The Electron
Topic#5
AMSAT Chem 1H

Student Edition

Electrons
Topic#5

Objectives

New Atomic Model

1) Explain the mathematical relationship among speed, wavelength, and frequency of electromagnetic radiation.
2) Discuss the dual wave-particle nature of light.
3) Discuss the significance of the photoelectric effect and line-emission spectrum of hydrogen to the development of the atomic model.
4) Describe the Bohr model of the hydrogen atom.

λ wavelength
ν frequency
E energy

$\lambda\uparrow$ $\nu\downarrow$ $E\downarrow$

$\lambda\downarrow$ $\nu\uparrow$ $E\uparrow$

inverse

$\uparrow\lambda \times \nu = c$

$\downarrow$

$E = h\nu$

$$\frac{E}{\nu} = h$$

direct

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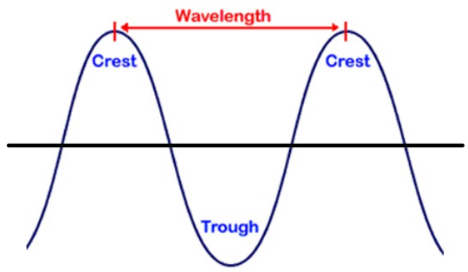
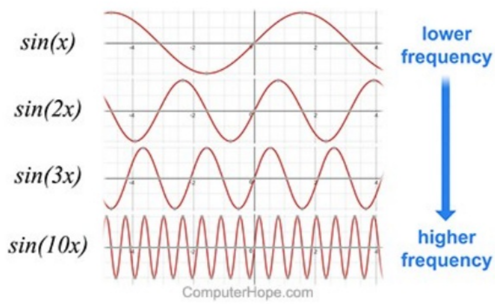
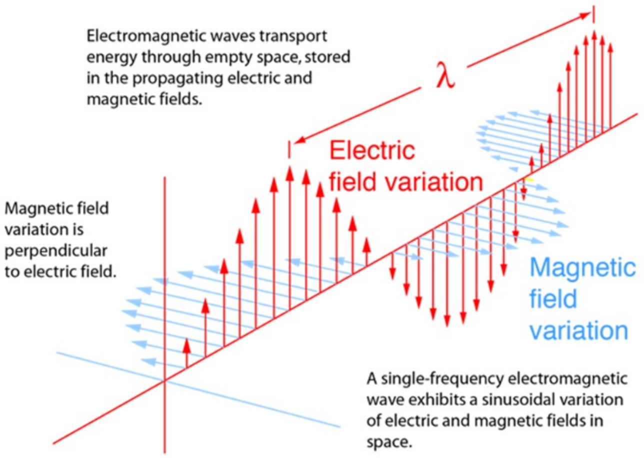
Waves, Photons, and Energy

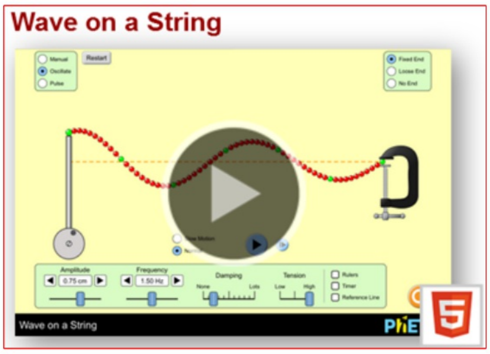
- Light is **electromagnetic radiation (EM)**
  - Exhibit wave-like behavior as it travels through space
- **EM Spectrum**
  - all light from gamma ray to radio waves
  - all forms have the same speed: **$3.00 x 10^8$ m/s = *c***
  - characterized by a **wavelength(*λ, lambda*)** and **frequency(ν, *nu*)**
  - each form has a different frequency and wavelength
    - Wavelength is the distance between two crest/troughs of a wave
      - Measured in meters (m)
    - Frequency is a measure of the number of items that pass a given point over a specific time, for light it is cycles per second.
      - Measured in 1/s or Hz (hertz)
- c = ***λ x* ν**, what is the label for speed?
- highest *E* and ***υ***, short ***λ*** lowest *E* and ***υ***, long ***λ***
  (gamma ray/x-ray/UV/VIB G YOR/IR/microwave/TV/radio)
- Energy of a photon of light: $E = h\nu = hc/\lambda$
  - Label for energy is joules (J)
  - Planck's constant (*h*) = $6.626 x 10^{-34}$ J•s





Waves, Photons, and Energy





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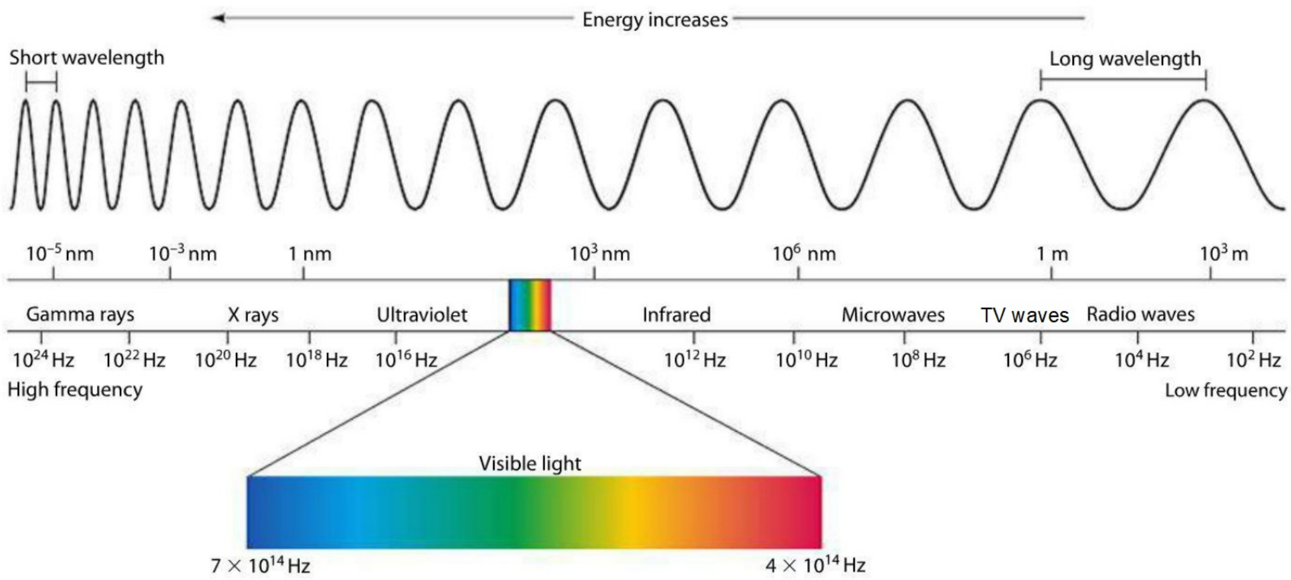

