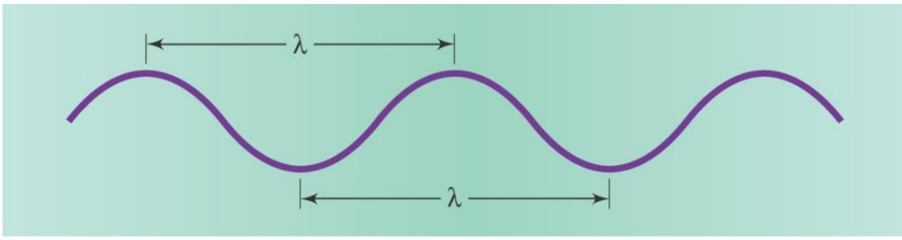
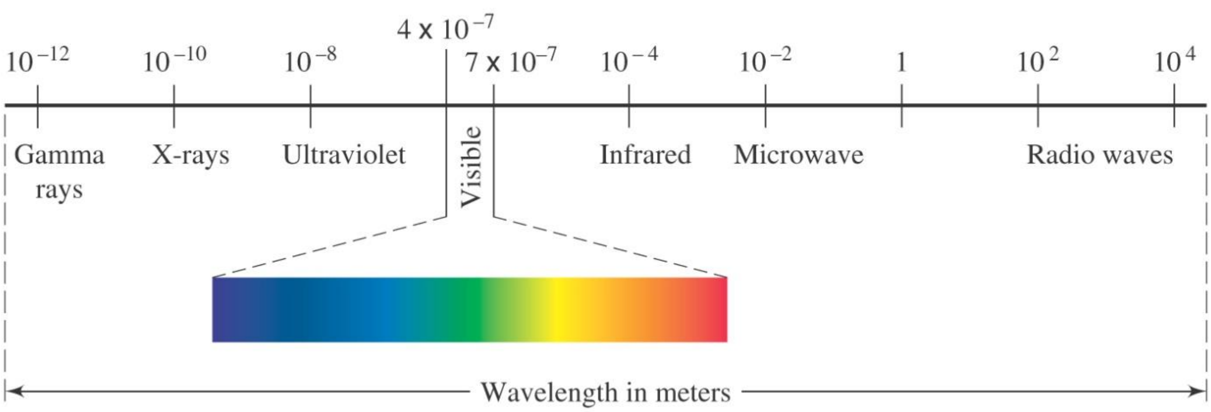
WAVES AND ELECTROMAGNETIC RADIATION

- All waves are characterized by their ***wavelength***, ***frequency*** and ***speed***.

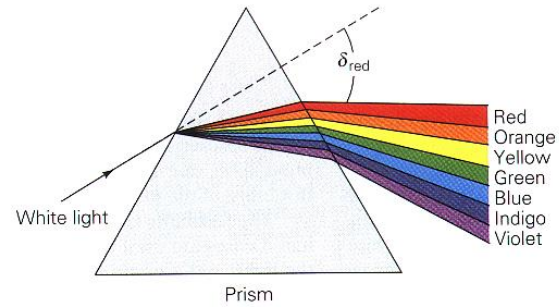


- ***Wavelength (lambda, λ):*** the distance between any 2 successive crests or troughs.
- ***Frequency (nu, ν):*** the number of waves produced per unit time.
- Wavelength and frequency are inversely proportional.
- ***Speed (c):*** tells how fast waves travel through space.
- ***Energy*** travels through space as ***electromagnetic radiation***. This radiation takes many forms, such as sunlight, microwaves, radio waves, etc.
- In ***vacuum***, all electromagnetic waves travel at the speed of light (3.00 x 10^8 m/s), and ***differ*** from each other in their ***frequency*** and ***wavelength***.
- The ***classification*** of electromagnetic waves according to their frequency is called ***electromagnetic spectrum***. These waves range from ***gamma rays*** (short λ, high *f*) to ***radio waves*** (long λ, low *f*).



