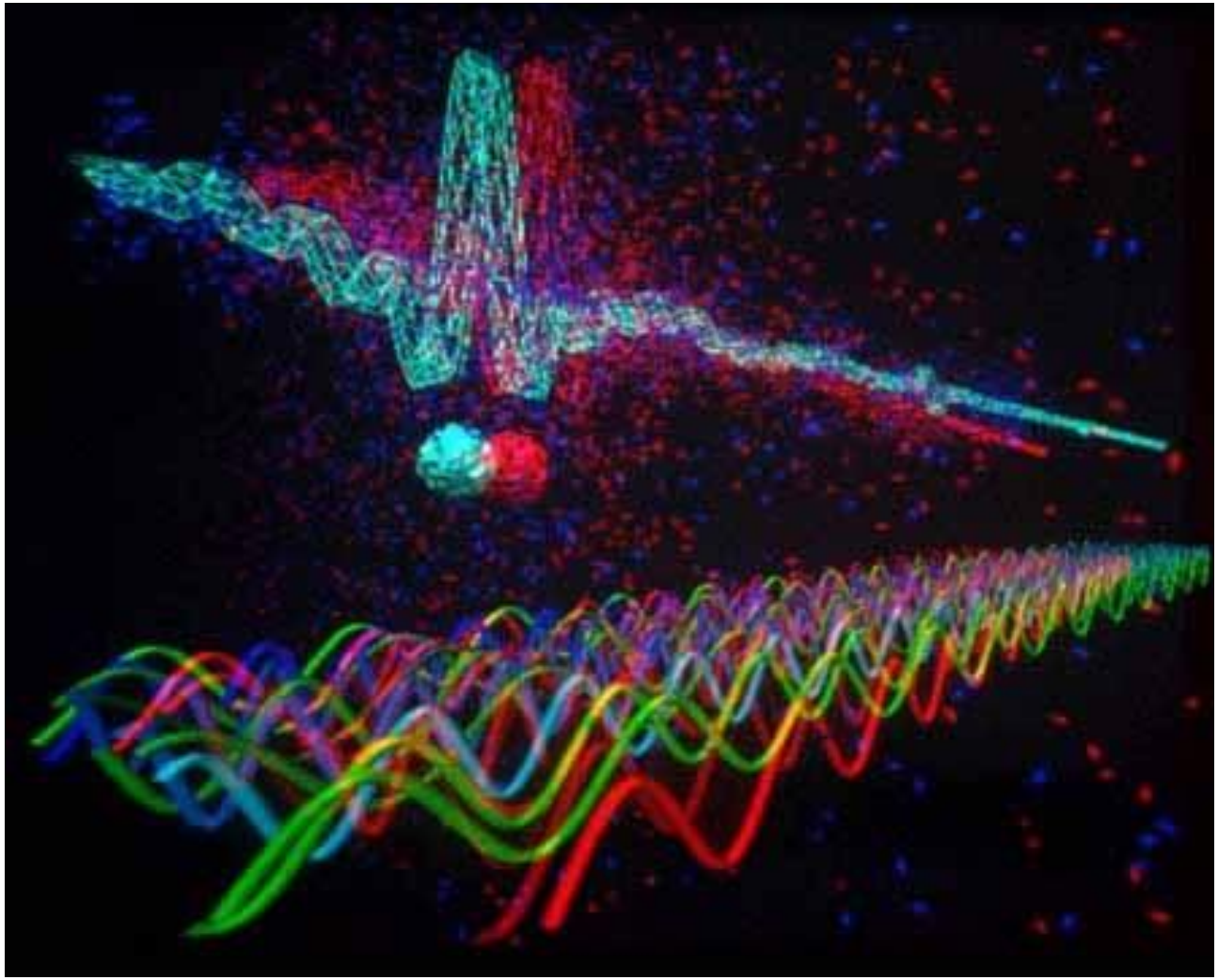
Atomic structure 2



Chapter 3

Arrangement of electrons in an atom

Learning Outcomes

- Emission and absorption spectra of the hydrogen atom .
- Balmer series in the emission spectrum as an example.
- Line spectra as evidence for energy levels.
- Energy sub-levels.
- Viewing of emission spectra of elements using a spectroscope or a spectrometer.

- Emission spectra known for hundreds of years
- Niels Bohr unlocked their secret
- Electrons orbiting in shells around the nucleus
- Won the 1922 Nobel Prize for Physics

Bohr

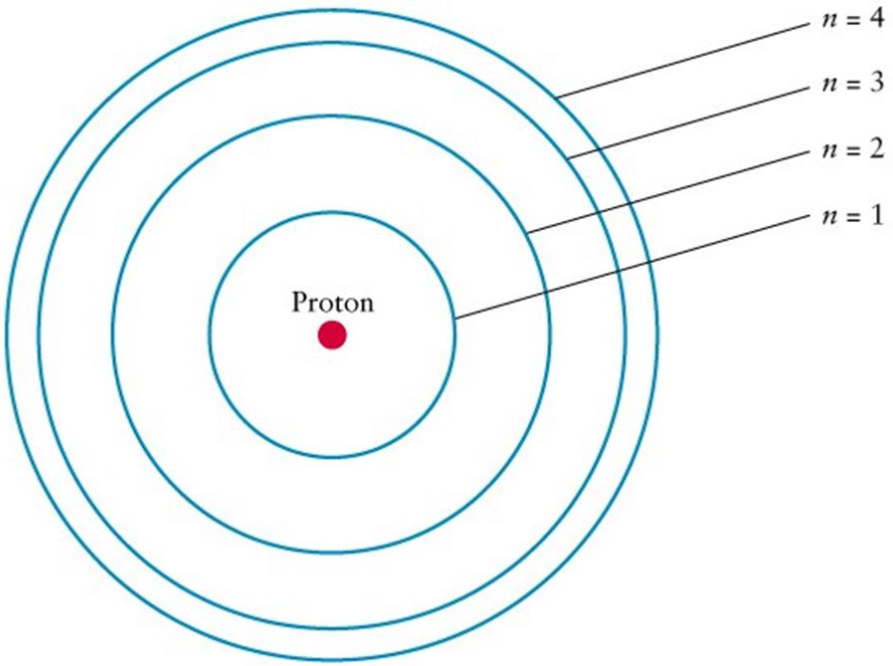
Niels Bohr

*Bohr's theory

- Electrons revolve around nucleus in fixed paths called orbits.
- Orbits are called energy levels (shells)
- Defn:
- **An energy level is the fixed energy of an electron**
- Or
- Shell which electrons of equal energy can occupy.

*

- Energy levels are represented by letter n
- - called the Principal quantum number



n=1 nearest nucleus, lowest energy
n=2
n=3
n=4 furthest from nucleus, highest energy

*quantisation

- Electrons can have only certain particular values of energy

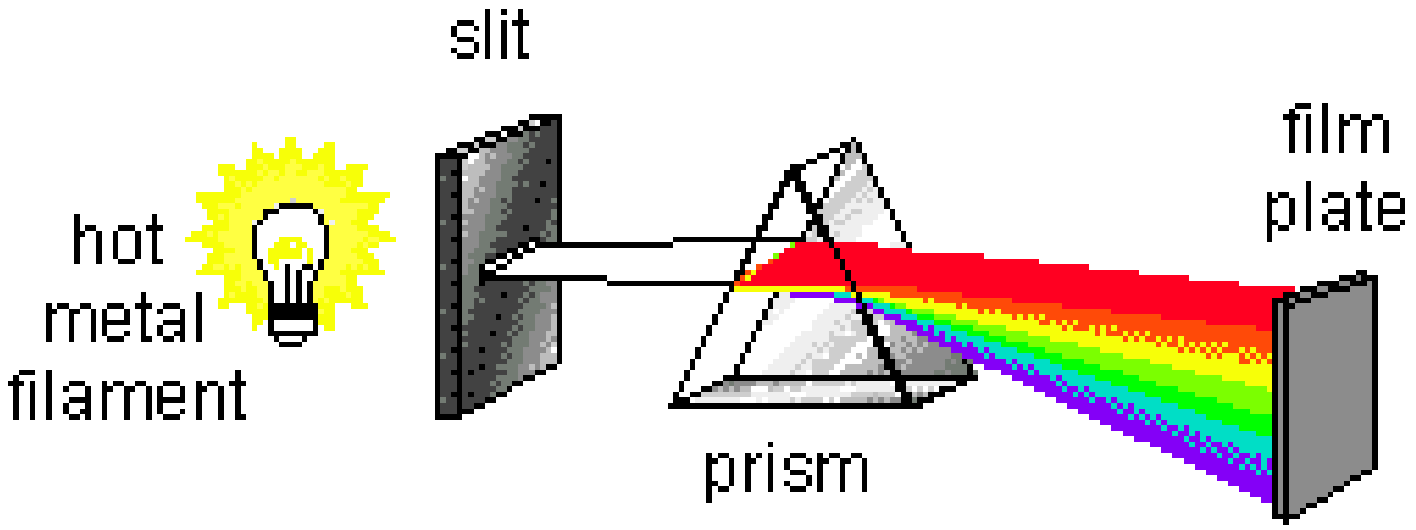
*Evidence for Energy Levels

- Came from Bohrs study of spectra
- There are 3 types of spectra

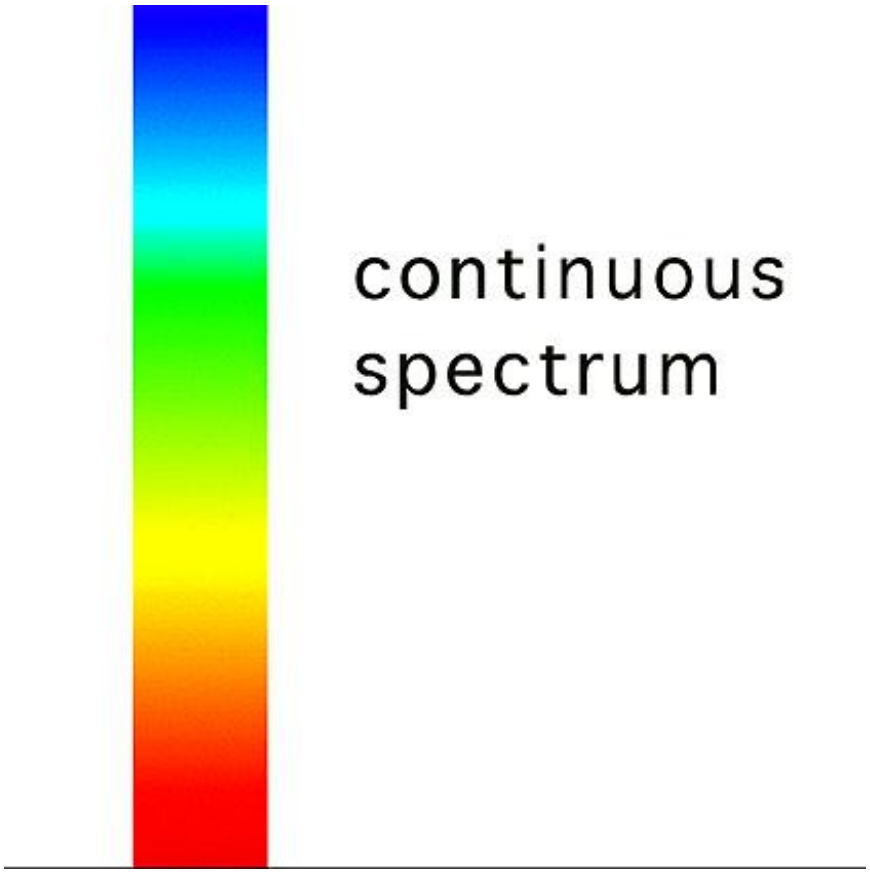
1. Continuous
2. Line Emission
3. Line Absorption

*1.Continuous

- A continuous spectrum is produced when white light is split by a prism into colours
- Eg a rainbow is a continuous spectrum



Continuous spectrum



A Spectrum in which all wavelengths are present between certain limits.

