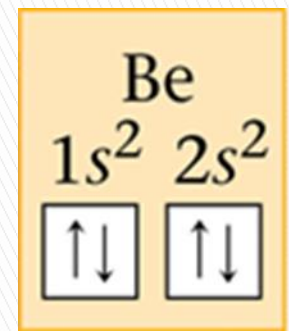- Lithium has an atomic number of 3, so it has 3 electrons available to place in the orbital diagram.

▶ Electron Configuration:
- Beryllium can fill the first Principal Energy Level labeled "1" and the second PEL labeled "2".
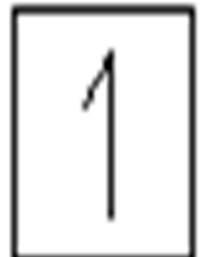- Beryllium can only fill the first sublevel labeled "s".

▶ Orbital Diagram
- Beryllium has an atomic number of 4, so it has 4 electrons available to place in the orbital diagram.

$$Li: \quad 1s^2 2s^1 \qquad \begin{array}{cc} 1s & 2s \\ \uparrow\downarrow & \uparrow \end{array}$$

Be: $1s^2 2s^2$

$$Be \\ 1s^2 \ 2s^2 \\ \uparrow\downarrow \quad \uparrow\downarrow$$

Hund's Rule

▸ When filling sublevels other than s, electrons are placed in individual orbitals first, before they are paired up.

▸ They must be placed singly before doubly.
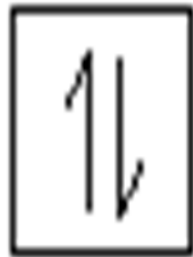
Carbon

- Electron Configuration:
  - Carbon can fill the first Principal Energy Level labeled "1" and the second PEL labeled "2".
  - Carbon can fill the first sublevel labeled "s" and the second sublevel labeled "p".
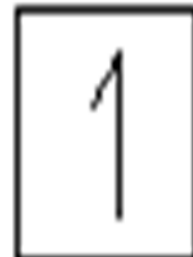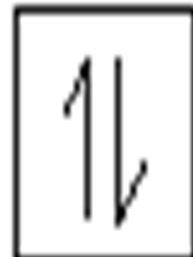- Orbital Diagram
  - Carbon has an atomic number of 6, so it has 6 electrons available to place in the orbital diagram.

C: $1s^2 2s^2 2p^2$

1s   2s   2p

Nitrogen

- Electron Configuration:
  - Nitrogen can fill the first Principal Energy Level labeled "1" and the second PEL labeled "2".
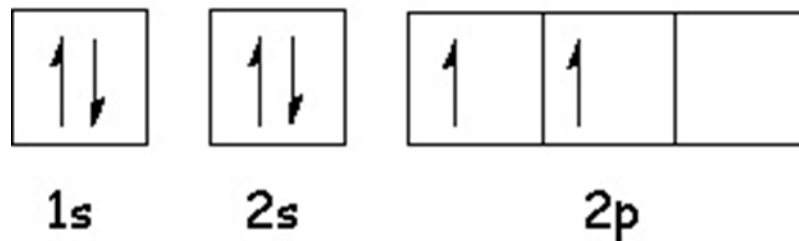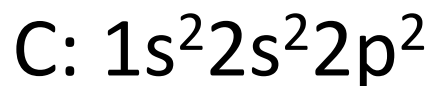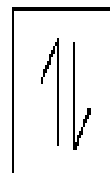 - Nitrogen can fill the first sublevel labeled "s" and the second sublevel labeled "p".
- Orbital Diagram
  - Nitrogen has an atomic number of 7, so it has 7 electrons available to place in the orbital diagram.

N: $1s^2 2s^2 2p^3$

Oxygen

- Electron Configuration:
  - Oxygen can fill the first Principal Energy Level labeled "1" and the second PEL labeled "2".
  - Oxygen can fill the first sublevel labeled "s" and the second sublevel labeled "p".
- Orbital Diagram
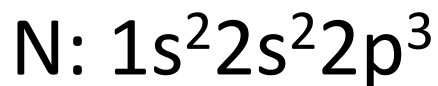 - Oxygen has an atomic number of 8, so it has 8 electrons available to place in the orbital diagram.

O: $1s^2 2s^2 2p^4$
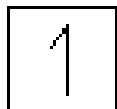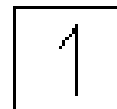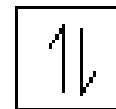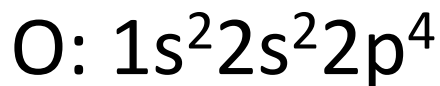
Noble Gas Configuration

- Is important because it shows the valence electrons present in an atom.

