

## Light Sources & Spectra

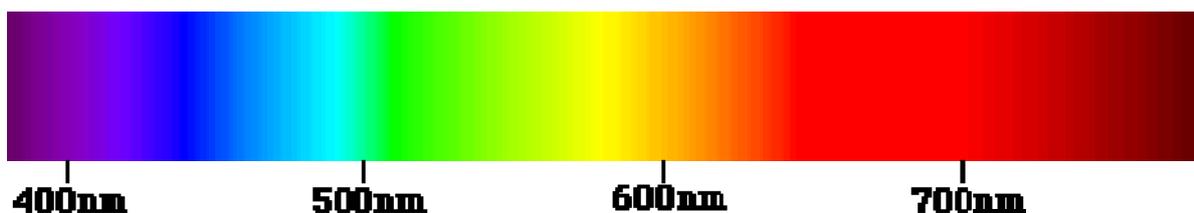
Light sources get taken for granted. It wasn't always so. Our ability to make light upon demand has transformed society from the dawn-to-dusk existence that it's been for most of our evolutionary past to the life we now have. This section is about making light, about the spectrum of light, about how we see the quantum nature of atoms in their spectra, and about how we exploit the quantum nature of matter to make lasers. There are a lot of topics in here and the lectures will cover a mixture of aspects practical, theoretical and historical.

### *The electromagnetic spectrum*

By direct sensing we experience a tiny part of the electromagnetic spectrum, that bit called light and heat radiation. You have already heard that the electromagnetic nature of light was deduced by James Clerk Maxwell only in the middle of the 19<sup>th</sup> century. The fact that there is a huge spectrum of further electromagnetic radiation that exists naturally is an even more recent discovery. Our ability to generate such radiation for our own uses is still being developed. Mankind has found the complete spectrum by a combination of experimentation and intellectual thought. This course is about the very small bit of the spectrum called light. Where is it? Sandwiched broadly between the longer wavelength radio waves and the shorter wavelength x-rays; more precisely, sandwiched between the comparatively extensive IR (infra-red) and UV (ultra-violet) regions.

The slide below shows the spectrum with increasing wavelength from left to right. The wavelengths of light in vacuum run between 400 nm and 750 nm. You should remember these figures because they define the visible spectrum. For the record, 1 nm (nanometre)  $\equiv 1.0 \times 10^{-9}$  m; 1 Å (Ångström)  $\equiv 1.0 \times 10^{-10}$  m; 1 μm (micron)  $\equiv 1.0 \times 10^{-6}$  m. You'll meet wavelengths in all these units. Nanometres are the current favourite.

### *Colour and wavelength*



Colour depends on the frequency of light. The frequency of light remains unchanged when light is reflected or refracted and so does the colour of each part of the spectrum. Wavelength does change with refraction, as we've seen. Given this, you may wonder why almost all graphs and data for colour related information are given in terms of the wavelength of light. The key point is that they are given in terms of the **vacuum wavelength**, though that is often left unsaid. Vacuum wavelength,  $\lambda$ , is related to frequency,  $f$ , by the relationship we've already met:

$$\lambda = \frac{c}{f} .$$

*Hot body emission*

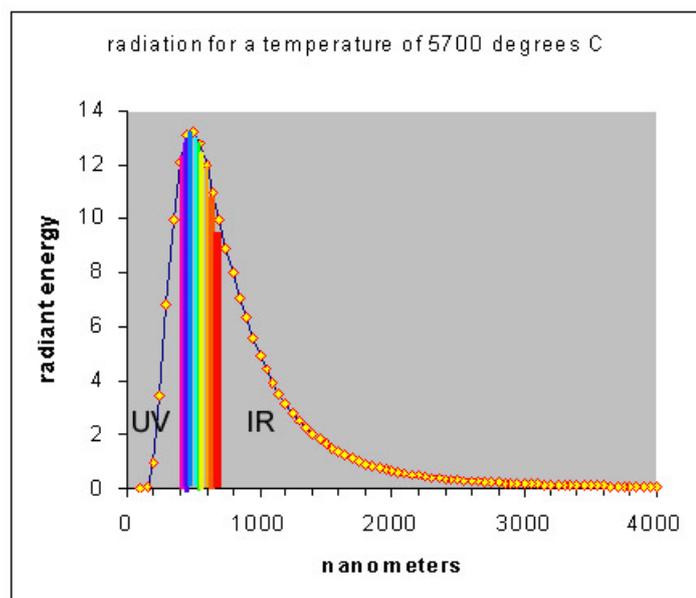
All bodies emit electromagnetic radiation over a wide range of wavelengths. It's a curious and fascinating fact that, broadly speaking, the spectrum of this radiation depends not on what the body is made of, whether it is metallic, ceramic, rough, smooth, how old it is or any such property but on the temperature of the body. Basically, just the temperature. As this fact became apparent towards the end of the nineteenth century, physicists set to work to try to explain what was going on. The link between radiation and temperature was apparent, because both are intimately related to motion of the constituents of matter. However, the physics of the time could not come up with an explanation for the variation of intensity with wavelength that was actually observed by careful measurements. The physical arguments from first principles that led to the wrong answer seemed cast-iron. The measurements were agreed by all who tried. It was apparently a case of immovable object meeting irresistible force.