ATOMIC SPECTRA

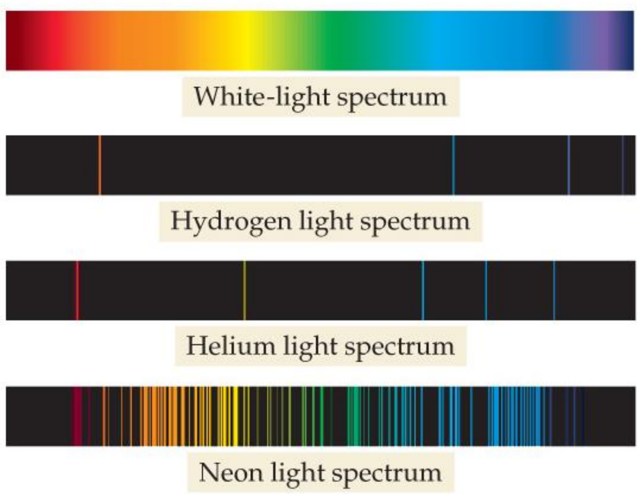
- When ***white light*** is passed through a glass ***prism***, it is ***dispersed*** into a spectrum of ***colors***. This is evidence of the ***wave nature*** of light.



- Scientists also have much evidence that light beams act as a ***stream of tiny particles***, called ***photons***.

- Scientists also discovered that when *atoms are energized* (by heat, light or electricity), they often *reemit the energy as light*. Neon lights are an example of this property of atoms.

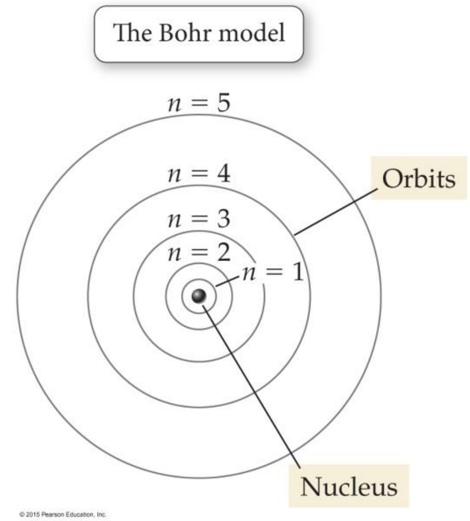
- The absorption and emission of light by atoms is due to the interaction of light with the electrons in the atom. As a result, atoms of different elements exhibit a unique color and wavelength of light. For example, mercury atoms emit light that appears blue, hydrogen atoms emit light that appears pink and helium atoms emit light that appears yellow-orange.



- These lines indicate that ***light is formed*** only at ***certain wavelengths and frequencies*** that correspond to ***specific colors***. Each ***element*** possesses a ***unique line spectrum*** that can be used to ***identify*** it.

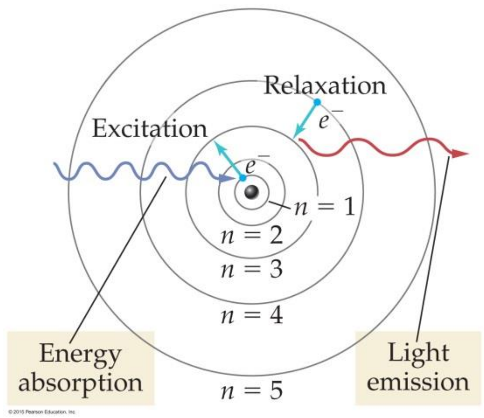
BOHR MODEL OF THE ATOM

- ***Neils Bohr***, a Danish physicist, studied the hydrogen atom extensively, and developed a model for the atom that was able to explain the line spectrum.



- ***Bohr's model*** of the atom consisted of electrons orbiting the nucleus at different distances from the nucleus, called ***energy levels***. In this model, the electrons could only occupy particular energy levels, and could ***"jump" to higher levels*** by ***absorbing energy***.