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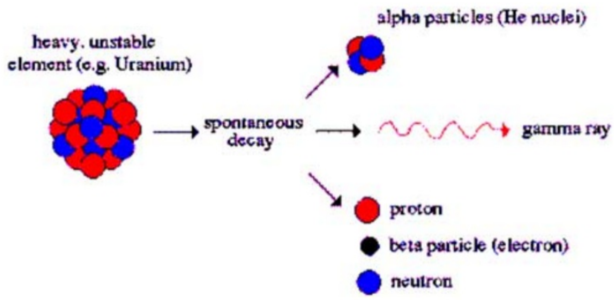
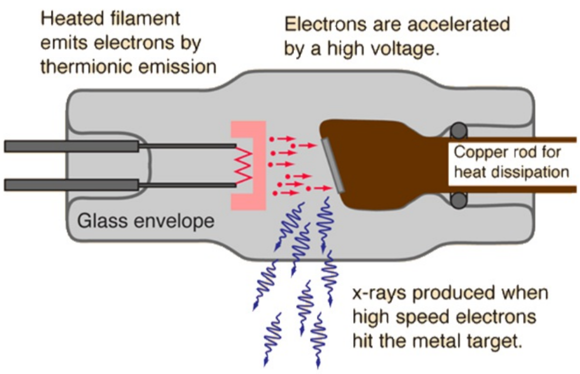
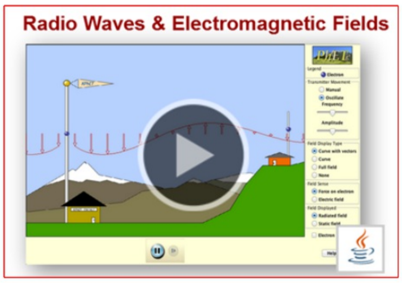


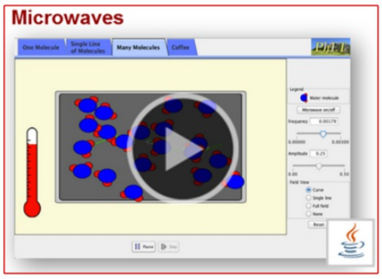
Gamma ray production



X-ray production

Waves, Photons, and Energy





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Waves, Photons, and Energy

$c = \lambda \times v$ $c = 3.00 \times 10^8$ m/s

$E_{ph} = hv = hc/\lambda$ $h = 6.626 \times 10^{-34}$ J•s

Electron Topic#4 - Sample WS#1 - Waves and Energy

1. Red light has a wavelength 675nm. What is the frequency and energy of this red light photon?
(Ans: 4.44×10^{14} 1/s, 2.94×10^{-19} J)

Given
$\lambda = 675$ nm
675×10^{-9} m

NTK
$c = \lambda * v = 3.00 \times 10^8$
(m/s) = (m) (1/s)
$v = c/\lambda$

Unk
$v =$ ______ 1/s or Hz
$E =$ ______ J

Solve:
$$v = \frac{3.00 \times 10^8 \text{ m/s}}{675 \times 10^{-9} \text{ m}} = 4.44 \times 10^{14} \text{ 1/s}$$

$$E_{ph} = \frac{hc}{\lambda} = \frac{(6.626 \times 10^{-34} \text{ J} \cdot \text{s})(3.00 \times 10^8 \text{ m/s})}{(675 \times 10^{-9} \text{ m})} = 2.94 \times 10^{-19} \text{ J}$$

$$E_{ph} = hv = (6.626 \times 10^{-34} \text{ J} \cdot \text{s})(4.44 \times 10^{14} \text{ 1/s}) = 2.94 \times 10^{-19} \text{ J}$$

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Waves, Photons, and Energy

$c = \lambda \times v$ $c = 3.00\text{x}10^{8}\text{m/s}$

$E_{ph} = hv = hc/\lambda$ $h = 6.626\text{x}10^{-34}\text{J}\bullet\text{s}$

2. Green light has a frequency of $6.00\text{x}10^{14}$Hz. What is the wavelength and energy of this green light photon?
(Ans:$5.00\text{x}10^{-7}$m/$3.98\text{x}10^{-19}$J

Gvn
$v = 6.00\times10^{14}$ Hz (1/s)

NTK
$\lambda = c/v$
$E = hv$

Unk
$\lambda =$ ______ m
$E =$ ______ J

Solve:

$\lambda = \frac{3.00\times10^{8}\,m/s}{6.00\times10^{14}\,1/s} = 5.00\times10^{-7}\,m$

$E = (6.626\times10^{-34}\,J\cdot s)(6.00\times10^{14}\,1/s) = 3.98\times10^{-19}\,J$

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Waves, Photons, and Energy

$c = \lambda \times \nu$ $c = 3.00 \times 10^{8}$ m/s

$E_{ph} = h\nu = hc/\lambda$ $h = 6.626 \times 10^{-34}$ J•s

3. (OYO) What is the energy and frequency of a violet light photon with a wavelength of 434 nm?

(Ans: 4.58×10^{-19} J/6.91×10^{14} Hz or (1/s))

Waves, Photons, and Energy

$c = \lambda \times \nu$ $\quad$ c = 3.00×10^{8}m/s

$E_{ph} = h\nu = hc/\lambda$ $\quad$ $h = 6.626\times10^{-34}$J•s

4. (OYO) What is the energy and wavelength of a orange light photon with a frequency of 4.95×10^{14}Hz?

(Ans: 3.28×10^{-19}J/6.06×10^{-7}m)

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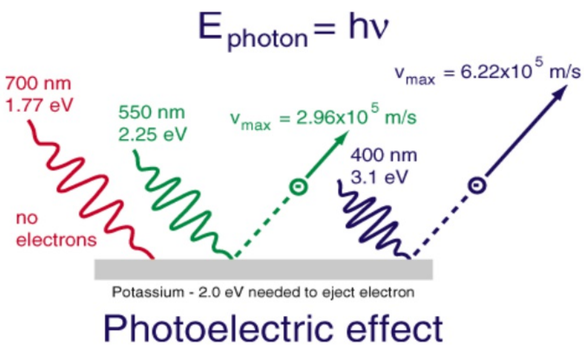
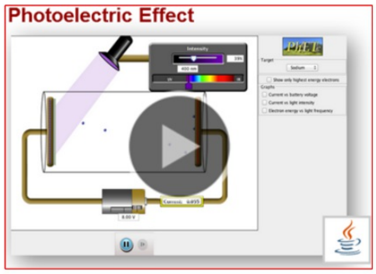
Photoelectric Effect

