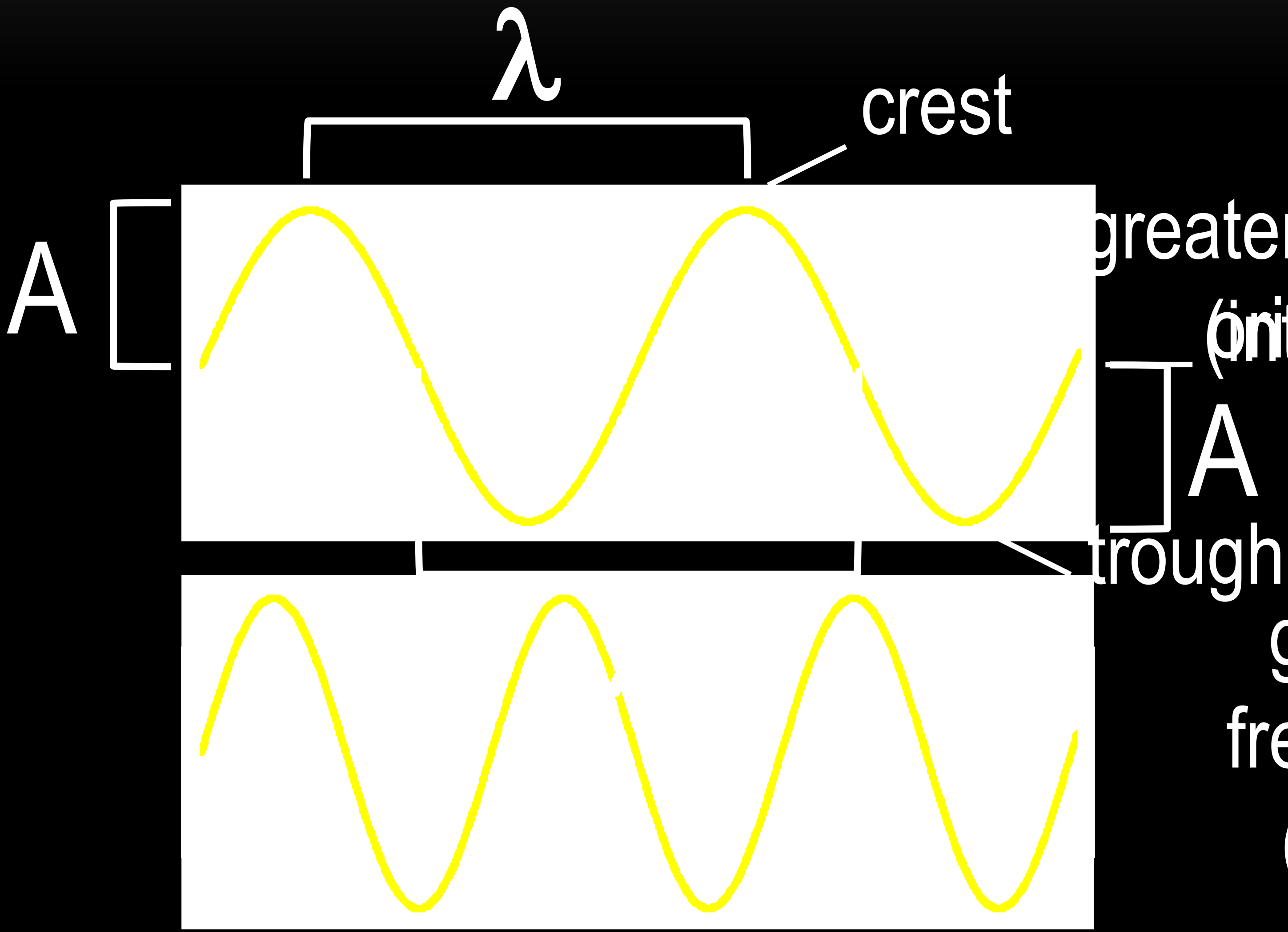
LIGHT AND THE QUANTUM MODEL

WAVES

- Wavelength (λ) - length of one complete wave
- Frequency (ν) - # of waves that pass a point during a certain time period
  - hertz (Hz) = 1/s
- Amplitude (A) - distance from the origin to the trough or crest

WAVES

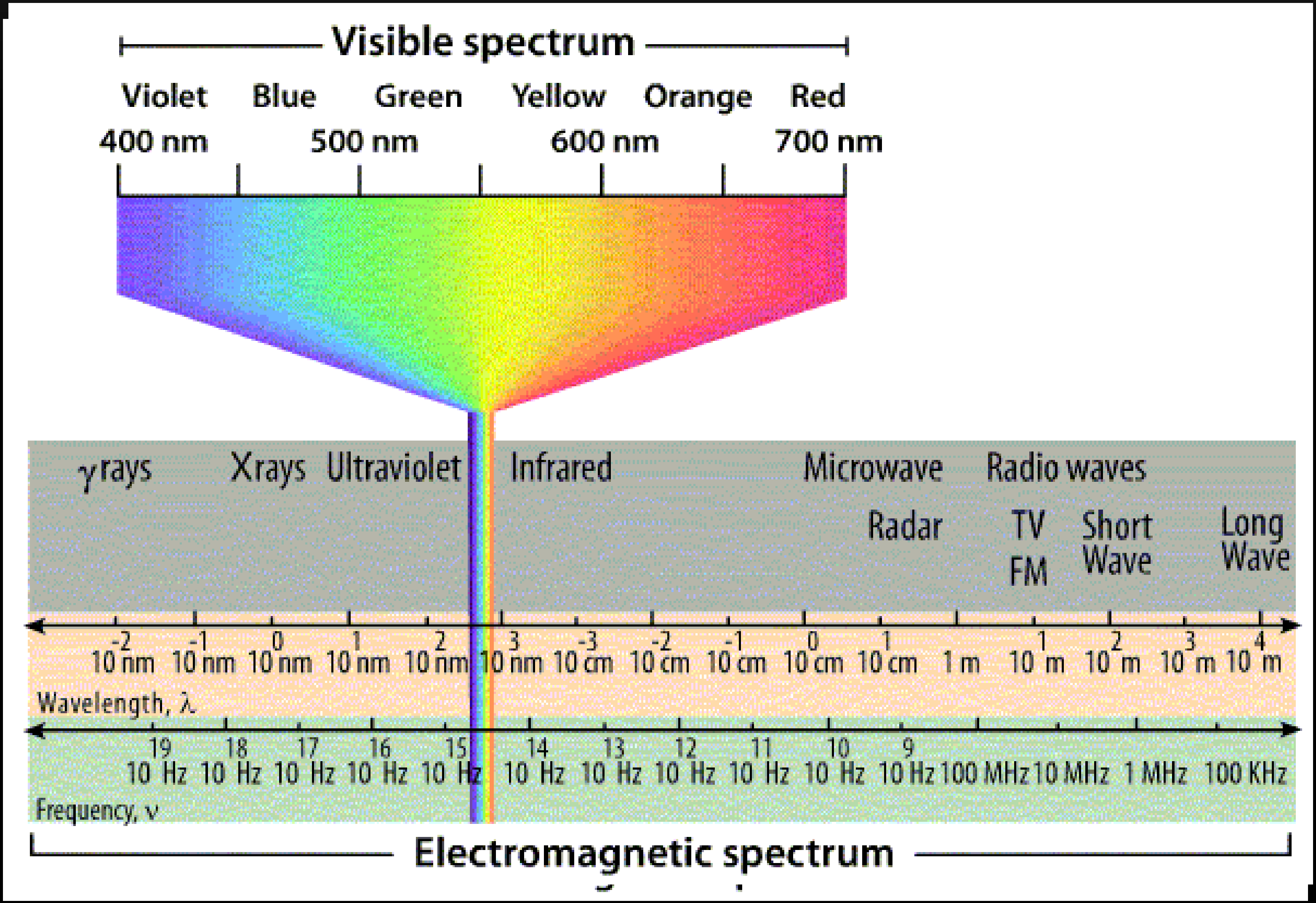


greater amplitude
(intensity)

greater
frequency
(color)

EM SPECTRUM

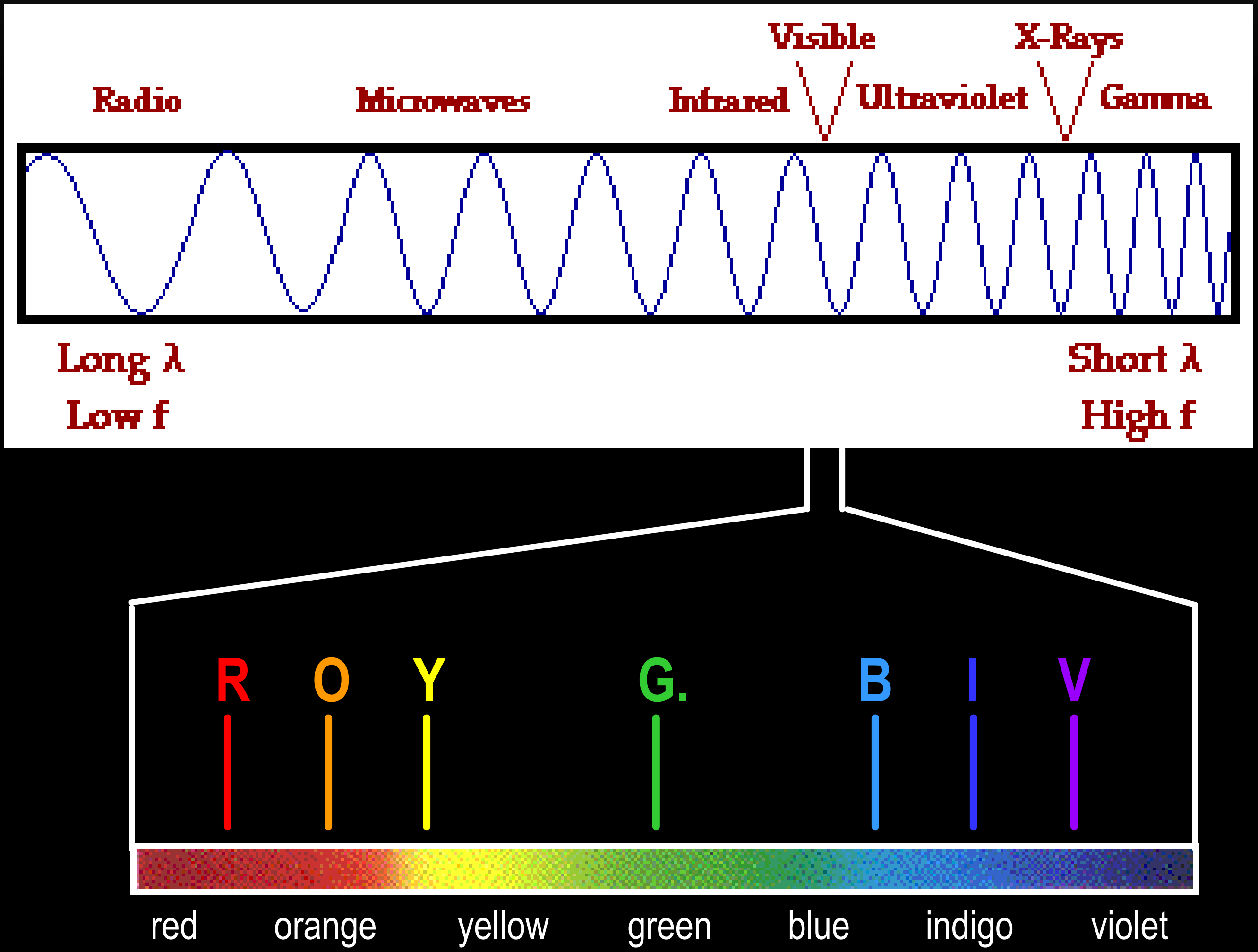
HIGH ENERGY



LOW ENERGY

EM SPECTRUM

LOW ENERGY



R O Y G. B I V

red orange yellow green blue indigo violet

HIGH ENERGY

EM SPECTRUM

- Frequency & wavelength are inversely proportional

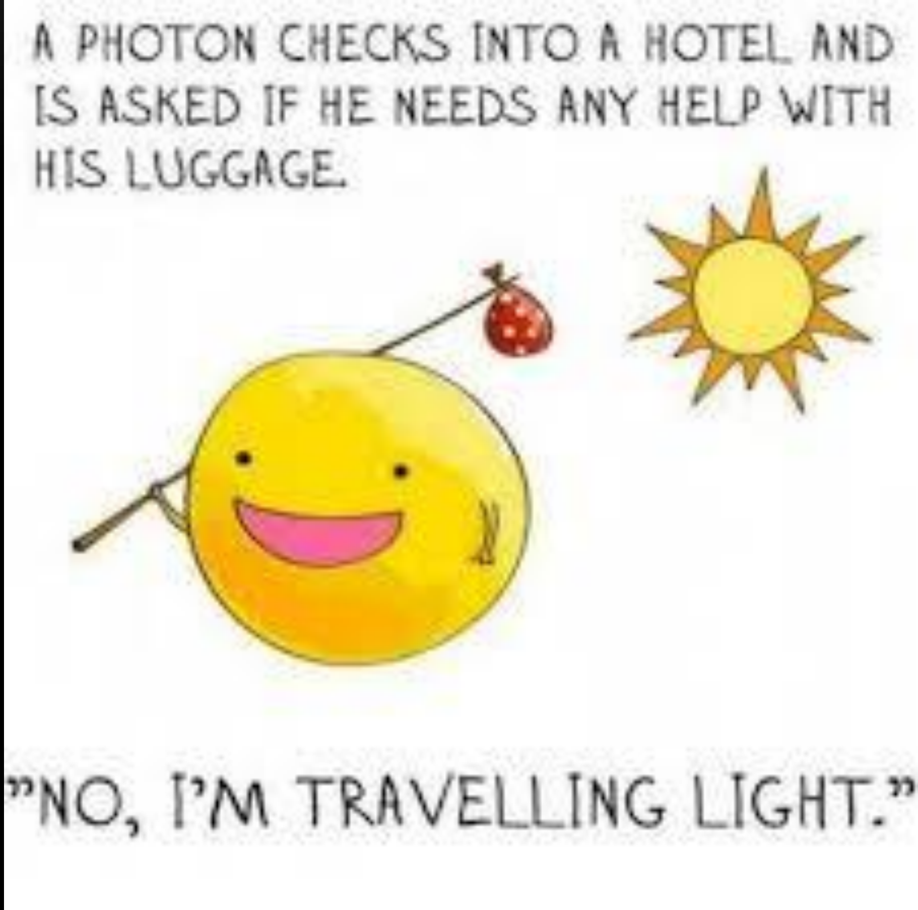
$$\boxed{c = \lambda \nu}$$

c: speed of light (3.00×10^8 m/s)
λ: wavelength (m, nm, etc.)
ν: frequency (Hz)

PHOTON

- A discrete particle of pure light, or electromagnetic radiation energy
- A quantum of energy (bundle, packet)

- Einstein postulated the existence of the **photon** to explain the "photoelectric effect" for which he obtained the Nobel prize in physics. In the photoelectric effect individual photons can liberate **electrons** and stimulate a current, demonstrating the particle-like nature of light.


A PHOTON CHECKS INTO A HOTEL AND IS ASKED IF HE NEEDS ANY HELP WITH HIS LUGGAGE.
"NO, I'M TRAVELLING LIGHT."

QUANTUM THEORY

The energy of a photon is proportional to its frequency.

