Heat of formation / enthalpy of formation!

34.086 g/mol

$$\underset{-20.50}{2H_2S(g)} + \underset{0}{3O_2(g)} \rightarrow \underset{-285.8}{2H_2O(l)} + \underset{-296.8}{2SO_2(g)} \quad \Big]\ \Delta H_f^\circ,\ kJ/mol\ \text{(Appendix C)}$$

What is the enthalpy change at standard conditions when 25.0 grams of hydrogen sulfide gas is reacted?

1 - Use Hess's Law to find the enthalpy change for the reaction as written.
2 - Convert 25.0 g hydrogen sulfide to moles using FORMULA WEIGHT.
3 - Convert moles hydrogen sulfide to enthalpy change using THERMOCHEMICAL EQUATION

(1) $\Delta H = \sum \Delta H_f^\circ,\text{products} - \sum \Delta H_f^\circ,\text{reactants}$

$= [2(-285.8) + 2(-296.8)] - [2(-20.50) + 3(0)]$

$= -1124.2\ kJ$

So the thermochemical equation is ...

$$2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(l) + 2SO_2(g);\ \Delta H = -1124.2\ kJ$$

(2) $34.086\ g\ H_2S = mol\ H_2S$ (3) $2\ mol\ H_2S = -1124.2\ kJ$

$$25.0\ g\ H_2S \times \frac{mol\ H_2S}{34.086\ g\ H_2S} \times \frac{-1124.2\ kJ}{2\ mol\ H_2S} = \boxed{-412\ kJ}$$

(2) (3)

propane

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)\ ;\ \Delta H = -2043\ kJ$$

Calculate the volume of propane gas at 25.0 C and 1.08 atm required to provide 565 kJ of heat using the reaction above.

1 - Convert the 565 kJ energy requirement to moles propane using THERMOCHEMICAL EQUATION
2 - Convert moles propane to volume using IDEAL GAS EQUATION

(1) $mol\ C_3H_8 = -2043\ kJ$

Since the reaction is the system here, the energy requirement gets a negatice sign. The reaction gives up this ampunt of energy!