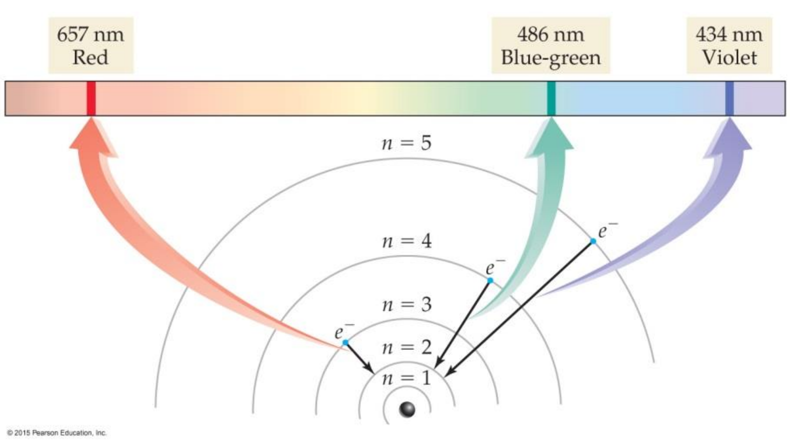
- The energy of each Bohr orbit, specified by a quantum number (n=1, 2, 3….) is fixed or quantized. Bohr's orbits are like steps on a ladder. It is possible to stand on one step or another, but impossible to stand between steps.



- The lowest energy level is called ***ground state***, and the higher energy levels are called ***excited states***. When electrons absorb energy through heating or electricity, they ***move to higher energy*** levels and become ***excited***.

- When ***excited electrons return to the ground*** state, ***energy is emitted*** as a photon of light is released. The ***color*** (wavelength) of the light emitted is determined by the ***difference in energy*** between the two states (excited and ground).

- The ***atomic spectrum*** is produced by many of these ***transitions between excited and ground states***.



Examples:

1. Rank the electromagnetic radiation listed below in order of (i) increasing wavelength, and (ii) increasing energy per photon.

   radio waves infrared microwaves ultraviolet

   increasing wavelength:

   increasing energy/photon:

2. In the Bohr model, what happens when an electron makes a transition between energy levels:

   a) Energy only increases

   b) Energy only decreases

   c) Energy increases or decreases

   d) Energy does not change

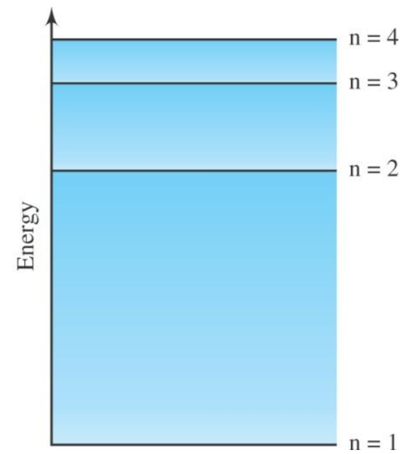
3. Two of the emission wavelengths in the hydrogen emission spectrum are 656 nm and 486 nm One of these is due to the n=4 to n=2 transition, and the other is due to the n=3 to n=2 transition. Which wavelength goes with which transition?

QUANTUM MECHANICAL MODEL OF THE ATOM

- In 1926 ***Erwin Shrodinger*** created a mathematical model that showed ***electrons as both particles and waves***. This model was called the ***quantum mechanical*** model.

- This model predicted electrons to be located in a probability region called ***orbitals***.

- An orbital is defined as a region around the nucleus where there is a high probability of finding an electron.

