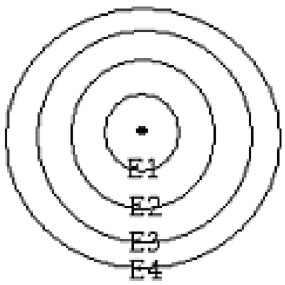
Electrons, Energy, & the Electromagnetic Spectrum Notes

Simplified, 2-D Bohr Model:

Figure 1



orbits, paths, shells, rings = ENERGY LEVELS

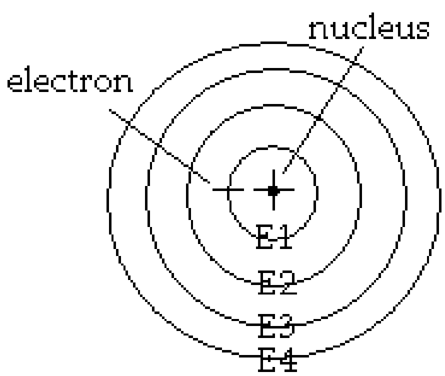
As the name "energy level" implies, there is a specific amount of energy associated with each energy level.

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

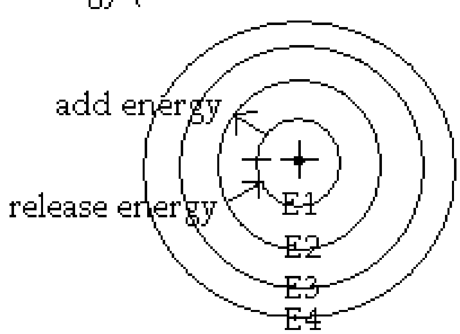
For hydrogen in the ground state...



~ 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

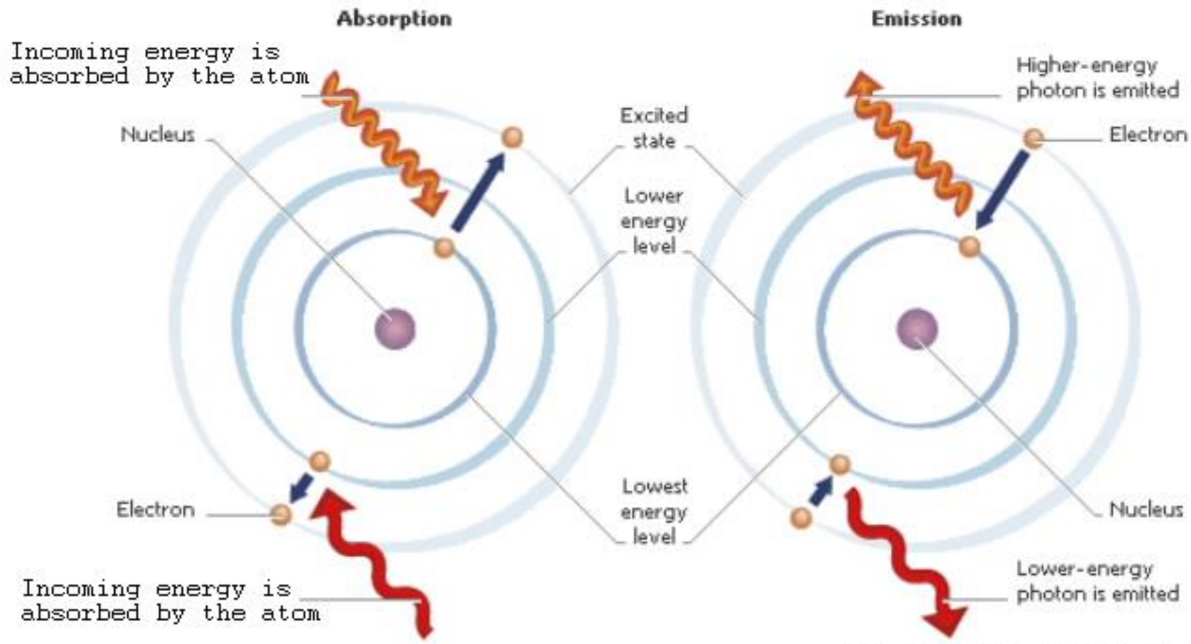
If energy (from an outside source) is added to the atom...