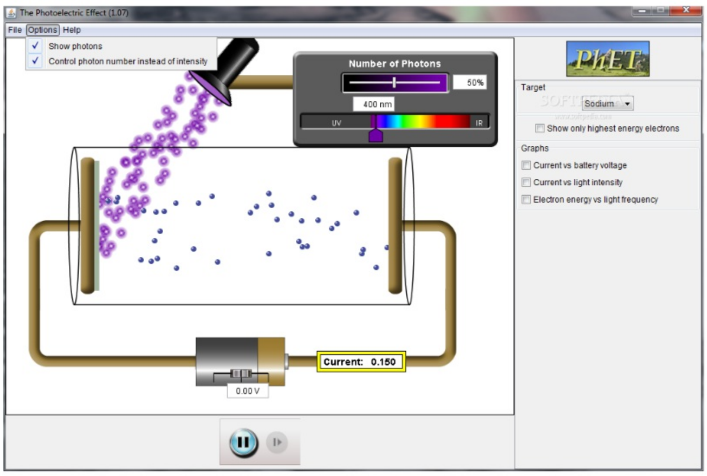
- The photoelectric effect describes the collision between a photon and an electron in a metal.
- If the photon has ENOUGH energy, it will eject the electron from the surface of the metal.
- At the moment of collision, the photon is considered a particle with no mass.
- Every metal needs a photon of specific energy for the effect to ocurr.
- Max Planck proposed hot objects release light in the form of packets called quanta.
  - quantum of energy - minimum amount of energy that can be gained or lost by an atom.
  - Planck's constant, 6.626×10^{-34} J●s
- Photon - a particle of EM radiation with no mass and carrying a quantum of *E*.
- Albert Einstein - light has a dual wave-particle nature. Used Planck's theory to explain the photoelectric effect.
- Arthur Compton (American physist) conclusively proved the particle nature of light when he demonstrated a photon hitting and moving an electron. This is called the Compton effect.

Photoelectric Effect



KE = 1/2mv²





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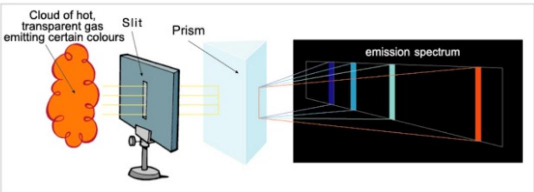
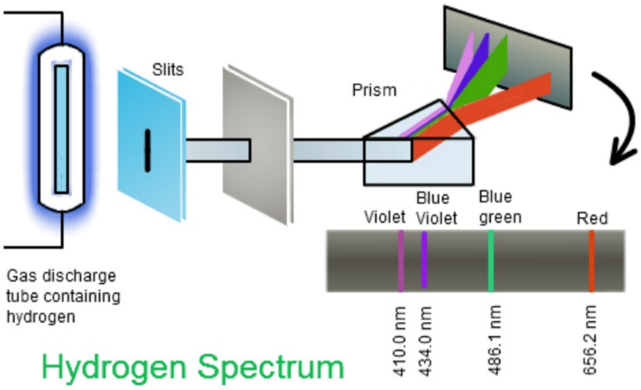
Bohr Model of Atom

- The hydrogen atom line emission spectrum is made from the excitation of the hydrogen atom
- Gaseous H is in its ground state, but when energy is added it becomes excited (excited state).
- Electrons in the excited state must release energy to move back to their ground state.
  - the energy released is in the form of a photon of a specific energy.
- Using a prism, the photons are separated into individual colors of light (line-emission spectrum).
- Each pure substance has a unique, identifying line-emission spectrum.
- White light produces a continuous spectrum (rainbow means happy!).

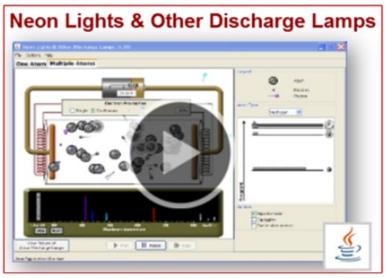


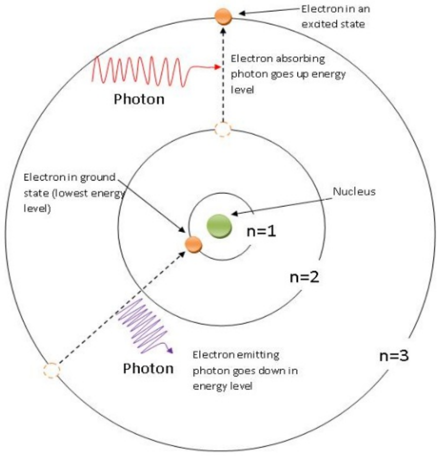


Bohr Model of Atom

Niels Bohr

- Linked the electrons in hydrogen to photon emission.
- Electrons were quantized (had specific energies) which means they could only exist in certain orbits around the nucleus.
- Labeled each orbit (energy level) by a quantum number (n). Lowest (ground state) $n = 1$, orbit closest to the nucleus.
- When a quantum of energy equivalent to the change in energy between energy levels is absorbed the electron "jumps" to a higher energy level ($n = 2, 3, 4, 5, 6$, or 7) where it is considered to be in an excited state (larger orbits further away from the nucleus).
- When the electron falls to a lower energy level a photon equivalent in energy is emitted.
- Only worked for hydrogen not for any other atom.
- Bohr used this model and Planck's equation to calculate the frequencies observed in the lines spectrum of hydrogen.





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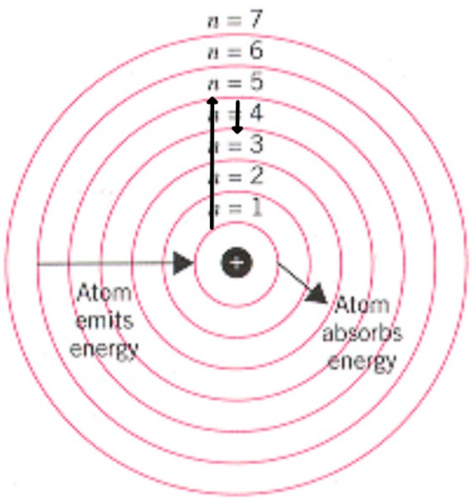
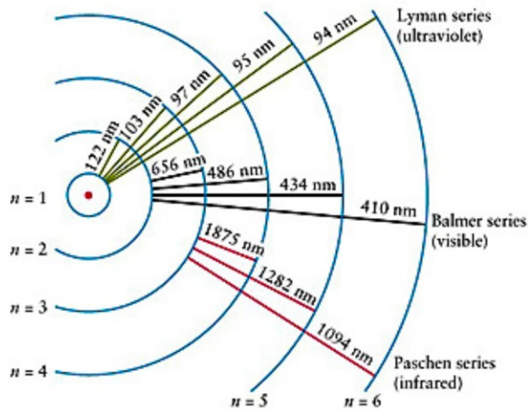
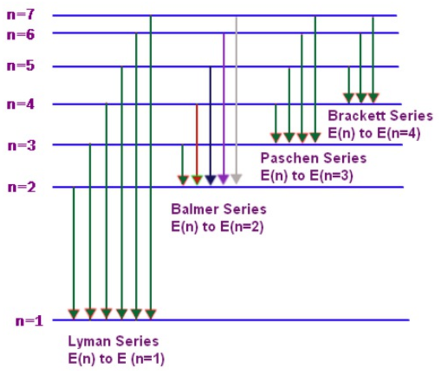
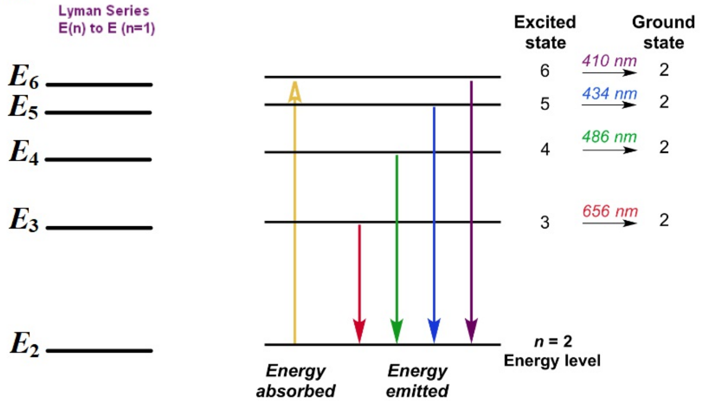
Bohr Model of Atom