*2. Line Emission spectra

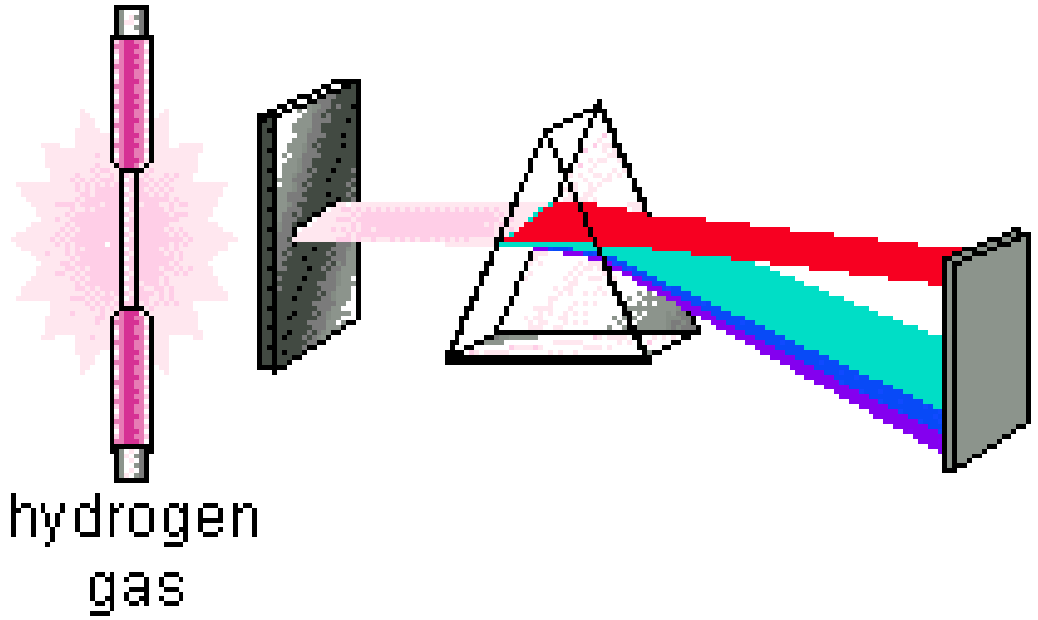
- Light is emitted from atoms when they are supplied with energy.
- This energy can be in the form of ***heat***
  – eg the flame tests) or

  electricity,

  (in special lamps containing the element called discharge tubes)
- If this light is viewed, after passing through a prism, a series of lines can be seen.

Line emission spectrum

Discharge tube →

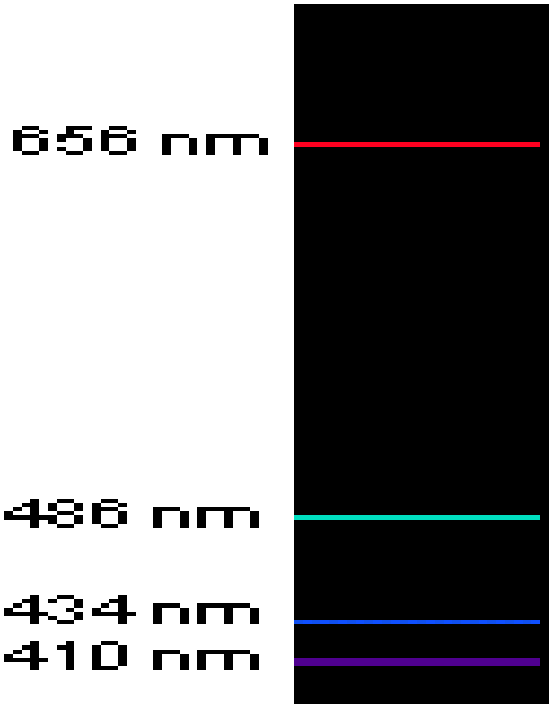


*

- **A line emission spectrum** is a series of coloured lines against a dark background, produced by passing light from a heated gas through a prism
- **A spectroscope** is an instrument used to analyse light emitted by elements.

*

- Each element has a unique line spectrum so they are used to identify elements in an unknown sample.



*Everday applications

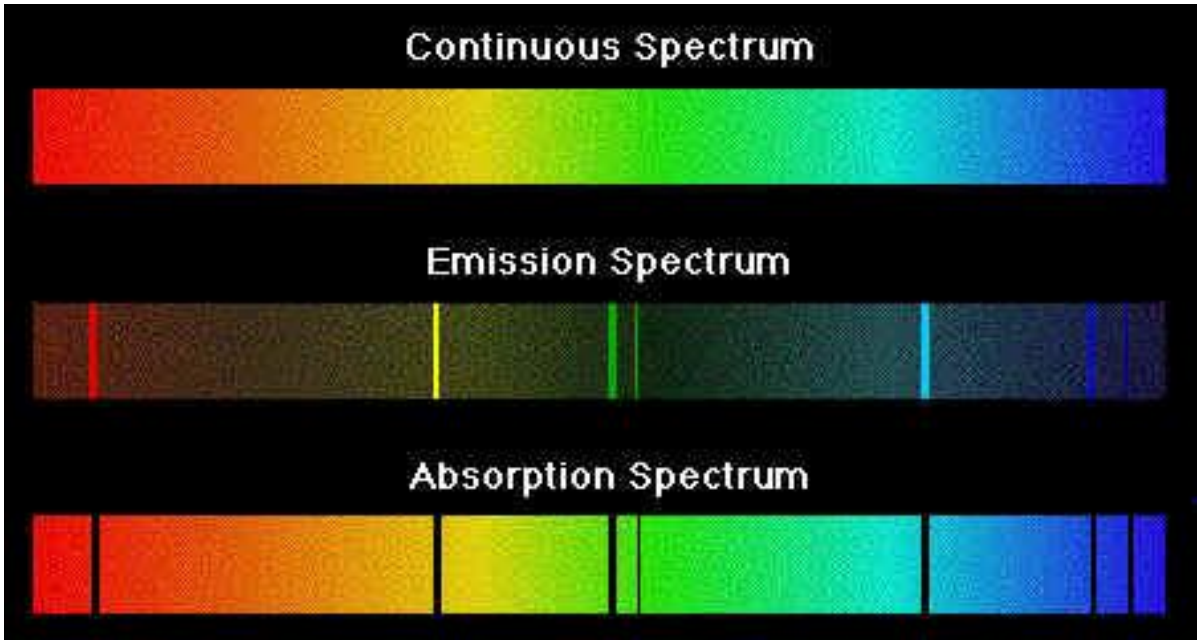
- Street lights are discharge tubes containing sodium, they emit yellow light
- Fireworks contain metals which emit different colours when heated.

*3. Line Absorption spectra

- Atoms can also absorb light.
- **A line absorption spectra** is produced when light passes through a gaseous element and some wavelengths of light are absorbed.
- It is a series of dark lines against a coloured background.

*

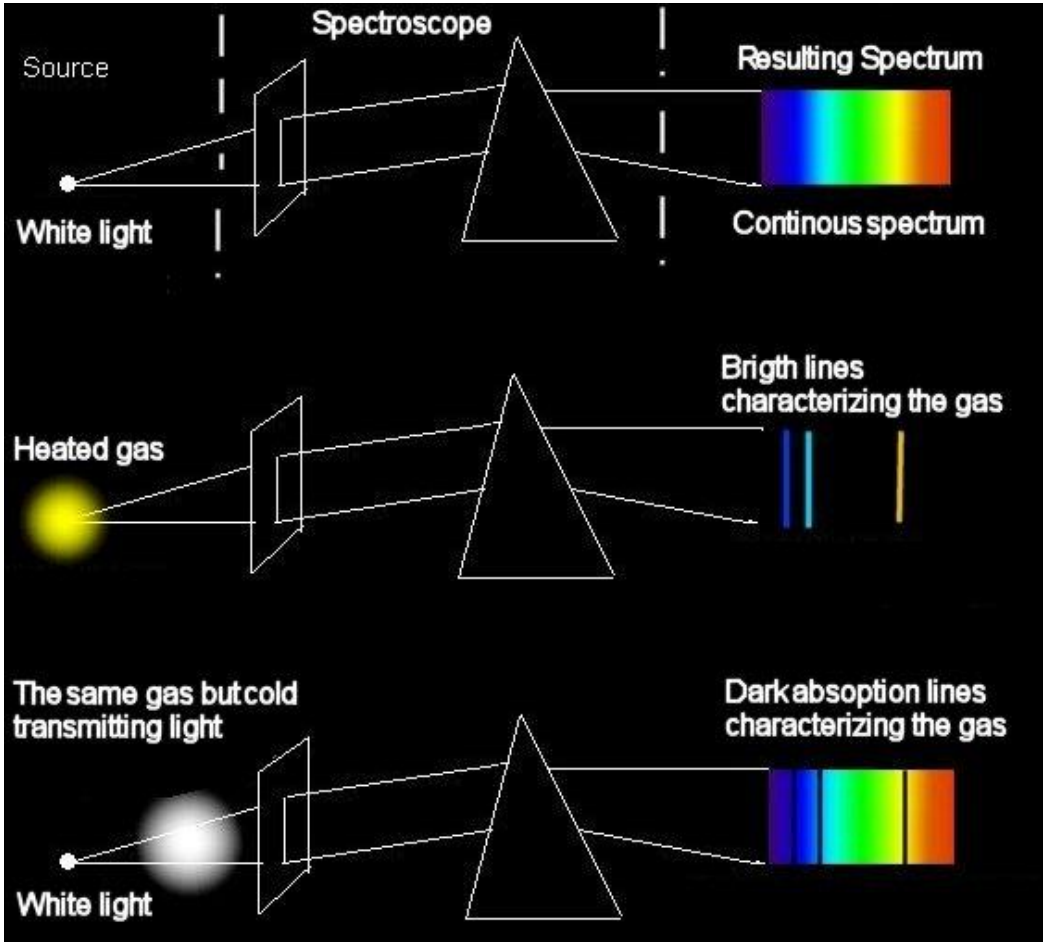
- The dark lines are at the same wavelengths as the coloured lines in the emission spectrum