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.



Why does the photon at the top (pictured in orange) have higher energy than the photon at the bottom (pictured in red)?

Light Calculation Notes

Here's how the type/form of EM radiation can be determined…

The amount of energy released when an electron falls from a higher to a lower energy level is directly proportional to its frequency.

The calculation follows the equation: $E = h \cdot \nu$

E = Energy (unit is J)

h = Planck's Constant (6.626×10^{-34} J·s)

ν = frequency (unit is Hz or $^1/_{\text{second}}$)

EXAMPLE 1: A particle of EM radiation has an energy of 1.15×10^{-16} J. What is its frequency?

$1.15 \times 10^{-16}\text{ J} = 6.626 \times 10^{-34}\text{ J·s} \cdot \nu$

$\nu = 1.74 \times 10^{17}$ Hz

The type of electromagnetic radiation can be determined if one knows the wavelength. The wavelength is inversely proportional to the frequency.

The calculation follows the equation: $c = \lambda \cdot \nu$

c = speed of light (3.00×10^8 m/s)

λ = wavelength (unit is m)

ν = frequency (unit is Hz or $^1/_s$)

EXAMPLE 2: What is the wavelength of the same particle from EXAMPLE 1?

$3.00 \times 10^8\text{ m/s} = \lambda \cdot 1.74 \times 10^{17}$ Hz

$\lambda = 1.72 \times 10^{-9}$ m

EXAMPLE 3: What type of electromagnetic radiation is the particle from EXAMPLE 1?

answer for wavelength is 10^{-9} so use chart below to determine…

x-rays or ultraviolet (either one is acceptable)