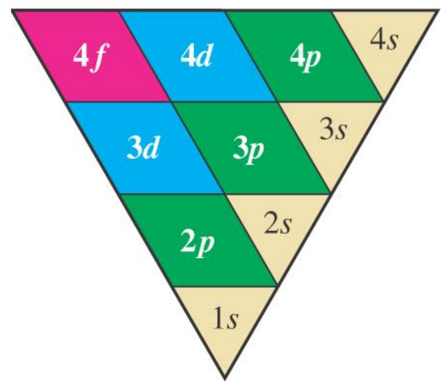


- Based on this model, there are ***discrete principal energy*** levels within the atom. Principal energy levels are designated by ***n***.

- The electrons in an atom can exist in any principal energy level. As ***n increases***, the ***energy of the electrons increases***.

- Each principal energy level is subdivided into sublevels.

Number of sublevels



- The ***sublevels*** are designated by the letters ***s, p, d and f***. As n increases, the number of sublevels increases.

- Within the sublevels, the electrons are located in ***orbitals***. The orbitals are also designated by the letters ***s, p, d and f***.

- The number of orbitals within the sublevels vary with their type.

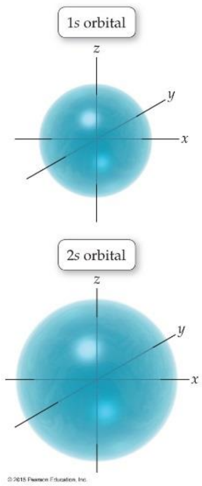
  s sublevel = 1 orbital
  p sublevel = 3 orbitals
  d sublevel = 5 orbitals
  f sublevel = 7 orbitals

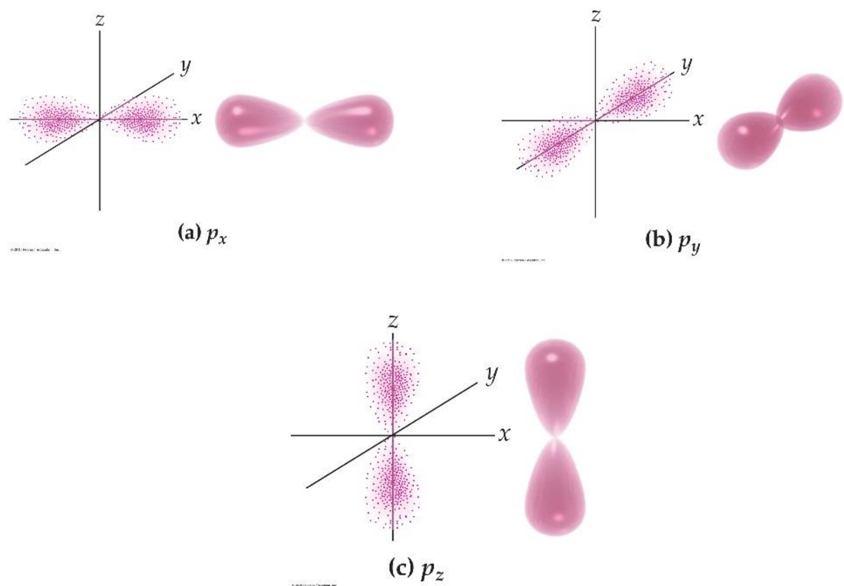
- An orbital can hold a maximum of 2 electrons.