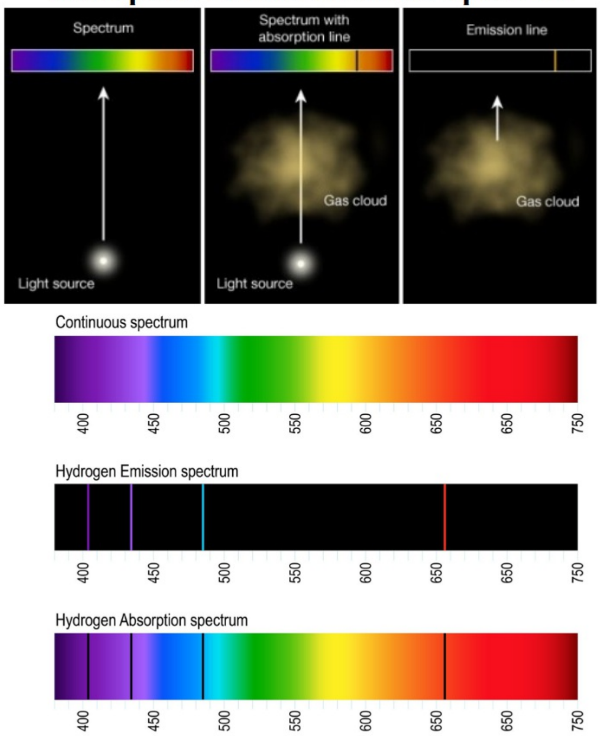


Line Emission Spectrum





Absorption vs. Line Emission Spectrum



Line Emission Spectrum

Laser - **L**ight **A**mplification by **S**timulated **E**mission of **R**adiation

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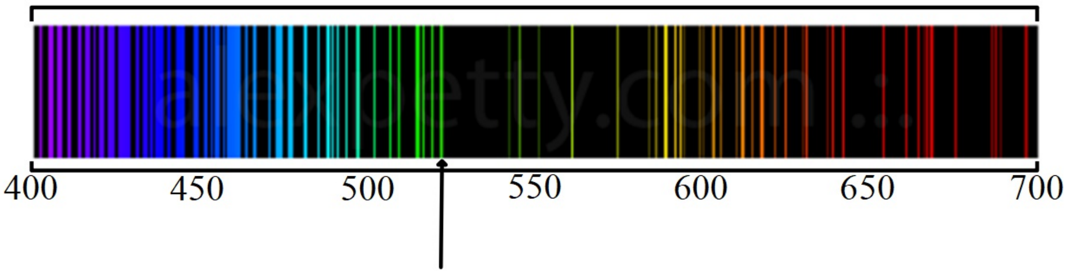
Line Emission Spectrum

Ar, Argon

$c = 3.00x10^8$ m/s $h = 6.626x10^{-34}$J•s



$\lambda =$

$\upsilon =$

$E =$

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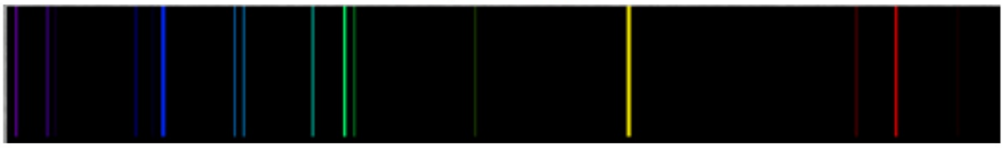
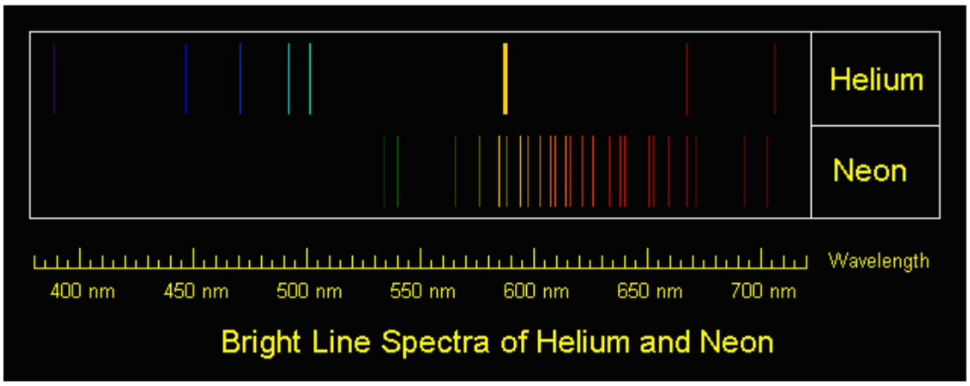
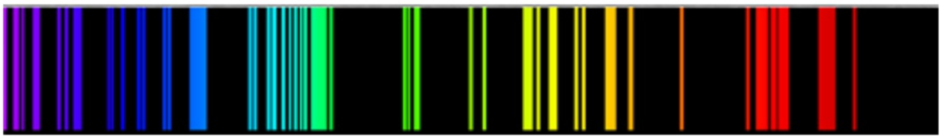
H_2, Hydrogen

He, Helium

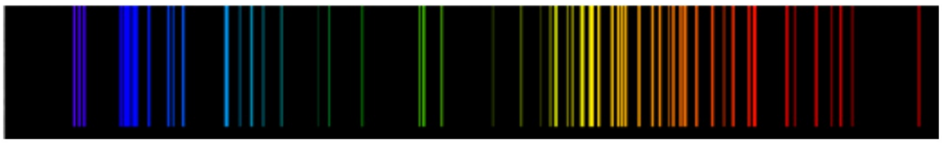




N_2 - Nitrogen



Ne - Neon



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Objectives

Quantum Model of Atom

1) Discuss Louis de Broglie's role in the development of the quantum model of the atom.
2) Compare and contrast the Bohr model and the quantum model of the atom.
3) Explain how the Heisenberg uncertainty principle and the Schrodinger wave equation led to the idea of atomic orbitals.
4) List the four quantum number and describe their significance.
5) Relate the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.

