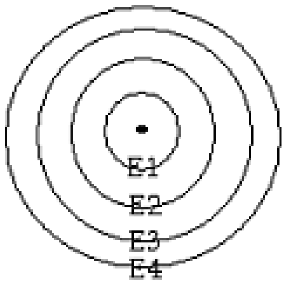
Electrons, Energy, & the Electromagnetic Spectrum Notes

Simplified, 2-D Bohr Model:

Figure 1



orbits, paths, shells, rings = ENERGY LEVELS

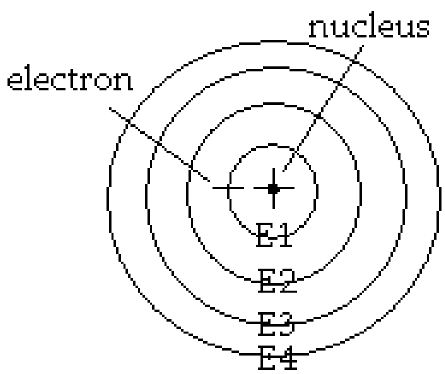
As the name "energy level" implies, there is a specific amount of energy associated with each energy level.

Electrons are lazy - they will occupy the location that requires the least amount of energy. This is called the AUFBAU principle.

lowest energy state = GROUND STATE

Figure 2

For hydrogen in the ground state...

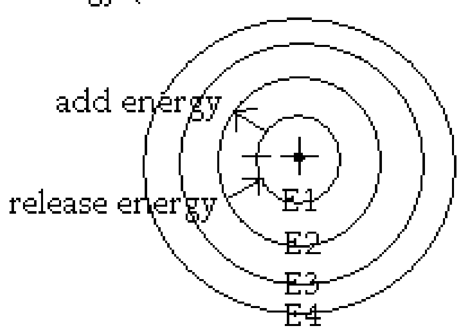


~ Electron is located in energy level 1 (E1) because E1 is closest to the nucleus. Nucleus is positively-charged & attracts negatively-charged electrons.

~ Therefore, low amount of energy needed to be in E1.

Figure 3

If energy (from an outside source) is added to the atom...



Electrons jump to a higher energy level because energy is absorbed. (Electron is now in the EXCITED STATE.)

As soon as the electron jumps to a higher level, it immediately falls back to a lower energy level. When the electron falls, energy is released.

The energy is released as electromagnetic radiation. The type of electromagnetic radiation produced depends on the difference in energy between the energy levels.

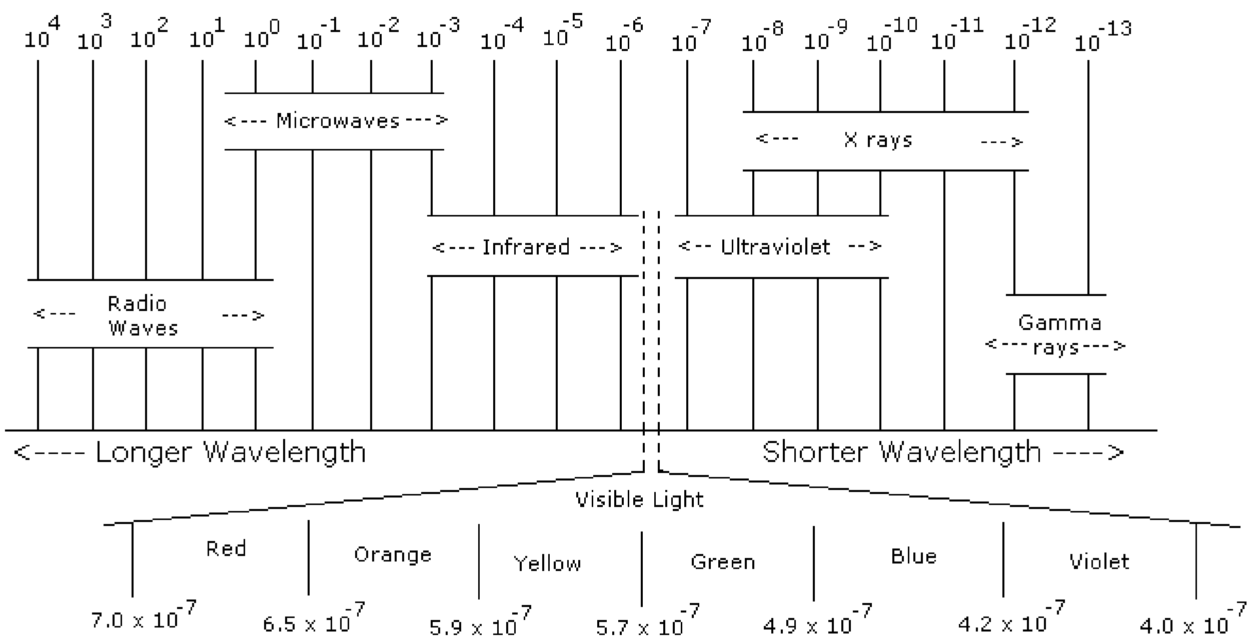
Light Calculations Notes:
* Frequency and wavelength are ______________________________ proportional
* Energy and frequency are _______________________________ proportional