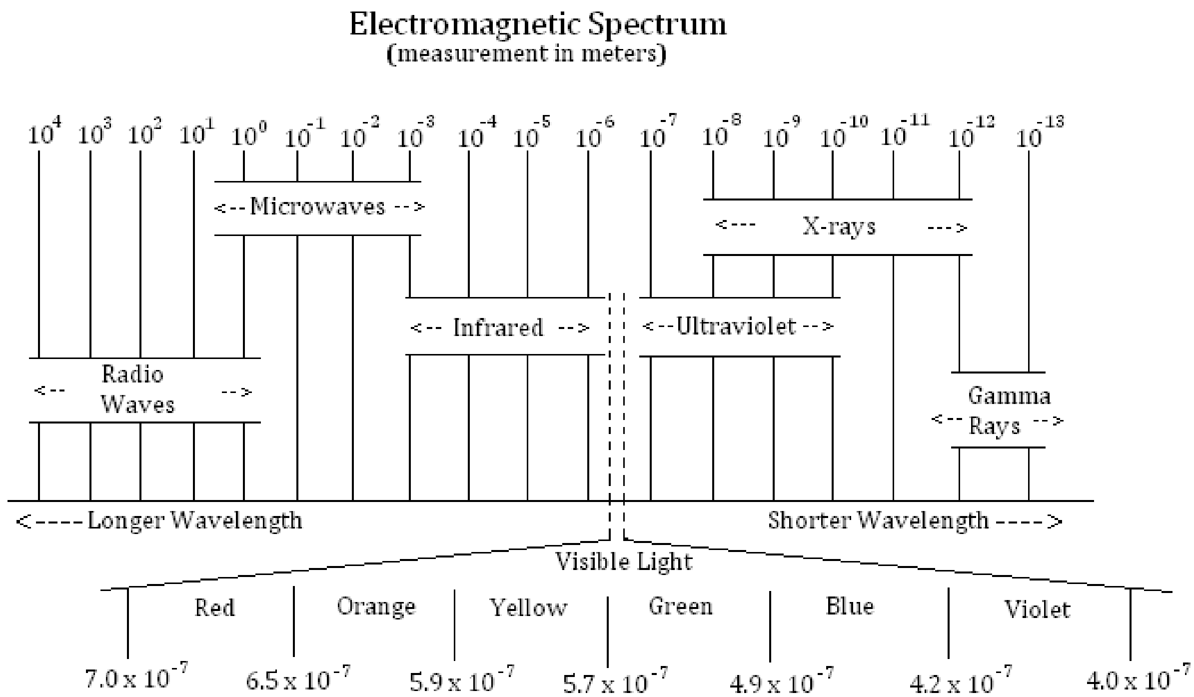
PROBLEMS FOR YOU TO TRY ON YOUR OWN…

1.) A particle of EM radiation has a frequency of 4.80×10^{14} Hz.

(A) How much energy does this particle have?

(B) What is the wavelength of this particle?
(C) What specific type of electromagnetic radiation does this particle represent?

2.) A particle of electromagnetic radiation has 2.39×10^{-13} Joules of energy.
(A) What is the wavelength of this particle?
(B) What type of electromagnetic radiation does this particle represent?



EM SPECTRUM, WAVELENGTH, FREQUENCY, AND ENERGY WORKSHEET

1.) Look at the EM spectrum below to answer this question.
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.) A beam of microwaves has a frequency of 1.0×10^9 Hz. A radar beam has a frequency of 5.0×10^{11} Hz.
Which type (microwave or radar)…
(A) has a longer wavelength?
(B) is closer to visible light on the EM spectrum?
(C) is closer to x-rays in frequency value?

3.) What is the frequency of an EM radiation wave if its wavelength is 3.6×10^{-9} meters?

4.) A beam of EM radiation has a wavelength of 4.257×10^{-7} cm. What is its frequency?

5.) A photon of light has a wavelength of 3.20×10^5 meters. Find…
(A) the frequency
(B) the energy
(C) the region of the EM spectrum/type of radiation

6.) A photon has an energy of 4.00×10^{-19} J. Find…
(A) the frequency
(B) the wavelength
(C) the region of the EM spectrum/type of radiation

7.) A bright line spectrum contains a line with a wavelength of 518 nm. Determine…
(A) the wavelength in meters
(B) the frequency
(C) the energy
(D) the color

*8.) Cobalt-60 is an artificial radioisotope that is produced in a nuclear reactor for use as a gamma ray source in the treatment of certain types of cancer. If the wavelength of the gamma radiation from a cobalt-60 source is 1.00×10^{-3} nm, calculate the energy of a photon of this radiation.

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 =

Example Questions:
1. Which has a longer wavelength - microwaves or x-rays?
2. Which has higher frequency - radio waves or ultraviolet?
3. Which has more energy - gamma rays or visible light?

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 (ν) of emitted light and Planck's constant (h). The equation is written as $E = h \cdot \nu$

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.

PROPERTIES OF LIGHT WORKSHEET

Part 1 - Select the best answer

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

Part 2 - Fill in the blanks

6. ______________ formed a theory to explain the structure of an atom by revising physical theories.
7. As the energy level increases, the amount of energy an electron will possess ______________.
8. Electrons give off energy in finite amounts called ______________ when returning to the ground state.
9. When this energy is released in the form of light it is called a ______________.
10. The speed of light = ______________ (give number and units)
11. The symbol for wavelength is ______________.
12. In the equation $c = \lambda \cdot \nu$, c represents ______________, ν represents ______________, and λ represents ______________.
13. In the equation $c = \lambda \cdot \nu$, λ and ν are ______________ proportional.
14. In the equation $E = h \cdot n$, h represents ______________ and E represents ______________.
15. In the equation $E = h \cdot \nu$, E and ν are ______________ proportional.
16. Bohr chose the element ______________ to prove his theory.