$$-565\ kJ \times \frac{mol\ C_3H_8}{-2043\ kJ} = 0.2765540871\ mol\ C_3H_8$$

(2) $PV = nRT$

$$V = \frac{nRT}{P}$$

$P = 1.08\ atm$ $\quad R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$

$T = 25.0\ °C = 298.2\ K$

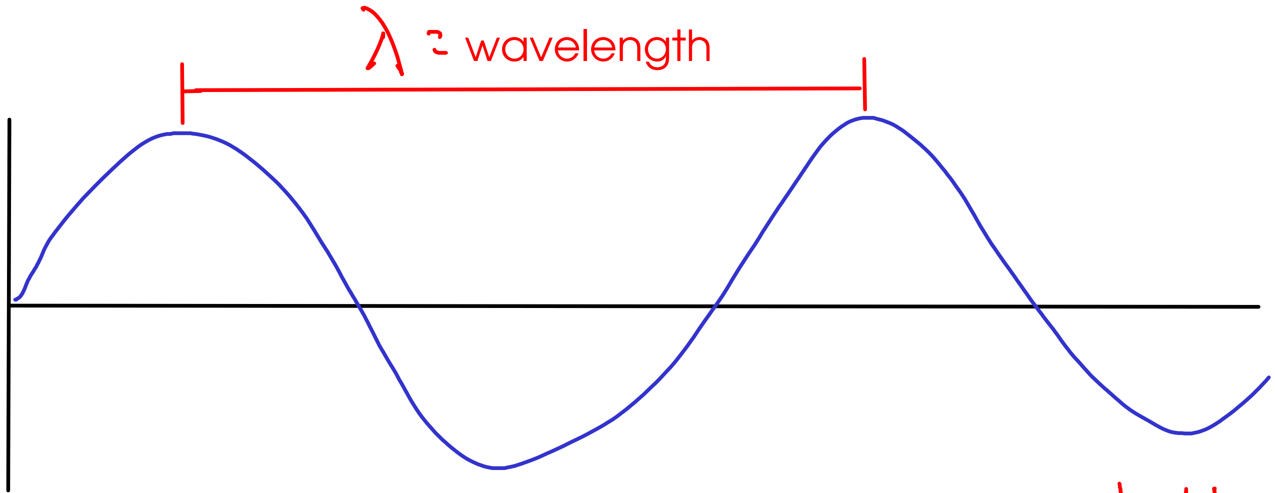
$n = 0.2765540871\ mol\ C_3H_8$

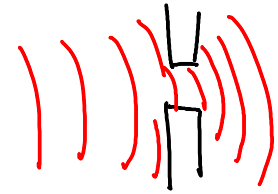
$$V = \frac{(0.2765540871\ mol\ C_3H_8)\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)}{(1.08\ atm)}$$

$= 6.27\ L$ of C_3H_8 @ $25.0\ °C$ + $1.08\ atm$

END OF CHAPTER 6

LIGHT





Diffraction

frequency = wavelengths / time = ν s^{-1} : Hertz, Hz

- Light has properties of WAVES such as DIFFRACTION (it bends around small obstructions).

- Einstein noted that viewing light as a particle that carried an energy proportional to the FREQUENCY could explain the PHOTOELECTRIC EFFECT!

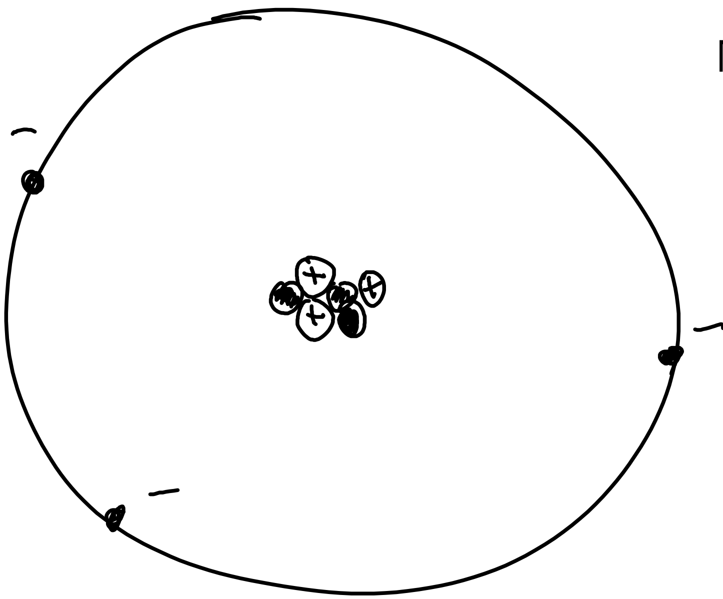
$$E_{photon} = h\nu$$

h — Planck's constant: 6.63×10^{-34} J·s

photon = particle or packet of light

(The photoelectric effect is the emission of electrons from a metal caused by exposure to light. Einstein discovered that if the light were not of the correct FREQUENCY, increasing the INTENSITY of the light would not cause electron emission. He concluded that individual photons must have enough energy to excite an electron - i.e. they must have the appropriate frequency.)

The photoelectric effect and Einstein's ideas about the energy content of light led us to discover a new model for the atom! How? Let's start with the nuclear model:



Nuclear model:

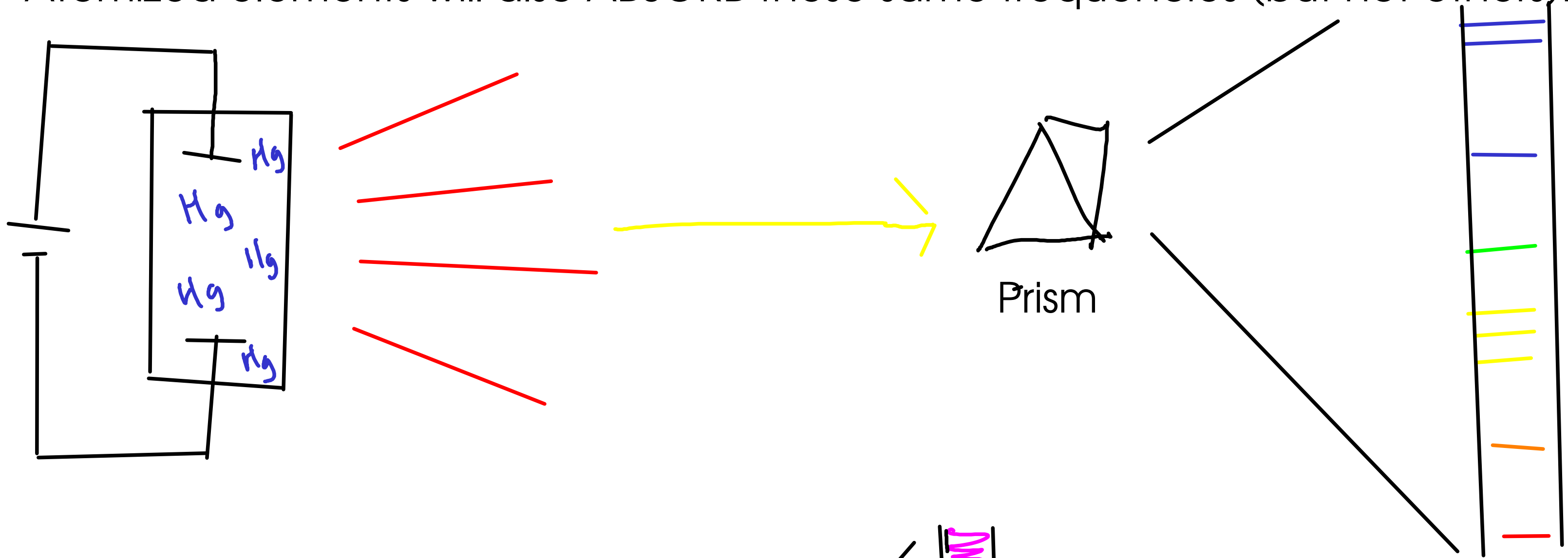
- Protons and neutrons in a dense NUCLEUS at center of atom
- Electrons in a diffuse (mostly empty) ELECTRON CLOUD surrounding NUCLEUS.

... so what's wrong with the nuclear model? Among other things, it doesn't explain ...

ATOMIC LINE SPECTRA

- if you take element and ATOMIZE it, if excited by energy it will emit light at unique frequencies. The set of emitted frequencies is called an ATOMIC LINE SPECTRUM.

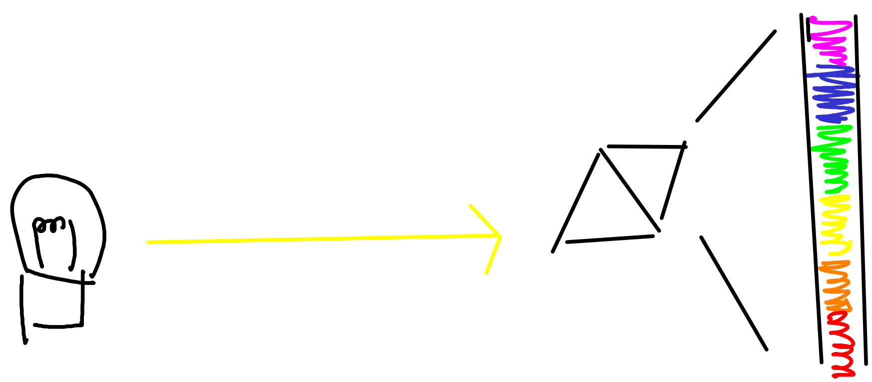
- Atomized elements will also ABSORB these same frequencies (but not others)!



p265 (9th)

p269 (10th)

Atomic line spectrum Unique to each element.