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

E: energy (J, joules)

h: Planck's constant (6.6262×10^{-34} J·s)

ν: frequency (Hz)

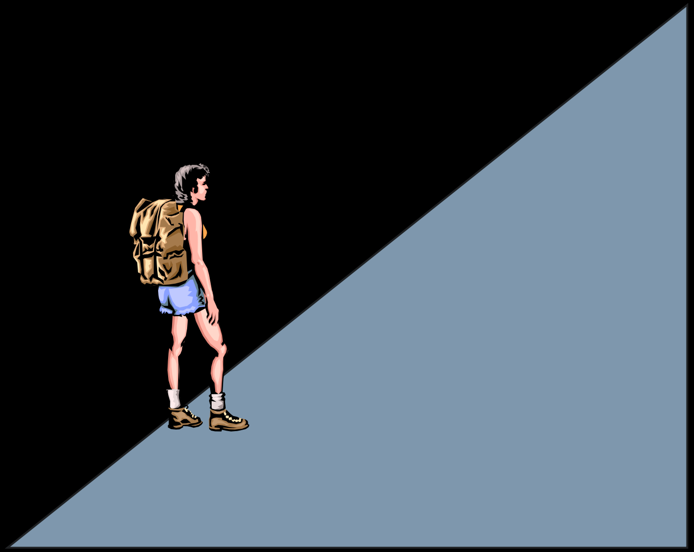
QUANTUM THEORY

- **Planck** (1900)
  - Observed - emission of light from hot objects
  - Concluded - energy is emitted in small, specific amounts (quanta)

  - Quantum - minimum amount of energy change

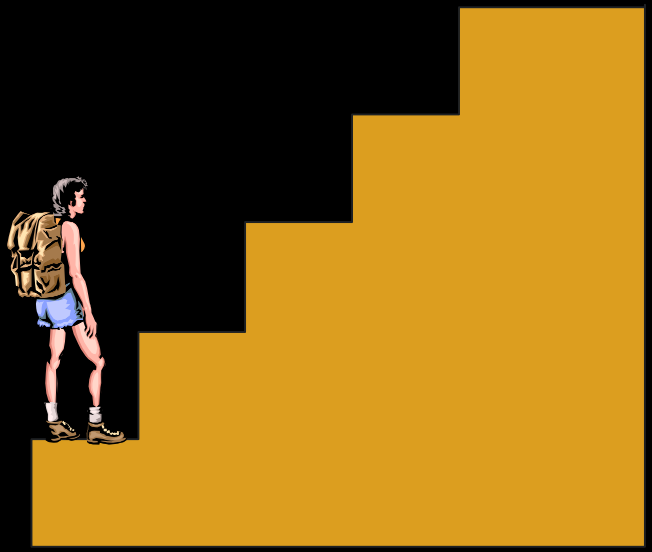
QUANTUM THEORY

- **Planck** (1900)



Classical Theory

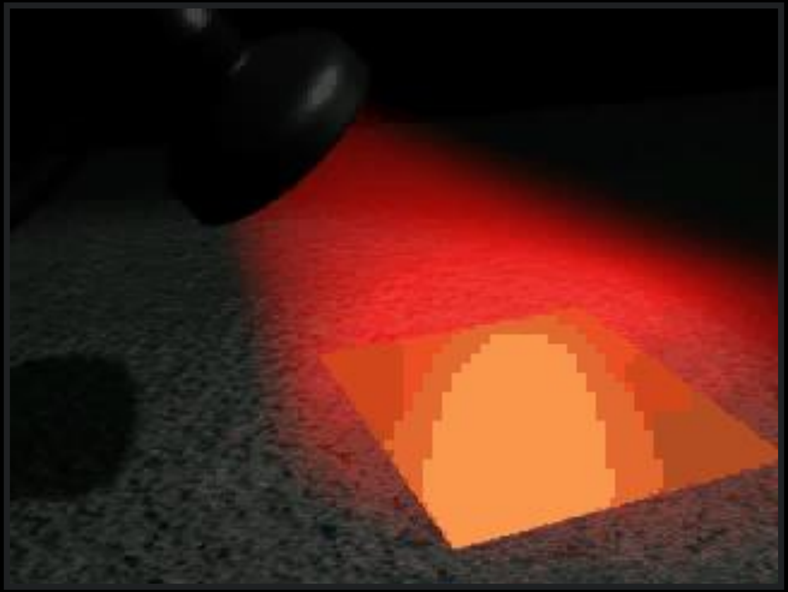
vs.



Quantum Theory

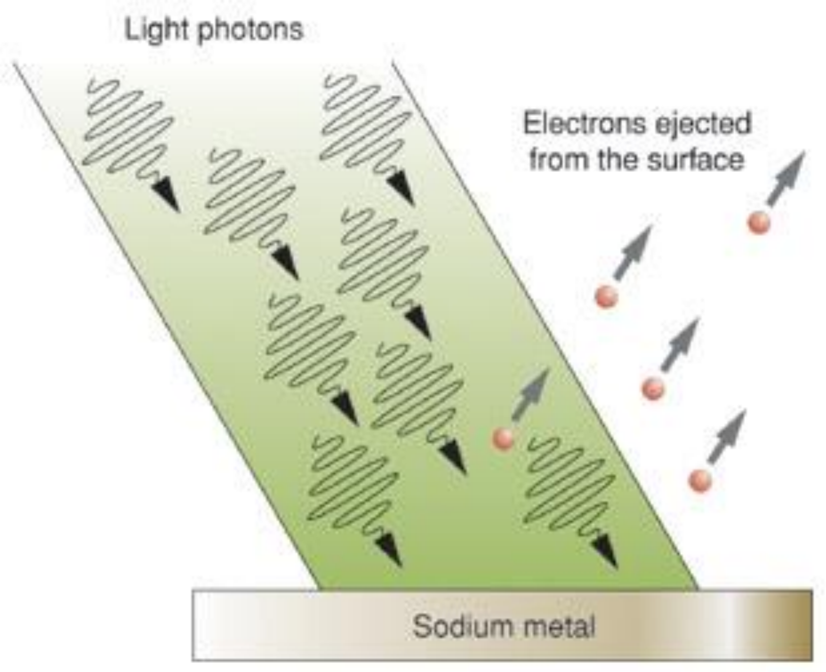
QUANTUM THEORY

- **Einstein** (1905)
  - Observed - photoelectric effect



PHOTOELECTRIC EFFECT

- The emission of electrons from a metal when light shines on it.



http://www.daviddarling.info/encyclopedia/E/Einstein_and_photoelectric_effect.html

QUANTUM THEORY

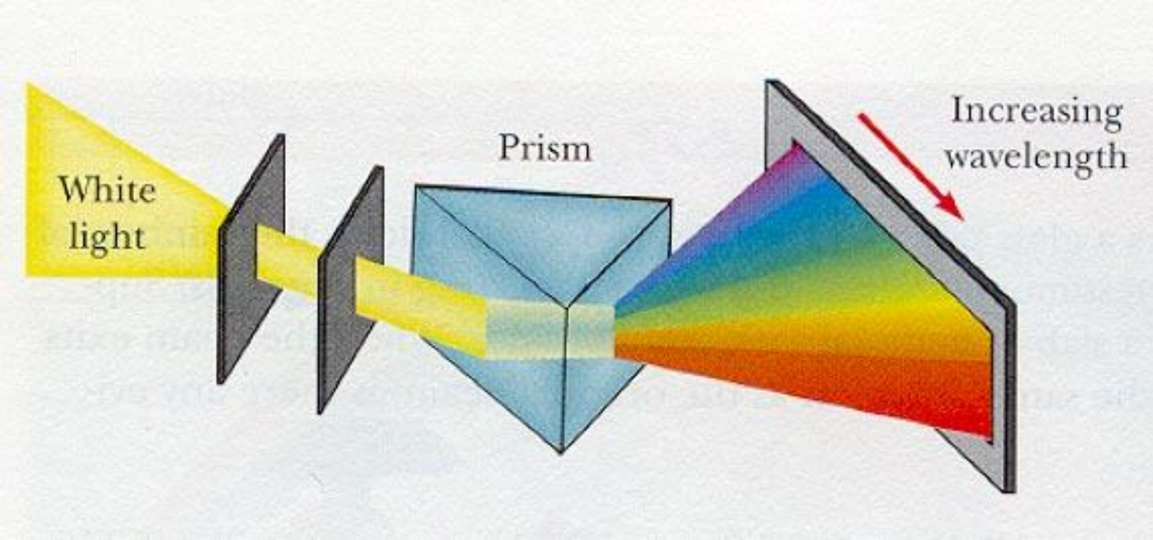
- **Einstein** (1905)
  - Concluded - light has properties of both waves and particles

“wave-particle duality”

  - Photon - particle of light that carries a quantum of energy

ATOMIC SPECTRA

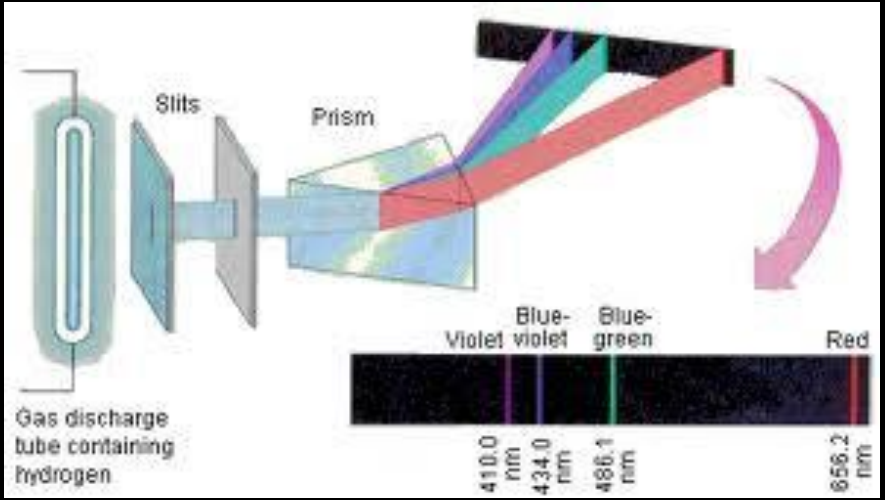
- White light is made up of all the colors of the visible spectrum.
- Passing it through a prism separates it.



http://www.astro.virginia.edu/class/oconnell/astr130/lec8-f02.html

NOT ALL LIGHT IS THE SAME

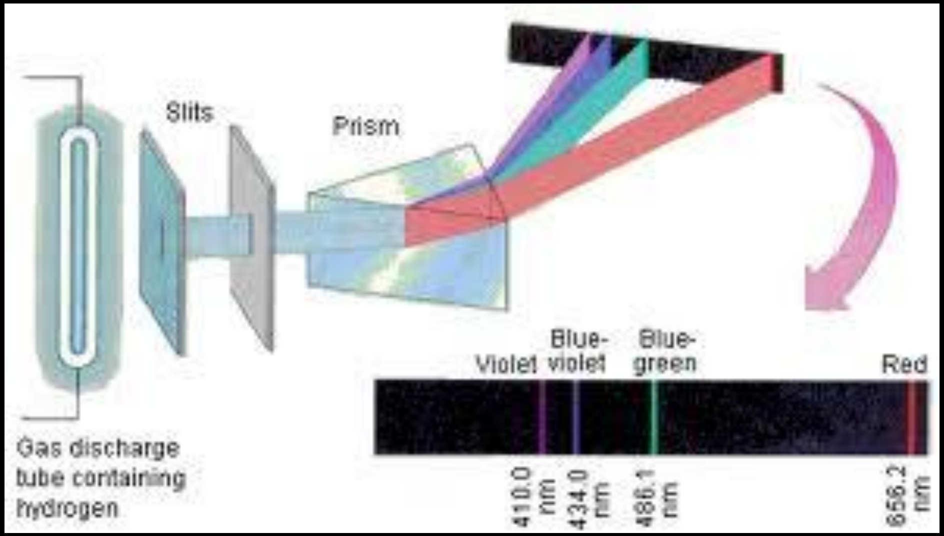
- By heating a gas with electricity we can get it to give off colors.
- Passing this light through a prism does something different.



http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch6/bohr.html

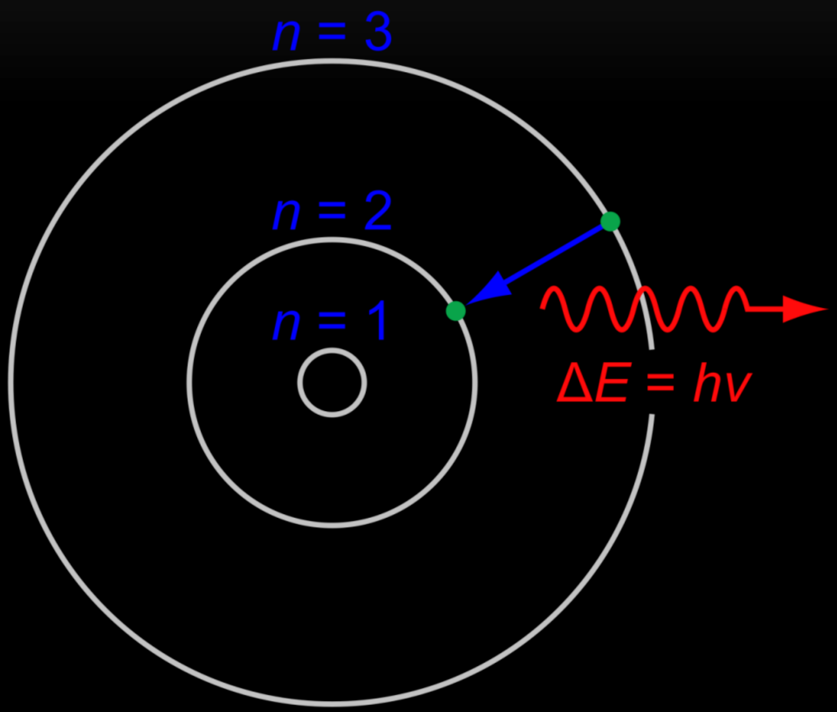
ATOMIC SPECTRUM

- Each element gives off its own characteristic colors.
- Can be used to identify the element.
- This is how we know what stars are made of.



http://chemed.chem.purdue.edu/genchem/topicreview/bp/ch6/bohr.html

WHAT CAUSES THE LINES?