Part 3 - True or False

17. Electrons may regularly occupy spaces between energy levels.
18. The varying wavelengths on the electromagnetic radiation spectrum travel at different speeds.
19. Atoms release energy when electrons jump to higher energy levels.

ELECTRON ARRANGEMENT NOTES

Heisenberg Uncertainty Principle:

GENERAL LOCATION -------------------- ______________________________

|
V

____

|
V

____

|
V

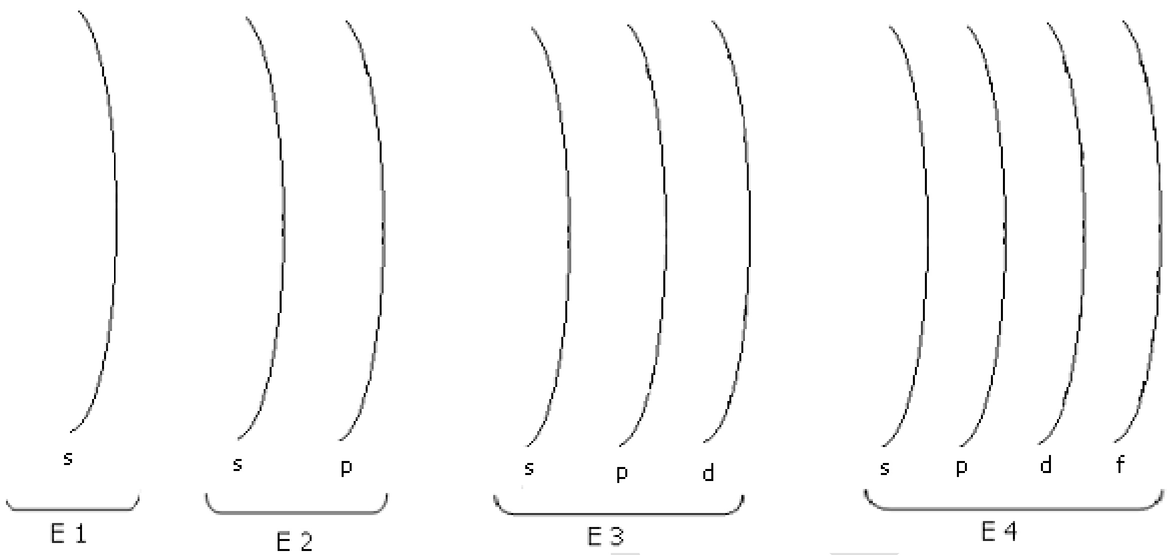
SPECIFIC LOCATION -------------------- ______________________________

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

.

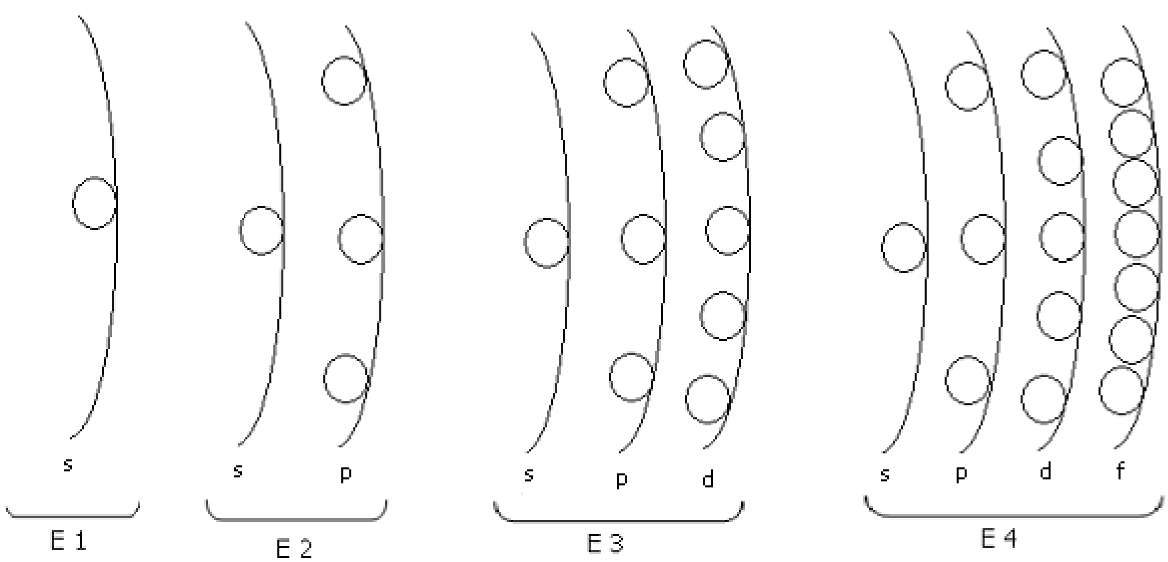
Electrons will occupy the location with the lowest amount of energy.

