QUANTUM THEORY AND THE BORH MODEL

The modern atomic theory simply revises Bohr's ideas about exact locations of the electrons and how they behave.

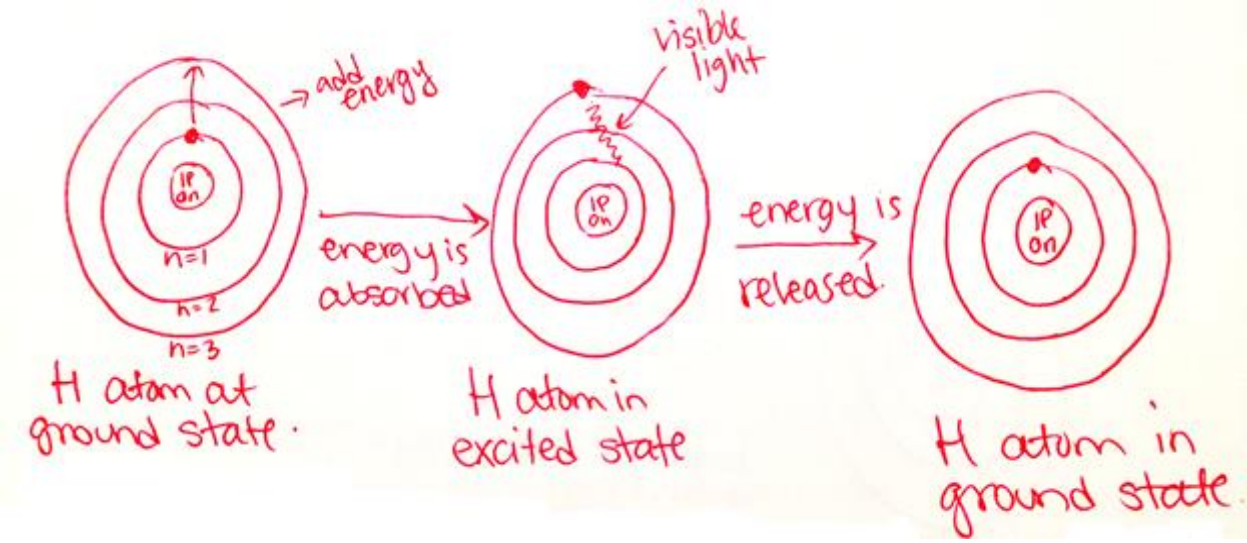
ELECTRONS CAN BE DESCRIBED AS **particles** OR **waves**.

Depending on the experiments that different scientists do, electrons can behave as very small particles or as waves.

Electrons usually exist in the lowest possible energy level. This situation is called the **ground** STATE of an atom. When an atom absorbs energy, the electrons absorb that energy and move to higher energy levels. The atom is then in an **excited** STATE.

However, an atom is not **stable** in its excited state. To become stable, its electrons must return to their ground state positions, **releasing** their absorbed energy as they do so. For many elements, this energy is released as **visible** light.

Explain:



Define quantum number: each energy state/level (n=1 → quantum number 1)

Thinking Question: Why do different elements show different colours?

- ➔ They have a different number of electrons – more transitions between energy levels
- ➔ Different colours have different energies, the further the electron has to fall, the higher light energy is given off

Lowest energy (red) –ROYGBV – highest energy (violet) $\Delta E = hc/\lambda$

Different elements have different numbers of **electrons** and they are at different energy levels. When they are excited, the different amounts of energy result in different **colours** (because different colours of light have different energies themselves).

When the characteristic colour of an element is viewed through a prism or an instrument called a **spectroscope**, its BRIGHT-LINE EMISSION SPECTRUM becomes obvious. This spectrum shows the different colours that are unique to that element. Each coloured **line** corresponds to the movement of a different electron. This bright-line emission spectrum is like the "**fingerprint**" of the element.

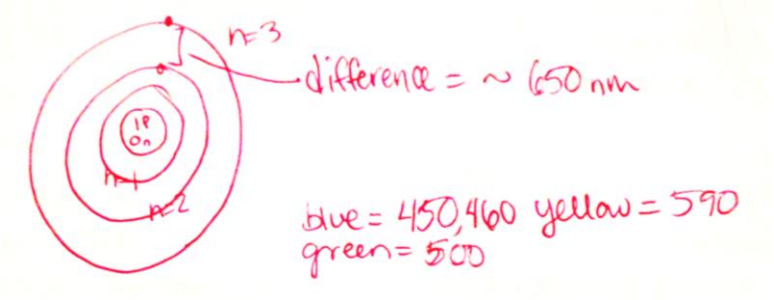
Spectroscope: **separates** light into its individual component wavelengths

You will be observing the BRIGHT-LINE EMISSION SPECTRA for four elements and two different types of light bulbs.

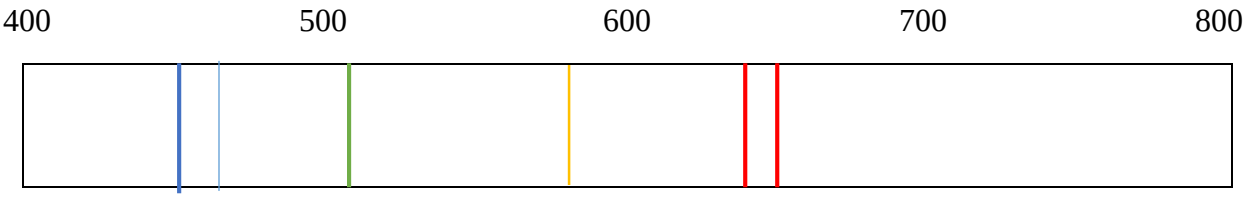
How do we draw emission spectra?

Each **electron** that drops from an excited state to a **lower** energy level will release energy. If this energy is **equivalent** to a wavelength of visible light, a particular colour will be shown.

For example, if the energy **difference** between the two orbits corresponds to the wavelength of red light, there will be a band of red light shown on the emission spectrum.



λ (nm)



Use your textbooks – complete the activity on page 257- 258