Arrangement of Electrons in the Atom

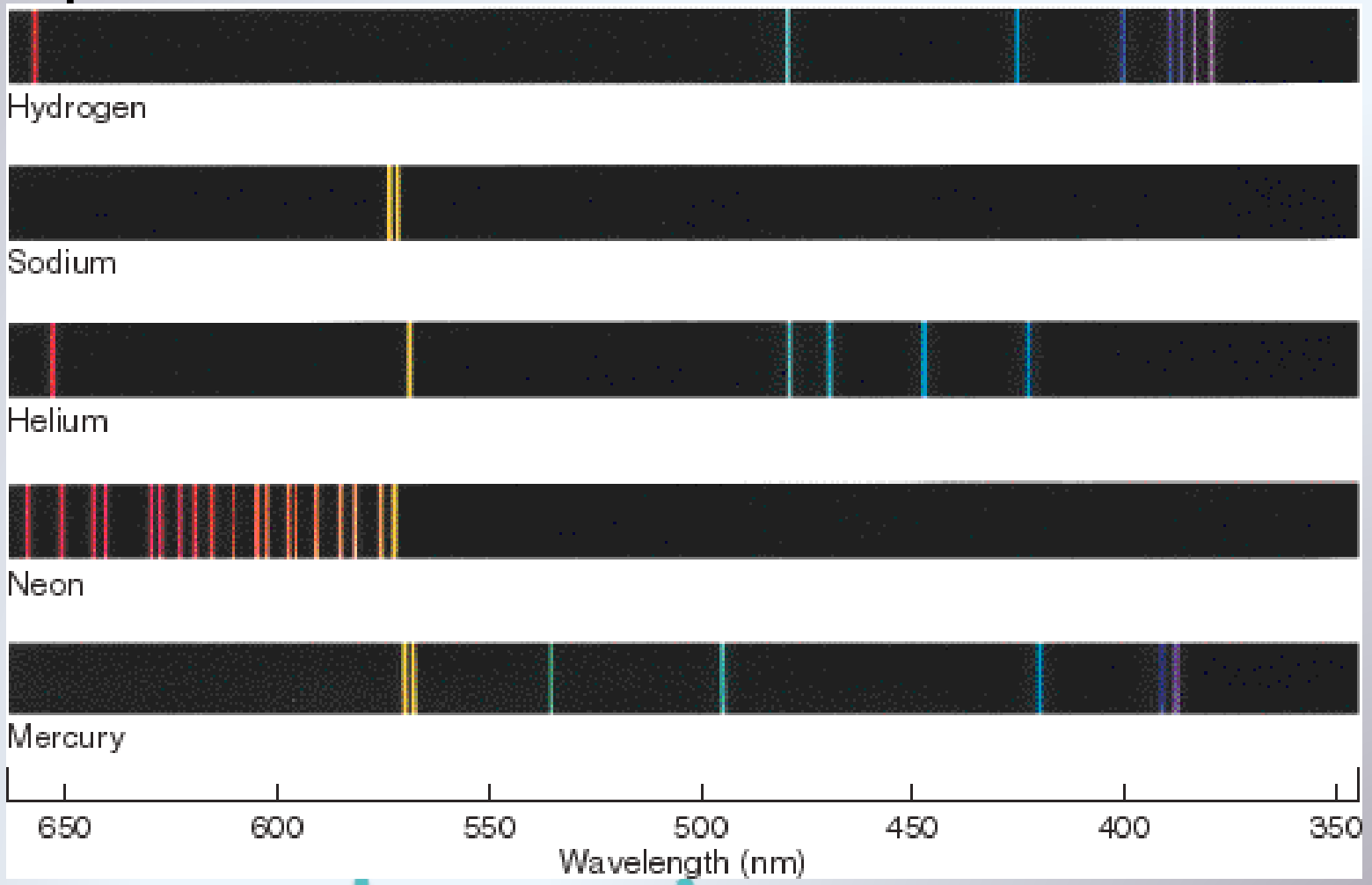
Bohr's Study of Atomic Spectra

- When white light is passed through a prism, the white light is broken up into an array of colours.