Light as a Particle Notes:
* Object emits energy in small, specific amounts (called ___________________________)
* _____________________________: particle of EM radiation carrying a quantum of energy
* Einstein suggested that light had properties of both waves and particles
Referred to as the ____ of light

Quantum Theory Notes:
* When atom falls from excited state to ground state, ____
* Energy of photon = difference ____
____
* Energy states of atoms are fixed

Bohr model of the hydrogen atom Notes:
* said that e- circled the nucleus in fixed paths
* when in path, has fixed amount of energy
* e- cannot exist in space between path
* drawback of Bohr's model =

Electromagnetic Spectrum
(measurement in meters)



Notice that the spectrum is shown from the longer wavelengths on the left to the shorter wavelengths on the right. Because wavelength and frequency are inversely proportional, the longer the wavelength, the lower the frequency. Because frequency and energy are directly proportional, the lower the frequency, the lower the energy.

Use the EM Spectrum diagram to answer these questions:

1.) As you move across the visible light spectrum from red to violet…
   (A) Does the wavelength increase or decrease?
   (B) Does the frequency increase or decrease?
   (C) Does the energy increase or decrease?

2.) Which has a longer wavelength, orange or violet light?
3.) Which has a higher energy, x-rays or gamma rays?
4.) Which has a lower frequency, radio waves or green light?
5.) Which has the shortest wavelength, violet or ultraviolet light?
6.) Which has lower energy, infrared light or x-rays?

EMISSION & ABSORPTION SPECTRA NOTES

According to the Bohr atomic model, electrons orbit the nucleus within specific energy levels. These levels are defined by unique amounts of energy. Electrons possessing the lowest energy are found in the levels closest to the nucleus. Electrons with higher energy are located in progressively more distant energy levels.

If an electron absorbs sufficient energy to bridge the "gap" between energy levels, the electron may jump to a higher level. Since this change results in a vacant lower orbital, this configuration is unstable. The "excited" electron releases its newly acquired energy as it falls back to its initial or ground state. Often, the excited electrons acquire sufficient energy to make several energy level transitions. When these electrons return to the ground state, several distinct energy emissions occur. The energy that the electrons absorb is often of a thermal or electrical nature, and the energy that an electron emits when returning to the ground state is called electromagnetic radiation.

In 1900, Max Planck studied visible emissions from hot glowing solids. He proposed that light was emitted in "packets" of energy called quanta, and that the energy of each packet was proportional to the frequency of the light wave. According to Einstein and Planck, the energy of the packet could be expressed as the product of the frequency (v) of emitted light and Planck's constant (h). The equation is written as $E = hv$

If white light passes through a prism or diffraction grating, its component wavelengths are bent at different angles. This process produces a rainbow of distinct colors known as a continuous spectrum. If, however, the light emitted from hot gases or energized ions is viewed in a similar manner, isolated bands of color are observed. These bands form characteristic patterns - unique to each element. They are known as bright line spectra or emission spectra.

By analyzing the emission spectrum of hydrogen gas, Bohr was able to calculate the energy content of the major electron levels. Although the electron structure as suggested by his planetary model has been modified according to modern quantum theory, his description and analysis of spectral emission lines are still valid.

In addition to the fundamental role of spectroscopy played in the development of today's atomic model, this technique can also be used in the identification of elements. Since the atoms of each element contain unique arrangements of electrons, emission lines can be used as spectral fingerprints. Even without a spectroscope, this type of identification is possible since the major spectral lines will alter the color.