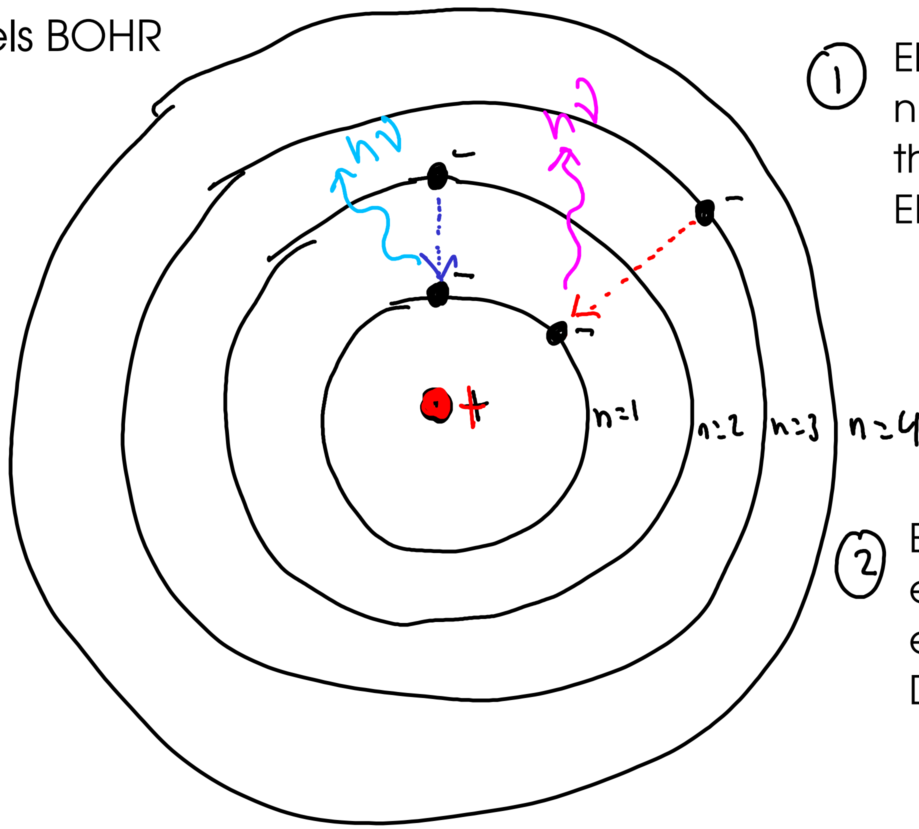


Continuous spectrum (from white light)

... so, why don't atoms by themselves emit continuous spectra like a flashlight would?

- The regular patterns of emission and absorption of light by atoms suggest that the electron cloud has some sort of regular structure. The specific frequencies of light emitted and abosrbed relate to specific values of ENERGY in the electron cloud.

Niels BOHR

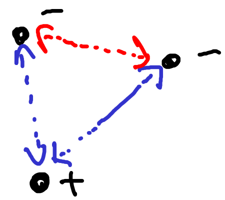


(1) Electrons can't be just ANYWHERE around a nucleus. They can exist only at certain distances from the nucleus. These distances correpsond to certain ENERGIES and are called ENERGY LEVELS!

(2) Electrons CAN move (transition) between different energy levels by gaining or losing exactly enough energy to get into the new energy level. This was a DIRECT transition .

Bohr's model was the first proposal that predicted the existence of atomic line spectra, and it exactly predicted the spectra of hydrogen and "hydrogen-like" (i.e. one-electron) species.

The spectra were "off" for multi-electron atoms.



Multi-electron atoms have interactions between electrons, not just interactions between electrons and nucleus!

- The additional interactions in multi-electron atoms introduced added complexity to the model of the atom! Bohr's model was too simple.

- Improvements in Bohr's model came from treating electrons as WAVES.

de Broglie relationship

Calculates the WAVELENGTH of ... matter?

$$\lambda = \frac{h}{m \times v}$$

- h: Planck's constant
- λ: wavelength
- m: mass
- v: velocity (m/s)

... for very large particles, the wavelength is very small.

Quantum mechanics treats the electrons as waves and models THAT behavior!

- To describe the electrons, we use WAVEFUNCTIONs - which are mathematical descriptions of the behavior or electrons.

- The wavefunction describes the probability of finding an electron in a given space

- For larger objects, the wave behavior isn't very important .... and quantum mechanics becomes traditional Newtonian physics.