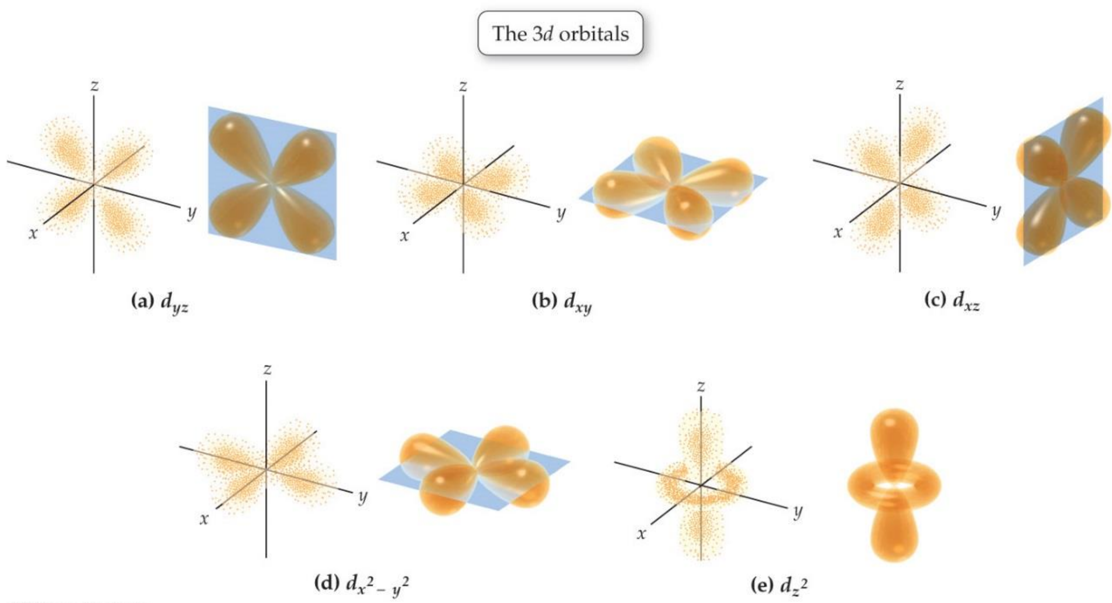
ORBITAL SHAPES

- Each orbital type can be represented by a geometric shape that encompasses the volume where the electron in likely to be found.

- These shapes for the s, p and d orbitals are shown below:







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Examples:

1. How would the 4d orbital be similar to the 3d orbital?

2. How would the 4d orbital differ from the 3d orbital?

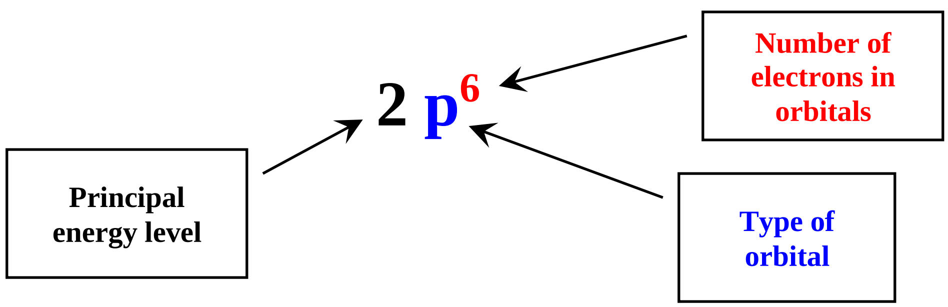
3. Which electron is, on average, farther from the nucleus: an electron in a 3p orbital or an electron in a 4p orbital?