ELECTRON ARRANGEMENT NOTES

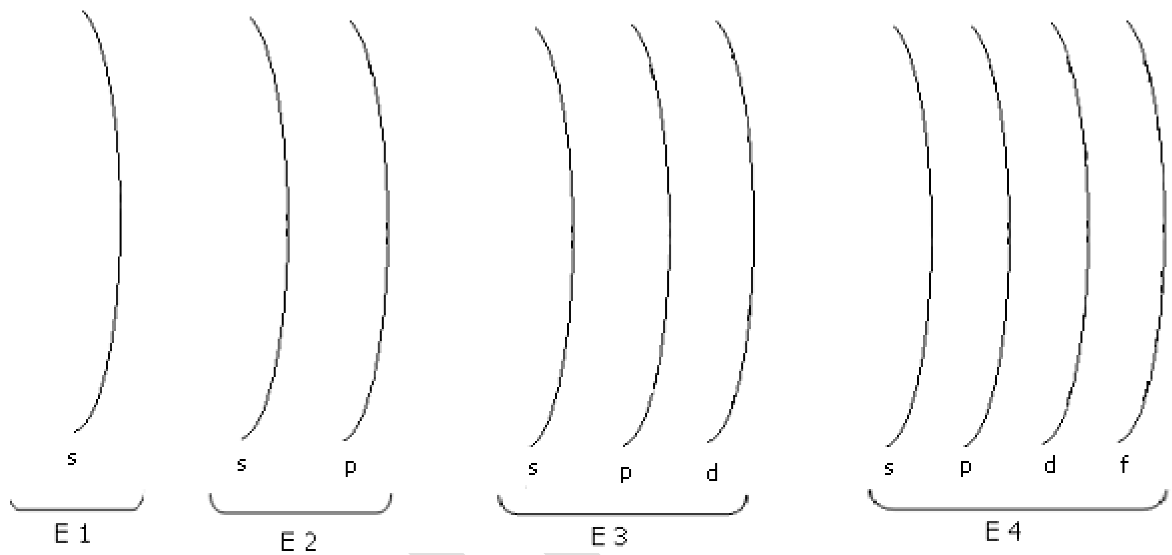
Heisenberg Uncertainty Principle:

GENERAL LOCATION -------------------- ____________________
|
V
____
|
V
____
|
V
SPECIFIC LOCATION -------------------- ____________________

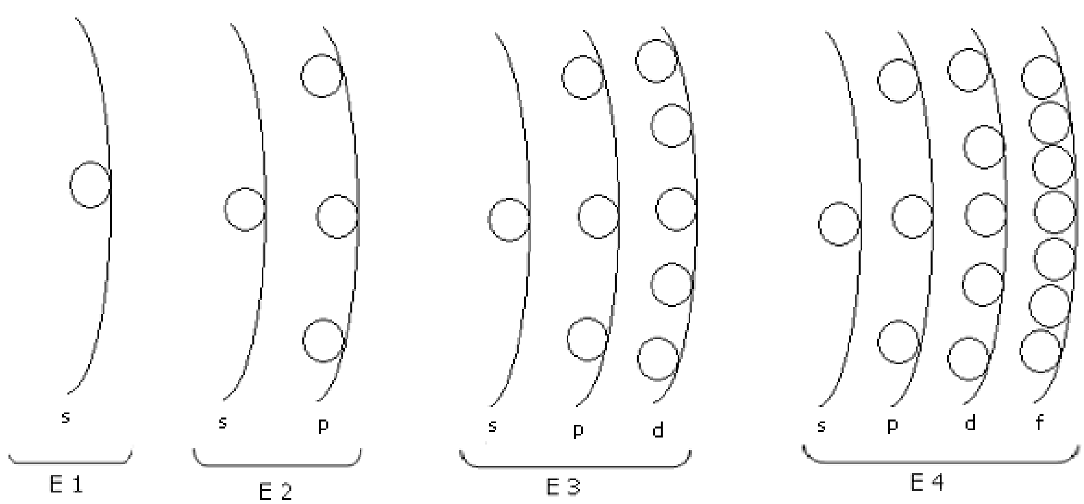
ENERGY LEVELS - divisions of the electron cloud
- numbered consecutively from closest to farthest away from nucleus

Electrons will occupy the location with the lowest amount of energy.

SUBLEVELS - divisions of energy levels
- designated by letters (s, p, d, f)
- number of sublevels in an energy level = # of the energy level



ORBITALS - divisions of sublevels
- number of orbitals in an energy level =
- "s" sublevel has 1 orbital; "p" sublevel has 3 orbitals; "d" sublevel has 5 orbitals; "f" sublevel has 7 orbitals

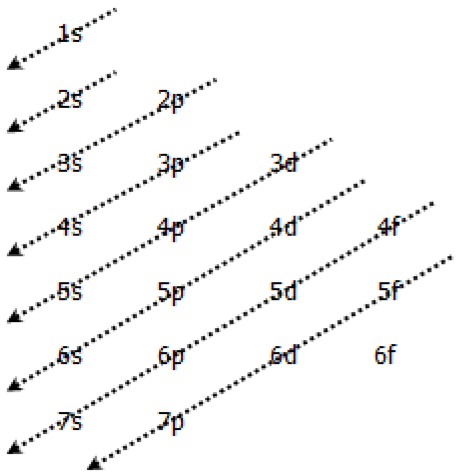


HOW MANY ELECTRONS CAN AN ORBITAL HOLD?