When we talk about describing electrons ... we will talk about the PARAMETERS that go into this WAVEFUNCTION ... without doing the actual math.

- There are FOUR of these parameters. (the Bohr model had only one!)

- The parameters are called "quantum numbers"

1. Principal quantum number
2. Angular momentum quantum number
3. Magnetic quantum number
4. Spin quantum number

- Giving the four parameters will uniquely identify an electron around an atom. No two electrons in the same atom can share all four. These parameters are called QUANTUM NUMBERS.

(1) PRINCIPAL QUANTUM NUMBER (n):

- "energy level", "shell"
- Represents two things:
  * The distance of the electron from the nucleus.
  * Energy. "n" is one factor that contributes to the energy of the electron.

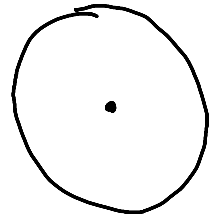
$n = 1, 2, 3, 4, \ldots$ (integers)

(2) ANGULAR MOMENTUM QUANTUM NUMBER: ℓ

- "subshell"
- Represents the SHAPE of the region of space where the electron is found.
  - (Bohr assumed CIRCULAR orbits for electrons ... but there are more possibilities.)
- "l" also contributes ENERGY. Higher values for "l" mean the electron has higher energy.

$\ell = 0$ to $n-1$, integers

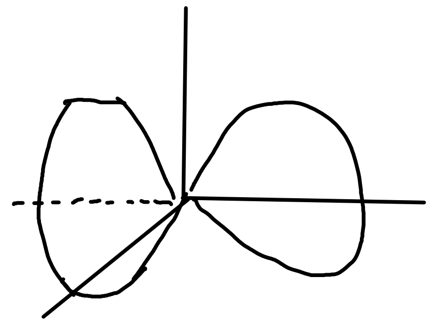
$n = 1$; $\ell = 0$



"l" = 0 ; spherical subshell

Also called an "s" subshell.

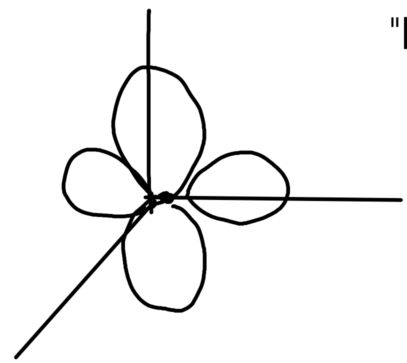
$n = 2$; $\ell = 0, 1$



"l"=1 ; dumbbell shaped

Also called a "p" subshell

$n = 3$, $\ell = 0, 1, 2$



"l"=2; flower-shaped

Also called a "d" subshell

(p 285, 3-D pictures of subshells)

Higher values for "l" translate to higher energies for the electron!

For convenience, and partially for historical reasons, we use letters to designate the different subshells. (p 290, 10th ed)

$\ell = 0$ "s"

$\ell = 1$ "p"

$\ell = 2$ "d"

$\ell = 3$ "f"

$\ell = 4$ "g"

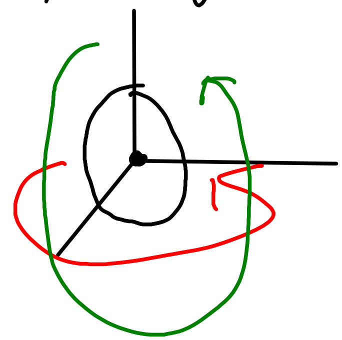
↓ The rest follow the alphabet

③ MAGNETIC QUANTUM NUMBER m_l

- Represents the ORIENTATION of a subshell in 3D space.