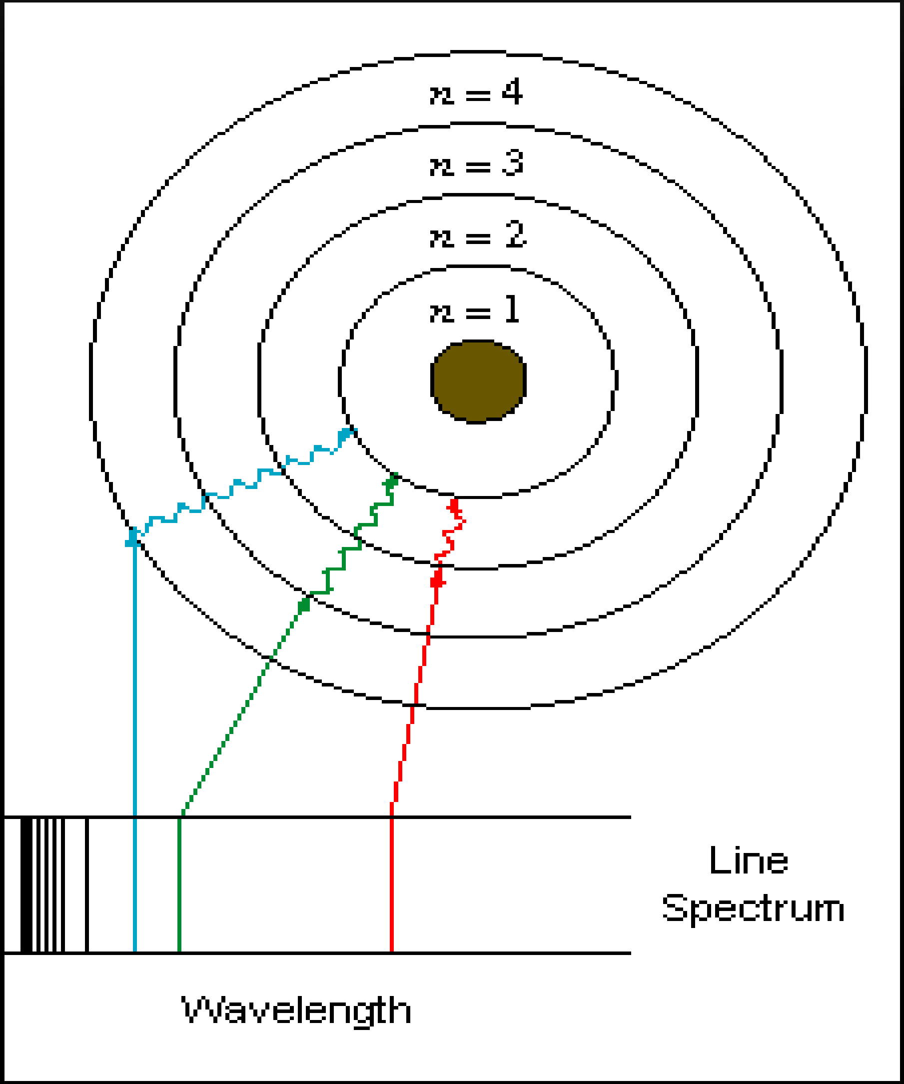


- As an electron drops from an excited state to a lower state, it releases energy in the form of light.

http://en.wikipedia.org/wiki/Bohr_model

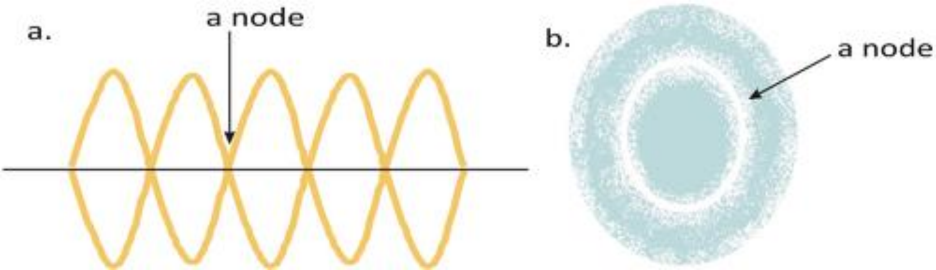
ATOMIC SPECTRUM

Line spectrum of the Hydrogen atom



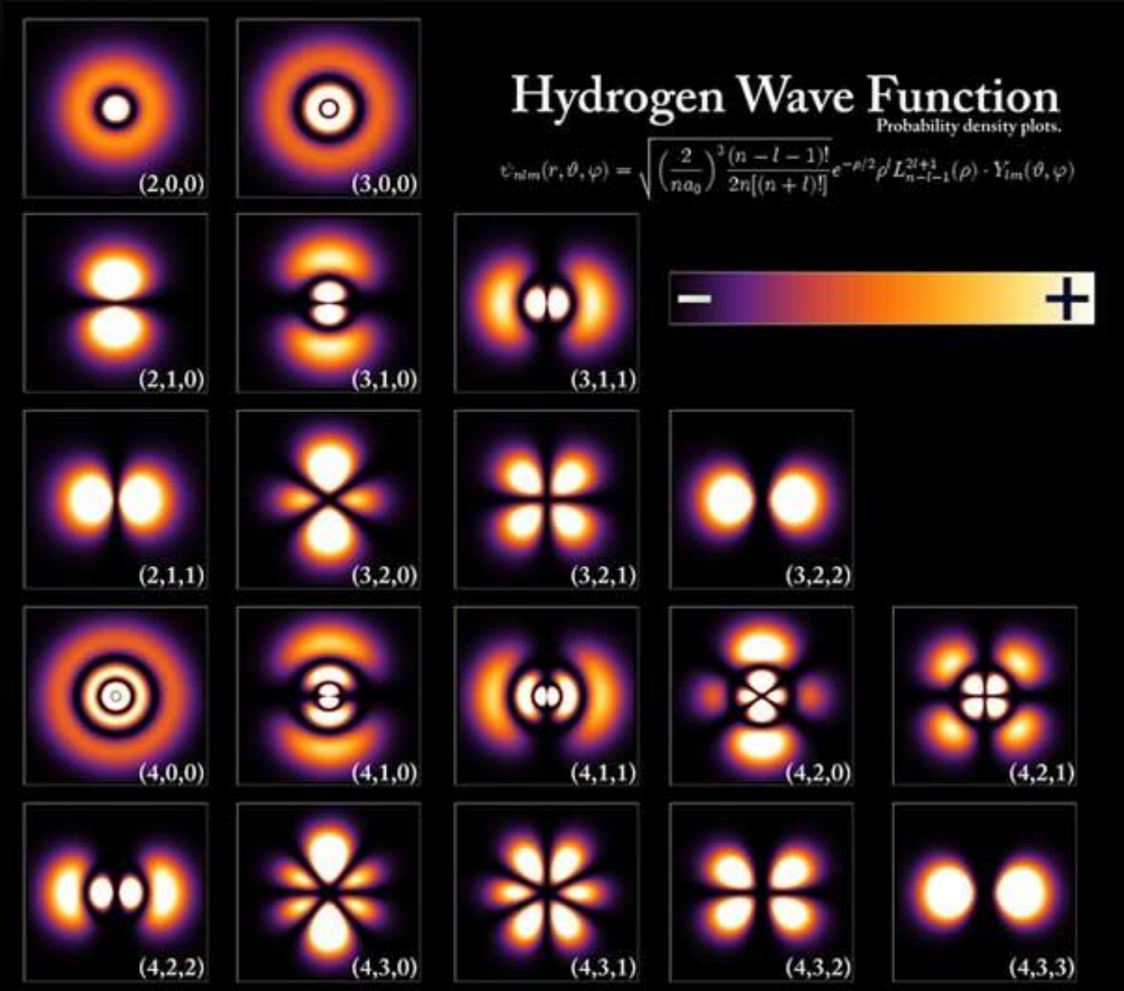
THE HEISENBERG UNCERTAINTY PRINCIPLE

- We know that:
  - Light is required to "see" electrons.
  - Photons are small "packets" of light.
  - Electrons are affected by photons.
    - When electrons absorb photons, they are physically moved to a new location.
- Therefore: *It is impossible to know (or determine) both the position and the velocity of an electron.*
  - This is known as *The Heisenberg Uncertainty Principle.*



THE SHRÖDINGER WAVE EQUATION

- Combining Bohr's model with de Broglie electron-wave theory, Shrödinger came up with an equation that predicted the *probability* of where an electron would be around the nucleus.
- Electrons were no longer in energy levels, but in complex patterns, or *clouds*.



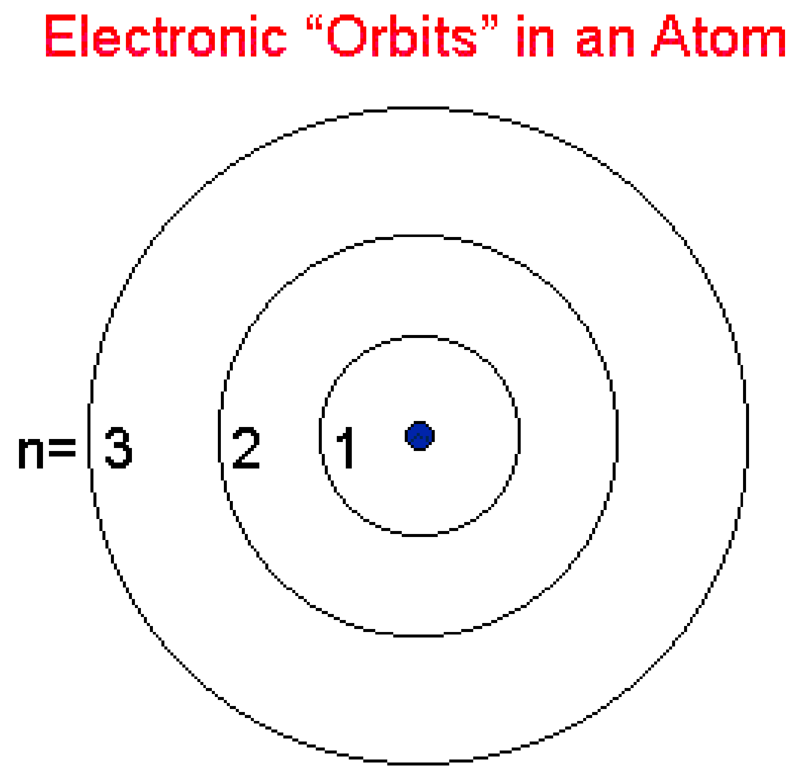
QUANTUM NUMBERS AND ORBITALS

- Quantum Numbers specify properties of atomic orbitals *and* the properties of the electrons in the orbitals.
- There are 4 different quantum numbers:

1. The Principal Quantum Number: Indicates the main energy level
2. Angular Momentum Quantum Number: Indicates the shape of the orbital
3. Magnetic Quantum Number: Indicates the orientation of the orbital around the nucleus (3-dimensional orientation)
4. Spin Quantum Number: Indicates the spin state of the electron.

PRINCIPAL QUANTUM NUMBER

- Energy Level occupied by the electron.
- Symbol: *n*
- Values: 1, 2, 3, …
- Example:
  - Electron in energy level 2 has an *n* = 2.

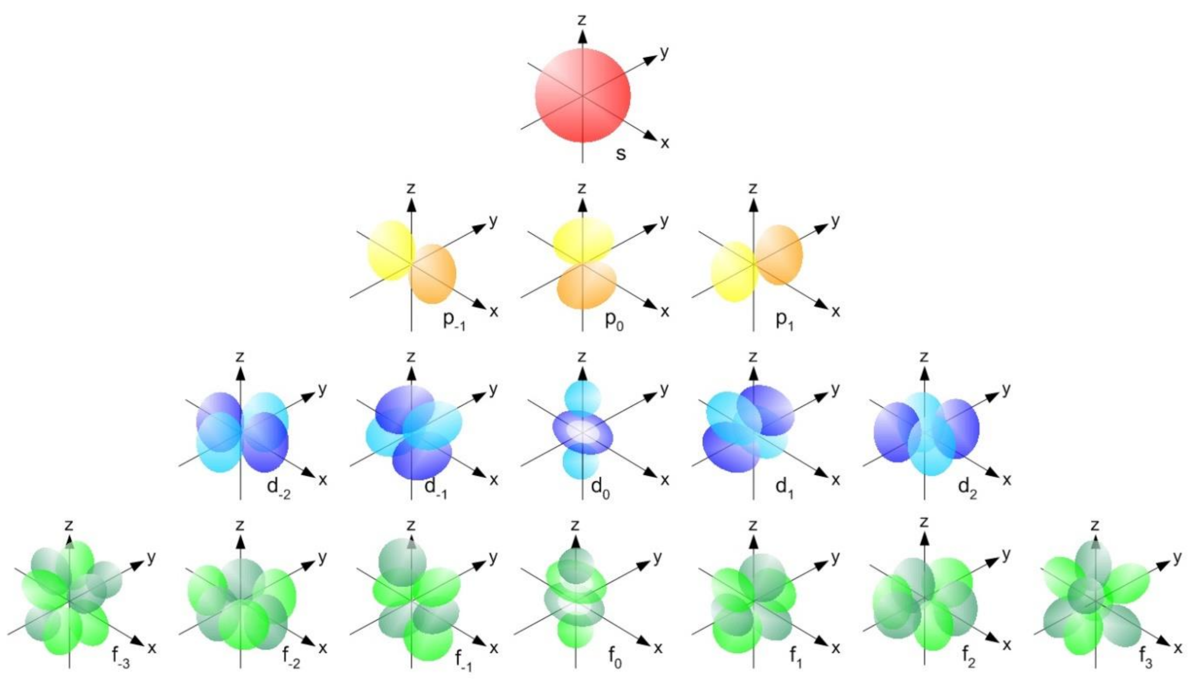


http://www.astro.cornell.edu/academics/courses/astro101/lectures/lec07.htm

ANGULAR MOMENTUM QUANTUM NUMBER

- Shape of the orbital
- Symbol: ℓ (italicized "l")
- Values: 0 through n-1
- Example:
  - An electron with an $n = 2$ can have $\ell = 0$ or 1
  - An electron with an $n = 5$ can have an $\ell = 0, 1, 2, 3,$ or 4

AN MBER



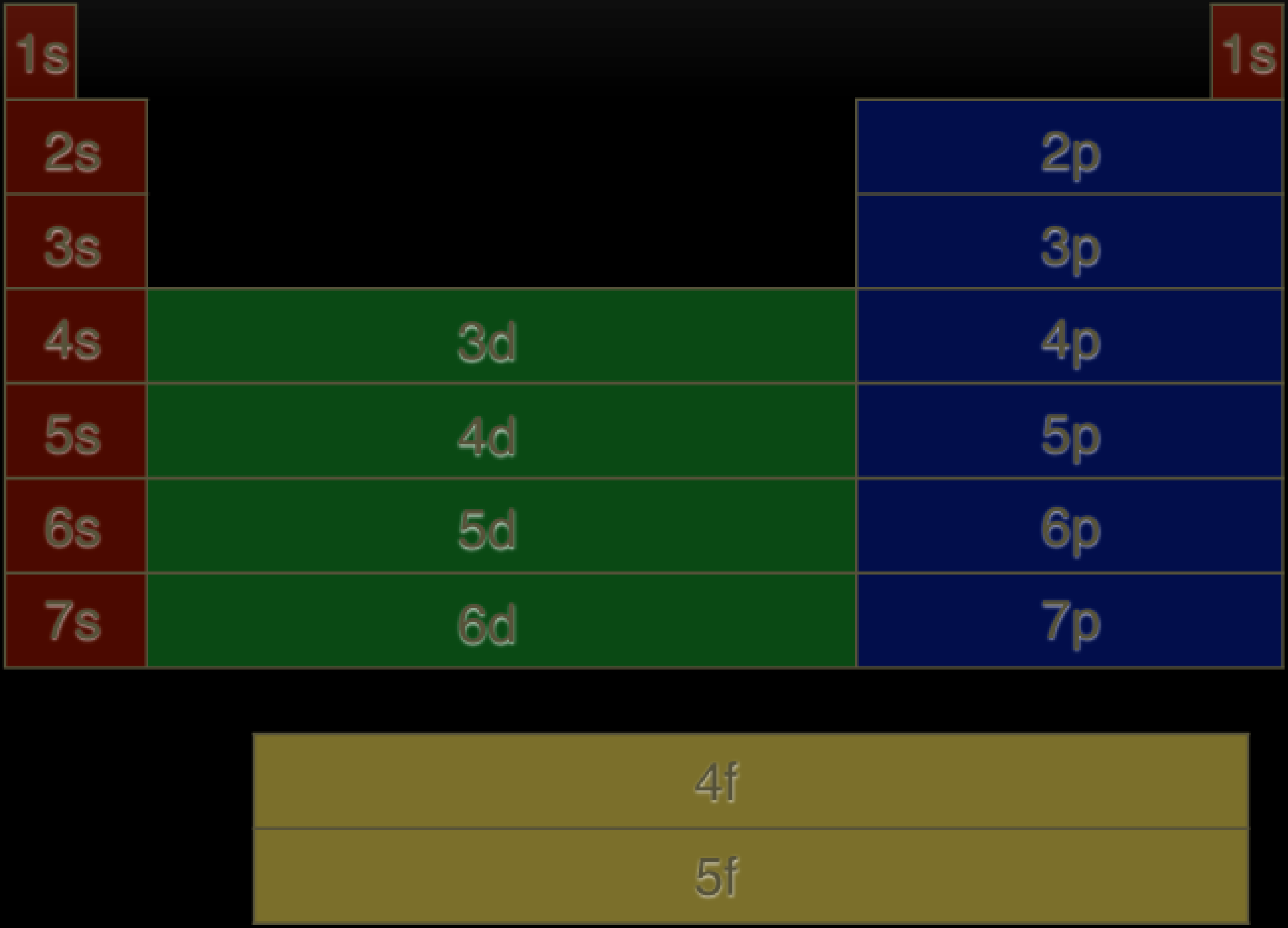
- Each orbital.
- ℓ = 0
- ℓ = 1 → p orbital
- ℓ = 2 → d orbital
- ℓ = 3 → f orbital

http://chemwiki.ucdavis.edu/Physical_Chemistry/Quantum_Mechanics/Atomic_Theory/Electrons_in_Atoms/Electronic_Orbitals