The force of experimental fact won. Fundamental physical principles which had been immensely successful in many fields were found wanting. The person who suggested what extra idea had to be included was Max Planck. His idea was the **quantum of energy**. Energy was absorbed and re-radiated from light in discrete bundles, called quanta, whose size, Planck said, depended only on the frequency of the light. I've mentioned Planck before and introduced the quantum of size  $hf$  (or  $h\nu$ , as its often written),  $h$  being Planck's constant. Planck's constant is now recognised as a fundamental constant of physics. With this extra idea, Planck was able to deduce what is now known as **Planck's radiation law** that predicted the shape of the emission of light from an ideal emitting hot body. His result agreed with the experimental evidence collected at the time. It still agrees with all experimental evidence, 100 years on.

A few years after Planck's bold step, Einstein went a step further. He realised that light itself was quantised in the units of  $hf$ , and these quanta became known as photons. This is explored in more detail in the final chapter of the course.

An ideal radiating body is called a **blackbody**. Such a body radiates by the maximum amount of any body and also absorbs incident radiation by the maximum amount possible. Hence if it isn't glowing itself in the visible wavelength range it appears black. The Sun, oddly enough, is a 'blackbody' on this definition, though obviously it is hot enough to glow. Any light from outside that falls onto the Sun isn't reflected off its surface but penetrates inside to mix with the light already there. The Sun is therefore close to being a perfect absorber.

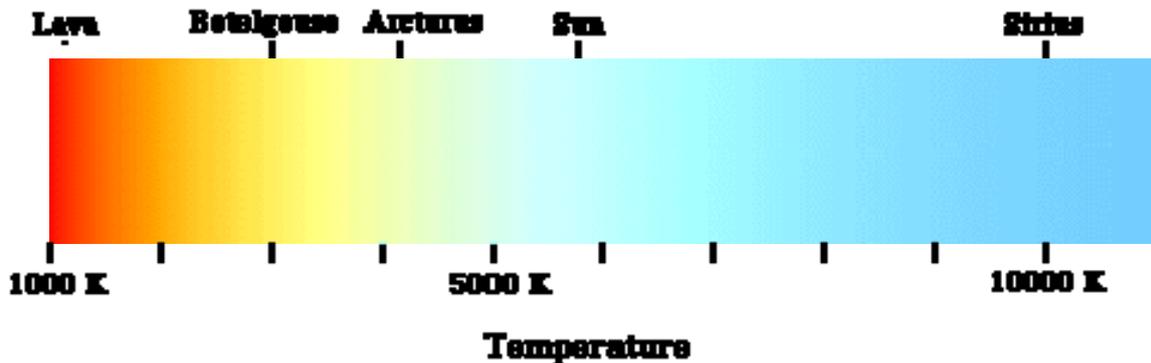
The slide shows the spectrum predicted for a body as hot as the Sun. 44% of the Sun's radiant



energy is light; 7% is UV and 49% is in the IR.

### *Appearance of hot bodies*

When a body is heated, it produces more radiation and the peak of its spectrum of radiation shifts to shorter wavelengths. The result is that the body appears brighter and its apparent colour changes. We all know this. If you put a poker in the fire, you can't see the radiation from it at first as it starts to heat; then it glows dull red; gradually it turns more yellow. The