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

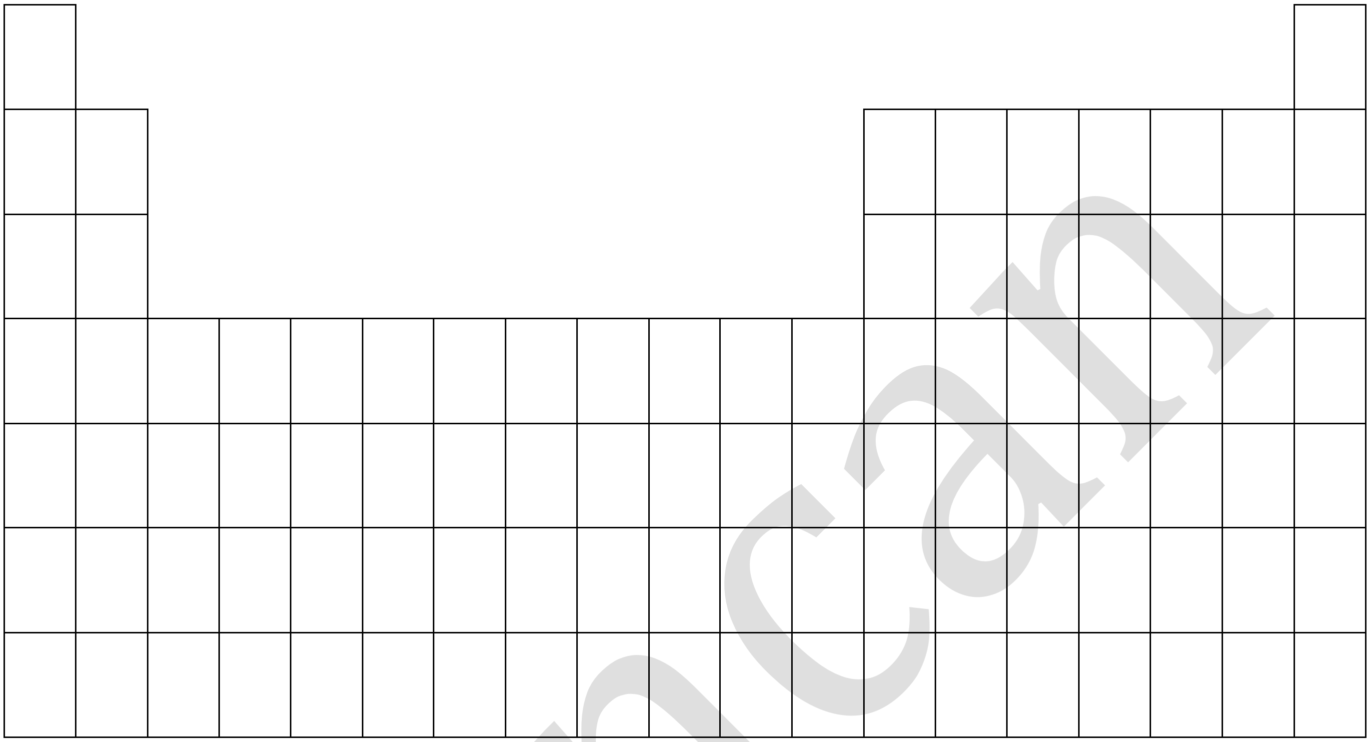


HOW MANY ELECTRONS CAN AN ORBITAL HOLD?