4. Which of the following orbitals in not permissible?

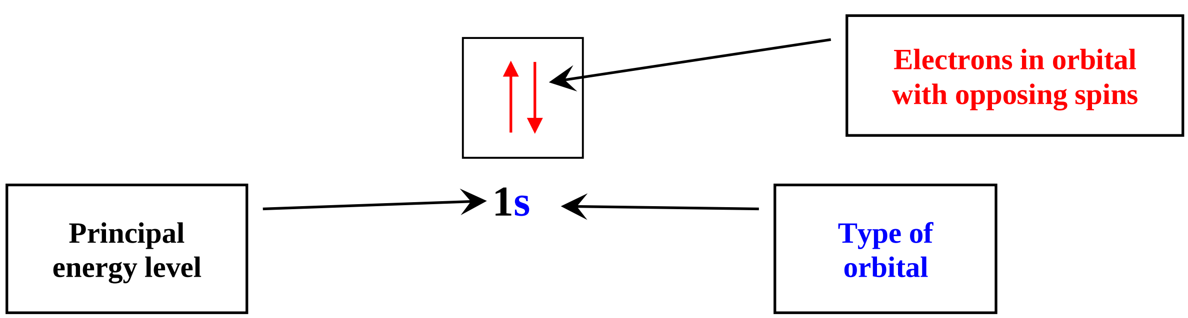
   3p 2d 4p 3s 3f

ELECTRON CONFIGURATION

- ***Similarities*** of behavior in the periodic table are due to the similarities in the ***electron arrangement*** of the atoms. This arrangement is called ***electron configuration***.
- The ***modern*** model of the atom describes the ***electron cloud*** consisting of ***separate energy levels***, each containing a ***fixed number of electrons***.
- Each orbital can be occupied by no more than 2 electrons, each with opposite spins. (Pauli exclusion principle)
- The electrons occupy the orbitals form the lowest energy level to the highest level. The energy of the orbitals on any level are in the following order: s < p < d < f.
- When orbitals of identical energy are available, each orbital is occupied by a single electron before a second electron enters the orbital (Hund's rule). For example, all three p orbitals must contain one electron before a second electron enters a p orbital.
- ***Electron configurations*** are written as shown below:



- Another notation, called the ***orbital notation***, is shown below:



Element	***Orbital Notation***	***Configuration***
Li	☐ ☐ 1s 2s	
Be	☐ ☐ 1s 2s	
B	☐ ☐ ☐☐☐ 1s 2s 2p	
C	☐ ☐ ☐☐☐ 1s 2s 2p	
N	☐ ☐ ☐☐☐ 1s 2s 2p	
O	☐ ☐ ☐☐☐ 1s 2s 2p	
F	☐ ☐ ☐☐☐ 1s 2s 2p	
Ne	☐ ☐ ☐☐☐ 1s 2s 2p	
Na	☐ ☐ ☐☐☐ ☐ 1s 2s 2p 3s	
Mg	☐ ☐ ☐☐☐ ☐ 1s 2s 2p 3s	
Al	☐ ☐ ☐☐☐ ☐ ☐☐☐ 1s 2s 2p 3s 3p	
Si	☐ ☐ ☐☐☐ ☐ ☐☐☐ 1s 2s 2p 3s 3p	
P	☐ ☐ ☐☐☐ ☐ ☐☐☐ 1s 2s 2p 3s 3p	
S	☐ ☐ ☐☐☐ ☐ ☐☐☐ 1s 2s 2p 3s 3p	

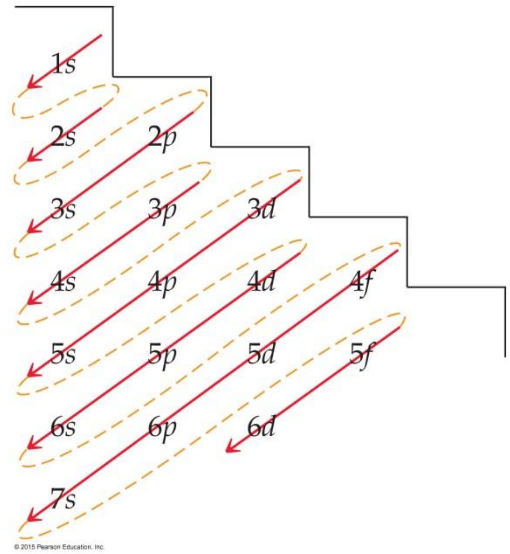
ELECTRON CONFIGURATION OF LARGER ATOMS

- As electrons occupy the 3rd energy level and higher, some anomalies occur in the order of the energy of the orbitals.

- Knowledge of these anomalies is important in order to determine the correct electron configuration for the atoms.

- The study aid shown on the right is used by beginning students to remember these exceptions to the order of orbital energies.