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

Noble-gas
abbreviated notation

↑↓   ↑↓   ↑ ↑ ↑

1s   2s   2p

Orbital diagram
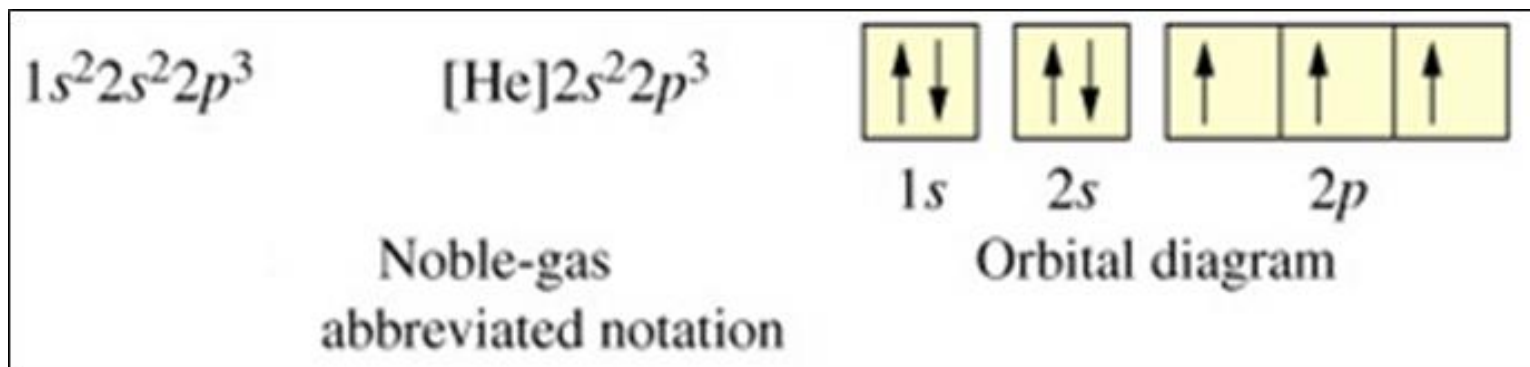
- Nitrogen has an atomic number of 7. It has 7 total electrons. If you look at the electron configuration, you can count 7 electrons.

Noble Gas Configuration

▸ But if you look at the Noble Gas Configuration, you can count 5 electrons.

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

$1s$	$2s$	$2p$		
↑↓	↑↓	↑	↑	↑

Noble-gas abbreviated notation   Orbital diagram
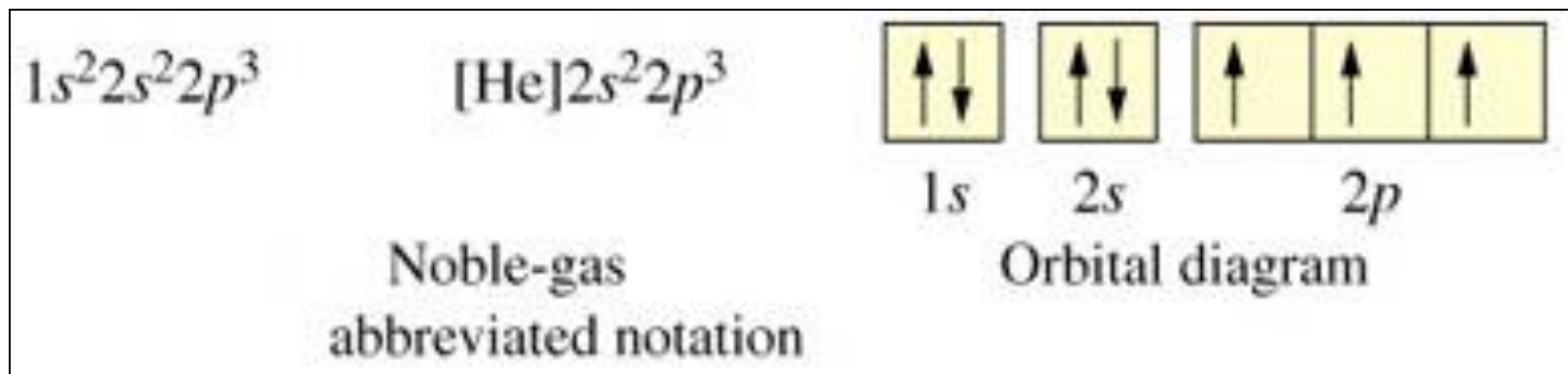
▸ These 5 electrons are the **valence electrons**, the electrons found in the outermost energy level. These are the electrons available for bonding.
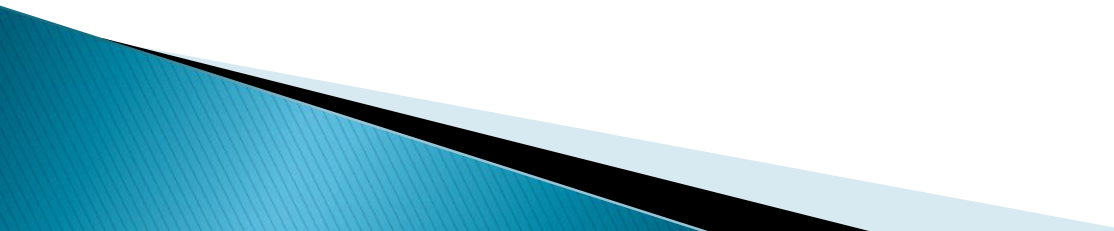
Organization of Orbitals

•The first row is Principal Energy Level1.
•The second row is Principal Energy Level 2.