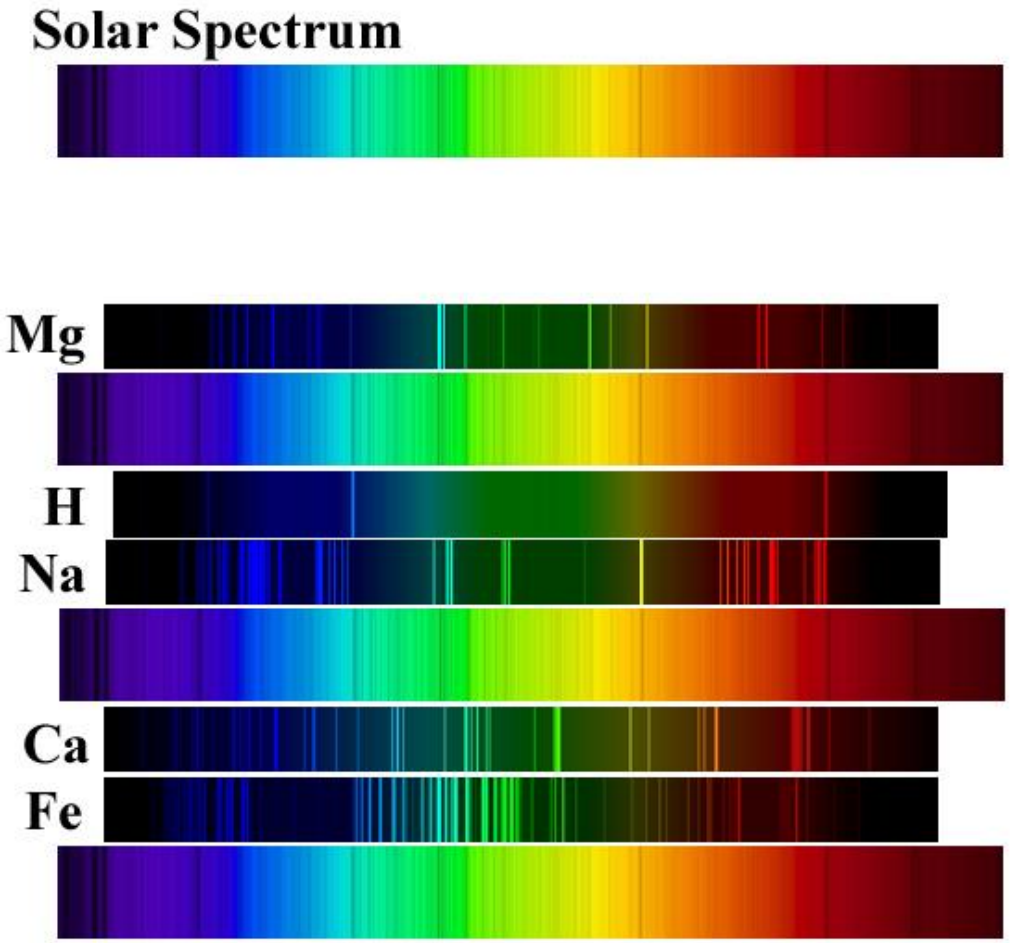
*Atomic Absorption Spectrometry

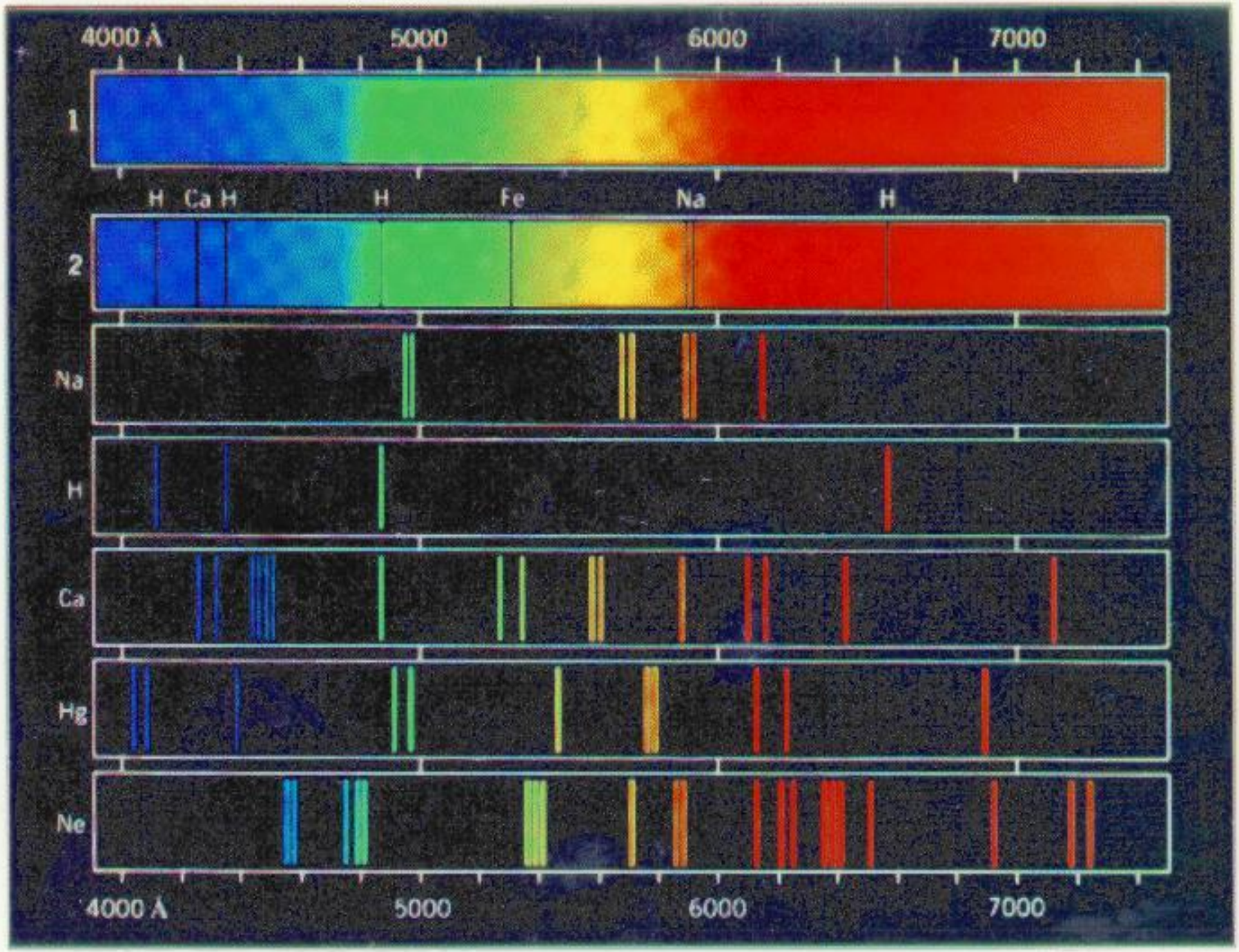
- Each element has an unique absorption spectrum.
- This is the basis of Atomic Absorption spectrometry. (AAS)
- AAS is used to detect heavy metals in soil or water. Eg Lead

Comparing the spectra



Emission spectrum of Sun



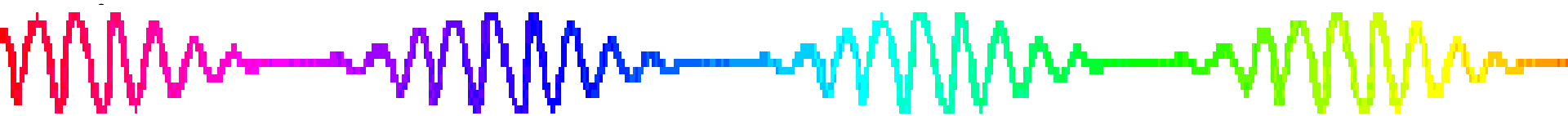

4000 Å
5000
6000
7000
1
H Ca H
H
Fe
Na
H
2
Na
H
Ca
Hg
Ne
4000 Å
5000
6000
7000

*

How Bohr explained line spectra of Hydrogen and used this as evidence for energy levels.

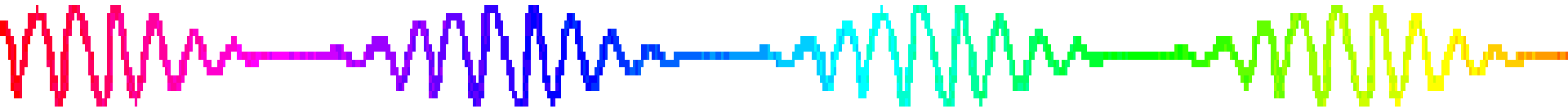
*Bohr's Theory

- Electrons occupy fixed energy levels.
- An electron is normally in its lowest available energy level- called the ground state (E_1)
- The electron can jump to a higher energy level, (E_2), if it absorbs energy in form of heat, light or electricity.
- The energy absorbed must equal the energy difference between the ground and excited state.
- This excited state is unstable.
- The electron falls back to a lower level.



*Bohr's Theory

- Energy is emitted as light of a specific frequency.
- The energy emitted (hf) corresponds to the energy difference between the 2 levels. (E_2- E_1)
- It can be calculated using E_2- E_1 = hf, where h is Planck's constant and f is frequency.
- This gives rise to light of a particular frequency and wavelength, which is seen as a line of a particular colour on the spectrum.