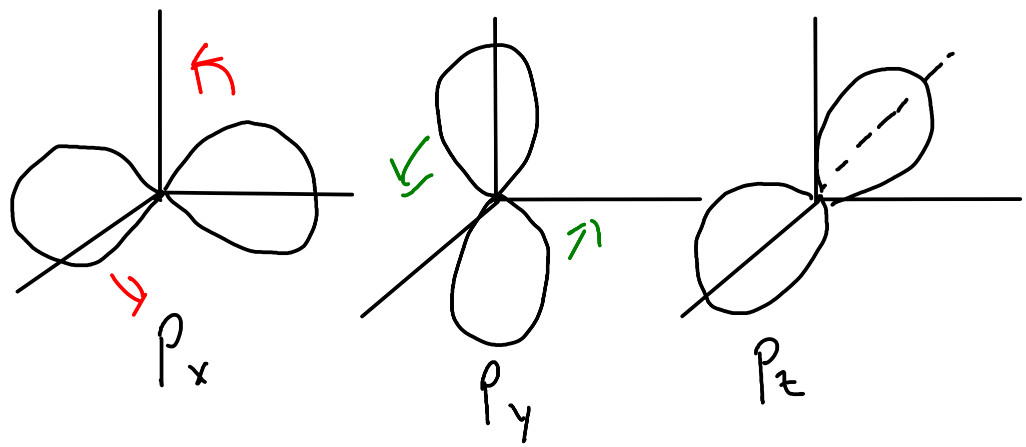
$m_l = -l$ to $+l$, integers

$l = 0$, $m_l = 0$
"s"



There is only one possible orientation for an "s" subshell!

$l = 1$; $m_l = -1, 0, 1$
"p"



There are THREE possible orientations for a "p" subshell!

$l = 2$, $m_l = -2, -1, 0, 1, 2$ (five orentations) ← [picture p285, p290 (tenth)]
"d"

$l = 3$, $m_l = -3, -2, -1, 0, 1, 2, 3$ (seven orentations)
"f"

... all the arrangements of a single subshell have the same energy. The magnetic quantum number DOESN'T contribute to the energy of an electron.