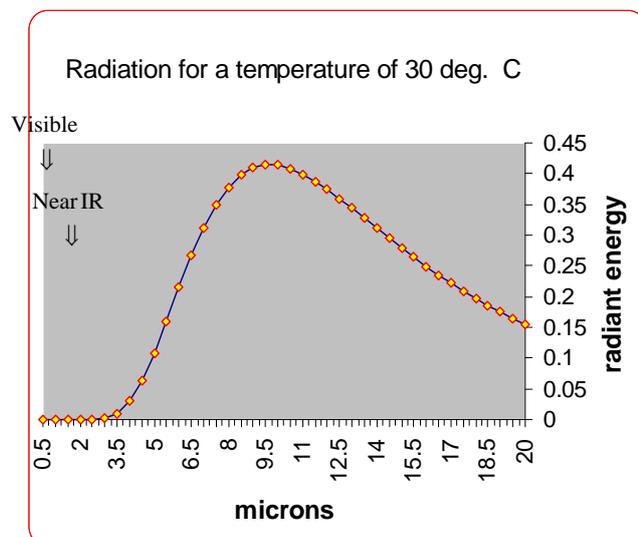
poker would melt before it got truly white hot but you can see the way the colour is changing. Higher temperatures than a poker reaches are found in every star we can see. The really hot stars have surface temperatures above 10,000 K and appear bluish, just because of Planck's radiation law.

This change of colour with temperature is taken over into the slightly woolly 'everyday' concept of **colour temperature**. Architects, for example, may talk about the light coming in a north-facing window as having a colour temperature of 12000°. What they mean is that it has typically a blueness similar to a hot body radiating at that temperature. In fact, most of the atmosphere supplying the light is below freezing. The blueness comes from Rayleigh scattering of sunlight, a mechanism that is quite different to blackbody radiation.

### *Emission from cooler bodies*

It isn't just hot bodies that emit electromagnetic radiation. We all do, at whatever temperature we are. The spectrum of radiation is predicted by Planck's law, with a modest modification to account for the fact that we aren't perfect radiators and absorbers. For a body at room temperature, all the radiation is in the infra-red or longer wavelength region of the electromagnetic spectrum.

Since radiation is energy, why don't we all cool down if we're radiating away energy? The answer is that we don't cool because we're all absorbing the radiation emitted by our surroundings at the same time. We will cool down only if the radiation we



receive is less than the radiation we emit. This happens if the temperature of our surroundings is less than our own temperature, as we know. At night the Earth cools down, at least it does on the side of the Earth where it is night. Meteorological satellites with IR sensors can easily image the Earth by means of emitted infrared radiation and send back cloud cover pictures, and much more, as easily by night as by day. Anyone who went to my meteorology lectures will have seen plenty of IR pictures of clouds.

### *Planck's radiation law*

The formula embodying Planck's radiation law is not too hard to appreciate. It's not very simple, but it's not too complex either. With a calculator, you should be able to evaluate it if you have to. The two previous pictures showing the radiation spectra for a hot and a cool body were drawn using Excel to evaluate Planck's law over the range of wavelengths shown.  $E_\lambda$  is the energy density of the radiation between wavelengths  $\lambda$  and  $\lambda+d\lambda$  (in units of  $\text{J m}^{-3} \text{nm}^{-1}$ ). Three fundamental physical constants are involved. 'c' because electromagnetic radiation is involved; 'h' because quanta are involved; 'k' because temperature is a fundamental concern.

$$E_\lambda = \frac{8\pi hc}{\lambda^5} \left[ \frac{1}{e^{hc/\lambda kT} - 1} \right]$$

### *The wavelength of maximum radiation, $\lambda_{\text{max}}$*

An earlier paragraph mentioned how the radiation spectrum shifted to shorter wavelengths as a body heated up. This result is essentially what **Wien's displacement law** says. The wavelength of maximum radiation,  $\lambda_{\text{max}}$ , is inversely proportional to the temperature of the body.

$$\lambda_{\text{max}} = \frac{3000}{T}, \lambda \text{ in } \mu\text{m}, T \text{ in K}$$

The lecture slide shows a couple of examples of the formula in action. If you've enough expertise in calculus, you can deduce it from Planck's radiation law. Wilhelm Wien's work measuring radiation was considered so important in providing a bridge that led from classical physics to the new quantum physics that he was awarded the Nobel Prize for Physics in 1911 "for his discoveries regarding the laws governing the radiation of heat".