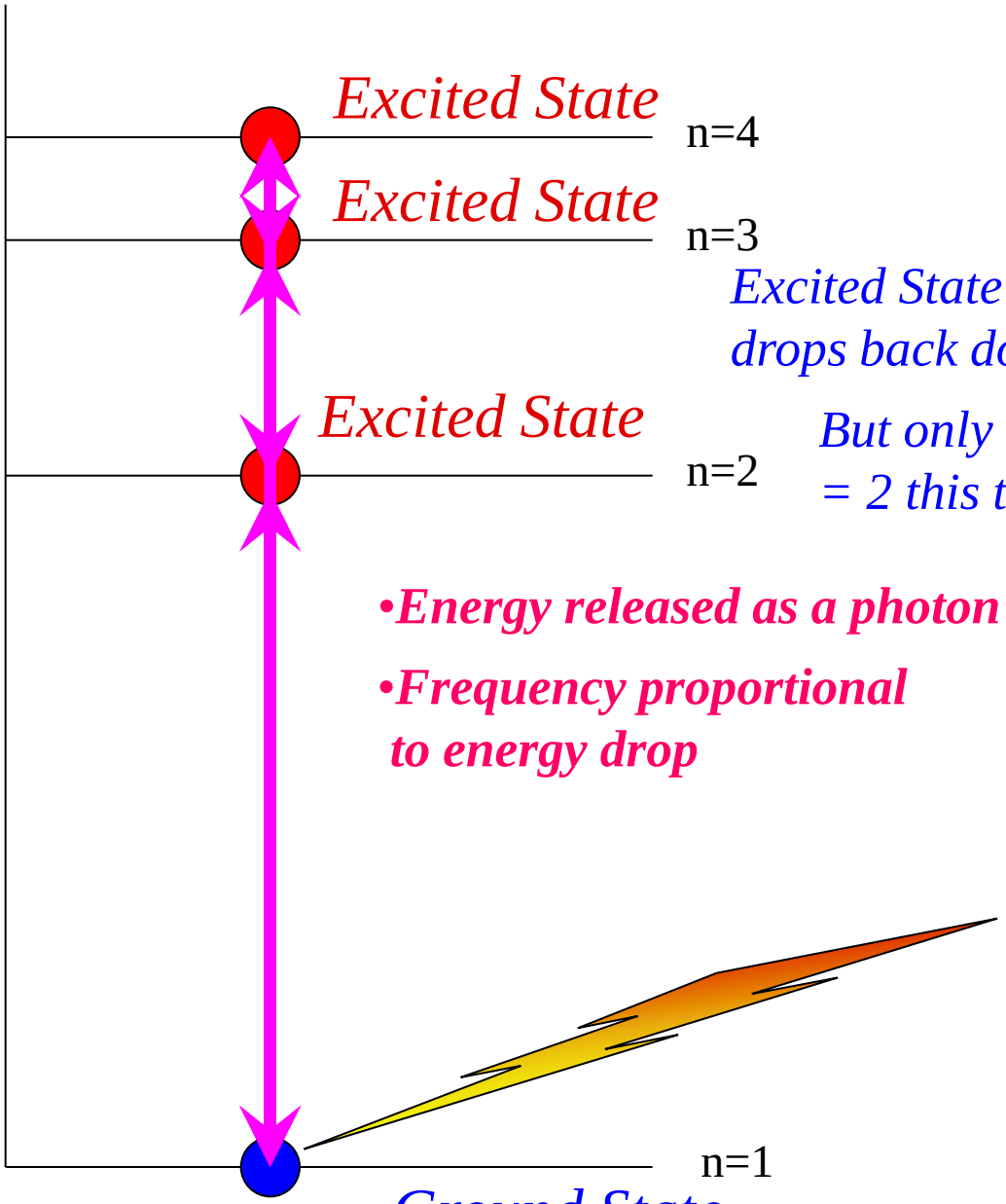


*

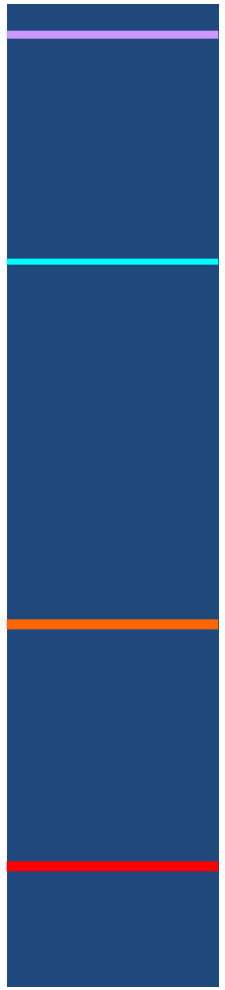
- *What evidence do Line spectra provide for the existence of energy levels?*
- This **is evidence for energy levels** since only specific frequencies are emitted so electrons must be restricted to specific energy levels.

Note:

- Ground state = lowest energy state
- Excited state = higher energy state


Excited State
n=4
Excited State
n=3
Excited State unstable and drops back down
Excited State
n=2
But only as far as n = 2 this time
•Energy released as a photon
•Frequency proportional to energy drop
n=1
Ground State

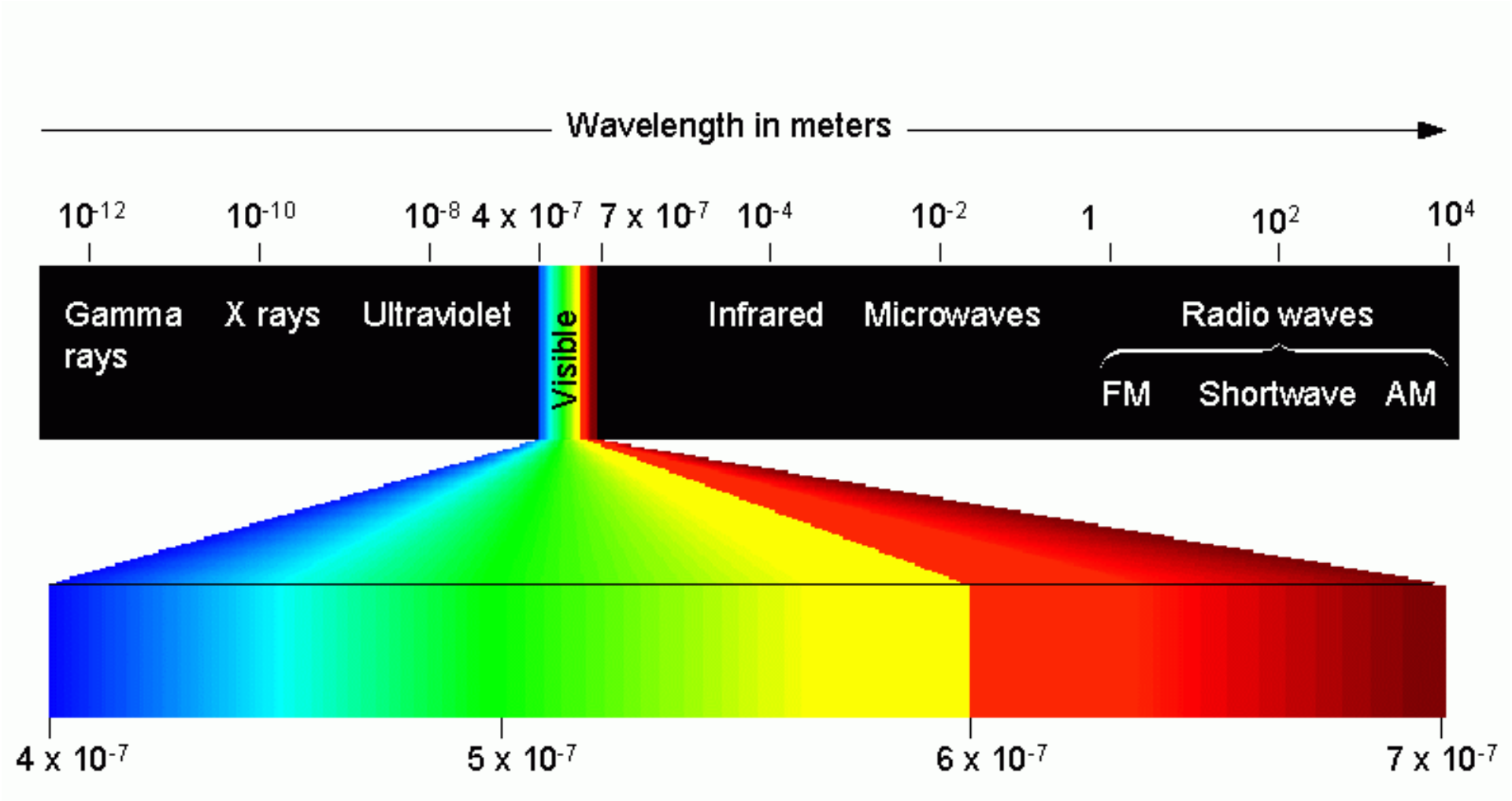
Spectrum


UV
Visible
IR

**Why do different elements have unique spectra?*

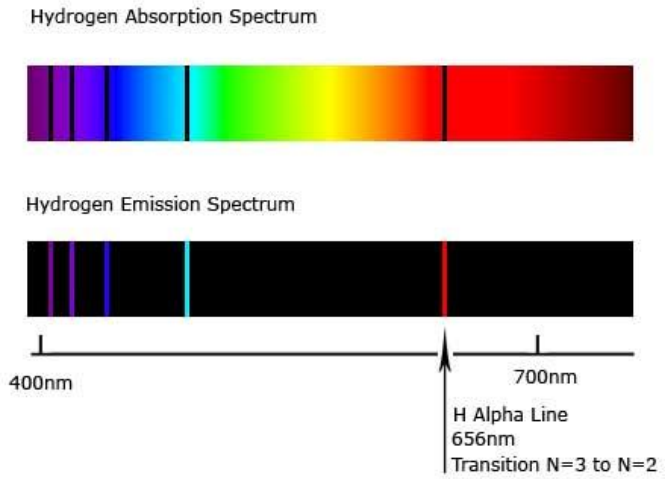
- Because each element has different electron configurations, giving rise to different electron jumps.

Types of electromagnetic radiation:

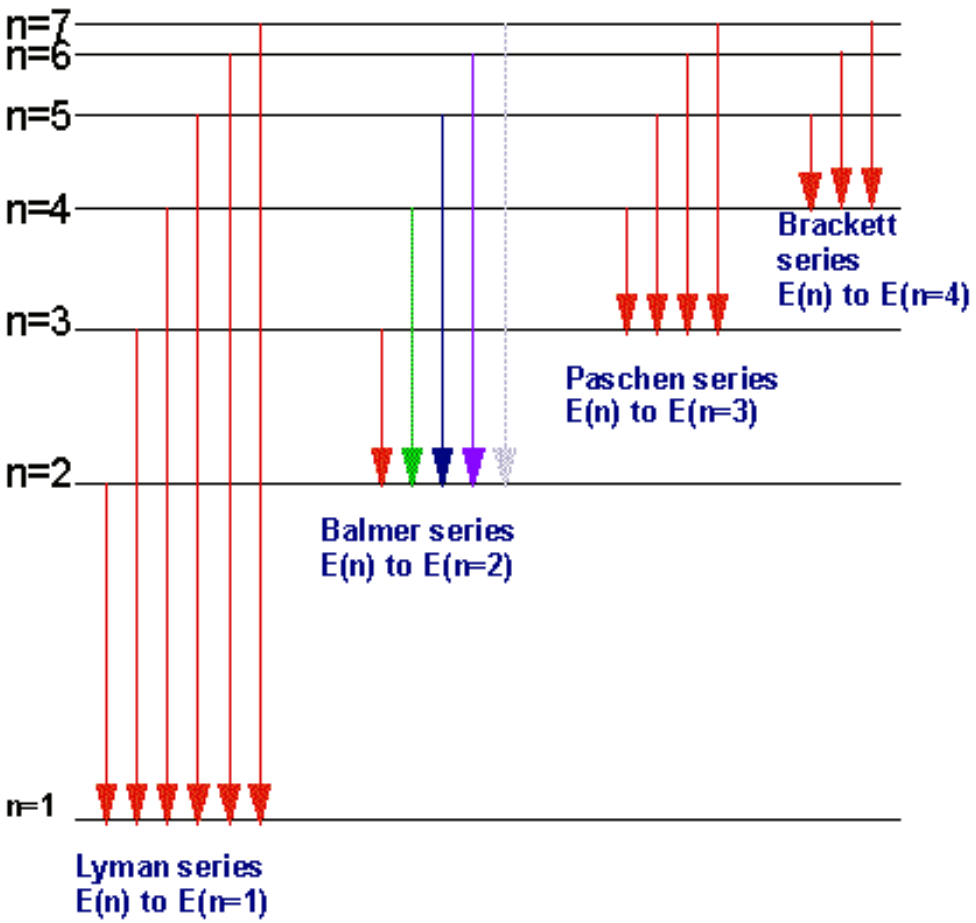


*Balmer Series

- In the emission spectrum of hydrogen,
- the series of lines in the <u>visible region</u> are called the **<u>Balmer series:</u>**
- They are due to electrons falling back to the n=2 shell from higher levels.