- This spread of colours is called a continuous spectrum.

- When a gas is heated, it emits light.
- If this light is passed through a spectrometer (contains a prism), the light is broken up to form a series of lines of different colours called an emission line spectra.

Each element has its own unique emission line spectrum… fingerprint

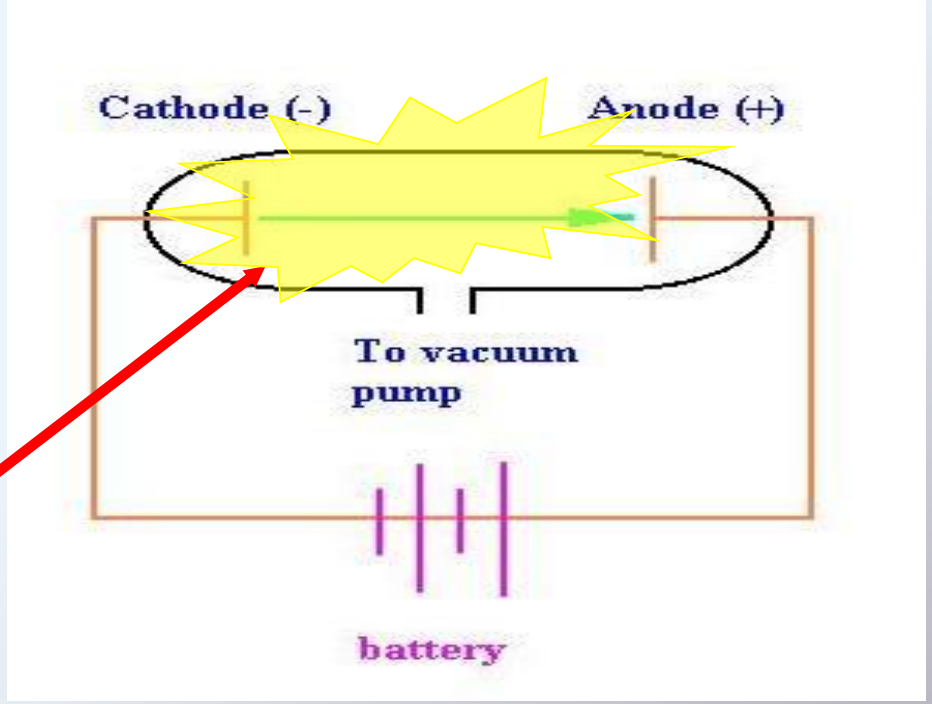


- The emission line spectrum of an element is often used by chemists to analyse materials for the presence of certain materials.

  e.g. analysis of steel for different metals.

- Different elements have different colours when electric current is passed through them.
- Sodium street lights are yellow
- Mercury vapour lamps are blue
- Neon advertising signs are red

Gas at low pressure glows when conducting electricity



Mandatory Experiment: Flame Tests

- When atoms of an element are supplied with energy under certain conditions they emit light.
- The light emitted has a colour characteristic of that particular element.

- The colours that strontium and barium emit are often seen in firework displays.
- Strontium nitrate gives a red colour.
- Barium nitrate gives a green colour.

- Elements that are gases or easily vaporised will emit light of a characteristic colour when placed in a discharge tube at low pressure and high voltage.

  e.g. sodium streetlights

Flame Test Results

Learn

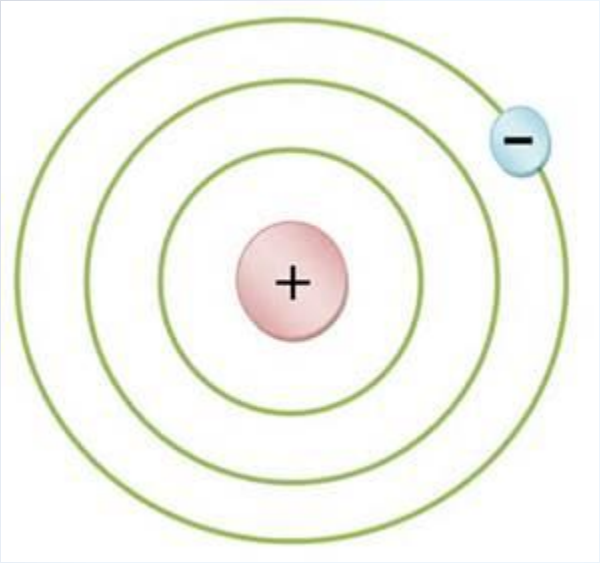
- Lithium…Crimson
- Sodium…Yellow
- Potassium…Lilac
- Barium…Green
- Copper…Blue/Green
- Strontium…Red