### *The Sun's spectrum*



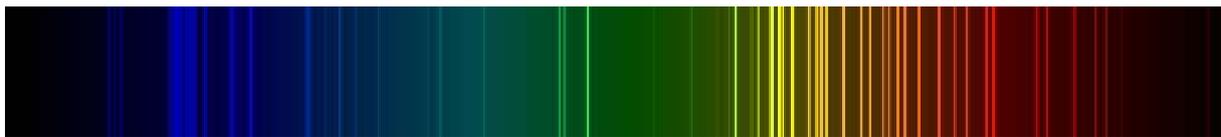
Newton saw the spectrum of the Sun as a continuous band of colour. With better prisms and optics, for example using a modern spectroscope, the spectrum clearly appears crossed by many dark bands. These bands were discovered by William Wollaston very early in the 19<sup>th</sup> century but investigated more fully by the German optician Joseph Fraunhofer (1787 – 1826) in 1814. Fraunhofer and his successors labelled the prominent lines by upper case letters and some less prominent ones by lower case letters. He located where they were in the spectrum and today they serve as natural spectrum markers.

Fraunhofer didn't live to find out that each line is associated with the absorption of one particular element in the Sun's outer atmosphere. Light is produced deeper within the Sun by blackbody radiation. Passing through the outer atmosphere, very specific wavelengths are absorbed. Each element present in the Sun absorbs at many different wavelengths, from the IR through into the UV. As a consequence there are a myriad of such absorption lines even across the visible spectrum. Several lecture slides show detail in this absorption. One slide shows just 10 nm of spectrum as it would appear on an analogue trace. On the following lecture slide, the same region is shown as it would appear on a black and white photograph.

The location and strength of each line is related to the element, its state of ionisation, its concentration, its temperature and, in fine detail, the pressure of its surroundings and the presence of a background magnetic field. Thus, a tremendous amount of information comes bundled in the spectrum. You just have to know how to read it, and that is the business of modern spectroscopy – reading the book of nature written in the spectral lines. Astronomy owes much of its knowledge to spectroscopy.

In the Sun, most atoms are single. There are no molecules because the temperature is too high to preserve molecular bonds.

#### *Emission spectra and absorption spectra*



If you excite atoms, giving them a lot more energy than usual by heating them up or energise them in an electrical discharge, then they emit spectral lines. This emission takes place at the same places in the spectrum as the atoms produce absorption lines. Gustav Kirchhoff (1824 - 1887) discovered this in 1859. The lecture slide shows the emission spectra of neon and helium, which those who took our first year lab saw, though not so well as shown here. Also illustrated is the spectrum of iron, which is commonly used by astronomers as a calibration spectrum because it has plenty of lines spread across the visible.

#### *The hydrogen spectrum*

Hydrogen is the most common element in the Sun, and in the universe. A hydrogen atom is the simplest of all atoms, normally consisting of one proton and one electron. Its spectrum is the simplest, having just four visible lines. These are part of the **Balmer series** of lines that extend slightly into the UV. These lines are named  $H_\alpha$ ,  $H_\beta$ ,  $H_\gamma$ , ..... . We'll see soon how they're formed. As Kirchhoff pointed out, the hydrogen emission spectrum has bright lines at just the same frequencies (i.e. wavelengths) as the hydrogen absorption lines.

The following lecture slide shows some images taken in hydrogen light. The two nebulae glow by hydrogen light because they are largely made of hydrogen. This hydrogen is excited by more energetic light coming from embedded stars, and by collisions. The two images of the Sun were taken through a narrow filter that just lets through the light of  $H_\alpha$ , thereby targeting a single species of atom in the Sun.