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Quantum Model

- Electrons as waves
  - Photoelectric effect and hydrogen line-emission spectrum - light both a particle and wave.
  - Louis de Broglie - electrons like waves (matter waves) in Bohr's quantized orbits. Electrons were waves confined to a space and had certain frequencies. Quantization of energy.
  - Electrons diffract (bend) (use wave nature of electrons for electron microscopes)
  - All moving objects have wavelike behavior. The larger the object, the smaller the wavelength
    - 50 gram golf ball has a wavelength of $3x10^{-34}$m (too small to measure)
- The Heisenberg Uncertainty Principle
  - Detection involved photons. Photons MOVED electrons
  - Uncertainty in trying to locate an electron.
  - Principle: Impossible to determine simultaneously the position/velocity of an electron (particle)

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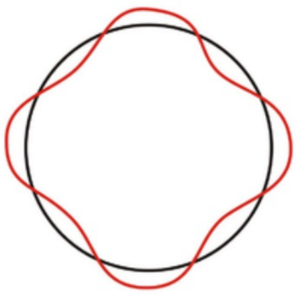
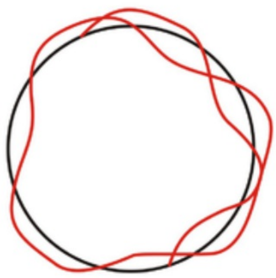
Quantum Model

- The Schrodinger Wave Equation (Erwin Schrodinger)
  - Developed an equation to quantify the electrons as waves
  - Only waves of specific energies (frequencies) satisfied the equation
  - Regions called orbitals (90% probability of finding electron in orbital)
  - Along with the Uncertainty Principle, the wave equation led to the Quantum theory
  - Quantum theory of the atom - mathematically describes wave properties of electrons (particles)
  - Electron Cloud - 90% probability of finding electron within this region



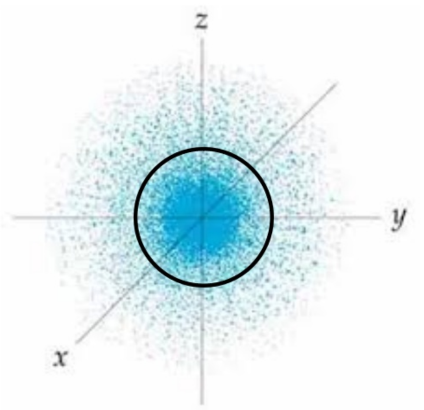
This wave fits perfectly on the circle

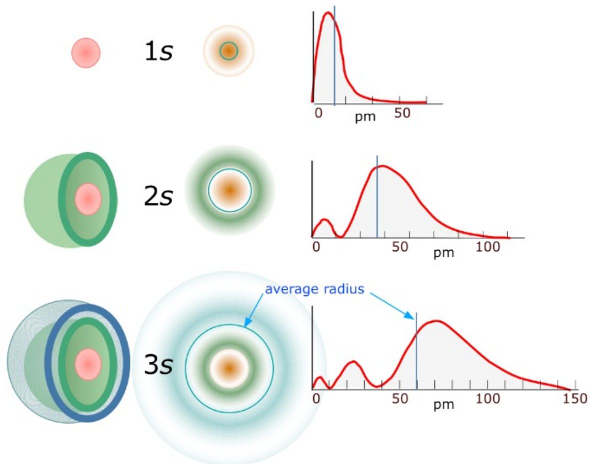


This wave, however, doubles over on itself

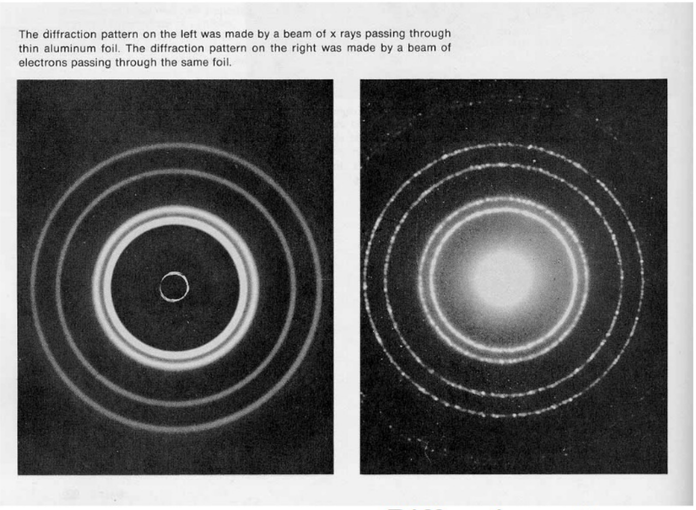
Quantum Model

Orbitals and Probability



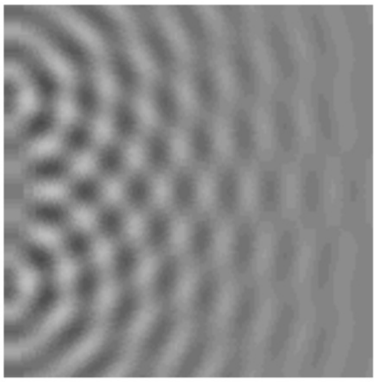


<u>Quantum Model</u>



Diffraction pattern of a beam of x-rays passing through Al foil.

Diffraction pattern of a beam of electrons passing through Al foil.



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Quantum Numbers

- specify the properties of atomic orbitals and the properties of the electrons in orbital

- Principal Quantum Number (*n*)
  - main energy level occupied by electron
  - values: 1-7 (n = 1: ground state (g.s))
  - as n increases, distance from electron to nucleus increases (energy)

- Angular momentum quantum number (*l*)
  - the shape of the orbital (sublevels)
  - types of sublevels (orbitals): *s*, *p*, *d*, and *f*
  - number of orbital shapes possible is equal to n
    - n = 1, one type of orbital (*s*)
    - n = 2, two types of orbitals (*s* and *p*)
    - n = 3, three types of orbitals (*s*, *p*, and *d*)
    - n = 4, four types of orbital (*s*, *p*, *d*, and *f*)
  - values for l are 0, 1, 2, and 3
    - l = 0: *s* type orbital (1 orbital)
    - l = 1: *p* type orbital (3 orbitals)
    - l = 2: *d* type orbital (5 orbitals)
    - l = 3: *f* type orbital (7 orbitals)

n = 2
s, p
8e⁻

n = 3
s, p, + d
9 orbital
$n^2 = 3^2 = 9$
18 electrons

n = 2 2 types orbitals
n^2 = 4 # of orbitals

- number of **orbitals per energy level**, n^2 $2n^2$ – # of e⁻
  - when n = 2 then $n^2 = 2^2$ = 4 orbitals in the second energy level
  - when n = 2 there are 2 types of sublevels (*s* and *p*) containing 4 orbitals (one from the *s* type sublevel and three from *p* type orbital).
- number of electrons per energy level is $2n^2$.