Evidence explaining the Bohr Theory

Niels Bohr realised that the model of the atom must explain…

- Why the emission spectra of the elements are line spectra rather than continuous spectra.
- Why the emission spectrum of each element is unique to that element.

Bohr Model of the Atom

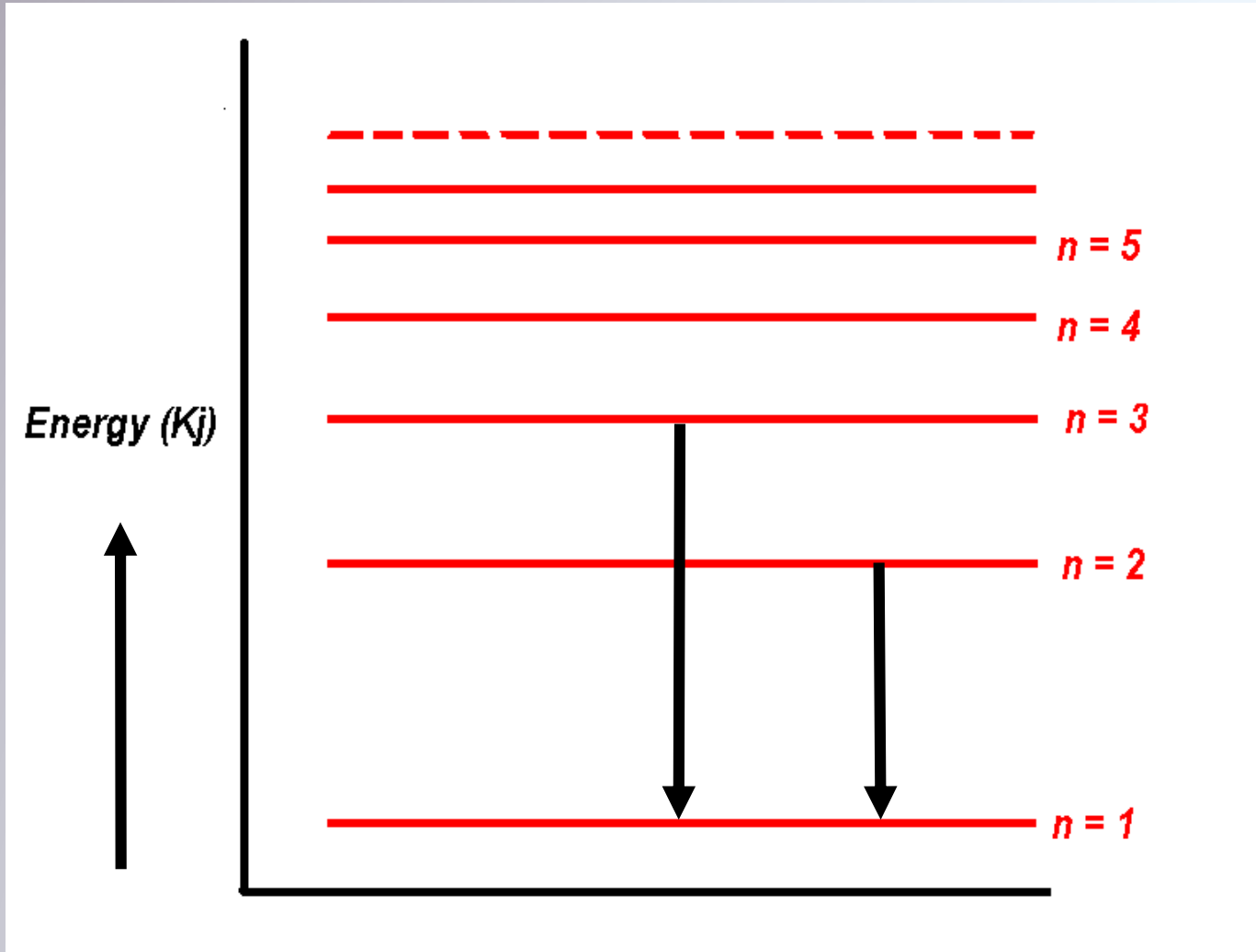


Bohr Model

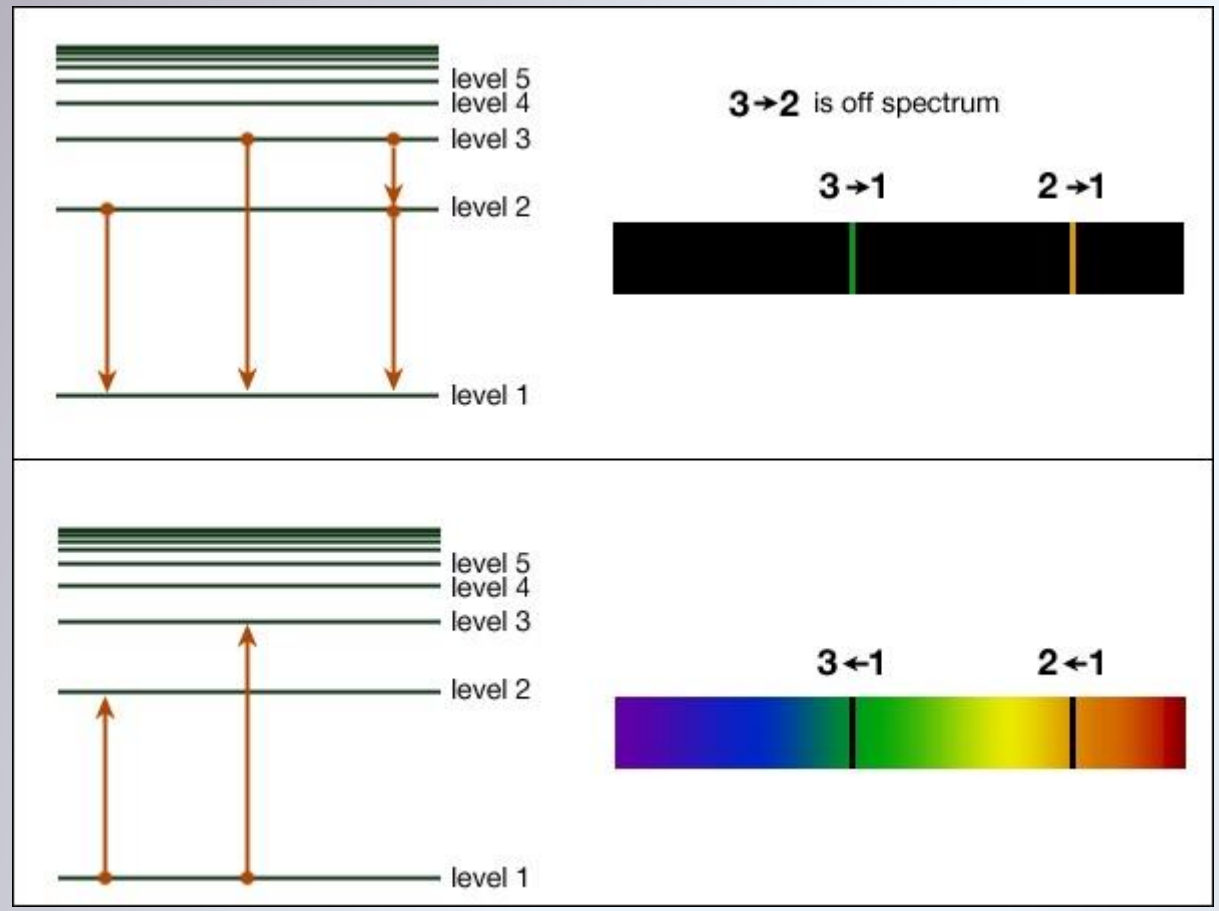
- Electrons revolve around the nucleus in fixed paths called orbits.
- Electrons in any one orbit have a fixed amount of energy. For this reason orbits are also called ENERGY LEVELS