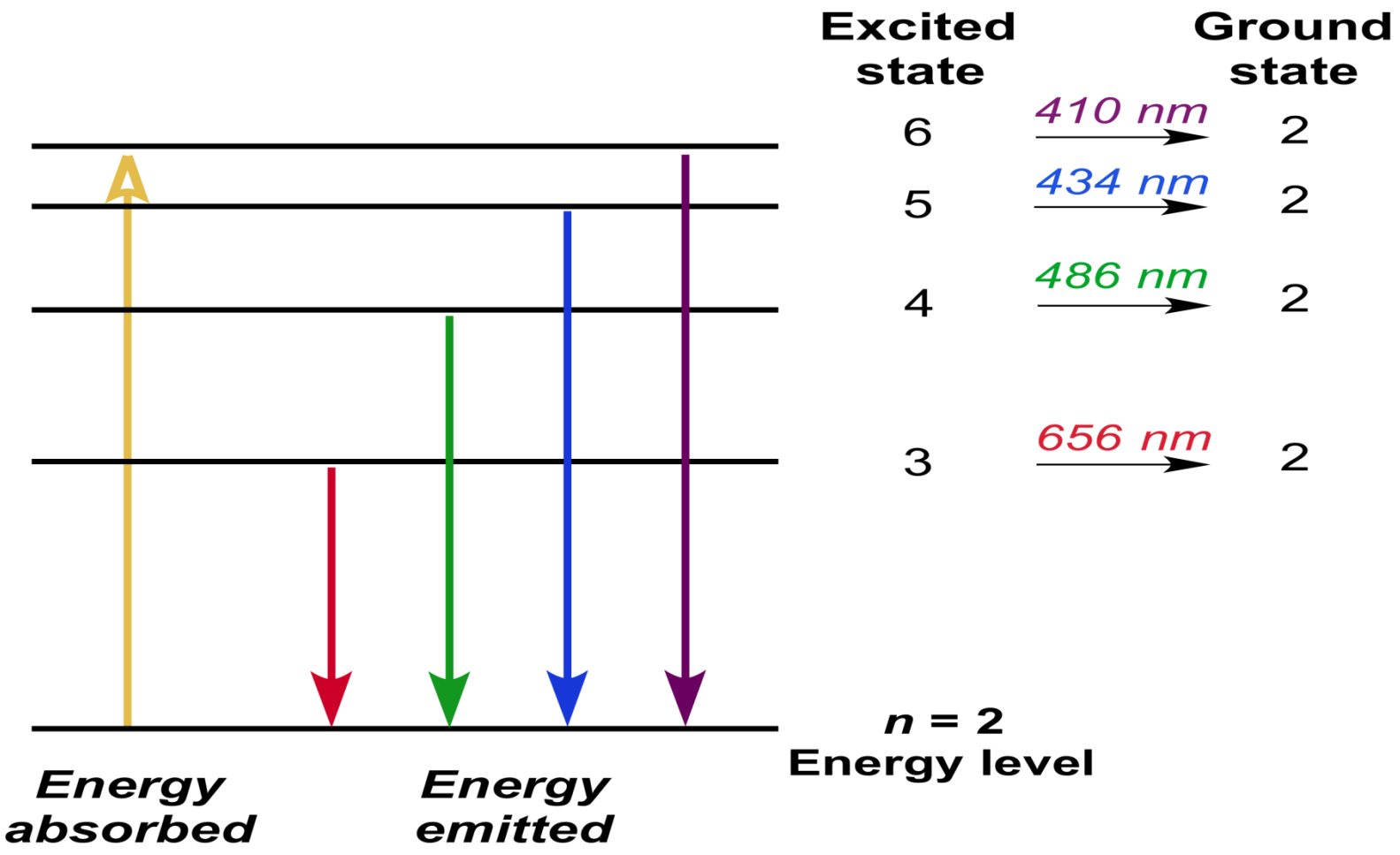
Electron transitions for the Hydrogen atom



*Balmer Series

Note that the first (red) line in the Balmer series is due to electrons falling from the n=3 to the n=2.

The second (green) line is due to electrons falling from the n=4 to the n=2 and so on.


Excited state
Ground state
6
5
4
3
2
2
2
2
410 nm
434 nm
486 nm
656 nm
n = 2
Energy level
Energy absorbed
Energy emitted

*Flame Test experiment

- **Metals emit characteristic colours of light when their atoms are excited.**
- **Method**
- Dip soaked wooden splint in the salt to be tested.
- Hold the splint in the blue flame.
- Note the colour produced.

*Flame Test results

Metal	Colour
Sodium (Na)	Yellow
Lithium (Li)	Pink/Red
Potassium (K)	Lilac
Copper (Cu)	Blue-Green
Calcium (Ca)	Pink
Barium (Ba)	Green
Strontium (Sr)	Red/Orange

*Why are different colours produced by metals in flame tests?

Different metal atoms have different electron arrangements so they emit different frequencies of light.

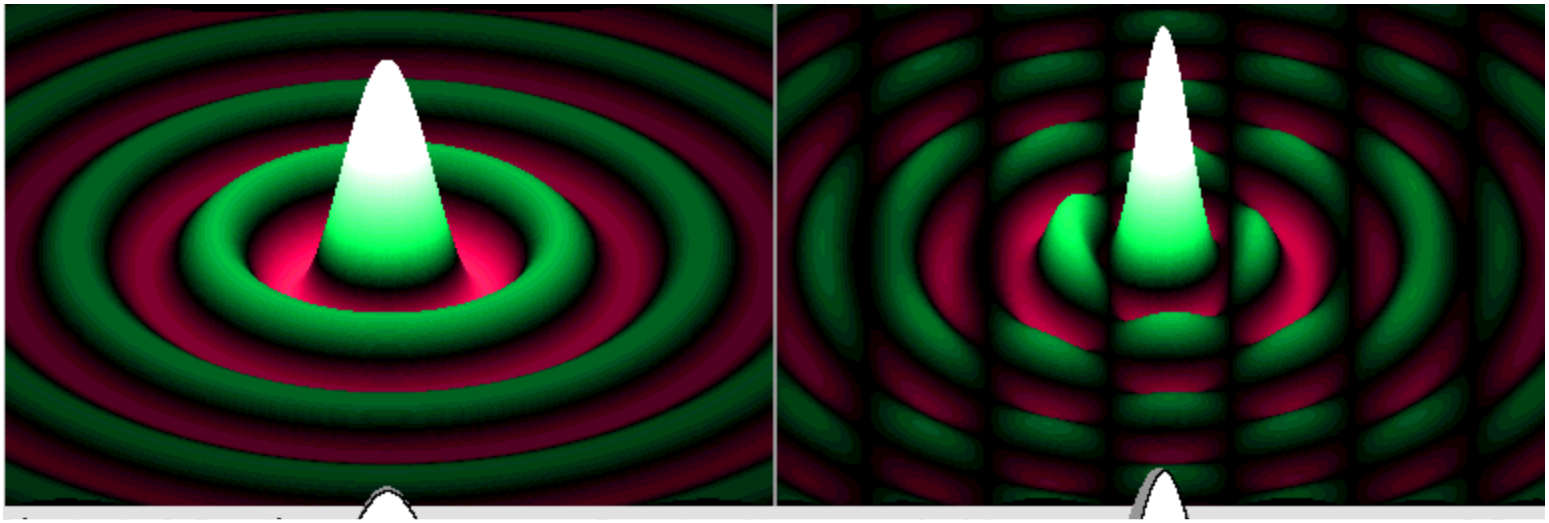
They have different spectra also.

*Limitations of Bohr's theory

- *Bohr's theory had to be modified because:*
- It only worked for simple atoms such as H.
- Electrons were found to have wave nature (by De Broglie).
- **Heisenberg's uncertainty principle** proposed that it is impossible to measure the exact position and velocity of an electron simultaneously.
- Energy sub-levels were discovered.

*De Broglie

- Discovered that electrons have wave characteristics.



Electrons were both particles and waves

Same for all sub-atomic particles

Matter exists as particles and waves at the same time.

Called wave particle duality.

Heisenberg Uncertainty Principle

Werner Heisenberg

"One cannot simultaneously determine both the position and momentum of an electron."

You can find out where the electron is, but not where it is going.

OR…

You can find out where the electron is going, but not where it is!

**Heisenberg's Uncertainty Principle*

- States that:
- It is not possible to measure the position and velocity of an electron at the same time.

*Discovery of sub-levels

- Magnification of lines in spectra revealed that some lines were made up of finer lines close together- called the Zeeman effect
- Each main level has sub-levels in it.
- Defn:
- **An energy sub-level** is a sub-division of an energy level which consists of orbitals of equal energy.