n	l	type (sub-shell)	# e⁻
1	0	s	2
2	0, 1	s, p	8
3	0, 1, 2	s, p, d	18
4	0, 1, 2, 3	s, p, d, f	32
5	0, 1, 2, 3	s, p, d, f	32

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Quantum Numbers

1, 0, 0, +½
n, ℓ, m_ℓ, m_s

s 1L / 0   1, 0, 0, −½

p – –1, 0, +1

d – –2, –1, 0, +1, +2

f – –3, –2, –1, 0, +1, +2, +3

- magnetic quantum number (m_l)
  - orientation of orbital around nucleus
    - number of orbitals (each holds 2 electrons) in a type
      - (s) ▪ $l = 0 \rightarrow m_l =$ 0; *s* type with 1 orbital
      - (p) ▪ $l = 1 \rightarrow m_l =$ -1,0,1; *p* type with 3 orbitals
      - (d) ▪ $l = 2 \rightarrow m_l =$ -2,-1,0,1,2; *d* type with 5 orbitals
      - (f) ▪ $l = 3 \rightarrow m_l =$ -3,-2,-1,0,1,2,3; *f* type with 7 orbitals
- spin quantum number (m_s)
  - (+1/2) clockwise spin and (-1/2)counterclockwise spin
  - only two values and an orbital has no more than 2 electrons
  - creates opposite magnetic fields, so an attraction exists between the two electrons in the orbital. Otherwise electrostatic repulsion would never allow the electrons to be in the same orbital.

+1/2 ↑   ↓ -1/2

n=2
s, p

	L: 3	n, ℓ, m_ℓ, m_s	
$1s^1$	1st electron:	1, 0, 0, +½	1L / 1s
$1s^2$	2nd elelectron:	1, 0, 0, −½	
$2s^1$	3rd electron:	2, 0, 0, +½	1 / 2s

orbital diagram: 1L / 1s   1 / 2s $1s^2 2s^1$

Quantum Numbers

Electron Topic#5 Sample Exercise WS#2 - Energy Levels

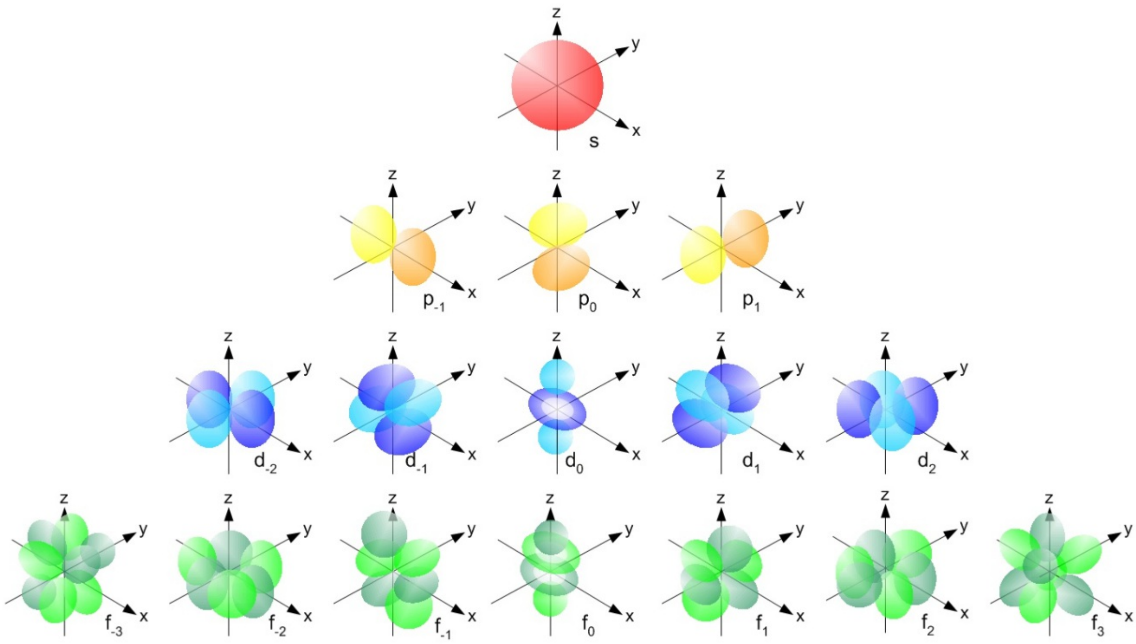
1. How many subshells (types of orbitals) are in shell two? $n = 2$, 2 subshells (s, p)
   How many orbitals can shell two have? $n^2 = 2^2 = 4$ orbitals
   How many electrons can shell two have? $2n^2 = 2(2)^2 = 8$ electrons
2. How many subshells (types of orbitals) are in shell five? $n = 5$, 5 subshells (s, p, d, f)
   * How many orbitals can shell five have? $n^2 = 5^2 = 25$ orbitals (1 + 3 + 5 + 7 = 16 orbitals
   * How many electrons can shell five have? $2n^2 = 2(5)^2 = 50$ electrons
   (32 electrons)

*When n = 4, 5, 6 or 7 number of sublevels is 4 (s, p, d, f) and electrons is 32.

s	p	d	f
1	2 3 4	5 6 7 8 9	10 11 12 13 14 15 16
$2e^-$	$6e^-$	$10e^-$	$14e^-$

$32e^-$

Quantum Numbers



Different Orbital Shapes

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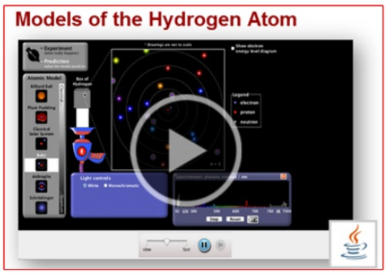
Quantum Numbers

Principal Energy Level (n)	Number of Orbitals Present s	p	d	f	Total Number of Orbitals (n^2)	Maximum Number of Electrons ($2n^2$)
1	1	-	-	-	1	2
2	1	3	-	-	4	8
3	1	3	5	-	9	18
4	1	3	5	7	16	32



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Electron Configuration/Orbital Diagrams

Electron configuration - an arrangement of electrons in an atom.
Orbital diagram - a graphical representation of the electrons in an atom obeying the rules governing placement of electrons.

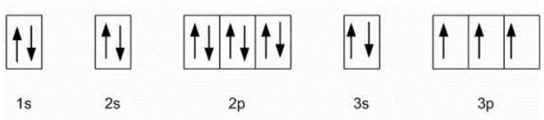
Rules

(1) **Aufbau Principle** - an electron must occupy the lowest energy orbital that can recieve it.

$$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^67s^25f^{14}6d^{10}$$

(2) **Pauli Exclusion Principle** - no two electrons in the same atom can have the same set of four quantum numbers (shothand: only two electrons per orbital with opposite spins.)