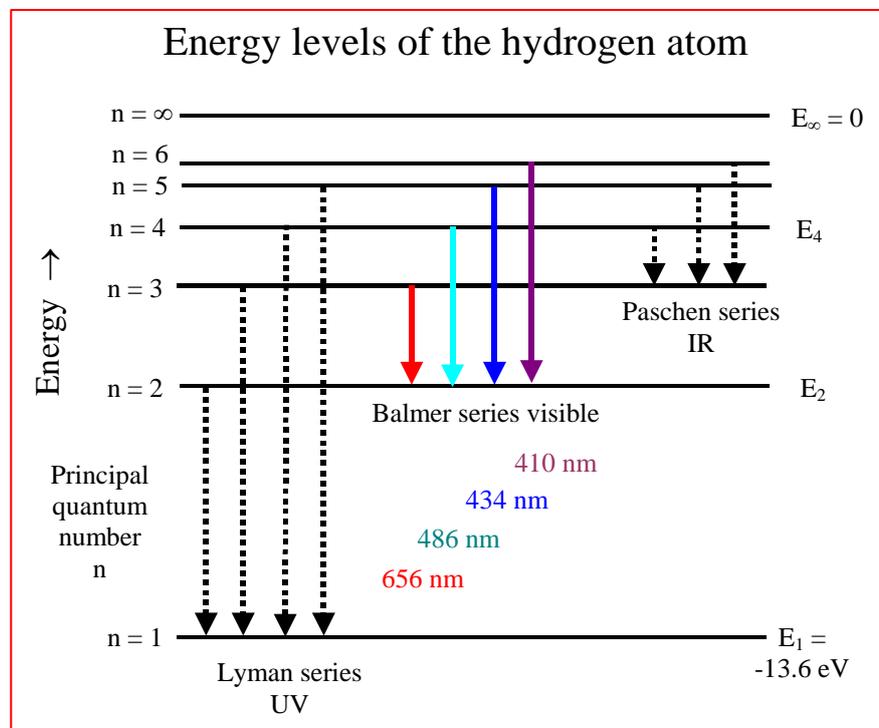
### Origin of the hydrogen spectrum

Electrons in atoms occupy discrete energy levels. These levels are labelled by quantum numbers. The **principal quantum number** is denoted by the symbol  $n$ , which takes on integer values 1, 2, 3, 4, .... . It is this quantum number that determines the energy of the electron in hydrogen.

Energy levels in atoms are negative. This means that if you take an electron at rest just outside a hydrogen atom, it loses energy by dropping into an energy level. The maximum energy it can lose is 13.6 eV (**electron volts**). When an electron drops from one level to a lower level, it emits the surplus energy as electromagnetic radiation. Niels Bohr (1885 - 1962) gave the crucial relationship that determines what frequency,  $f$ , of radiation is emitted. He realised it is determined by Planck's constant:

$$hf = E_i - E_k, \text{ } f \text{ is radiation frequency}$$

$E_i$  is the energy level at which the electron starts, called **the initial energy level**.  $E_k$  is the **final energy level**. The 4 visible Balmer lines are produced by **electronic transitions** finishing at the  $n = 2$  level. Transitions ending at  $n = 1$  all cover a sufficiently big energy gap that the resulting radiation is in the UV. This series is called the **Lyman series**. A third series ending  $n = 3$  is the Paschen series, which is all in the IR. The diagram above takes a



little studying if you are not used to such things. The energy gaps between levels are not drawn quite to scale in the diagram, in order to show the transitions better.

### The quantised atom

Niels Bohr was the first to think up a model for a quantised atom. His picture is often called **the Bohr atom**. His **quantisation** doesn't appear directly in the energy levels. His fundamental hypothesis was that the angular momentum of the electron is quantised into integral multiples of  $h/2\pi$ , a simple combination of constants that is given the special symbol  $\hbar$  and called the **Dirac h**. Therefore the angular momentum of an electron is  $n\hbar$ . This introduces the quantum number  $n$ . Bohr applied the basic methods of mechanics to Rutherford's simple nuclear model of the hydrogen atom, with this added quantisation. He quickly found that the electron energy levels are then given by:

$$E_n = -\frac{13.6}{n^2} \text{ eV}$$

The units are electron volts, mentioned above and familiar to all chemists. You can use this equation to calculate any energy level you want to, and hence the difference between any two energy levels. When you convert eV difference into electromagnetic frequency (remember that  $hf = E_i - E_k = \text{photon energy}$ ), then you can predict where any spectral line of hydrogen is located. Using  $\lambda = c/f$ , yields the corresponding vacuum wavelength.

A typical calculation runs like this. You want to find the wavelength,  $\lambda$ , of the beginning of the Paschen series that starts with a transition between energy levels  $E_4$  and  $E_3$ . The energy difference is therefore  $E_4 - E_3 = -13.6\left(\frac{1}{4^2} - \frac{1}{3^2}\right) = 0.661 \text{ eV}$ . Now  $1 \text{ eV} \equiv 1.602 \times 10^{-19} \text{ J}$  and hence  $E_4 - E_3 = 1.06 \times 10^{-19} \text{ J} = hf = hc/\lambda$ . Using the values of the constants  $h$  and  $c$  gives the wavelength required as  $1.88 \text{ }\mu\text{m}$ , which is in the infrared as expected from the previous discussion.