MAGNETIC QUANTUM NUMBER

- Each individual orbital can have different orientations around the nucleus
- Symbol: m_ℓ
- Values: - ℓ to + ℓ
- Example:
  - An electron with an $\ell = 0$ can only have $m_\ell = 0$
  - An electron with an $\ell = 2$ can have $m_\ell = -2, -1, 0, 1, 2$

ORBITAL LOCATION ON PERIODIC TABLE



1s			1s
2s		2p	
3s		3p	
4s	3d	4p	
5s	4d	5p	
6s	5d	6p	
7s	6d	7p	

4f

5f

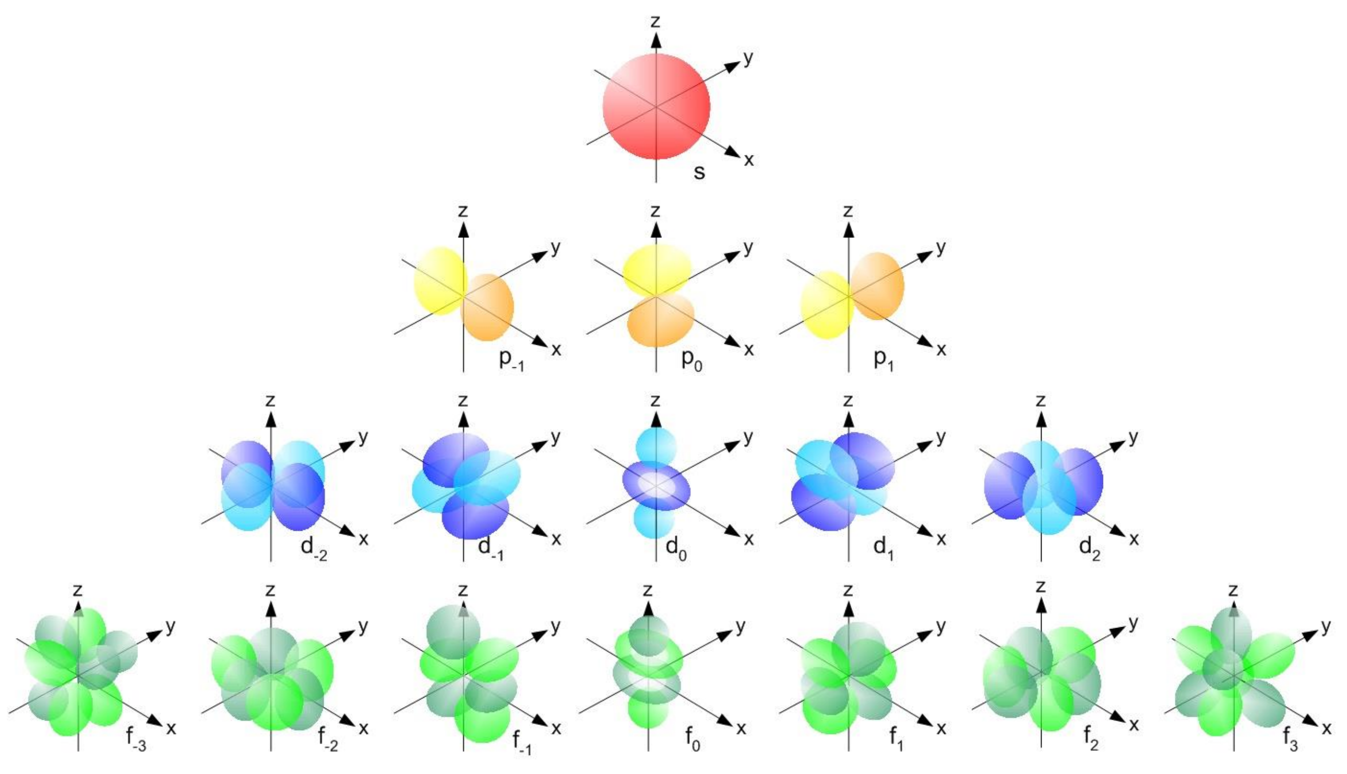
MAGNETIC QUANTUM NUMBER

$\ell = 0$; s orbital

$\ell = 1$; p orbital

$\ell = 2$; d orbital

$\ell = 3$; f orbital



http://chemwiki.ucdavis.edu/Physical_Chemistry/Quantum_Mechanics/Atomic_Theory/Electrons_in_Atoms/Electronic_Orbitals

SPIN QUANTUM NUMBER

- Spin state of the electron
- Symbol: *N/A*
- Values: -½, +½ for *any* orbital or value of ℓ .

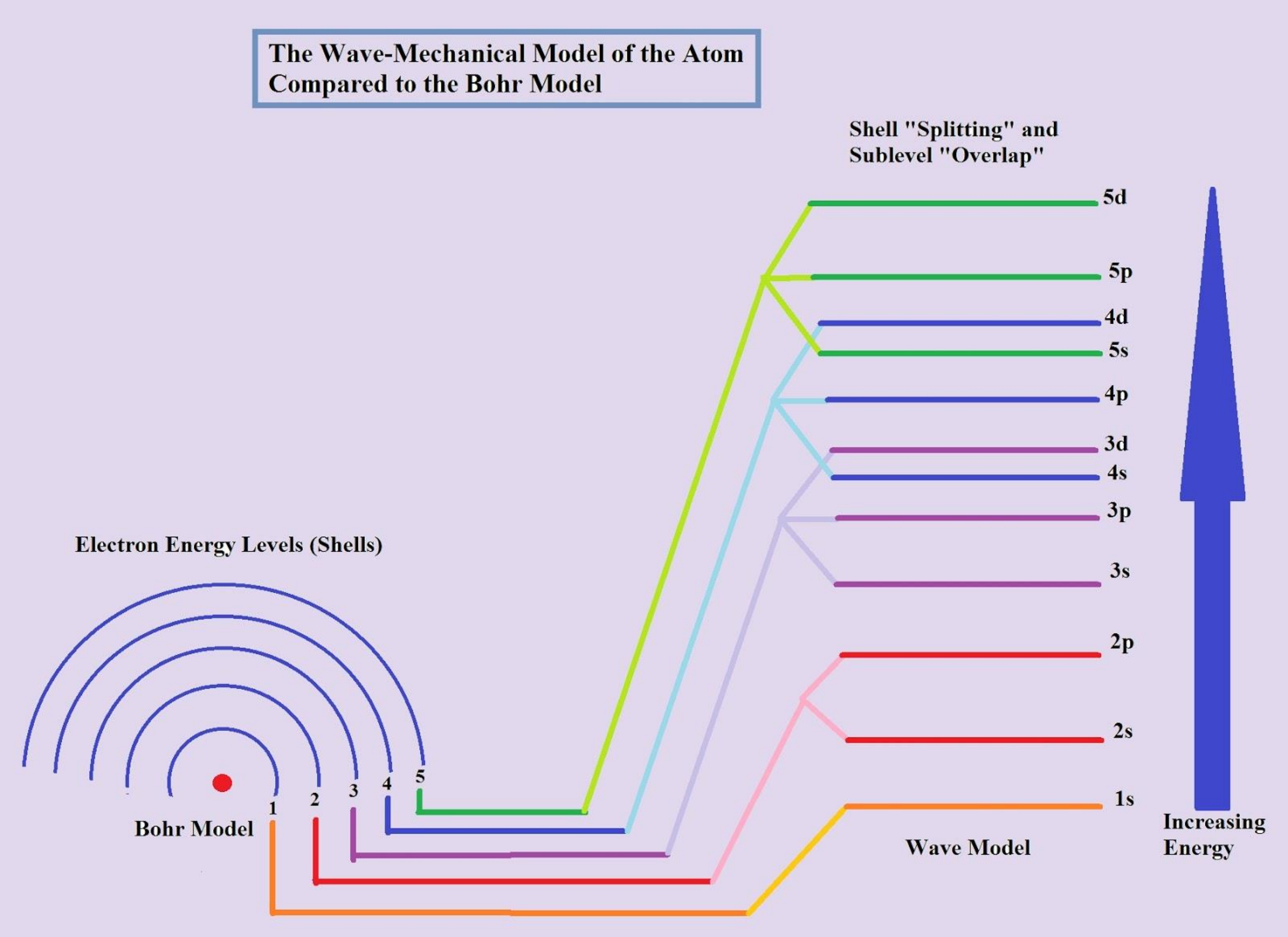
QUANTUM NUMBERS FOR THE FIRST FOUR LEVELS IN THE HYDROGEN ATOM

n	ℓ	Orbital Designation	m_ℓ	Number of Orbitals
1	0	1s	0	1
2	0	2s	0	1
	1	2p	-1,0,+1	3
3	0	3s	O	1
	1	3p	-1,0,+1	3
	2	3d	-2,-1,0,+1,+2	5
4	0	4s	0	1
	1	4p	-1,0,+1	3
	2	4d	-2,-1,0,+1,+2	5
	3	4f	-3,-2,-1,0,+1,+2,+3	7

ELECTRON CONFIGURATIONS

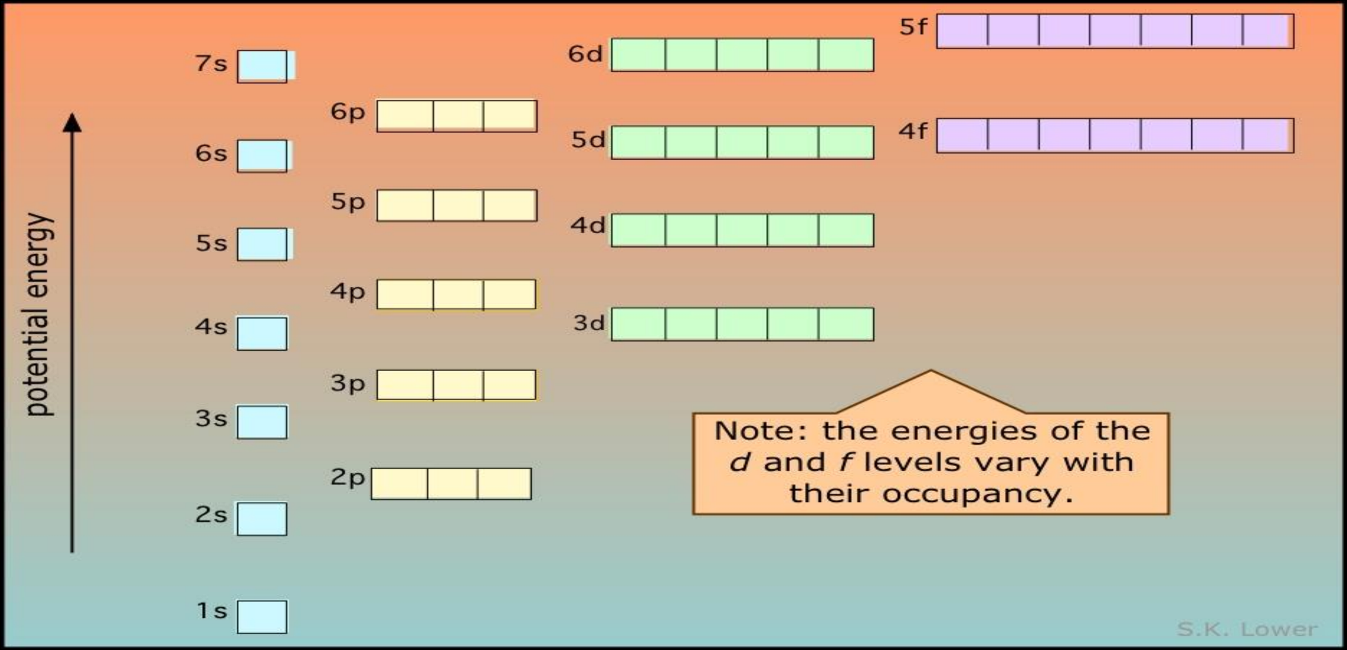
- The arrangement of electrons in their orbitals around the nucleus of an atom is called the electron configuration.
- There are 3 Rules for determining the electron configuration:
  - The Aufbau Principle states that electrons will always occupy the lowest energy orbital that is available (ground state).
  - The Pauli Exclusion Principle also states that any two electrons in the same atom cannot have the same set of quantum numbers.
  - Hund’s Rule states that, in orbitals of equal energy, electrons will first occupy different orbitals before pairing up.

BOHR MODEL VS. WAVE MECHANICAL MODEL



ORBITAL ENERGIES

- In general, the energy of the different energy levels increase as *n* increases, so electrons will start at level 1 and move up.
- However, different orbitals within an energy level can overlap other energy levels.



http://www.chem1.com/acad/webtext/atoms/atpt-5.html

ORBITALS FILL IN AN ORDER

- Lowest energy to higher energy.
- Adding electrons can change the energy of the orbital. <u>Full orbitals</u> are the absolute best situation.
- However, **<u>half filled</u>** orbitals have a lower energy, and are next best
  - Makes them more stable.
  - Changes the filling order

BY ENERGY LEVEL

- First Energy Level
- Has only s orbital
- only 2 electrons
- $1s^2$

- Second Energy Level
- Has s and p orbitals available
- 2 in s, 6 in p
- $2s^22p^6$
- 8 total electrons

BY ENERGY LEVEL

- Third energy level
- Has s, p, and d orbitals
- 2 in s, 6 in p, and 10 in d
- $3s^2 3p^6 3d^{10}$
- 18 total electrons

- Fourth energy level
- Has s, p, d, and f orbitals
- 2 in s, 6 in p, 10 in d, and 14 in f
- $4s^2 4p^6 4d^{10} 4f^{14}$
- 32 total electrons

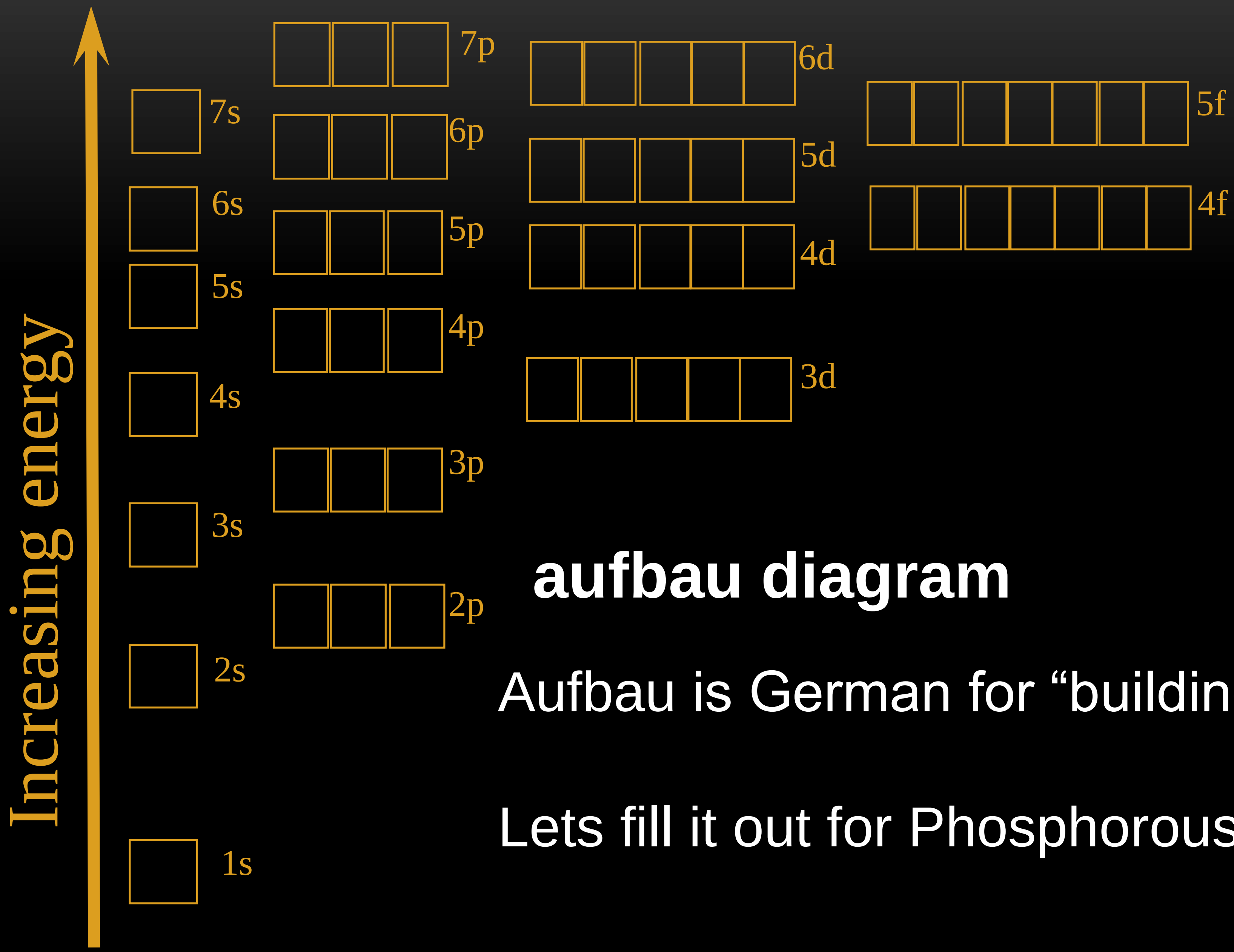
EXCEPTIONS

- The d orbital “likes” to be full with either 5 or 10 electrons.
  - Cr should have an electron configuration of

    $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4$

    Instead, one electron is transferred from the 4s orbital to the 3d orbital so it is ½ full.