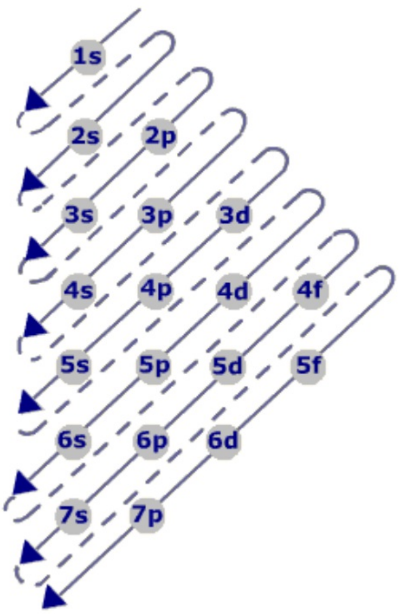
(3) **Hund's Rule** - orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all electrons in singly occupied orbitals must have the same spin state. (shorthand: in *p*, *d*,and *f* type orbitals, each orbital must have one electron in it before a second electron can be placed there.)

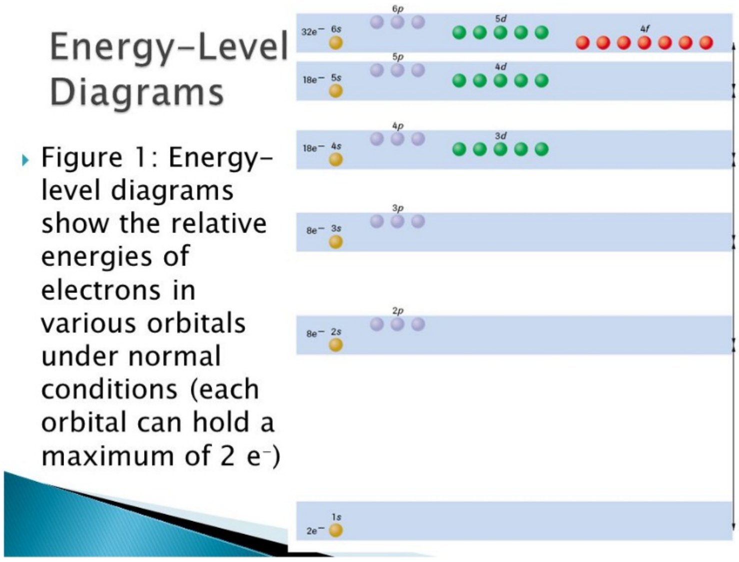


The Electron
Topic#5

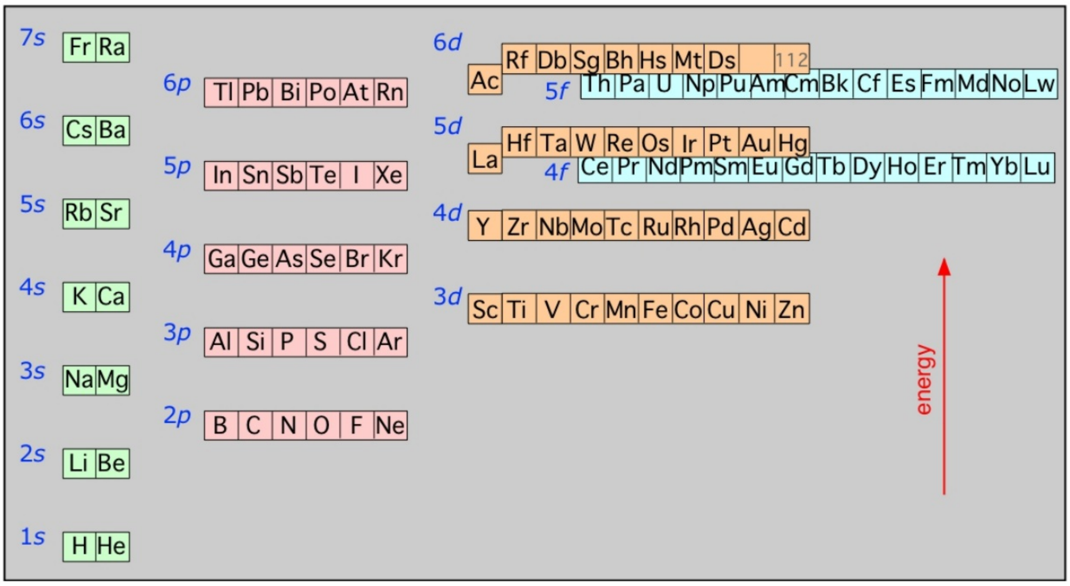
Electron Configuration/Orbital Diagrams

$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^67s^25f^{14}6d^{10}$





Electron Configuration/Orbital Diagrams



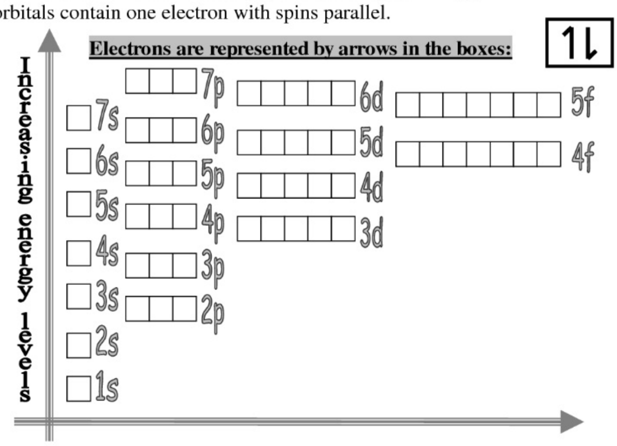
Electron Configuration/Orbital Diagrams

Electronic Configuration – Aufbau Diagrams

Task: To show how the electronic configuration of an element can be represented in a graphical form. Complete an Aufbau Digram for an element from each of period 2, 3, 4 and 5.

Info: There are three principles which govern the filling of orbitals by electrons:

1) Aufbau Principle: Electrons enter orbitals of lowest energy first.
2) Pauli's exclusion principle: An atomic orbital contains a maximum of two electrons.
3) Hund's Rule: When electrons occupy orbitals of equal energy, one electron enters each orbital until all the orbitals contain one electron with spins parallel.