Bohr won the Nobel Prize for Physics in 1922, the citation saying "*for his services in the investigation of the structure of atoms and of the radiation emanating from them*". This was only 9 years after his pioneering paper, a short time in Nobel terms and indicative of how important Bohr's work was seen by his contemporaries. [Bohr had a son, Aage Bohr, who won a share of the Nobel Prize in Physics 53 years later for his work on the structure of the atomic nucleus. There have not been many father and son Nobel winners.]

### *Schrödinger's wave equation*

Bohr's model is quoted in every introductory textbook on atomic structure and those who take our course PX2510 will find out how the details come together to give the energy expression I've quoted above. In reality, Bohr's model was only top dog for about a dozen years. It was then replaced by **Schrödinger's wave equation**. This has stood the test of time. (Erwin Schrödinger (1887 - 1961) received the Nobel Prize for his work, hearing about it while he had a visiting post in Oxford in 1933. He spent from 1939 to 1956 in Dublin, becoming a naturalised Irishman).

Schrödinger built his idea upon de Broglie's perception in 1923 that particles, and electrons in particular, could have a wavelength associated with them, as mentioned in the 'fundamentals' section of this course. In 1926, Schrödinger built a wave picture of electrons in atoms. 'Picture' is the right word because visualization of the atom was one of Schrödinger's goals. However, it wasn't long before Schrödinger's wave equation was found to be applicable more generally and his initial idea of creating a visual picture had to be abandoned in terms of a more general interpretation of what the equation meant. This generality was realised by another famous physicist, Max Born (who some time later was based in Edinburgh). Schrödinger's equation now brings with it a more subtle philosophy for describing nature. In this philosophy, everything has a wavefunction, often called  $\psi$ .  $\psi$  depends on both space and time. You find out what  $\psi$  is by solving the **time dependent Schrödinger equation**. When you know what  $\psi$  is, then you can work out the properties of the system you're talking about. This may sound a bit like a theory of everything that promises much and probably doesn't deliver anything. Schrödinger's theory does deliver.

Let's home in on an atom. As Bohr said, the atom has discrete energy states. Every system has discrete energy states. If a system is put into one of these states and left completely isolated, it stays in that state. In the language of wavemechanics, these states are called **the energy eigenstates of the system**. The  $\psi$  values for these states are special. They can be described as the product of a space dependent part and a rather simple time dependence (written  $e^{iEt/\hbar}$ , for those who'd like to know, where  $E$  is the energy of the state). The space dependent part of the wavefunction shows how the system you are talking about is spread over space, for example how the electron distribution of the atom varies in space. The space dependent part obeys the **time independent Schrödinger equation**, which looks like this in one dimension:

$$-\frac{\hbar^2}{2m} \frac{d^2\psi}{dx^2} + U(x)\psi = E\psi$$

$U(x)$  is the potential energy derived from the forces that act on the system being discussed. This equation both defines the energy states  $E$  of the system **and** defines the wavefunctions that describe the system when it has these energies. If you put an energy value into the right-hand-side that is not an energy level of the system, then the equation has no solutions. If you make  $E$  one of the energy eigenvalues, then the equation has a solution that is called the corresponding eigenstate. There are techniques for finding the  $E$  values and the corresponding  $\psi$  values, but they will not be part of this course. We'll just look at a few results.

In summary, if you want to know what the energy levels are for an atom **and** what the electron distribution looks like around the atom for each state, solve a time independent equation like the above, but allow coordinates in 3 dimensions. There is one final important point. The value of  $\psi$  itself at a point in space does not on its own describe what the electron distribution looks like in an atom. The value of  $\psi\psi^* = |\psi|^2$  gives the probability of finding an electron at each point in space. The next two or three lecture slides show how it all works.

#### *Application to the hydrogen atom*

A normal hydrogen atom is simply one electron and one proton. The electron is in an **electrostatic potential**,  $U$ , that is given by the ordinary law of electrostatic attraction, discover by Coulomb over 200 years ago. Honours physics students will recognise the equation:

$$U(r) = -\frac{e^2}{4\pi\epsilon_0 r}$$

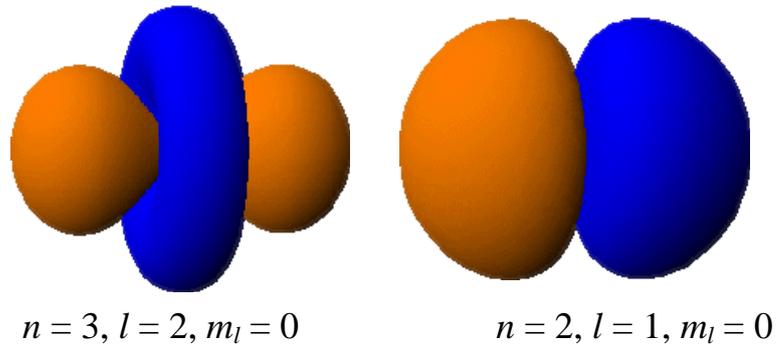
Here 'e' is the magnitude of the charge on both electron and proton and  $r$  is their separation. Using this in the time independent Schrödinger equation generates solutions  $\psi$  that are described by 3 quantum numbers:

1.  $n$  the principal quantum number, which takes on values 1, 2, 3, 4, .... . It's not obvious but for a hydrogen atom the energy depends only on  $n$ , as in the Bohr model.
2.  $l$  the angular momentum quantum number, which takes on values 0, 1, ..  $n-1$ . For historical reasons, the  $l$  quantum numbers are denoted by letters  $s, p, d, f, g, \dots$  for  $l$  values of 0, 1, 2, 3, 4, ... . Thus you can have  $2s$  and  $2p$  states, but not  $2d$  states,

because with  $n = 2$ , the largest value of  $l$  that is allowed is  $l = 1$ .  $l$  specifies the total angular momentum of the atom in units of  $\hbar$ .

- $m_l$  the quantum number describing the orientation of angular momentum, which takes on values  $-l, (-l+1), (-l+2) \dots, 0, \dots, (l-1), +l$ . Thus  $2p$  states exist with  $m_l = -1, 0, +1$ .

This is a course on optics and not on quantum theory so I don't want to spend too much time labouring the details. However, spectral light emission is intimately tied in with the quantum nature of the atom and hence I'll say a little more. Notice that the Schrödinger picture is different from the Bohr one in that the  $l$  quantum number is now the one associated with the electron's angular momentum, not the  $n$  number. Moreover, the lowest  $l$  state is the state with  $l = 0$ , which has no angular momentum at all. Bohr talked about electron orbits and chemists still like the description 'orbitals' but in Schrödinger's picture we don't have orbits, we have probability distributions of electrons around the nucleus of the atom. Remember that the probability distribution is given by  $|\psi|^2$ . Variant shapes of probability distributions for a selection of quantum states are clear on the next lecture slide. What are shown are surfaces of constant probability.



In order to show how the probability changes in space, cutaway shapes or other plots are necessary. One final point to make is that the different colours above represent opposite phases in different parts of the wavefunction. Phase information is lost in forming  $|\psi|^2$  but it is relevant when working out the strengths of energy level transitions.

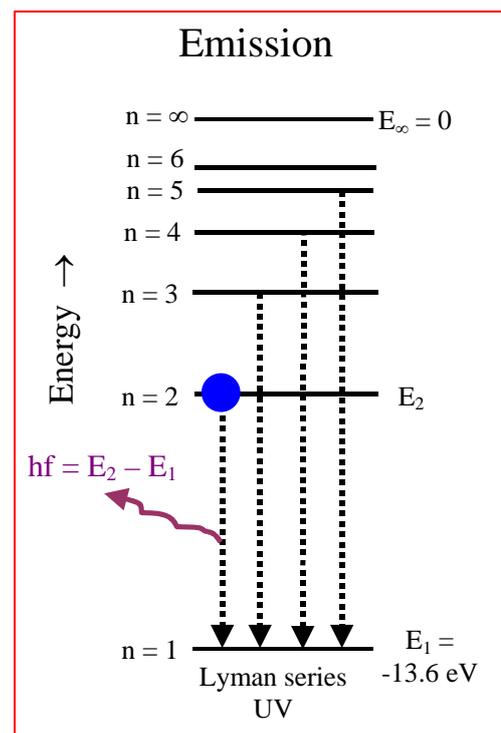
### Energy level transitions

Emission lines are caused by electrons dropping down to a lower energy state. The wavefunction of the electron will change shape. During the transition, the wavefunction satisfies the time dependent Schrödinger equation and oscillates with a frequency  $(E_i - E_k)/h$ . The animation on the lecture slide shows the behaviour. The transition is often represented by an energy level diagram like the one here. In this case the emission highlighted produces the longest wavelength line in the Lyman series, namely one with an energy of 10.2 eV or a wavelength of about 121 nm.