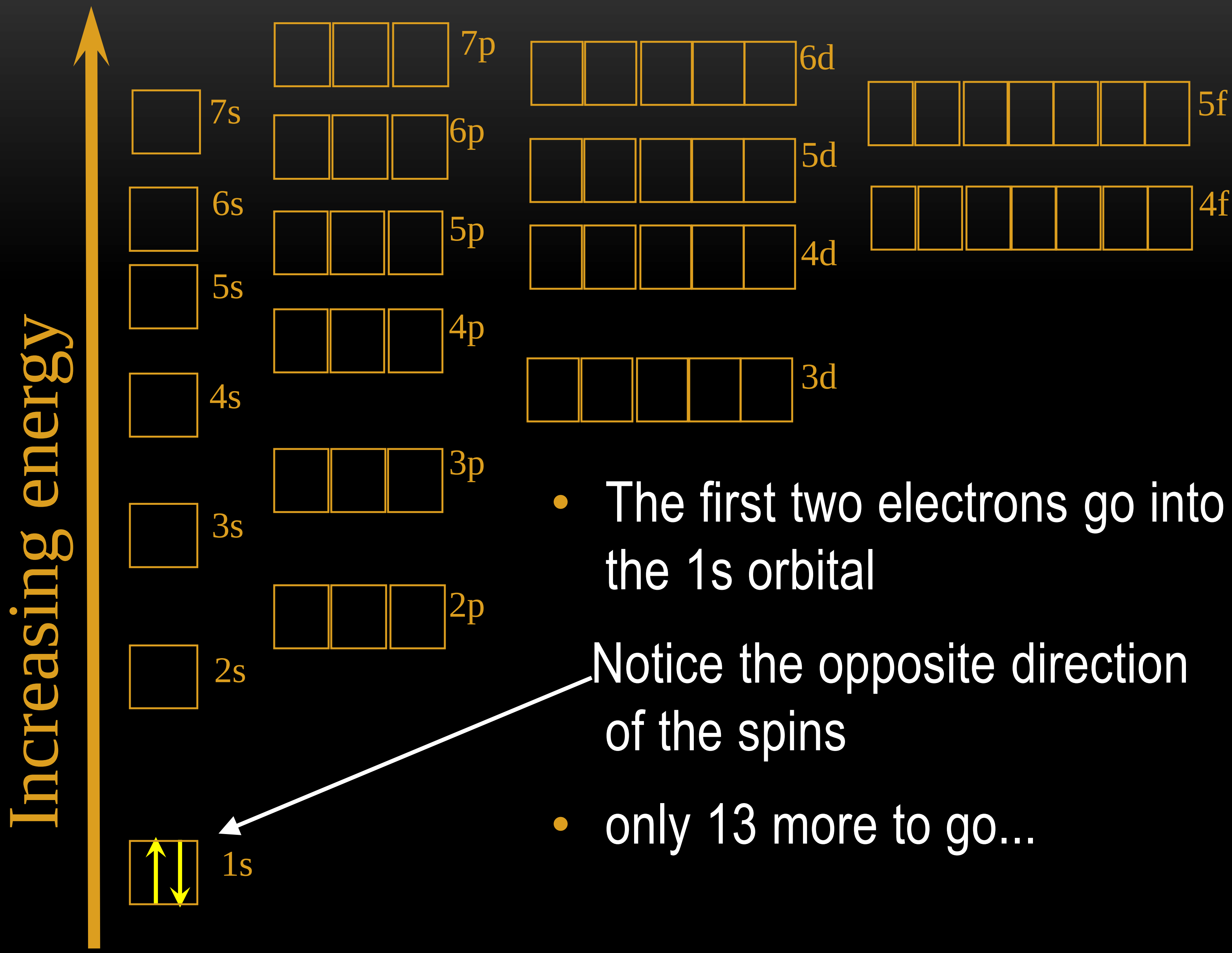
    $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$



aufbau diagram

Aufbau is German for “building up”

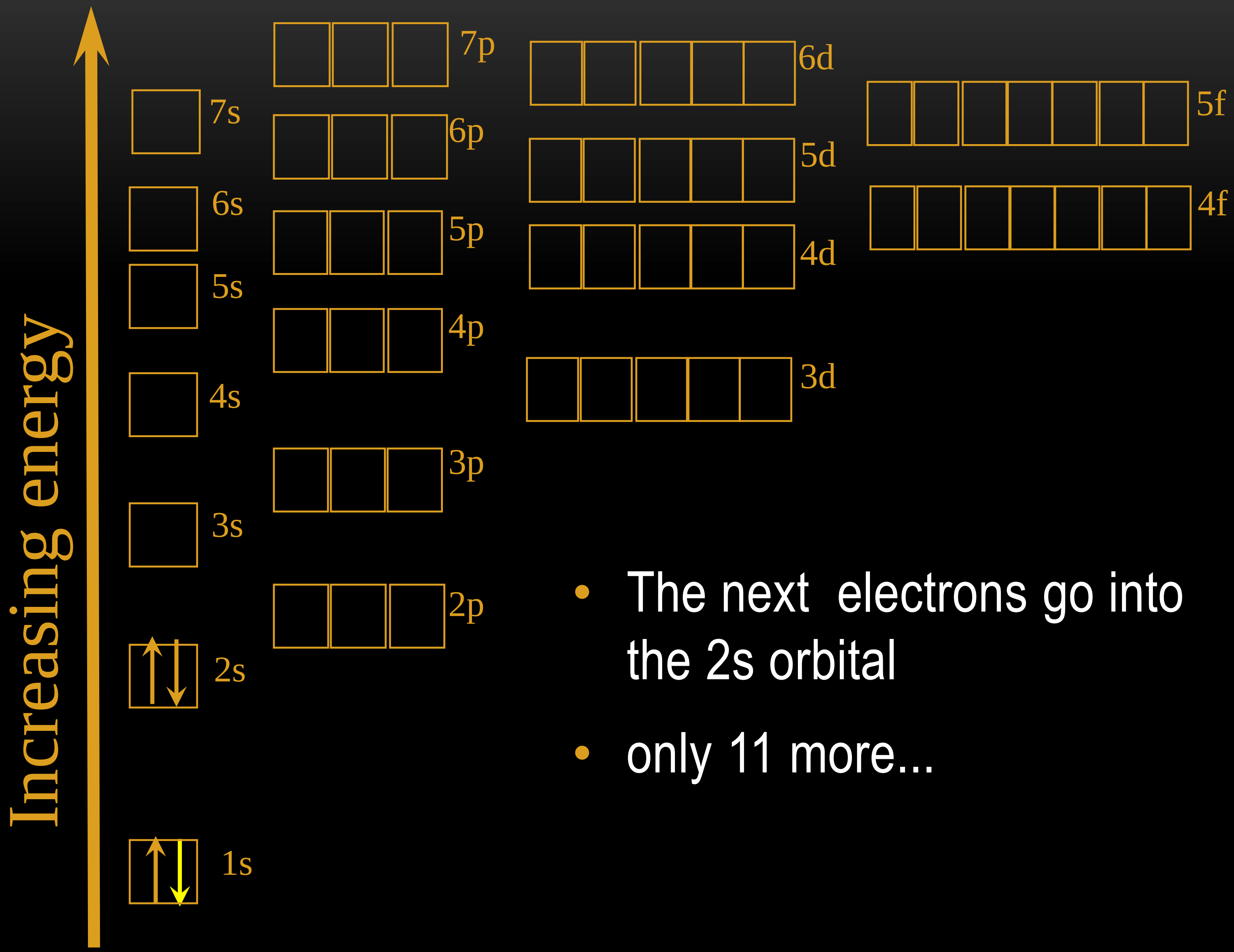
Lets fill it out for Phosphorous



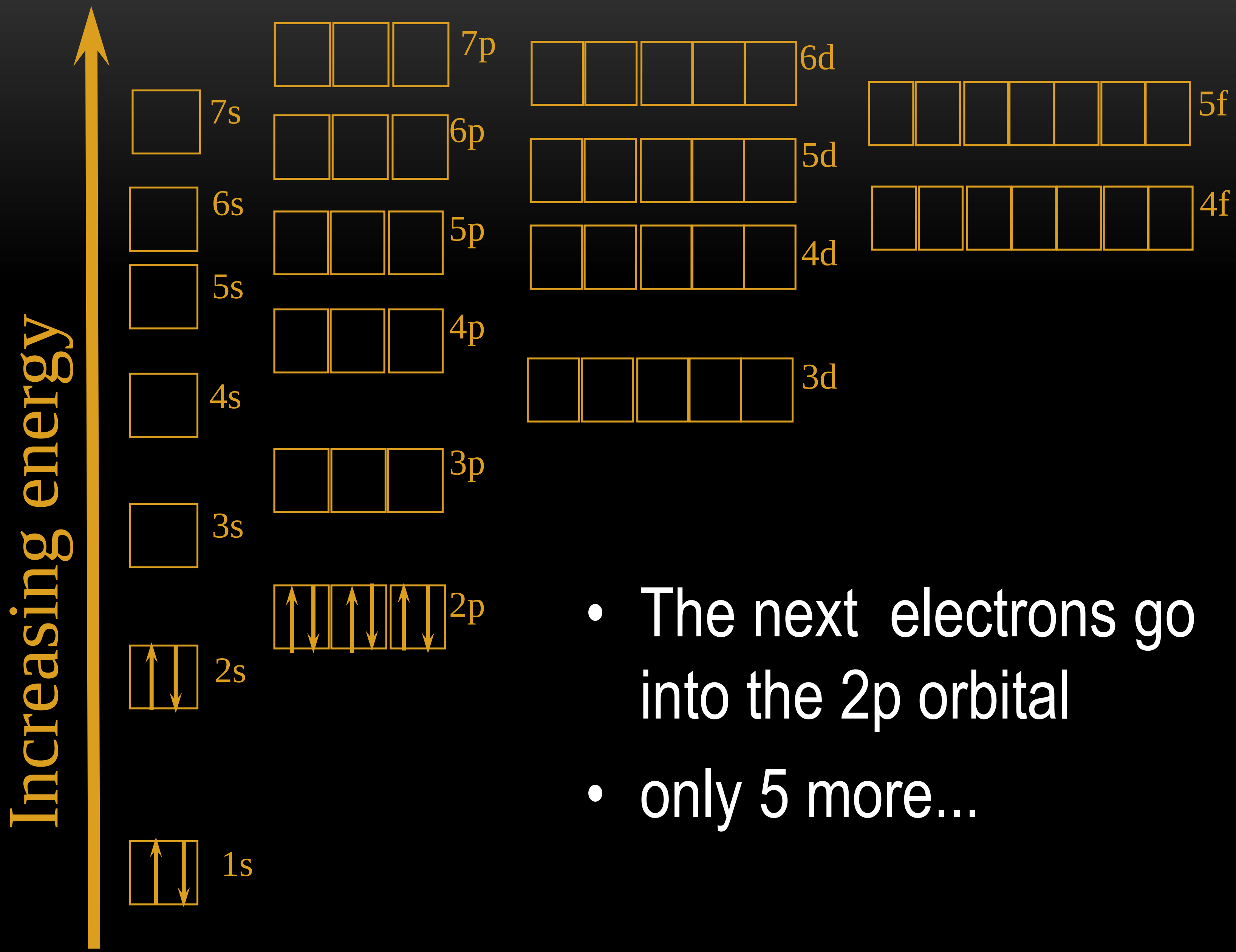
- The first two electrons go into the 1s orbital

Notice the opposite direction of the spins

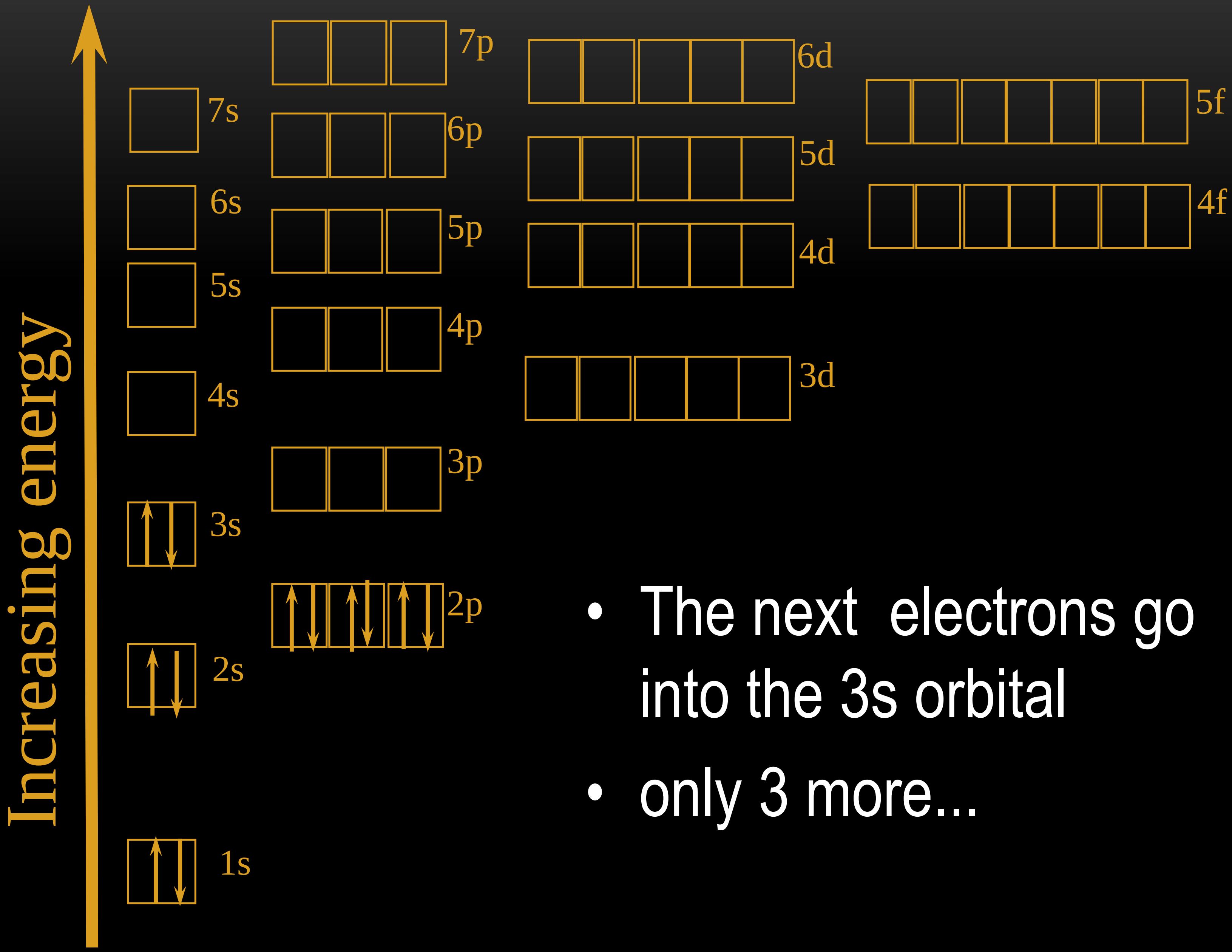
- only 13 more to go...