- The energy of the electron in an energy level is fixed and at a definite value.
- As long as an electron remains in one particular energy level, it neither gains nor loses energy.

- When an atom absorbs energy, electrons jump from a lower energy level to higher energy level.
- Electrons are less stable in the higher energy level and do not stay here for long.
- **Energy is lost** when an electro falls from a **higher energy level** to **lower energy level**. Since an electron can only fall back to a certain definite energy levels, only fixed amounts of **light energy** can be given off.



Movement of Electrons



Emission Line Spectrum

Absorption Line Spectrum

What spectroscopic evidence is there for the existence of energy levels?

Consider the Hydrogen Atom:

- The electron occupies the lowest energy level available to it.
- This is the **ground state.**
- When energy is supplied to the atom, the electron absorbs the energy and moves out to a higher energy level – the **excited state.**
- The electron falls back to lower energy levels emitting definite amounts of energy in the form of light.

Continued

- The energy difference between levels gives a definite frequency of light in the spectrum. i.e. $E2 – E1 = hf$

  h = Plank's constant, f = frequency of light.

- Since only definite amounts of energy are emitted, this implies that electrons can occupy only definite energy levels.

  So… energy levels must exist in the atom.

Balmer Series…Visible Spectrum

- The set of electron transitions that give rise to visible spectrum is called the **Balmer Series.**
- The other series of lines (Lyman, Paschen, etc.) appear in the ultraviolet and infrared of the spectrum and hence are invisible.

Summary

- Electron normally in the **ground state**.
- Energy supplied as (heat of electricity).
- Electron jumps to a higher energy level, **excited state**.
- Unstable in excited state.
- Drops back to a lower level, emitting energy in the form of light.

Atomic Absorption Spectrometry

- Atoms can absorb light.
- If white light is passed through a gas sample of an element, it is found that the light which comes out has some wavelengths missing i.e. dark lines observed on the spectrum.

- Such a spectrum is called an Atomic Absorption Spectrum.
- The dark lines correspond to the missing wavelengths of light.
- Certain wavelengths of light are absorbed by certain elements .

- The absorption spectrum is like a negative photograph of the emission line spectrum.
- This is because atoms in the ground state absorb the same radiation as they emit in the excited state.

Uses of the two types of spectra

Emission Line Spectrum

- Identify elements.
- Gives particular colours to fireworks.

Absorption Line Spectrum

- Identify elements.
- Detect concentrations of elements.

Sample Exam Questions

Explain how the expression E2 – E1 = hf links the occurrence of the visible lines in the hydrogen spectrum to energy levels in a hydrogen atom.

- Energy difference between higher (E2) and lower (E1) level.
- f: frequency of line in spectrum, each line produced electrons falling from particular higher level to particular lower level.
  h: is Plank's constant.
- The energy difference (E2 – E1) is a constant (h) multiplied by the frequency (f).

Principle of AAS

- Atoms in the ground state absorb light of a particular wavelength characteristic of that element.
- Absorption is directly proportional to concentration.

Processes of AAS

- The sample solution is sprayed into a flame.
- The same element is converted into the atoms of the element.
- Ground state atoms absorb radiation from a source made of the element.
- Abundance is measured and identified.

Sample Exam Questions

Name the series of lines in the visible part of the line emission spectrum of hydrogen.

- Balmer Series.

What evidence do line emission spectra provide for the existence of energy levels in the atom?

- Only specific frequencies of light are emitted.
- Therefore the electrons must be restricted to specific energy values.

Why is it possible for line spectra to be used to distinguish between different elements?