*

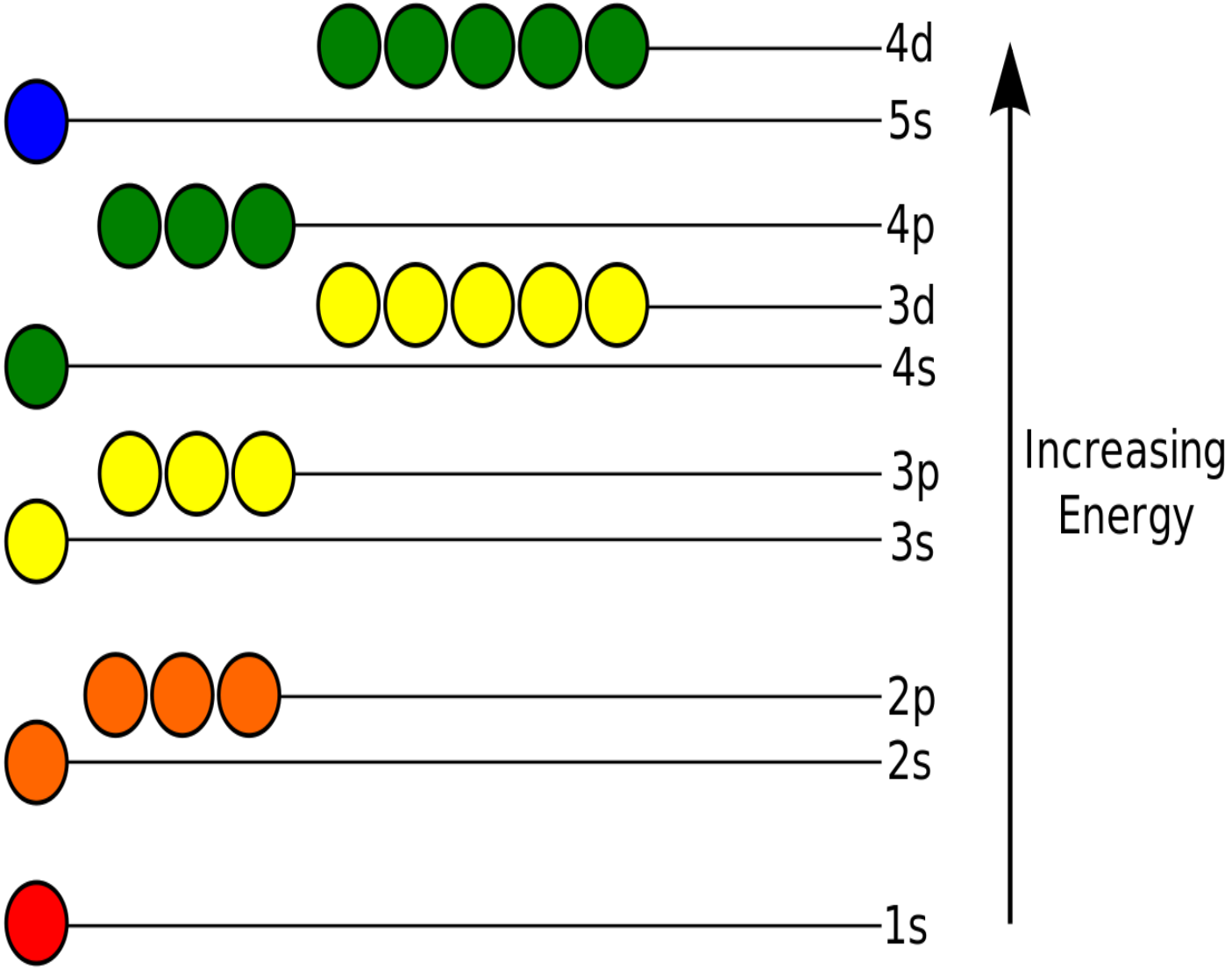
- There are 4 types of sublevels, **s, p , d and f.**
- **The number of sublevels in a main energy level is the same as the value of n for the main energy level.**
- n = 1 so there is 1 sub-level only, an s

  n = 2 so there are 2 sub-levels, an s and p
- n = 3 so there are 3 sub-levels, an s, p and d
- n = 4 so there are 4 sub-levels, an s, p, d and f

MAIN ENERGY LEVEL	KINDS OF SUBLEVELS
1	1s
2	2s, 2p
3	3s, 3p, 3d
4	4s, 4p, 4d, 4f
5	5s, 5p, 5d, 5f

*

- *Energy of sub-levels s→p→d*
- *S is lower in energy than p and so on.*

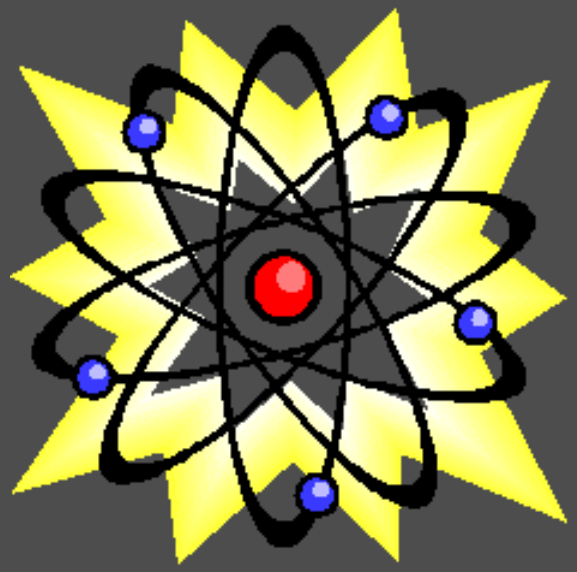

4d
5s
4p
3d
4s
3p
3s
2p
2s
1s
Increasing
Energy

The Bohr Model of the Atom

Niels Bohr

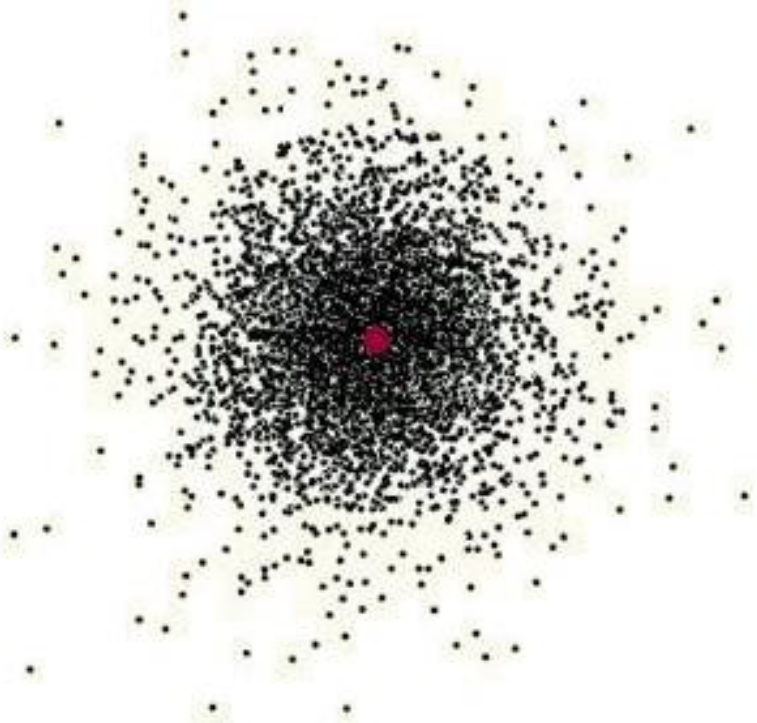
I pictured electrons orbiting the nucleus much like planets orbiting the sun.

But I was wrong! They're more like bees around a hive.



WRONG!!!

Electron cloud model- Schrodinger

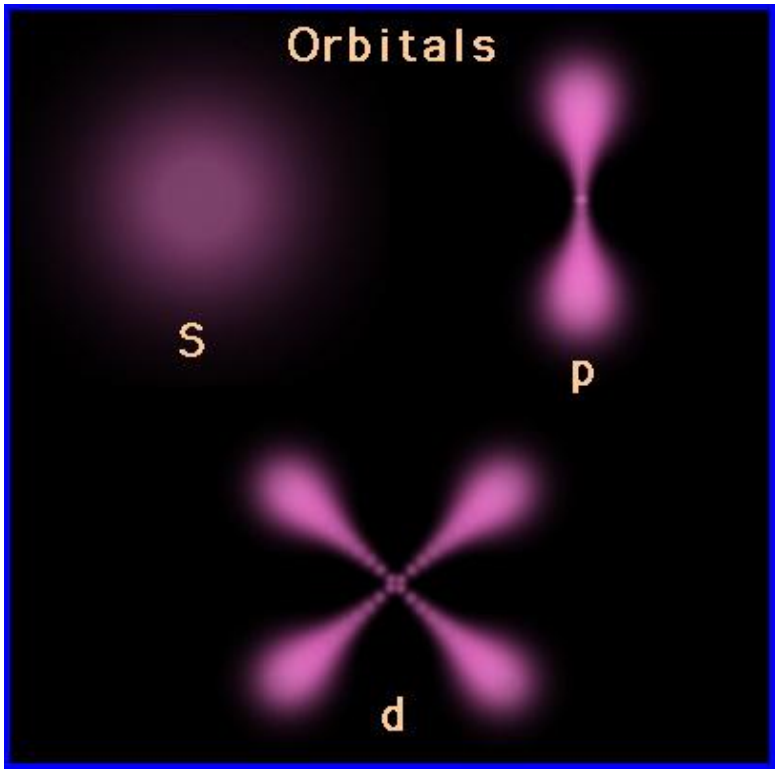


The Modern Model

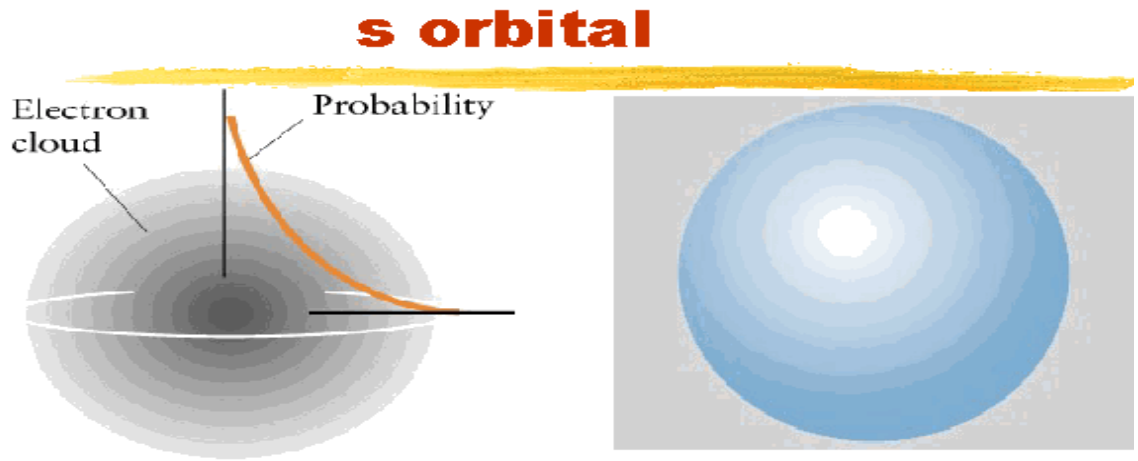
- Heisenberg's uncertainty principle meant that electrons could not be moving with a certain speed in fixed orbits from the nucleus.
- The location of the electron could only be given as a **probability** of it being within a certain space inside the atom.
- The notion of orbits was replaced by **orbitals**

*Atomic Orbitals

- Defn:
- A region around the nucleus where there is a high probability of finding an electron.