•Principal Energy Level
3 begins in the 3rd row.

•Principal Energy Level 4
•begins in the 4th row.

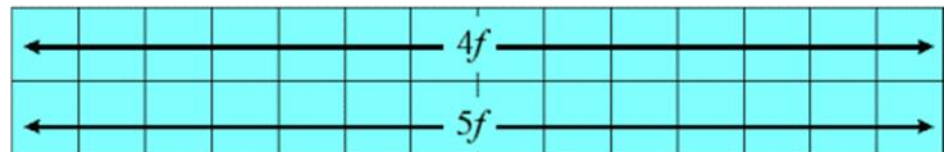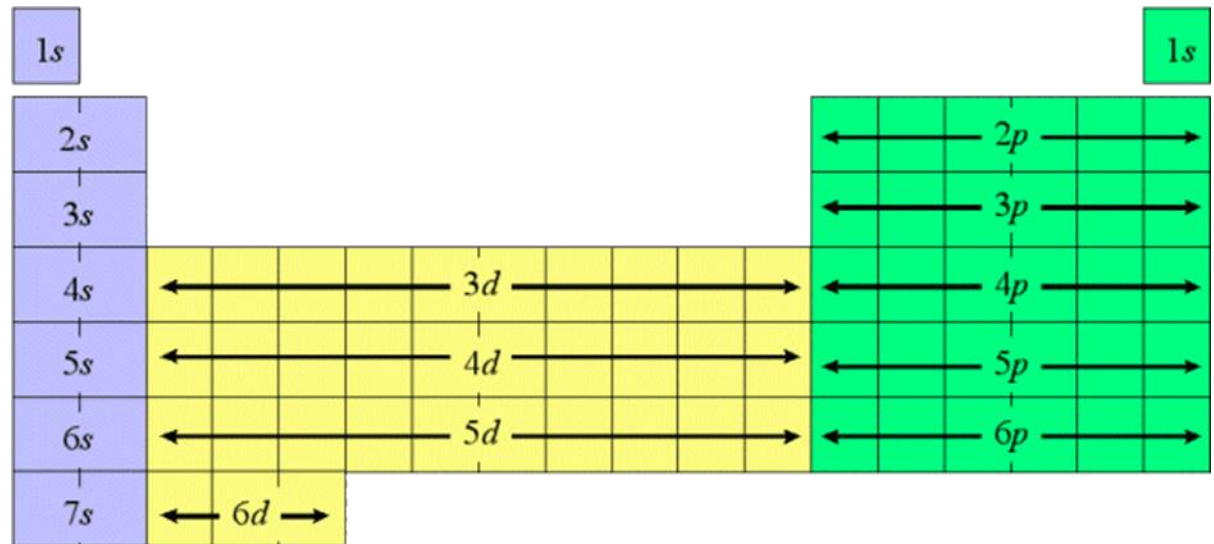|																	
$1s$																	$1s$
$2s$												$\longleftarrow$ $2p$ $\longrightarrow$					
$3s$												$\longleftarrow$ $3p$ $\longrightarrow$					
$4s$	$\longleftarrow$ $3d$ $\longrightarrow$									$\longleftarrow$ $4p$ $\longrightarrow$					
$5s$	$\longleftarrow$ $4d$ $\longrightarrow$									$\longleftarrow$ $5p$ $\longrightarrow$					
$6s$	$\longleftarrow$ $5d$ $\longrightarrow$									$\longleftarrow$ $6p$ $\longrightarrow$					
$7s$	$\longleftarrow$ $6d$ $\longrightarrow$																

$\longleftarrow$ $4f$ $\longrightarrow$
$\longleftarrow$ $5f$ $\longrightarrow$

Sub Levels

▸ The "s" sublevel can hold 2 electrons.

▸ The "p" sublevel can hold 6 electrons.
  ◦ 2 electrons in each of the 3 orbitals (x, y, z)

▸ The "d" sublevel can hold 10 electrons.
  ◦ 2 electrons in each of the 5 orbitals.

▸ The "f" sublevel can hold 14 electrons.
  ◦ 2 electrons in each of the 7 orbitals.

Practice Problems

▸ Write the electron configuration and the orbital diagram for Fluorine.

Fluorine, F $1s^2 2s^2 2p^5$

| ↑↓ | | ↑↓ | | ↑↓ | ↑↓ | ↑ |

▸ Write the electron configuration and the orbital diagram for Magnesium.

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

▸ Write the electron configuration and the orbital diagram for Sulfur.

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