Each element has a unique line spectrum because

- Atoms of each element have a different number of electrons.
- This leads to different arrangement of energy levels.
- This gives rise to different electronic transitions.

Why do different metals produce different flame colours?

- They have different numbers of electrons.
- The electrons will be moving between different energy levels in each metal.
- This will cause the emission of light of different frequencies.
- Different colours of light will be produced in the visible region.

Describe how Bohr used line emission spectra to explain the existence of energy levels in atoms.

- Electrons in ground state jump/move to higher level (excited state).
- Excited state is unstable so electrons fall back to lower levels emitting energy as light.
- $E2 – E1 = hf$

Why does each element have a unique line emission spectrum?

- Each element has a different arrangement of energy levels / difference electron configuration. This gives rise to different electron transitions.

The fact that each element has a unique line spectrum forms the basis for which instrumental technique?

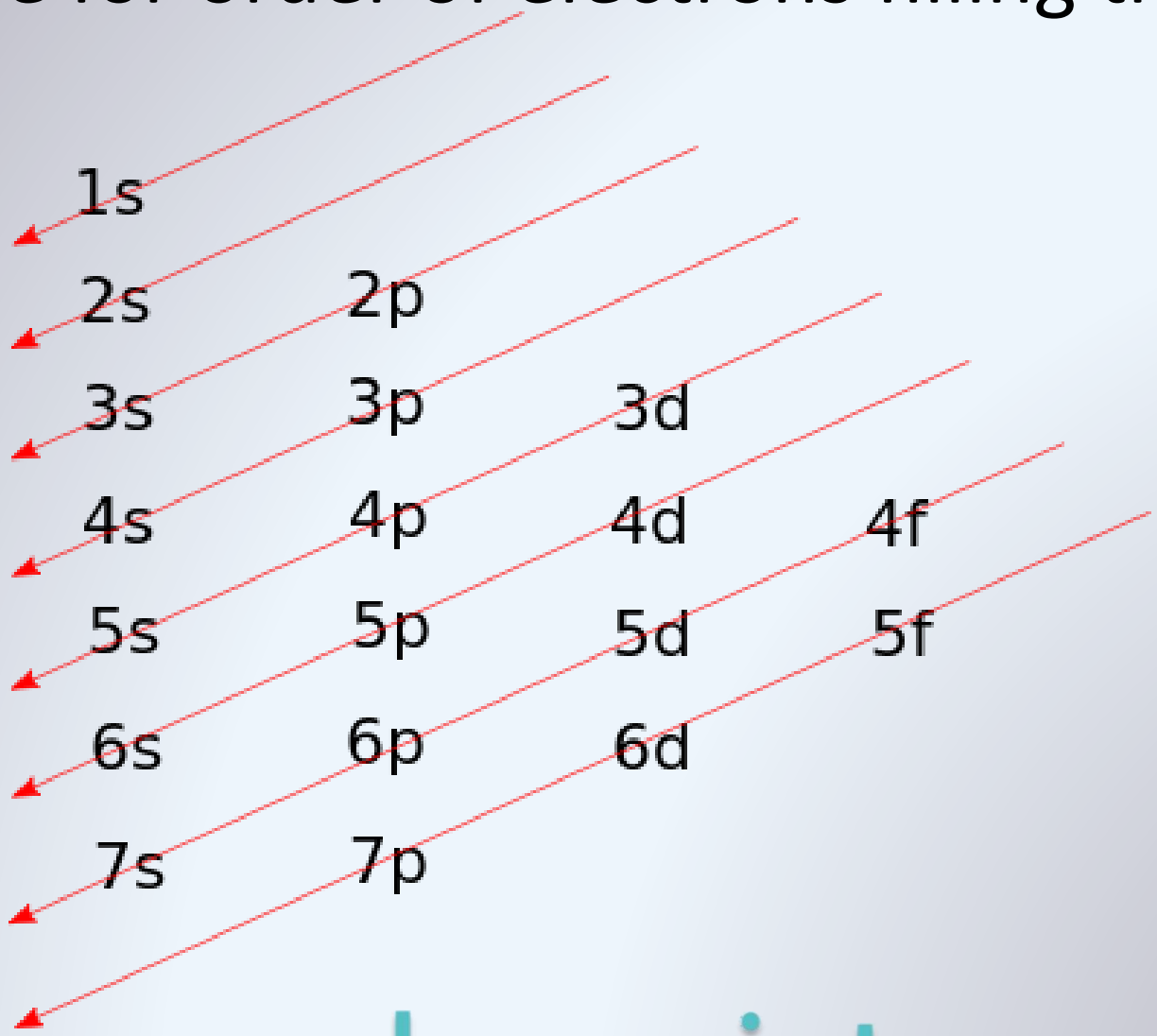
- Atomic Absorption Spectrometry

Bohr Model of the Atom

- Electrons move around the nucleus in fixed orbits called energy levels (shells).
- The number of electrons that each energy level can hold is determined by the formula $2n^2$ (n = the energy level).
- If n = 1…electron capacity = $2(1^2) = 2$
- If n =2… electron capacity = $2(2^2) = 8$

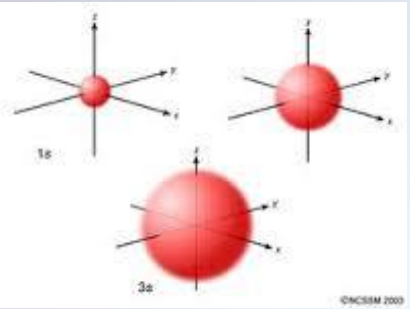
- Scientist found on closer inspection of the bonds of light in an emission line spectrum that some consisted of two or more lines very close together.
- This is evidence for the existence of energy sub-levels and are denoted by letters; s, p, d, f, etc.