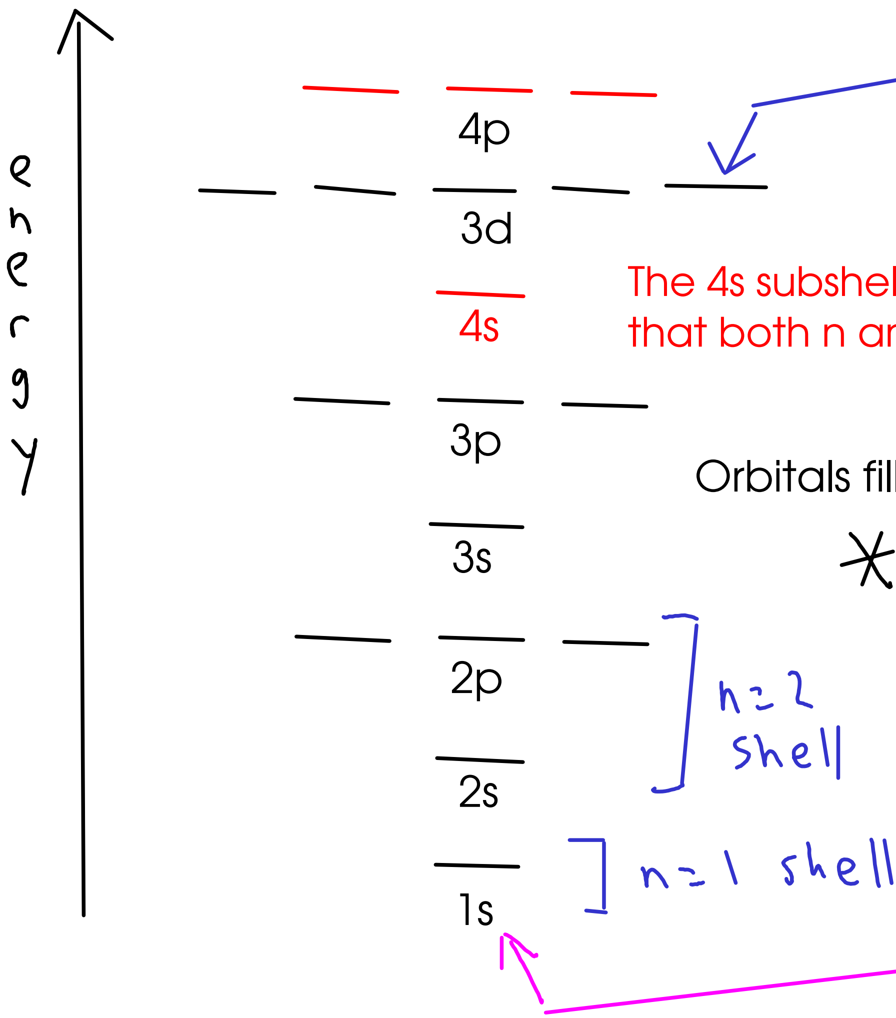
④ (MAGNETIC) SPIN QUANTUM NUMBER: m_s

$m_s = -1/2$ or $+1/2$ "spin down" or "spin up"

- An ORBITAL (region with fixed "n", "l" and "ml" values) can hold TWO electrons.

ORBITAL DIAGRAM

- A graphical representation of the quantum number "map" of electrons around an atom.



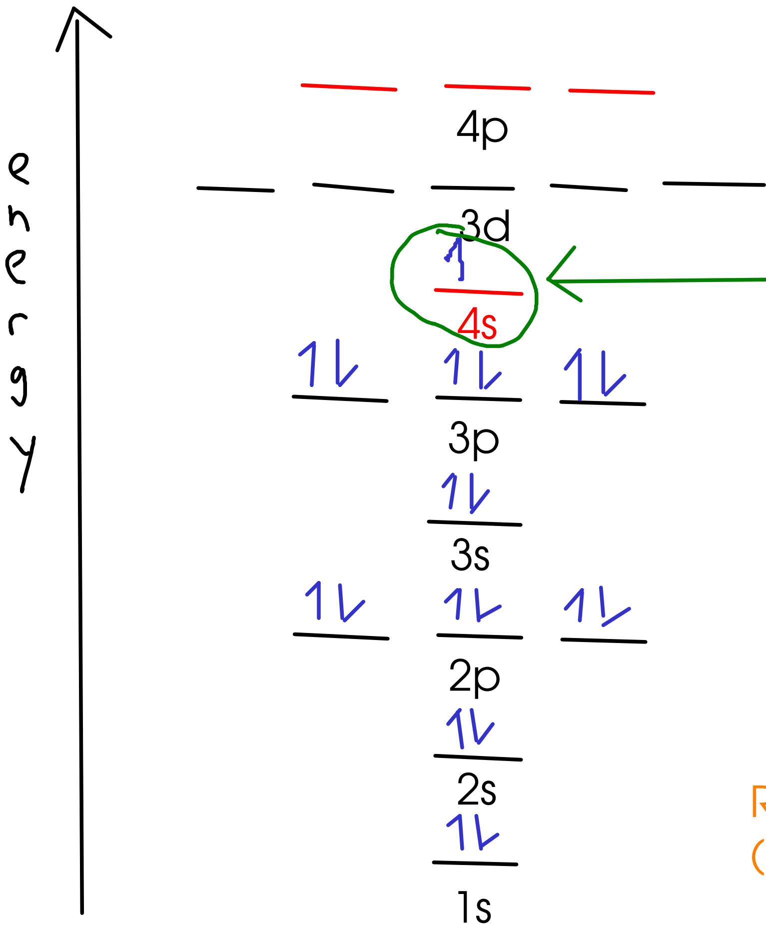
Each blank represents an ORBITAL, and can hold two electrons.

The 4s subshell is lower energy than the 3d subshells .... remember that both n and "l" contribute to energy!

Orbitals fill in order: The lower energy orbitals fill first.

✱ Where there's more than one orbital with the same energy, electrons don't pair up until each one has one electron in it!

1s: This means the first shell (n=1) and the "s" subshell ("l"=0)



How would an orbital diagram for the element POTASSIUM look?

K, Z = 19

atomic number

Electrons in the outermost shell of an atom are called VALENCE electrons. THESE electrons are normally involved in chemical bonding.

Remember: Potassium tends to lose a single electron (forming a cation) in chemical reactions. K^+

A note on chemical bonding and electron arrangement:

- Filled and half-filled subshells seem to be preferred by atoms.