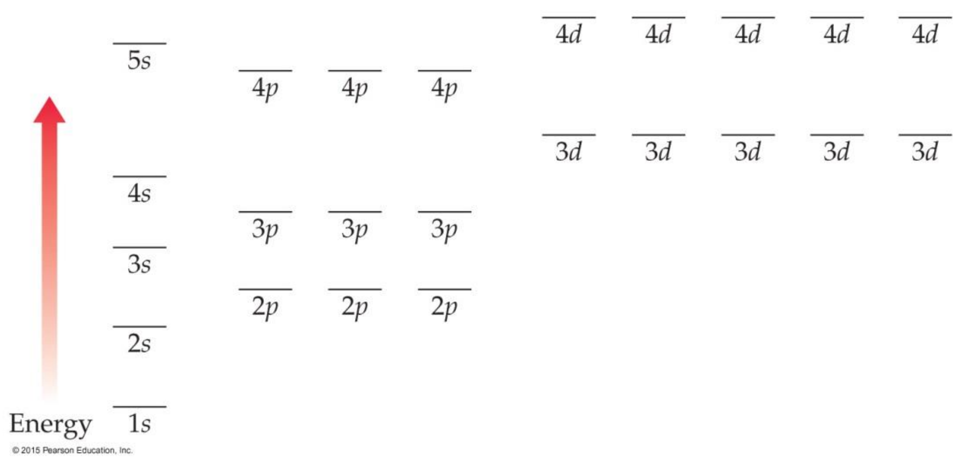
- The order of the energy of the orbitals is determined by following the tail of each arrow to the head and continuing to the next arrow in the same manner. Listed below is the order of energy of the orbitals found in this manner:



1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s

- The electrons in an atom fill from the lowest to the highest orbitals. The knowledge of the location of the orbitals on the periodic table can greatly help the writing of electron configurations for large atoms.

- The energy order of the sublevels are shown below. Note that different sublevels within the same principle shell have different energies.