- **Erwin Schrodinger** used mathematics to calculate regions in space where the probability of finding an electron is high… called orbitals.

Diagonal rule for order of electrons filling the energy sub levels.



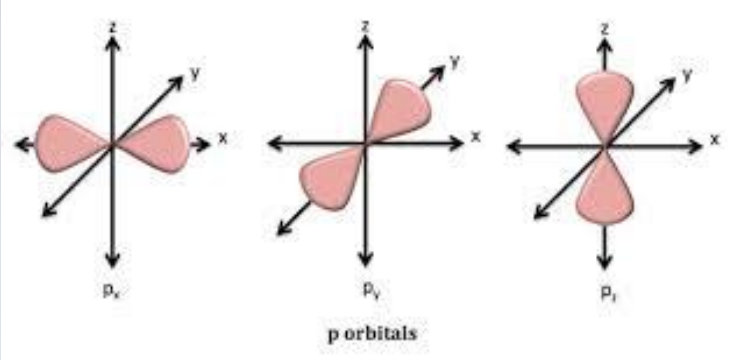
S sub-level

- Spherical Shape
- Holds two electrons



P sub-level

- Dumb-bell shape.
- Holds 6 electrons



D sub-level

- Holds 10 electrons
- 5 type (diagram not needed)

Wave Nature of Electron

- In 1924 De Broglie (a French scientist) proposed that all moving electrons had a wave motion associated with them i.e. it can be shown to have a wavelength.
- **It does not travel on a fixed path** at a definite distance from the nucleus.
- In 1927 he was proved correct.

- Werner Heisenburg (German scientist) considered the wave motion of electrons mathematically.
- He put forward a famous principle called the **Heisenburg Uncertainty Principle** which states that it is **impossible to measure at the same time both the velocity and the position of an electron.**
- Heisenburg reasoned that when a beam of light is used to detect the presence of an electron, the position of the electron is determined by because the electron has such a small mass, its velocity is immediately changed by the beam of light

What this discovery lead to:

- The idea of electrons travelling in fixed orbits had to be modified.
- If they are not travelling in fixed orbits then we can only talk about the probability of finding electrons at a particular position inside the atom.
- The Bohr Theory only worked for simple species of atoms e.g. hydrogen.
- The Bohr Theory was updated due to the discovery of sub-levels and the development of the Heisenburg Uncertainly Principle.

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