HOW MANY ELECTRONS CAN EACH SUBLEVEL HOLD?
"s" = ___ e-   "p" = ___ e-   "d" = ___ e-   "f" = ___ e-

HOW MANY ELECTRONS CAN AN ENERGY LEVEL HOLD?

What is the order in which the sublevels fill with electrons?   Use the DIAGONAL RULE.



ELECTRON ARRANGEMENT WORKSHEET

1. What is an electron cloud?
2. Name the three major divisions within an electron cloud with respect to the energy of an electron.
3. What letter represents the principal quantum number?
4. What does the principal quantum number tell about an electron?
5. What formula is used to determine the maximum number of electrons that can occupy any energy level?
6. What is the maximum number of electrons for each of the following?
   (A) 1st energy level   (B) 4th energy level   (C) n = 3   (D) n = 5
7. Energy levels are divided into _______________.
8. How can we determine the possible number of sublevels in any energy level?
9. Name the four primary sublevels in order of increasing energy.
10. Circle the sublevel that represents the lowest energy in each pair.
    (A) 1s or 2s   (B) 2s or 2p   (C) 4f or 4d   (D) 3d or 4s   (E) 7s or 5d
    (F) 6s or 4s   (G) 4p or 5p   (H) 3s or 3d   (I) 2p or 3s
11. Sublevels are divided into _______________.
12. Each orbital can hold up to _______________ electrons.
13. Sketch the shapes of the orbitals for the sublevels listed.
    (A) s: (B) p_x: (C) p_y: (D) p_z:

14. How many orbitals are in each sublevel?
    (A) s _______________   (B) p _______________   (C) d _______________   (D) f _______________

More Electron Arrangement NOTES

Examples:

Se $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$

Sn $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^2$

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

HOEL (Highest Occupied Energy Level): energy level furthest from the nucleus that contains at least one electron
How to determine this using electron configuration?
~ largest non-exponent number

Se $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4}s^2\ 3d^{10}\ \underline{4}p^4$ HOEL = 4

Sn $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ \underline{5}s^2\ 4d^{10}\ \underline{5}p^2$ HOEL = 5

Hg $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ \underline{6}s^2\ 4f^{14}\ 5d^{10}$ HOEL = 6

Valence Electrons: electrons in the HOEL
How to determine this using electron configuration?
~ add up exponents of terms in HOEL

Se $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ \underline{4}s^{\mathbf{2}}\ 3d^{10}\ \underline{4}p^{\mathbf{4}}$
HOEL = 4 Valence electrons = 2 + 4 = 6

Sn $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ \underline{5}s^{\mathbf{2}}\ 4d^{10}\ \underline{5}p^{\mathbf{2}}$
HOEL = 5 Valence electrons = 2 + 2 = 4

Hg $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ \underline{6}s^{\mathbf{2}}\ 4f^{14}\ 5d^{10}$
HOEL = 6 Valence electrons = 2

Noble Gas Configuration: shortcut for electron configuration
How is it written?
~ [symbol for noble gas closest to element with lower atomic #]
~ [after brackets] next number is the period that the element is located in
~ after that number, write "s"
~ continue electron configuration in diagonal rule order until appropriate # of electrons is reached
NOTE: ending of electron configuration and noble gas configuration should be the same

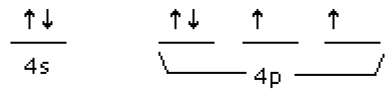
Se $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$
$\underset{18}{[Ar]}\ \underset{20}{4s^2}\ \underset{30}{3d^{10}}\ \underset{34}{4p^4}$

Sn $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^2$
$\underset{36}{[Kr]}\ \underset{38}{5s^2}\ \underset{48}{4d^{10}}\ \underset{50}{5p^2}$

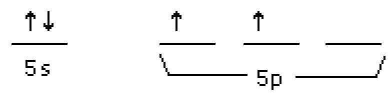
Hg $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}$
$\underset{54}{[Xe]}\ \underset{56}{6s^2}\ \underset{70}{4f^{14}}\ \underset{80}{5d^{10}}$