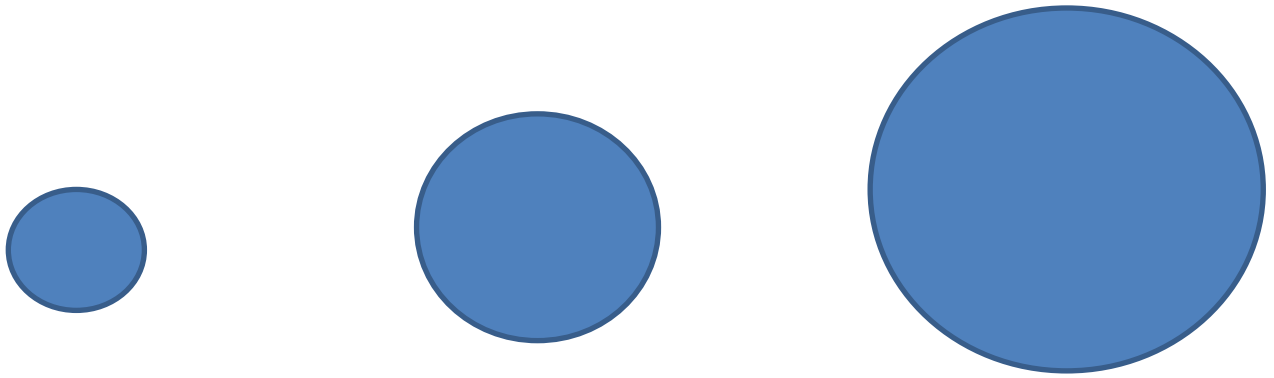


*Schrodinger solved wave equations and discovered shapes of orbitals.



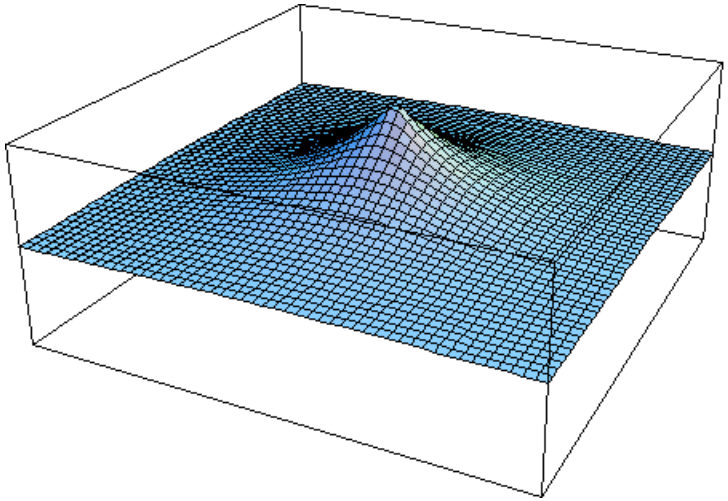
*Shapes of Orbitals

- S orbitals are Spherical.
- 1s is closer to the nucleus and lower in energy than 2s.

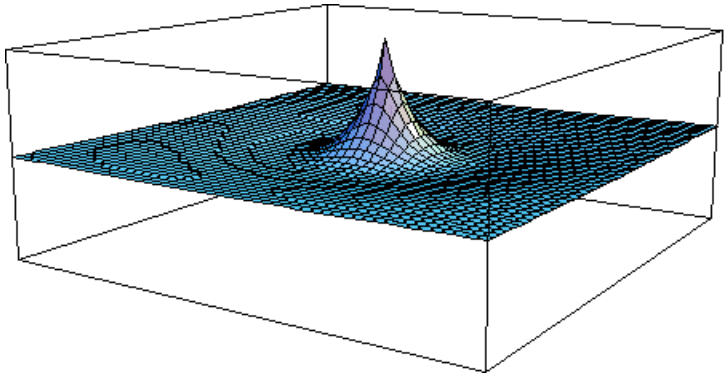


1s 2s 3s

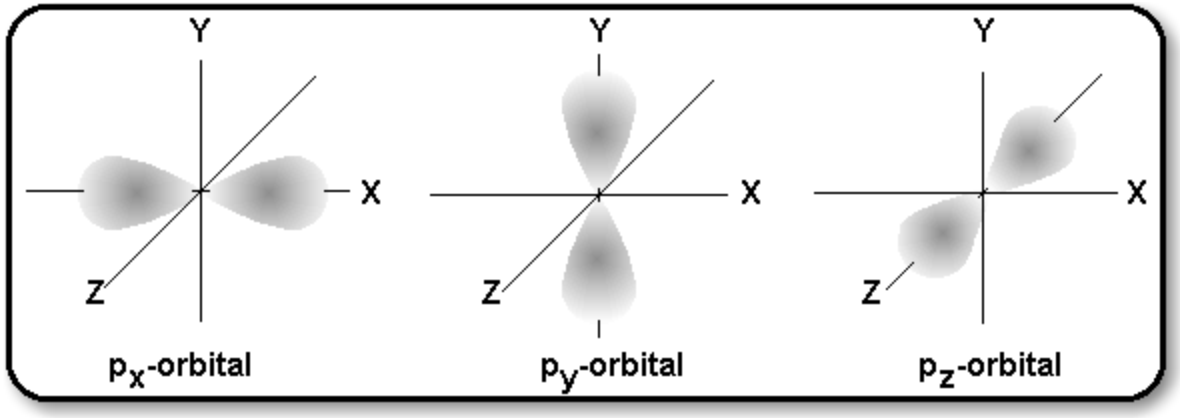
1s



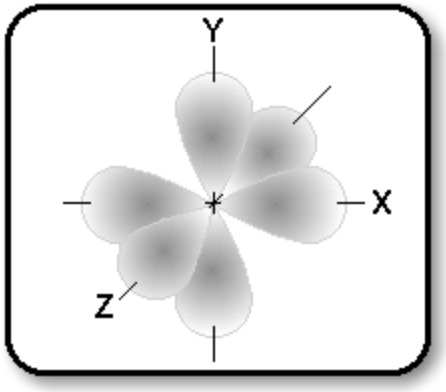
2s



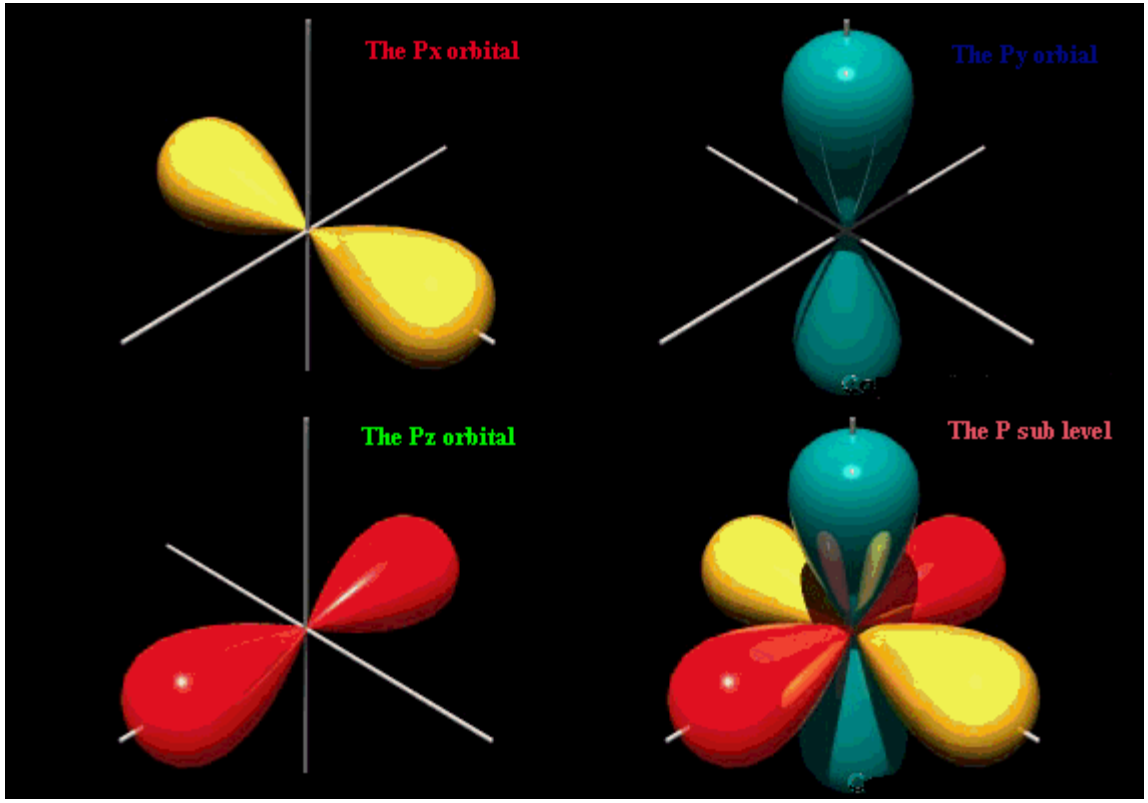
*P-orbitals are dumb-bell shaped
3 types exist- Px, Py, Pz