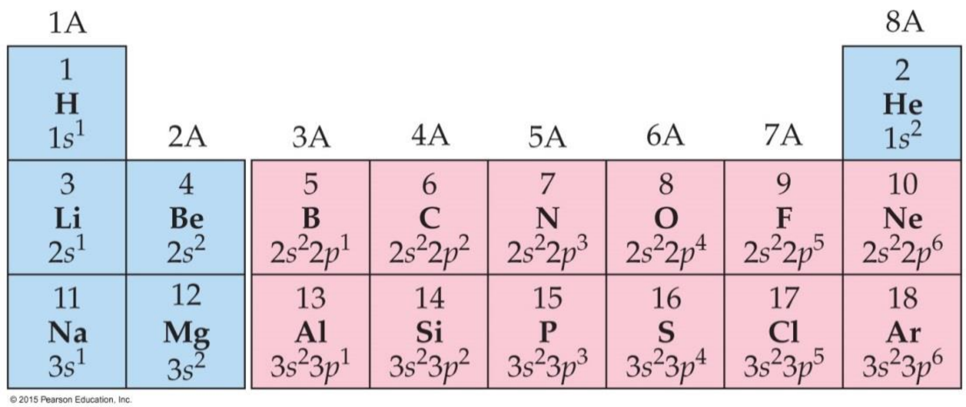


ELECTRON CONFIGURATION AND PERIODIC TABLE

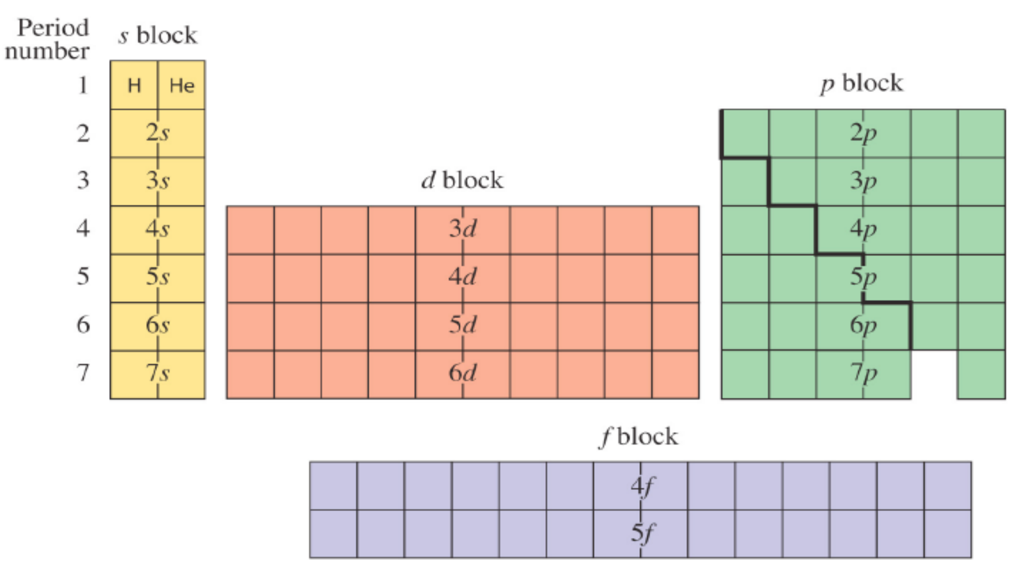
- The horizontal rows in the periodic table are called ***periods***. The period number corresponds to the number of energy levels that are occupied in that atom.

- The vertical columns in the periodic table are called ***groups*** or ***families***. For the main-group elements, the group number corresponds to the number of electrons in the outermost filled energy level (***valence electrons***).

- The ***valence electrons configuration*** for the elements in periods 1-3 are shown below. Note that elements in the same group have similar electron configurations.



- The location of the different orbital types in the periodic table is shown below:



ELECTRON CONFIGURATION SUMMARY

- Electrons occupy orbitals so as to minimize the energy of the atom. Therefore, low energy orbitals fill before higher energy orbitals.

- Orbitals can hold no more than two electrons each. When two electrons occupy the same orbital, they must have opposite spins. This is known as the ***Pauli exclusion principle***.

- When orbitals of identical energy are available, they are first occupied singly with parallel spins rather than in pairs. This is known as ***Hund's rule***.

Examples:

1. Use the periodic table to write complete electron configuration for phosphorus.

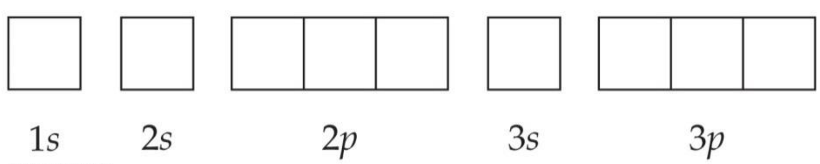
   phosphorous, Z =

   electron configuration:

2. Draw an orbital diagram for silicon (Si) and determine the number of unpaired electrons. The available orbitals are listed below.

   Si, Z =

   orbital notation:



ABBREVIATED ELECTRON CONFIGURATION

- When writing electron configurations for larger atoms, an ***abbreviated*** configuration is used.

- In writing this configuration, the non-valence (***core***) electrons are summarized by writing the symbol of the noble gas prior to the element in brackets followed by configuration of the valence electrons. For example:

K	$\mathbf{1s^2 2s^2 2p^6 3s^2 3p^6 4s^1}$	or	$\mathbf{[Ar]\ 4s^1}$
	complete configuration		abbreviated configuration
Br	$\mathbf{1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5}$	or	$\mathbf{[Ar]\ 4s^2 3d^{10} 4p^5}$
	complete configuration		abbreviated configuration

Examples:

1. Write abbreviated electron configurations for each element listed below:

   a) Fe (Z=26):

   b) Sb (Z=51):

2. Give the symbol of the element with each of the following electron configurations:

   a) [Ne] $3s^2\ 3p^1$

   b) [Ar] $4s^2\ 3d^8$