Orbital Notation: drawing of how electrons are arranged in orbitals; will only need to do this for the HOEL
*NOTE: ___ = orbital ↑ or ↓ = electrons

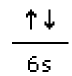
Se $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$



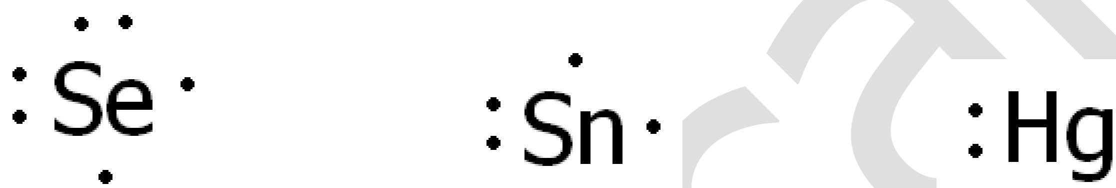
Sn $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^2$

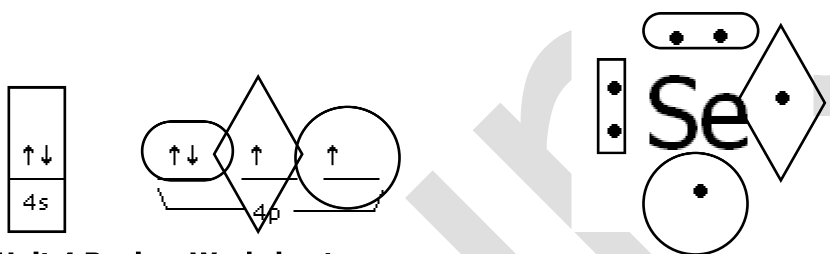


Hg $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}$



Dot Diagrams: symbol represents nucleus and non-valence ("inner-shell") electrons; dots around symbol represent valence electrons





Unit 4 Review Worksheet

Section I - Electromagnetic Spectrum

1. Label both ends of the spectrum with high/low frequency, high/low energy, and long/short wavelength

radio waves microwaves infrared light ROYGBIV ultraviolet light x-rays gamma rays

2. Which has a higher energy, gamma or x-rays?
3. Which has a shorter wavelength, radio or ultraviolet?
4. Which has a lower frequency, yellow or green light?
5. In the equation $E = h \cdot v$, energy and frequency are ____________________ proportional.
6. In the equation $c = \lambda \cdot v$, wavelength and frequency are __________________ proportional.
7. The symbol for wavelength is _____.
8. Electrons give off energy in the form of a ____________________ when returning to the ground state.
9. Which scientist proposed the idea that electrons travel around the nucleus in fixed paths?
10. When an electron moves from the ground state to the excited state, energy is ____________________.
11. Bohr chose the element ____________________ to prove his theory.
12. The dual wave-particle nature of electrons describes how the electrons in atoms can behave as ____________________ and ____________________.

Section II - Electrons

13. What is an electron cloud?
14. Who proposed the uncertainty principle?
15. Who is credited with the idea that electrons are placed in the lowest energy level first?
16. What rule requires that each of the "p" orbitals (at a particular energy level) receive one electron before any of the orbitals can have two electrons?
16. What is the maximum number of electrons in any orbital?
17. The principal quantum number, n, indicates the ____________________.
18. The maximum number of electrons in an energy level can be determined by the equation _______________ That means the maximum number of electrons in the 3rd energy level is ____________________.
19. The number of sublevels in any energy level can be determined by ____________________.
20. The number of orbitals in an energy level can be determined by the equation ____________________. So, the 3rd energy level has _____ orbitals. (_____ is/are "s" orbitals, _____ is/are "p" orbitals, and _____ is/are "d" orbitals)
21. List the four sublevels according to increasing energy.
22. The "s" sublevel is shaped like a ____________________ and has _____ orbitals.
23. A "p" sublevel is shaped like a ____________________ and has _____ orbitals.
24. The "d" sublevel has _____ orbitals and the "f" sublevel has _____ orbitals.

Section III - Electron configuration, noble gas configuration, valence electrons, orbital notations

25. What is the electron configuration for phosphorus?
26. How many total electrons are in a neutral atom of phosphorus?
27. Write the noble gas configuration for phosphorus.
28. What is the highest occupied energy level for phosphorus?
29. What is the atomic number of phosphorus?
30. Draw the orbital notation for phosphorus.
31. How many electrons are in the highest occupied energy level of phosphorus?
32. How many inner-shell electrons does phosphorus have?
33. Draw the electron dot diagram for phosphorus.