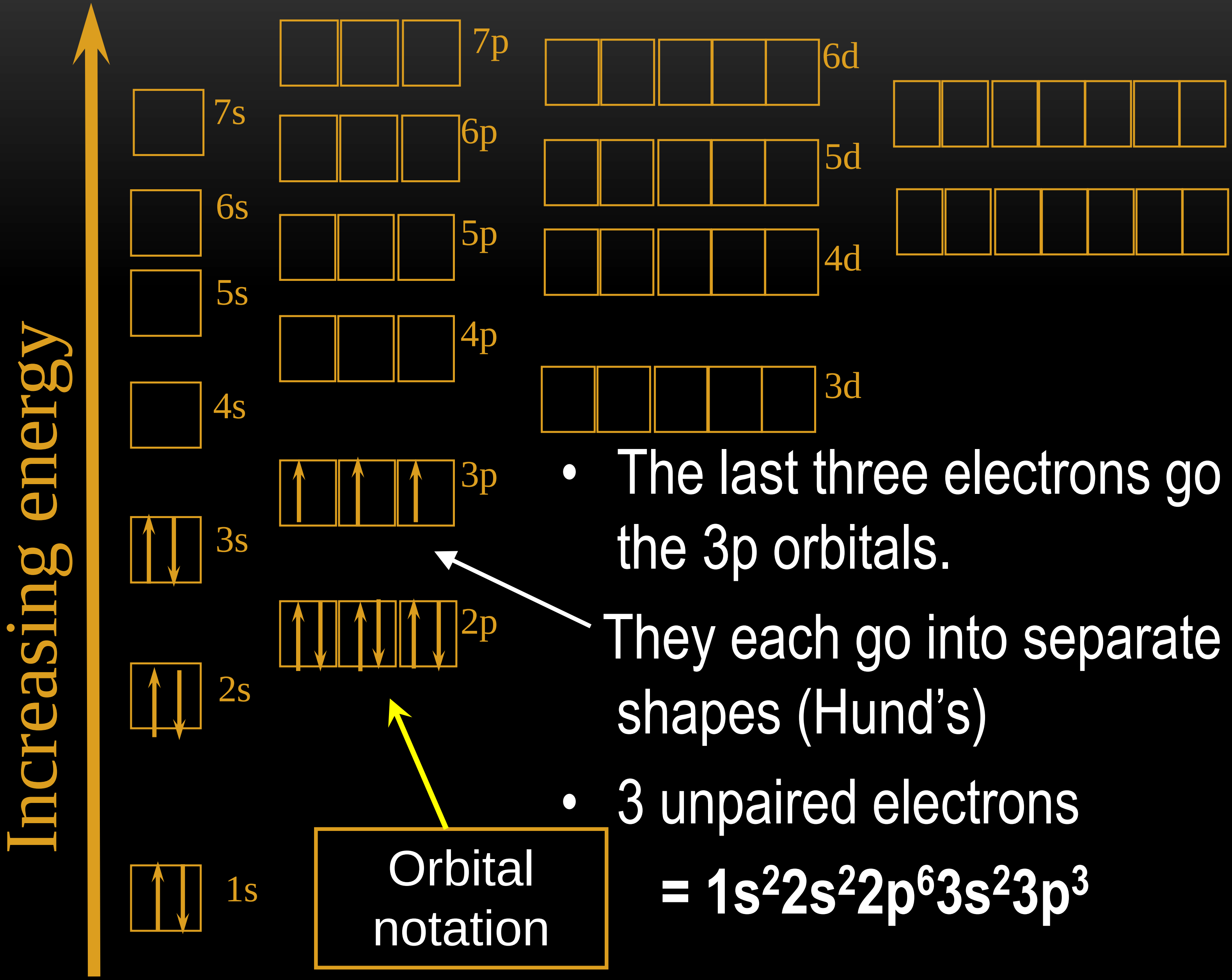
- The next electrons go into the 2s orbital
- only 11 more…



- The next electrons go into the 2p orbital
- only 5 more...



- The next electrons go into the 3s orbital
- only 3 more...



- The last three electrons go into the 3p orbitals.

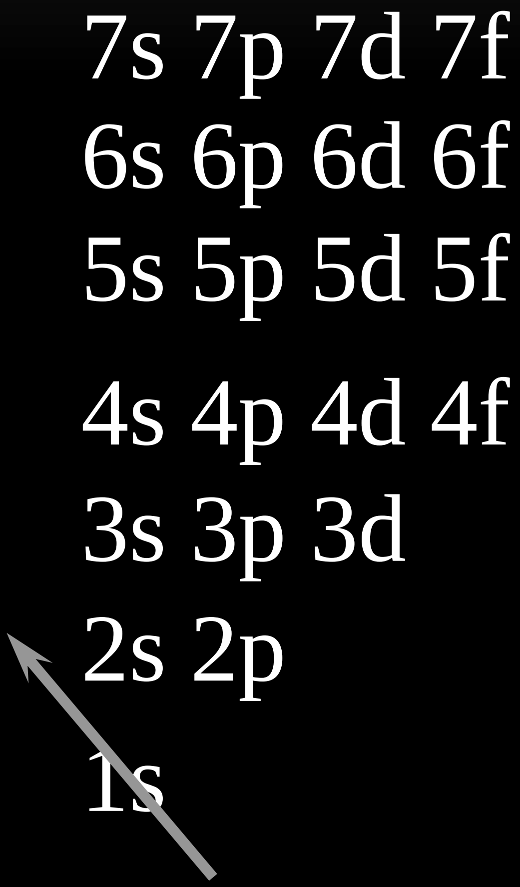
They each go into separate shapes (Hund’s)

- 3 unpaired electrons

= $1s^22s^22p^63s^23p^3$

Orbital notation

THE EASY WAY TO REMEMBER



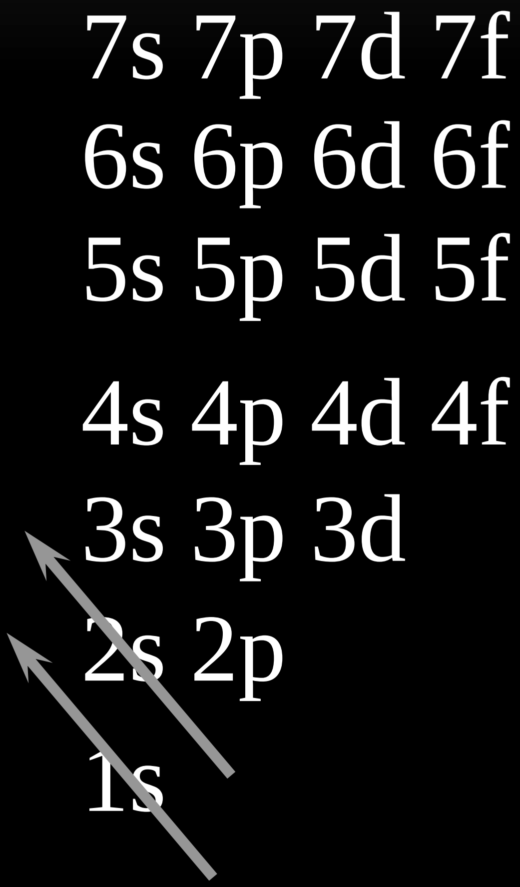
7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

Lets do this for Hassium (atomic Number 108)

- $1s^2$

- 2 electrons

FILL FROM THE BOTTOM UP FOLLOWING THE ARROWS



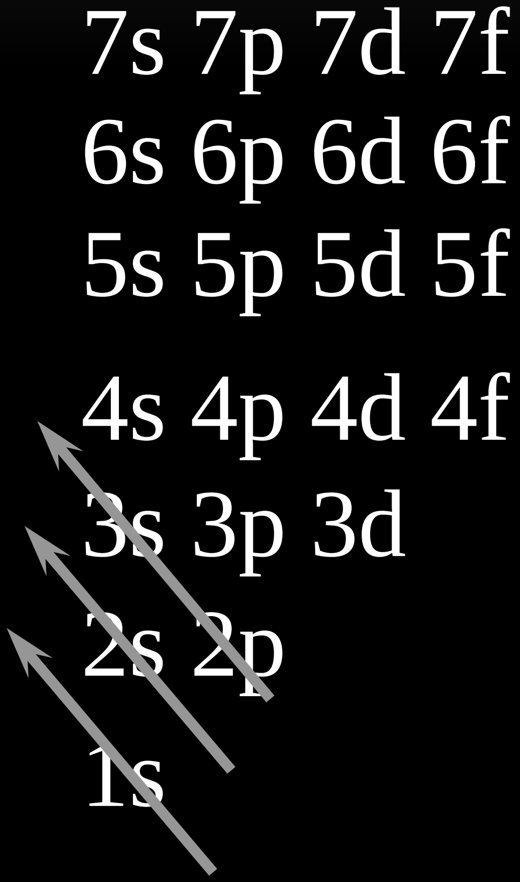
7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

Lets do this for Hassium (atomic Number 108)

- $1s^2\ 2s^2$

- 4 electrons

FILL FROM THE BOTTOM UP FOLLOWING THE ARROWS



7s 7p 7d 7f

6s 6p 6d 6f

5s 5p 5d 5f

4s 4p 4d 4f

3s 3p 3d

2s 2p

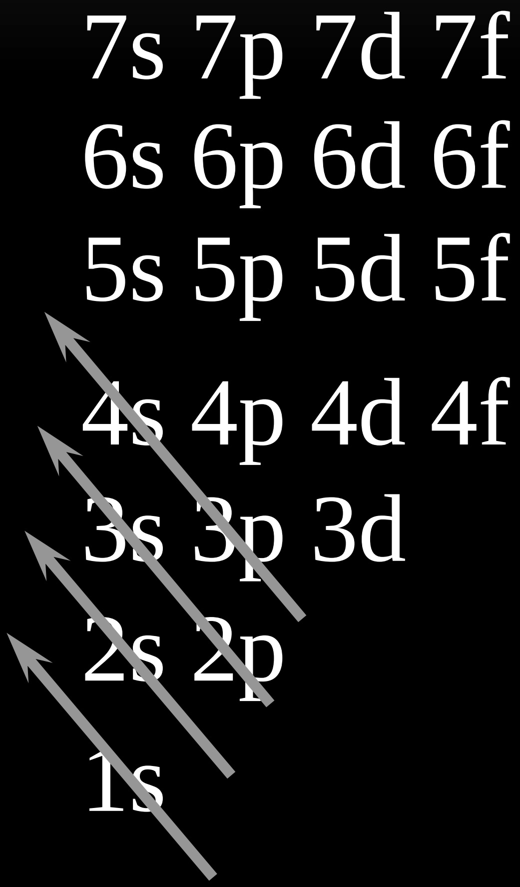
1s

Lets do this for Hassium (atomic Number 108)

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

- 12 electrons

FILL FROM THE BOTTOM UP FOLLOWING THE ARROWS



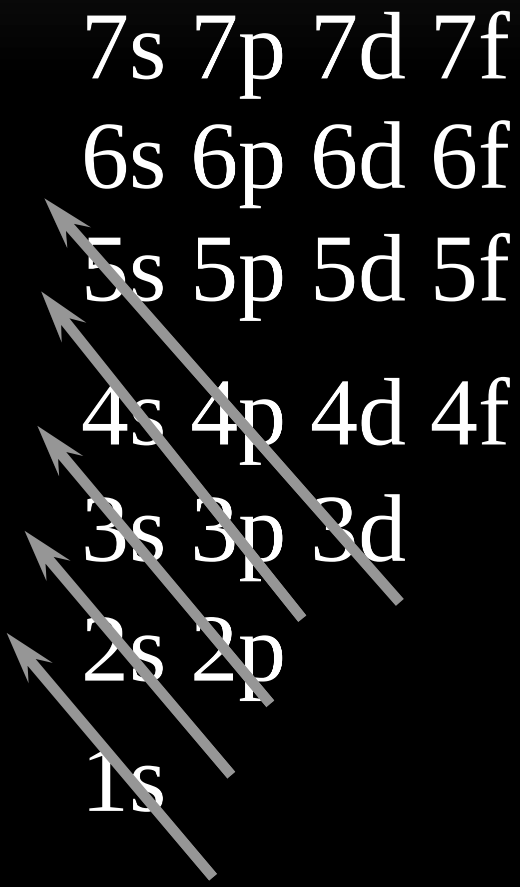
7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

Lets do this for Hassium (atomic Number 108)

- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

- 20 electrons

FILL FROM THE BOTTOM UP FOLLOWING THE ARROWS

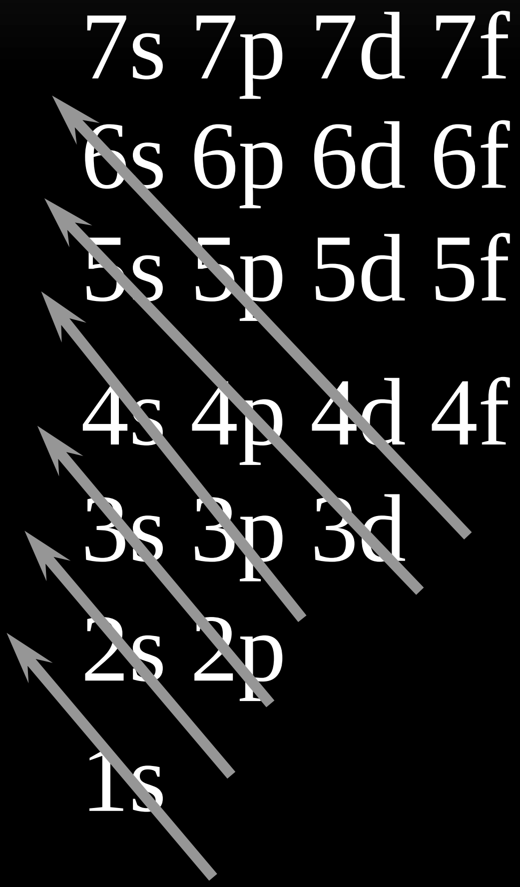

7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

Lets do this for Hassium (atomic Number 108)