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 PERIODIC TABLE.



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

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

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

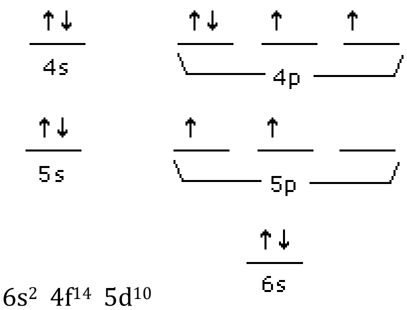
Noble Gas Configuration: shortcut for electron configuration
How is it written?
~ [symbol for noble gas closest to element with lower atomic #]
~ [after brackets] write the rest of the configuration as if you had written the configuration for the noble gas
NOTE: ending of electron configuration and noble gas configuration should be the same

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

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

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

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

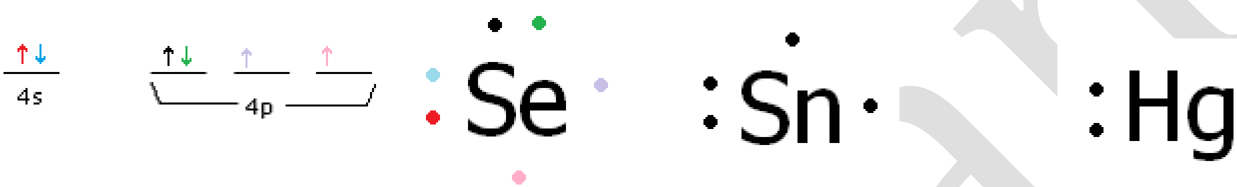


Se $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$

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 - Problems Given: $E = h \cdot \nu$ $h = 6.626 \times 10^{-34}$ J·s
$c = \lambda \cdot \nu$ $c = 3.00 \times 10^8$ m/s

1. What is the frequency of a wave with a wavelength of 3.5×10^{-4} m?
2. What is the energy of a photon with a frequency of 5.41×10^{17} Hz?
3. What type of electromagnetic radiation is described in question 2?

Section II - Electromagnetic Spectrum

4. 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

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

Section III - Electrons

16. What is an electron cloud?
17. Who proposed the uncertainty principle?
18. Who is credited with the idea that electrons are placed in the lowest energy level first?

19. 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?

20. What is the maximum number of electrons in any orbital?

21. The principal quantum number, n, indicates the ___________________.

22. 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 ___________________.

23. The number of sublevels in any energy level can be determined by ___________________.

24. 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)

25. List the four sublevels according to increasing energy.

26. The "s" sublevel is shaped like a ___________________ and has _____ orbitals.

27. A "p" sublevel is shaped like a ___________________ and has _____ orbitals.

28. The "d" sublevel has _____ orbitals and the "f" sublevel has _____ orbitals.

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

29. What is the electron configuration for phosphorus?

30. How many total electrons are in a neutral atom of phosphorus?

31. Write the noble gas configuration for phosphorus.

32. What is the highest occupied energy level for phosphorus?

33. What is the atomic number of phosphorus?

34. Draw the orbital notation for phosphorus.

35. How many electrons are in the highest occupied energy level of phosphorus?

36. How many inner-shell electrons does phosphorus have?

37. In which orbitals are the inner-shell electrons located?

38. Draw the electron dot diagram for phosphorus.