Electron Configuration/Orbital Diagrams

Sample Problems WS#2 - Electron Configuration/Orbital Diagram:

3. (a) C *Z* = ___ ___ ___ ___ ___ ___
1s 2s 2p

electron config: ______________________________ # of unpaired e^- : _____

(b) Mg *Z* = ___ ___ ___ ___ ___ ___ ___
1s 2s 2p 3s

electron config: ______________________________ # of unpaired e^- : _____

(c) S *Z* = ___ ___ ___ ___ ___ ___ ___ ___ ___ ___
1s 2s 2p 3s 3p

electron config: ______________________________ # of unpaired e^- : _____

(d) Cr *Z* = ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___
1s 2s 2p 3s 3p 4s 3d

electron config: ______________________________ # of unpaired e^- : _____

Electron Configuration/Orbital Diagrams

(e) Ta Z = ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___

1s 2s 2p 3s 3p 4s 3d

___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___

4p 5s 4d 5p 6s

___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___

4f 5d

electron config: ______________________________ # of unpaired e^- : _____

(f) Sm Z = ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___

1s 2s 2p 3s 3p 4s 3d

___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___

4p 5s 4d 5p 6s

___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___

4f 5d

electron config: ______________________________ # of unpaired e^- : _____

4. What element has the electron configuration:
   a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$?
   b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{10}$?

The Electron

Topic#5

<u>*s,p,d,* and *f*-block Periodic Table Activity</u>

1) Select **4** colors and create a key with those four colors
   a) **1st** color: ***s*-block** b) **2nd** color: ***p*-block** c) **3rd** color: ***d*-block** d) **4th** color: ***f*-block**
2) Number the columns (groups) 1-18 and the rows (energy levels) 1-7.
3) Color columns **1-2** the ***s*-block** color. Color columns **3-12** the ***d*-block** color. Color columns **13-18** the ***p*-block** color.
4) Color the **lanthanide** and **actinide series** the ***f*-block** color.
5) Put **[He]** to the left of **Li**, **[Ne]** to the left of **Na**, **[Ar]** to the left of **K**, put **[Kr]** to the left of **Rb**, **[Xe]** to the left of **Cs**, and **[Rn]** to the left of **Fr**.
6) Create a key for the Noble gas electron configurations listed below:

[He] = $1s^2$	**[Kr]** = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
[Ne] = $1s^2 2s^2 2p^6$	**[Xe]** = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$
[Ar] = $1s^2 2s^2 2p^6 3s^2 3p^6$	**[Rn]** = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$

7) Element boxes: put a **1** in **H/He**, **2** in **Li/B, 3** in **Na/Sc/Al**, **4** in **K/Y/Ga/Ce**, **5** in **Rb/La/In/Th**, **6** in **Cs/Ac/Tl**, and **7** in **Fr.** Circle all of the numbers you just put in the element boxes.
8) Put a **1** above **H**, **2** above **Be**, **3** above **B**, **4** above **C**, **5** above **N**, **6** above **O**, **7** above **F**, **2** above **He**, and an **8** to the right with an arrow towards **Ne**.
9) These are the <u>valence electrons</u> (outer *s* and *p* electrons) for the groups under the numbers. So, **B** has a valence of **3** and **Ne 8**.
10) The elements in the ***s*-block** and ***p*-block** are called main-group or representative elements.

The Electron

Electron Configuration/Orbital Diagrams **Topic#5**

[Ne] = $1s^2 2s^2 2p^6$
[Ar] = $1s^2 2s^2 2p^6 3s^2 3p^6$
[Kr] = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
[Xe] = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$
[Rn] = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$

5. Draw the electron configuration, orbital diagram, Noble gas electron configuration, and valence electrons:

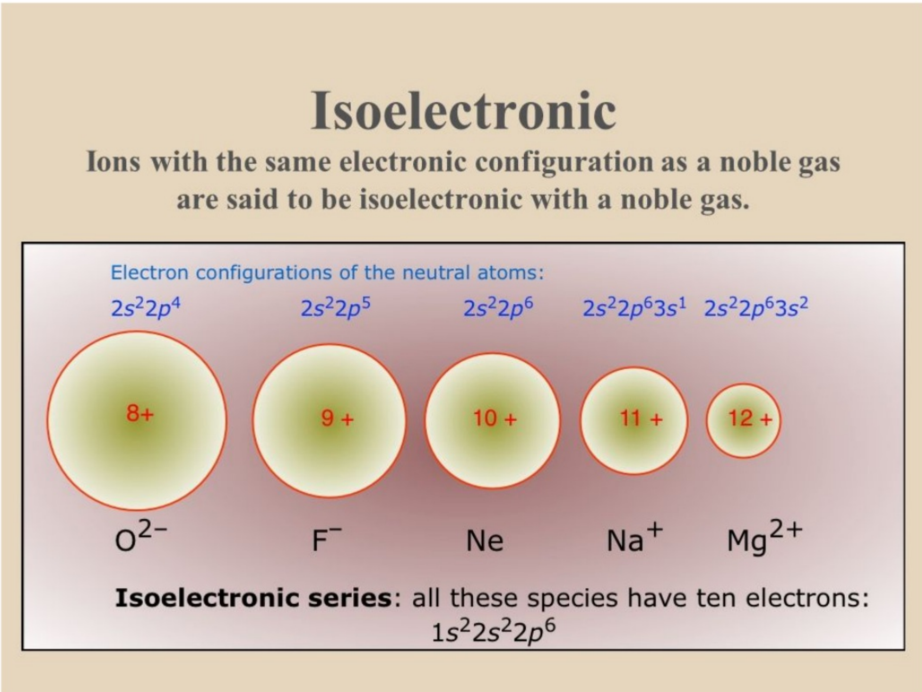
a. Ca

b. P

c. Co

d. Pd

e. Pb


Isoelectronic
Ions with the same electronic configuration as a noble gas are said to be isoelectronic with a noble gas.
Electron configurations of the neutral atoms:
$2s^2 2p^4$
$2s^2 2p^5$
$2s^2 2p^6$
$2s^2 2p^6 3s^1$
$2s^2 2p^6 3s^2$
8+
9 +
10 +
11 +
12 +
O^{2-}
F^-
Ne
Na^+
Mg^{2+}
Isoelectronic series: all these species have ten electrons:
$1s^2 2s^2 2p^6$

The Electron
Topic#5
Line Emission

(Use in periodic table)

Na - $1s^2 2s^2 2p^6 3s^1$

Valence Electrons - outer **s** and **p** electrons

Valence Electrons: ____________

How many electrons would this atom need to gain/lose to become an ion?__________

Calculation for charge: ______________________

What is the charge and symbol for the ion? ______ ______

The Electron
Topic#5
Line Emission

(Use in periodic table)

<u>Sample Problem 4.6</u> – Valence Electrons and Electron Dot Diagram

Determine the valence electrons and show the electron diagram of each of the following elements.

	Val Electrons	e^- Dot		Val Electrons	e^- Dot		Val Electrons	e^- Dot
(1) Na	__________	_______	(5) Ra	__________	_______	(8) Si	__________	_______
(2) Cs	__________	_______	(6) Rn	__________	_______	(9) Cl	__________	_______
(3) Mg	__________	_______	(7) O	__________	_______	(10) C	__________	_______
(4) Al	__________	_______						