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

- 38 electrons

FILL FROM THE BOTTOM UP FOLLOWING THE ARROWS



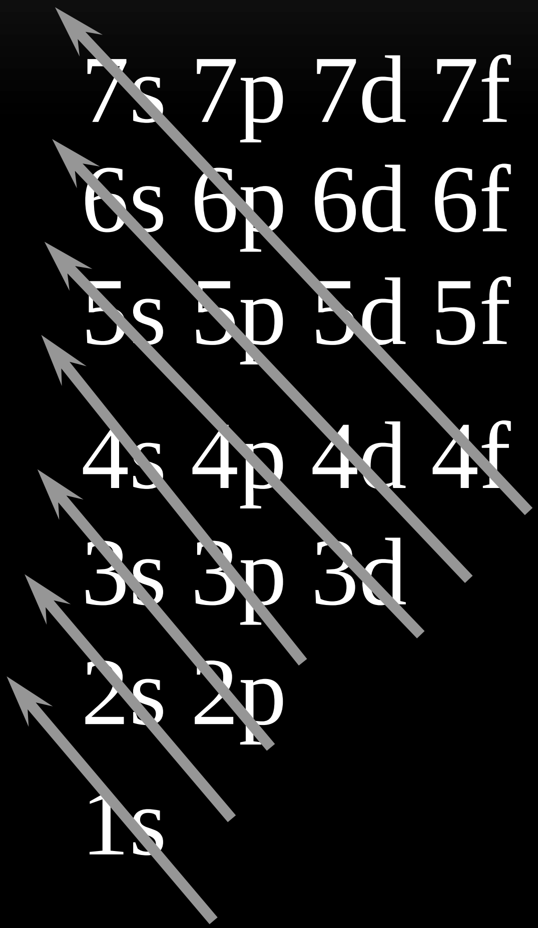
7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

Lets do this for Hassium (atomic Number 108)

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

- 56 electrons

FILL FROM THE BOTTOM UP FOLLOWING THE ARROWS

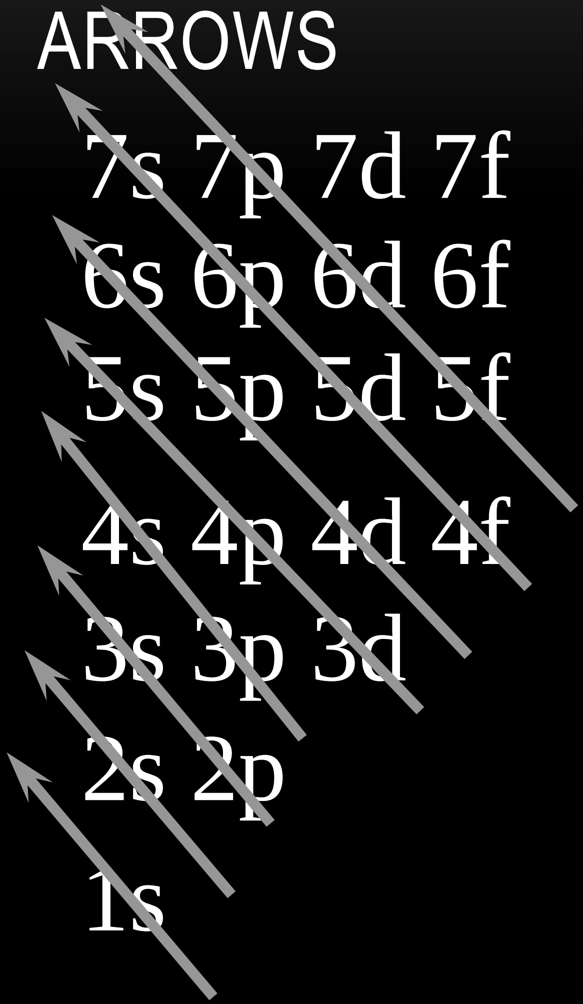


7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

Lets do this for Hassium (atomic Number 108)

- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2$
- 88 electrons

FILL FROM THE BOTTOM UP FOLLOWING THE ARROWS

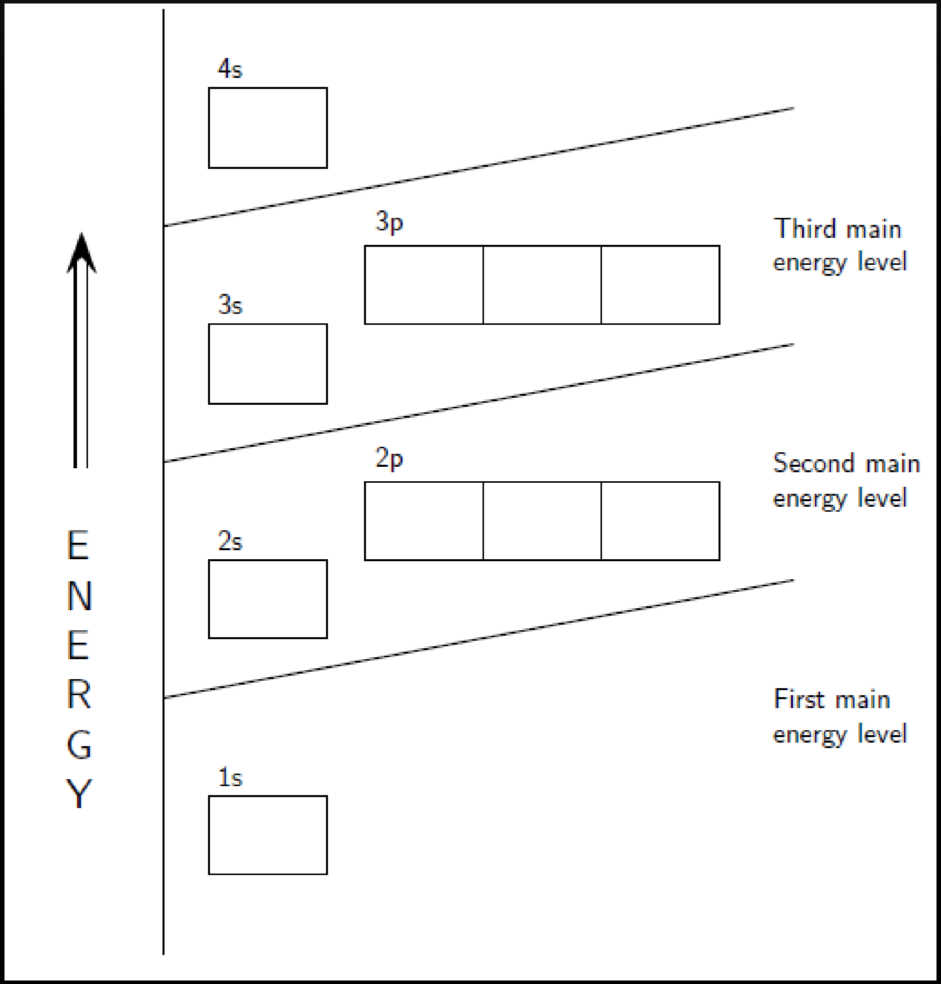


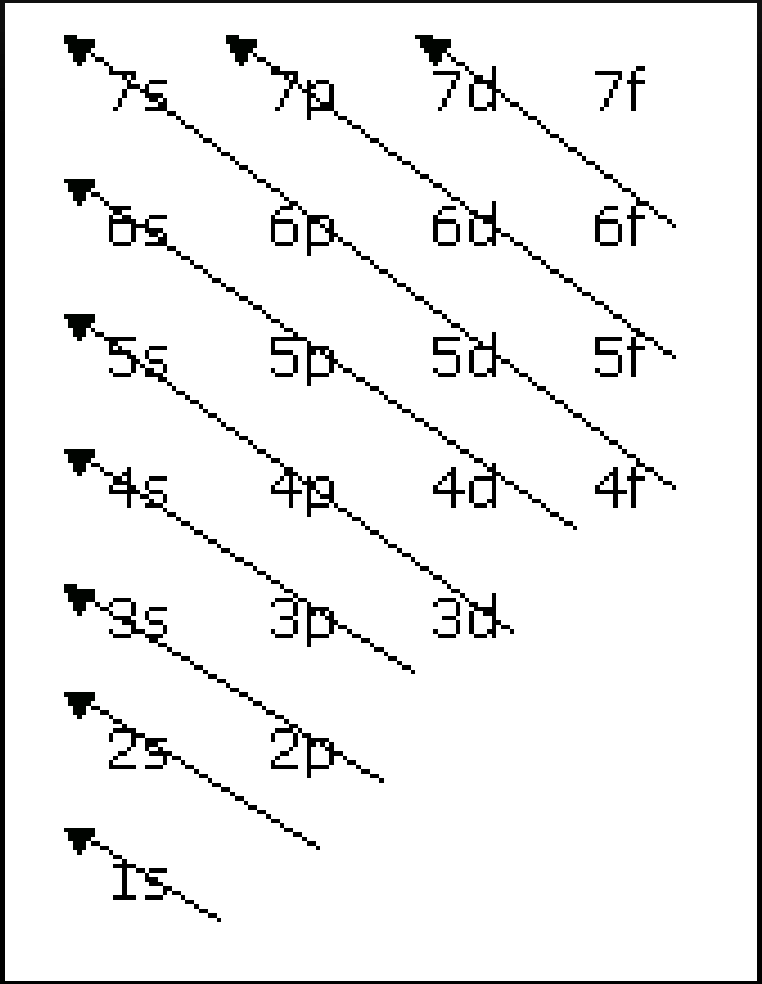
7s 7p 7d 7f
6s 6p 6d 6f
5s 5p 5d 5f
4s 4p 4d 4f
3s 3p 3d
2s 2p
1s

Lets do this for Hassium (atomic Number 108)

- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 5d^{10}\ 6p^6\ 7s^2\ 5f^{14}\ 6d^{10}\ 7p^6$
- 108 electrons

ORBITAL FILLING DIAGRAM





NOTATION

- Electron Configuration

S $16e^-$ $1s^2 2s^2 2p^6$ $3s^2 3p^4$

Core e^-

Inner Shell

Valence e^-

Outer Shell

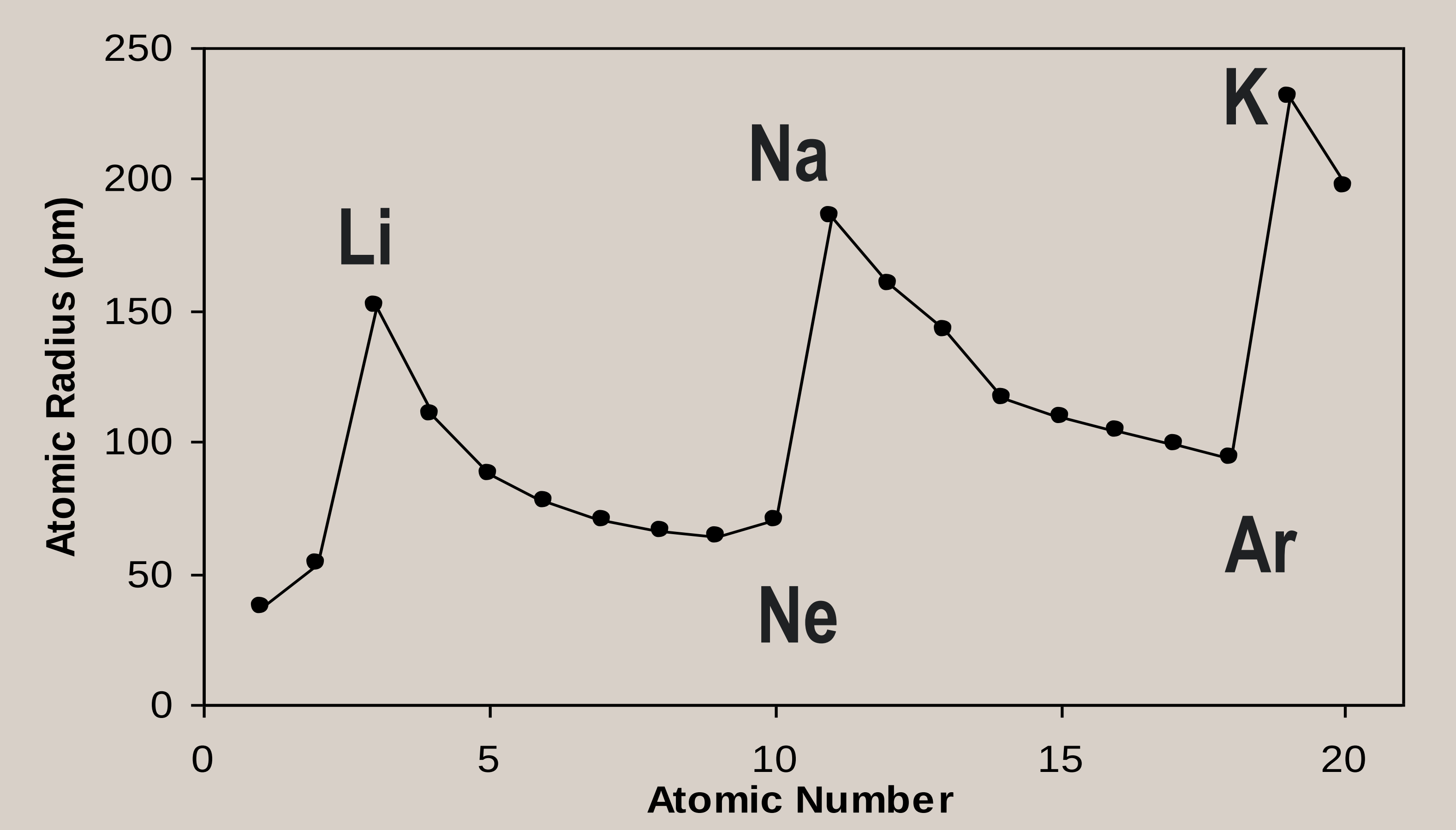
NOTATION

- **Noble Gas Configuration**
  - Keep only valence e$^-$
  - Use previous Noble Gas

