Name:

Periodic Table and Atomic Structure	Objectives
3. Arrangement of Electrons in the Atom	-define and explain energy levels in atoms **-describe and explain the emission spectrum of the hydrogen atom using the Balmer series in the emission spectrum as an example** **-describe and explain the absorption spectrum** -use flame tests to provide evidence that energy is absorbed or released in -discrete units when electrons move from one energy level to another **-explain how flame tests provide evidence that energy is absorbed or released in discrete units when electrons move from one energy level to another** -relate energy levels in atoms to everyday applications such as sodium street lights and fireworks **-discuss the uses of atomic absorption spectrometry (AAS) as an analytical technique** **-illustrate how line spectra provide evidence for energy levels** **-use a spectroscope or a spectrometer to view emission spectra of elements** **-define and explain energy sub-levels** **-state the Heisenberg uncertainty principle** **-state the dual wave-particle nature of the electron (mathematical treatment not required)** **-define and explain atomic orbitals** **-describe the shapes of s and p orbitals**

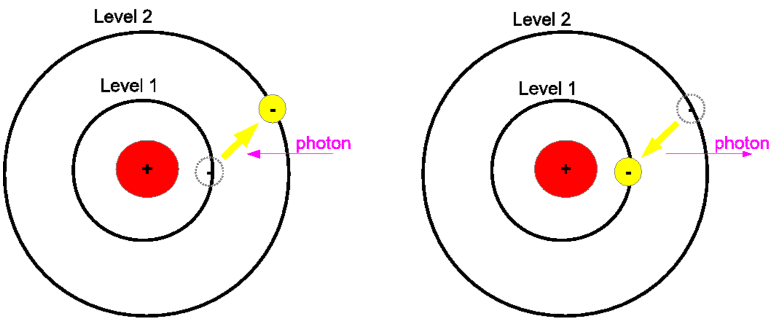
Bohr's Study of Spectra

- White light passed through a prism forms a **continuous spectrum** (no gaps).
- Bohr passed the light from a Hydrogen Gas Discharge Tube through a prism. The spectrum observed consisted of some narrow, coloured lines. He called this an **emission spectrum**.
- Spectrometers allow measurement of the frequency of each band of light.
- Spectroscopes allow the spectrum to be viewed, but not measured.
- Every element has its own unique emission spectrum, which can be used to identify the element.

Flame Tests
Method:
1. Dip damp wooden splint into sample of salt.
2. Place sample into blue Bunsen flame.
3. Record the colour of the flame.

Metal Present	Colour
Lithium	Crimson
Potassium	Lilac
Barium	Green
Strontium	Strontium
Copper	Blue-Green
Sodium (street lights)	Yellow

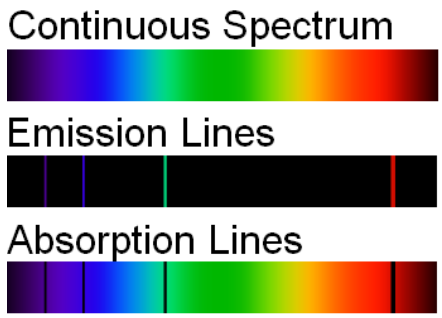
Bohr's Theory: Explaining the Evidence for the Existence of Energy Levels (Hydrogen only)
Def^n: An **energy level** is defined as the fixed energy value that an electron in an atom may have.
Def^n: The **ground state** of an atom is one in which the electrons occupy the lowest available energy levels.
Def^n: The **excited state** of an atom is one in which the electrons occupy higher energy levels than those available in the ground state.



1. In the ground state the electron occupies the lowest available energy level.
2. The electron can jump to a higher energy level (excited state) if it absorbs energy.
3. The excited state is instable.
4. Electron falls back to a lower energy level.
5. Energy given off as a photon. The energy of the photon emitted corresponds to the difference between the two energy levels (E_2-E_1=hf)
6. Each frequency of light emitted appears as a coloured line on the spectrum.
7. The separate lines obtained in the spectrum show that the electron can only have particular values of energy.
8. The emission of visible light is due to electrons falling to the n=2 energy level. This is known as the **Balmer Series** of emissions lines.
9.

Atomic Absorption Spectroscopy (AAS)

Principles:



1. White light is passed through a gaseous sample of an element.
2. Certain frequencies of light are missing from the light that emerges from the other side of the sample, resulting in dark lines in the spectrum. This spectrum is an **atomic absorption spectrum**.
3. These dark lines correspond exactly to the coloured lines in the **emission spectrum**. The dark lines are due to the electrons in the element absorbing certain frequencies of the white light to become excited.

Uses:

Instrument used: **Atomic Absorption Spectrometer**

Detection of the **presence and concentration** of certain elements (e.g. lead, chlorine) dissolved in water.

1. Atoms of an element absorb light of a particular wavelength of light.
2. The amount of light absorbed is proportional to its concentration (low absorption means low concentration, etc.)

Energy Sublevels

Every energy level (except n=1) consists of a number of **sublevels**.

Def^n: A **sublevel** is a subdivision of a main energy level and consists of one or more orbitals of the same energy.

The sublevels for the first 4 energy levels are shown below:

n=1: s *n=2*: s, p *n=3*: s, p, d *n=4*: s, … etc. (this is far as the course needs)

Wave Nature of the Electron:

De Broglie said that electrons have a wave motion as well as being particles.

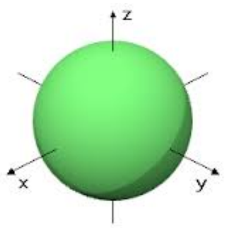
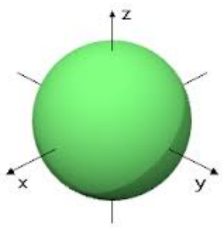
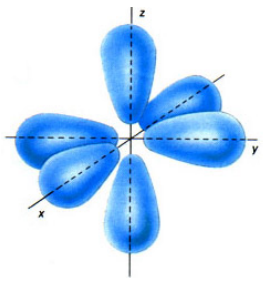
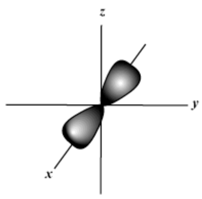
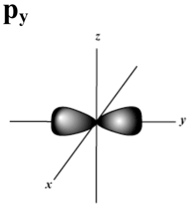
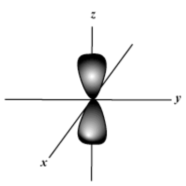
Heisenberg's Uncertainty Principle states that it is impossible to measure the velocity and speed of an electron simultaneously.

Limitations of Bohr's Theory:

1. Only accounted for the emission spectrum of Hydrogen (simple case – only 1 e^-)
2. Doesn't take the wave nature of the electron into account.
3. Did not take the existence of sublevels into account.

Atomic Orbitals:

Def^n: An **atomic orbital** is a region in space where there is a high probability of finding an electron.

Sublevel	**Orbitals**	**No. of orbitals**	**Shape**	**No. of e^-s per orbital**	**Total no. of e^-s in sublevel**
s 	**s** 	1	spherical	2	2
p 	**p_x**  **p_y**  **p_z** 	3	Dumbbell	2	6
d	Not needed	5	Not needed	2	10