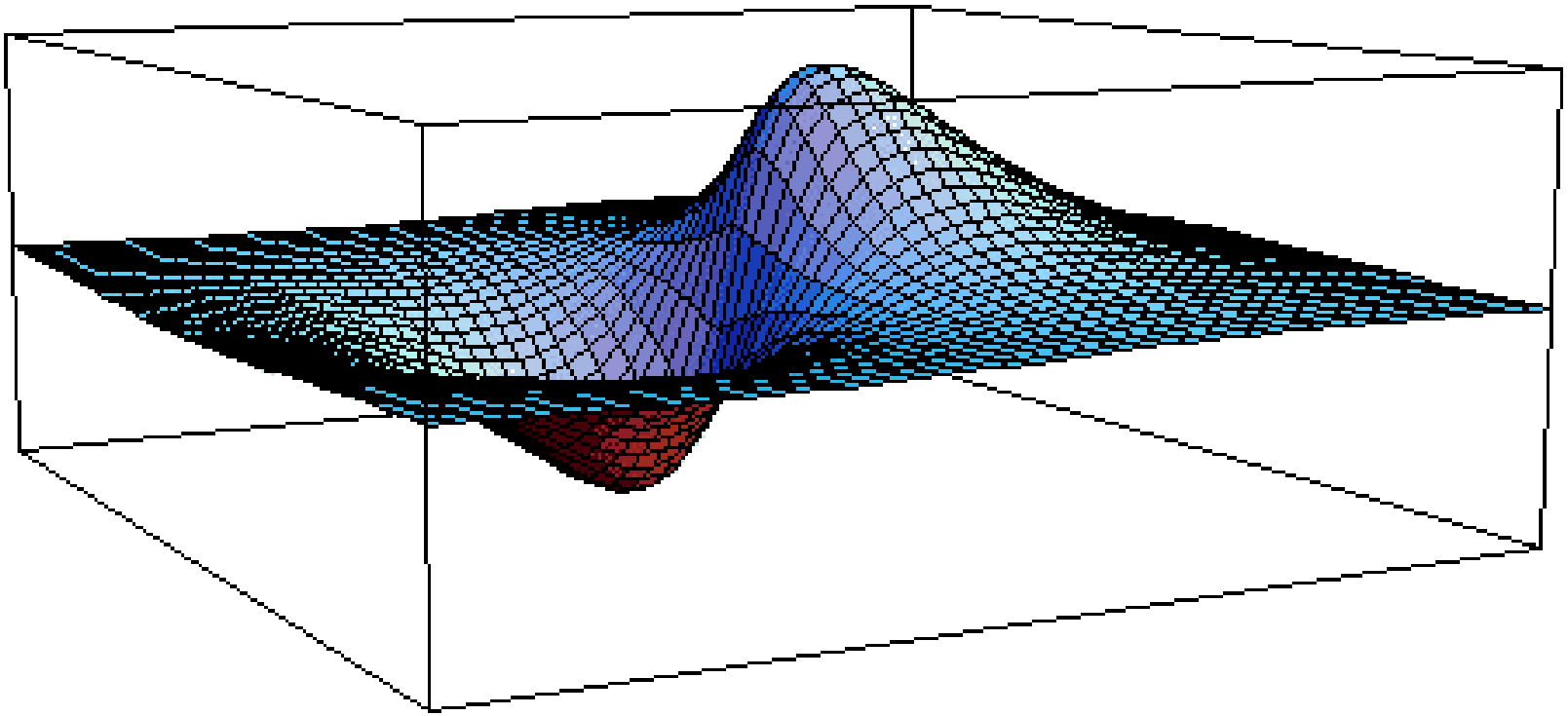
P-orbitals



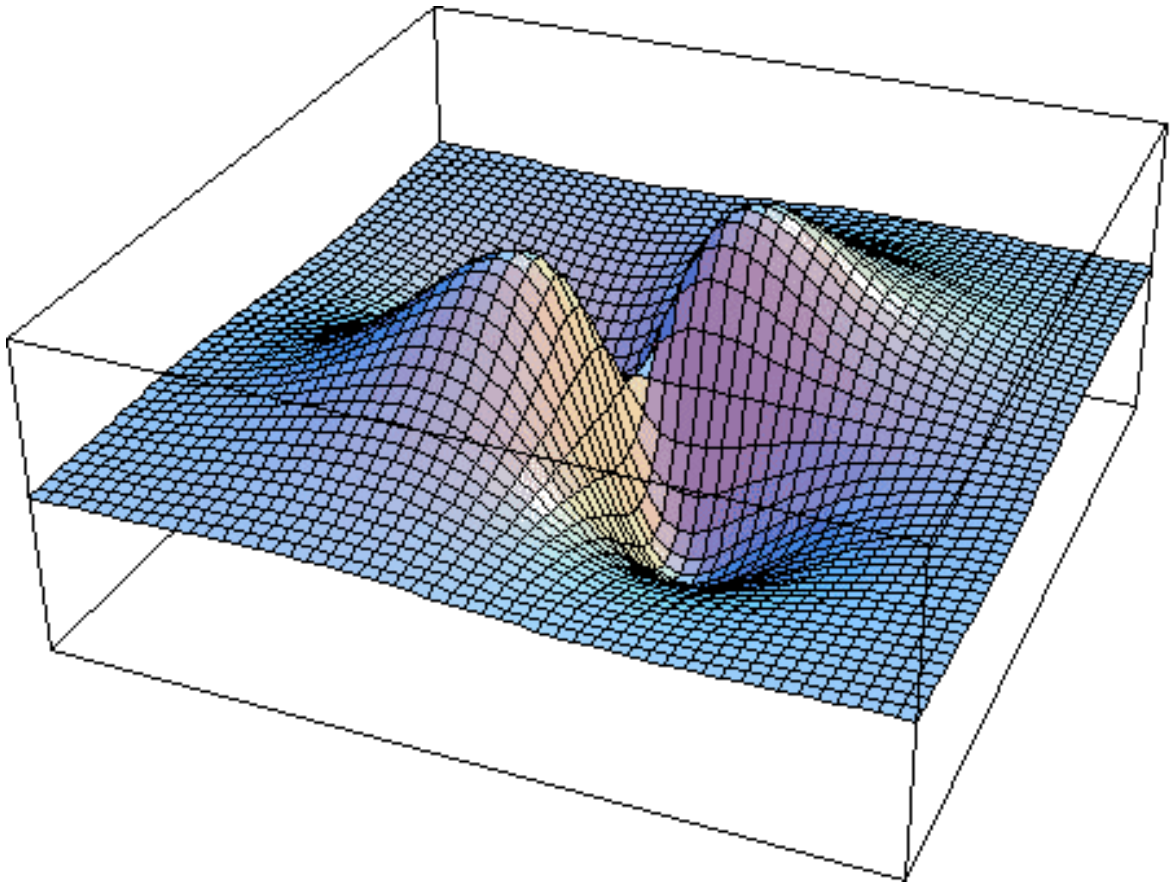
The "p" orbital is dumb belled shaped and each P sub level is made of three "p" orbitals.



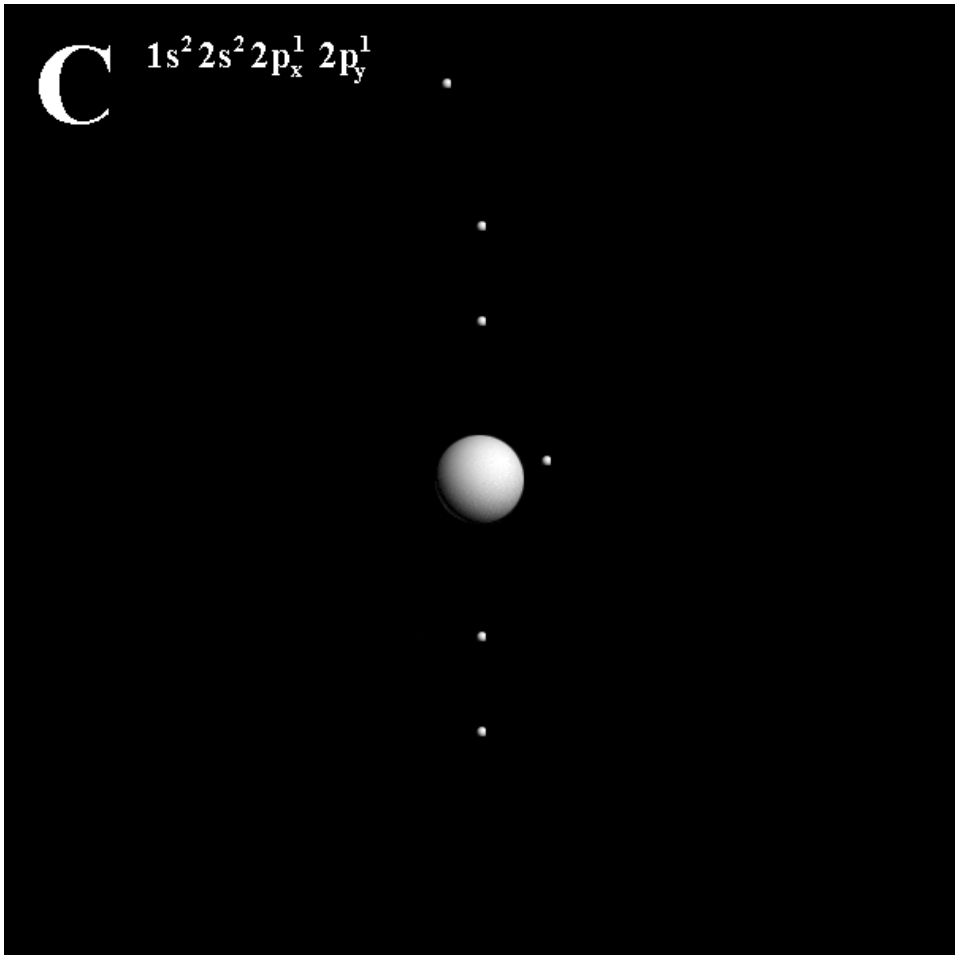
2p



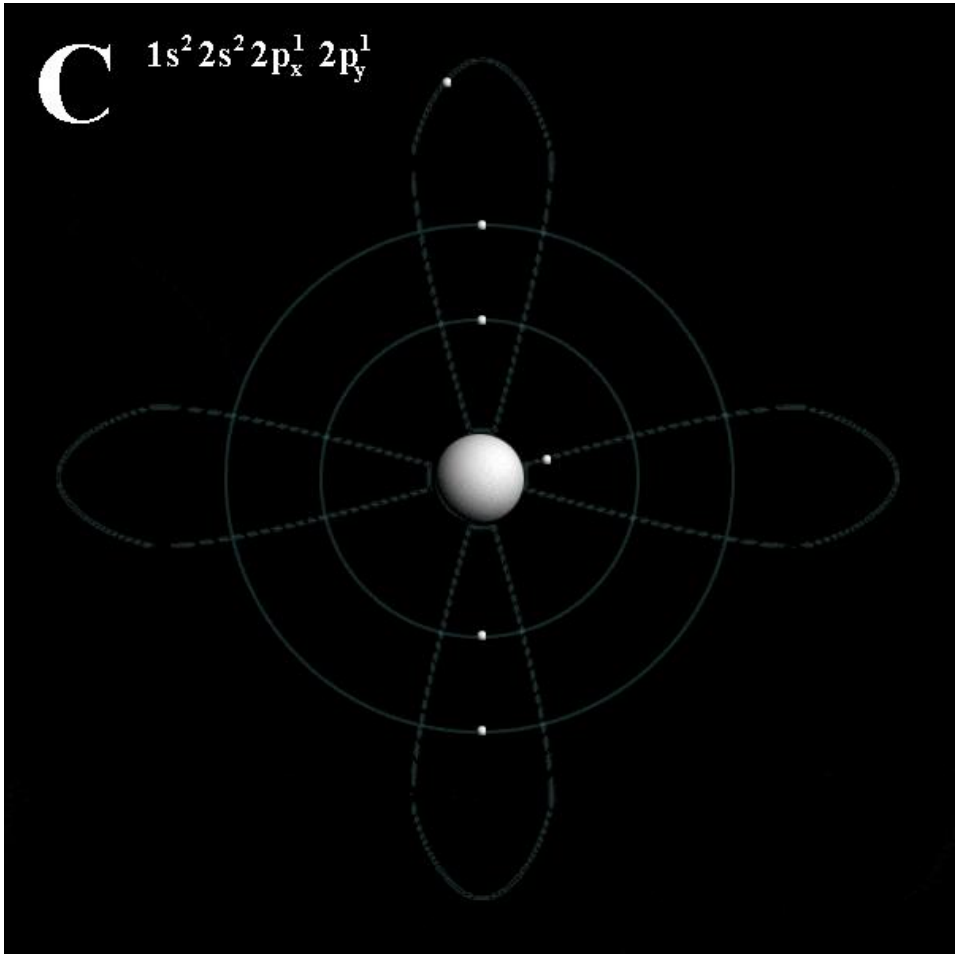
3d



Electrons moving



Electron paths



Number of Sublevels