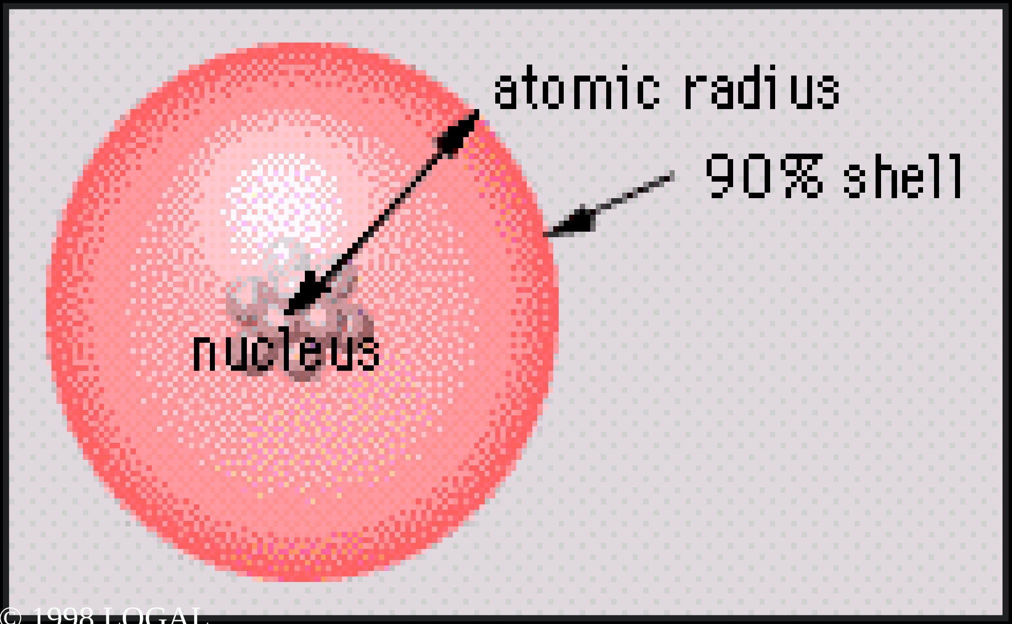
S 16e$^-$ [Ne] $3s^2\,3p^4$

ATOMIC RADIUS

- **Atomic Radius**



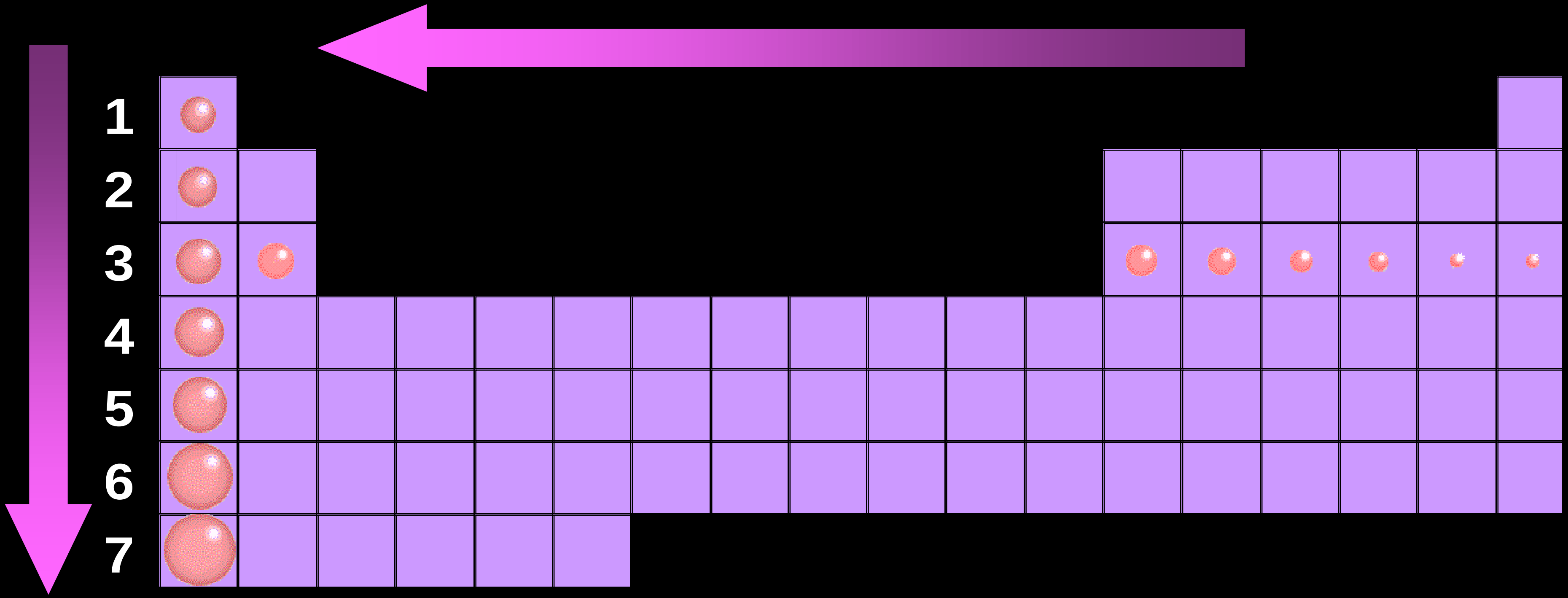
ATOMIC RADIUS



ATOMIC RADIUS

- **Atomic Radius**

Increases to the LEFT and DOWN



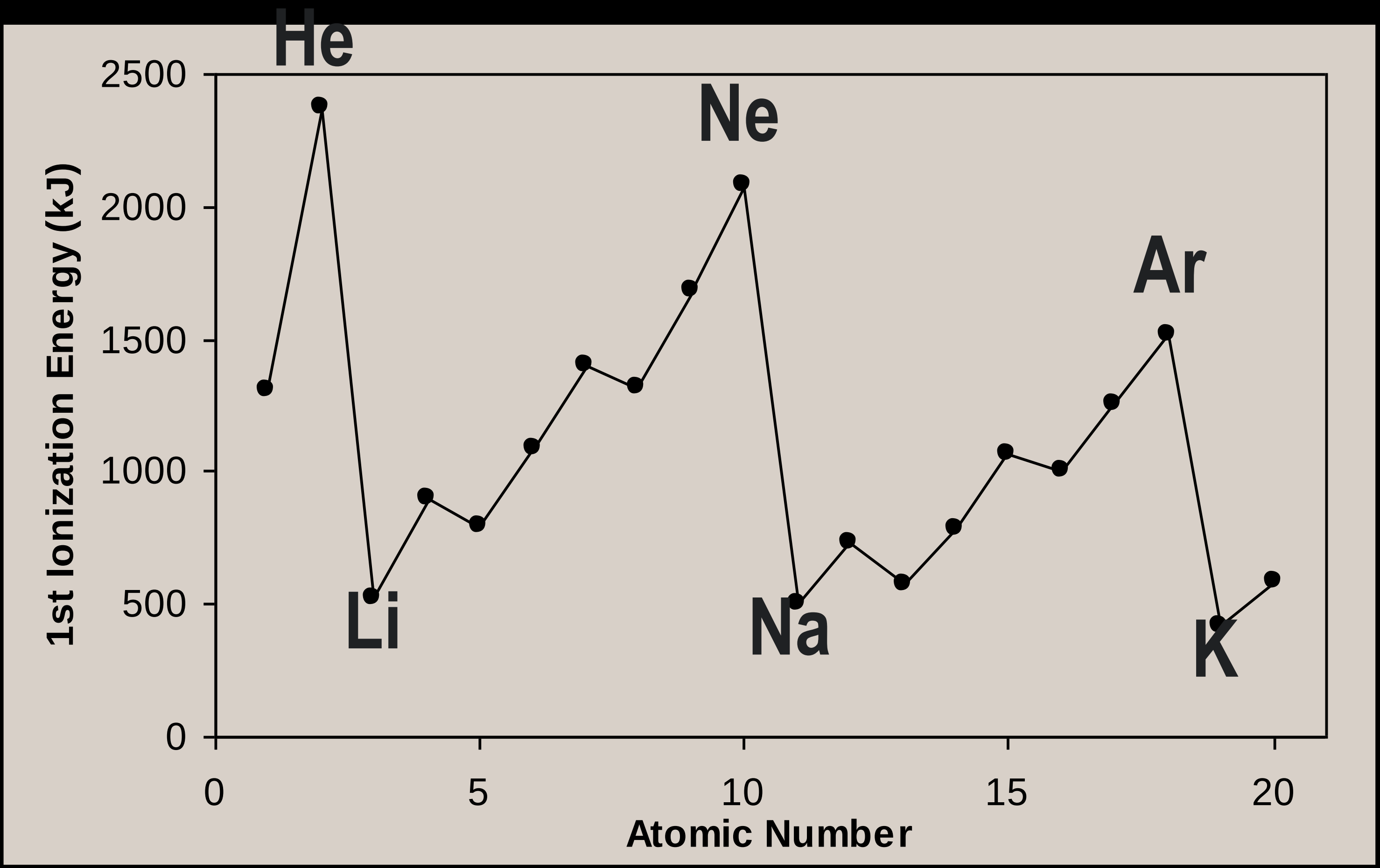
ATOMIC RADIUS

- **Why larger going down?**
  - Higher energy levels have larger orbitals
  - Shielding - core e^- block the attraction between the nucleus and the valence e^-

- **Why smaller to the right?**
  - Increased nuclear charge without additional shielding pulls e^- in tighter

IONIZATION ENERGY

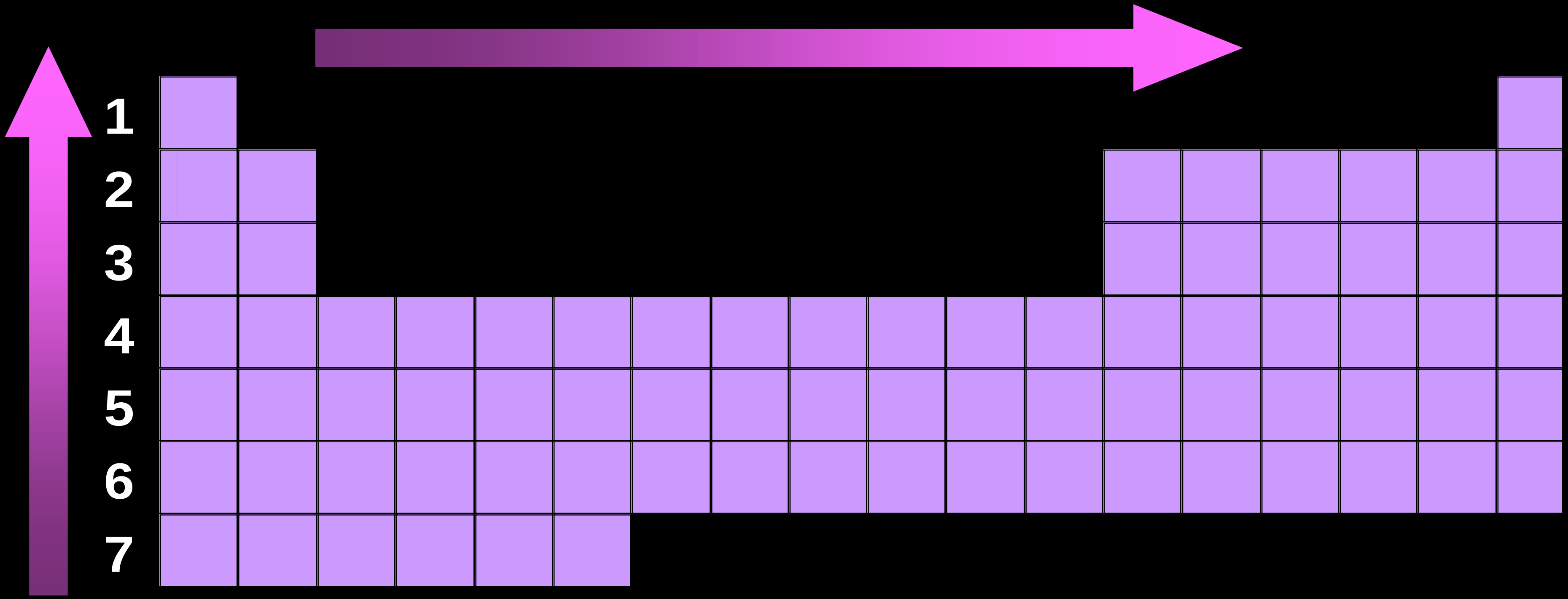
- **First Ionization Energy:** energy required to remove one e$^-$ from a neutral atom.



IONIZATION ENERGY

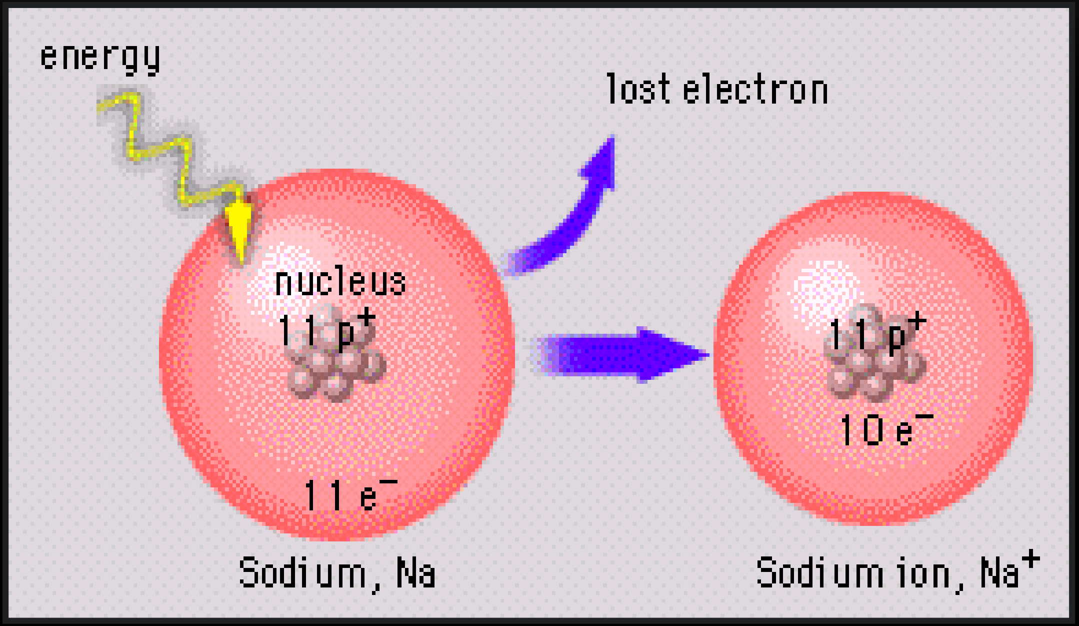
- **First Ionization Energy**

Increases UP and to the RIGHT



IONIZATION ENERGY

- **Why opposite of atomic radius?**
  - In small atoms, e$^-$ are close to the nucleus where the attraction is stronger
- **Why small jumps within each group?**
  - Stable e$^-$ configurations don't want to lose e$^-$



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IONIZATION ENERGY

- **Successive Ionization Energies**

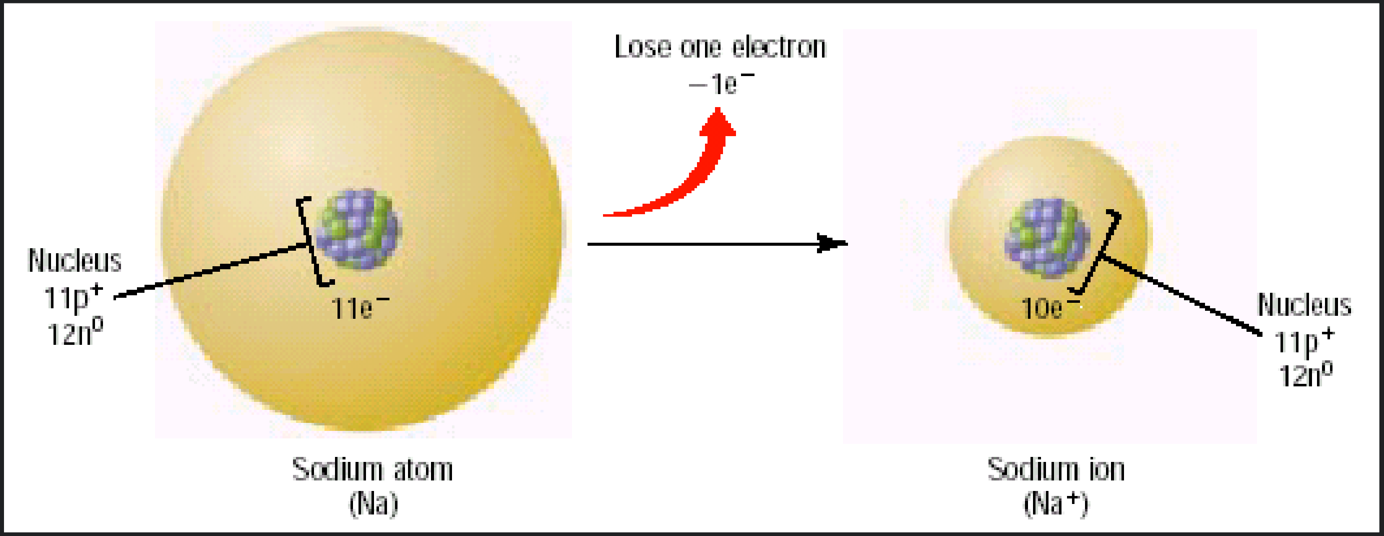
Large jump in I.E. occurs when a CORE e^- is removed.

Mg		1st I.E.	736 kJ
		2nd I.E.	1,445 kJ
Core e^-	→	**3rd I.E.**	**7,730 kJ**

IONIC RADIUS

Cations (+)

- ☒lose e^-
- ☒smaller



Anions (–)

- ☒gain e^-
- ☒larger