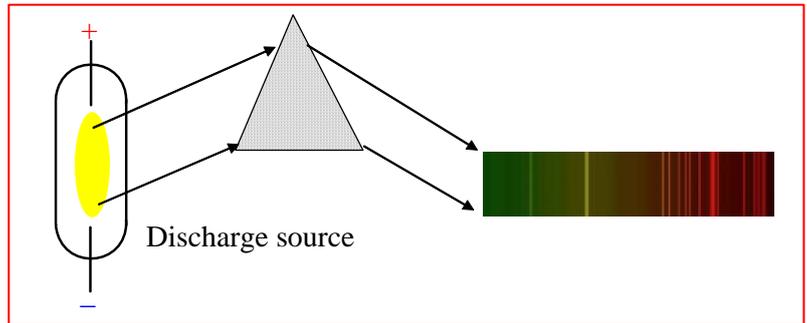
### Factors affecting the spectrum

Four factors are highlighted here:

- the number of excited atoms;



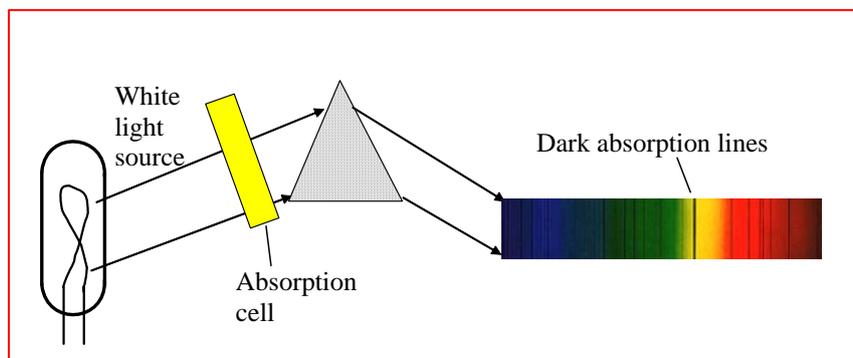
- which excited states are populated when you have a whole collection of atoms, and by how much each state is populated. You will meet the Boltzmann distribution in later courses. This describes how the population of states depends on temperature.
- the *transition probability* between states. This is the likelihood that electrons will fall from one given state to another. Relative values for different transitions can be calculated from the wavefunctions for the initial and final states. Many energetically possible transitions cannot in fact take place and this is partly accounted for by the so called 'selection rules' that govern transitions. The most obvious selection rule is that  $\Delta l = \pm 1$ . This comes about quite simply because the angular momentum of a photon is  $\hbar$  and from the law of conservation of angular momentum if a photon is created with this angular momentum, then an atom must change its angular momentum by this amount.
- the pressure of the emitting substance. This effect is more subtle: higher pressure produces atoms closer together, which spreads out their energy levels, making the lines broader.



The picture above shows suitable conditions for producing an emission spectrum. One variant way of creating emission spectra is by putting a little of the material in a flame and using the heat of the flame to excite the atomic electrons. It was in this way that Kirchhoff and Bunsen discovered the new elements caesium (Cs) and rubidium (Rb) spectroscopically in the early 1860s. A score of other elements were discovered spectroscopically before the end of the nineteenth century. One of these was helium (He), discovered by the astronomer Norman Lockyer in the corona of the Sun during his spectroscopic observation of the 1868 total eclipse. The conspicuous yellow line in the helium spectrum occurs near the sodium yellow lines and Lockyer realised that no element he had seen on Earth had a line at the observed position. He named the element after the Sun (helios) but Lockyer made a mistake in thinking it might be a metal and added the metallic 'ium' ending.

### Creating absorption spectra

To create a dark line spectrum, you need a source of white light behind the absorbing medium. Those frequencies that are just right to give an electron enough energy to rise up the energy level diagram will absorb radiation and

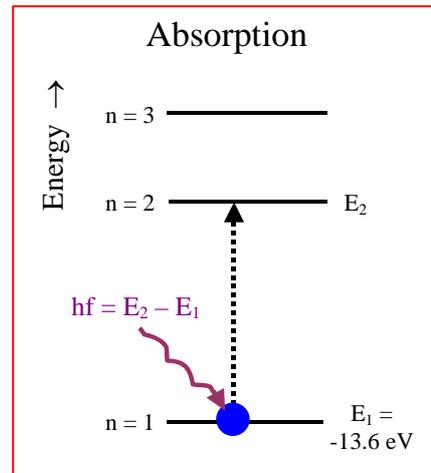


**excite the electron.** The sketch shows what would happen with incident light of wavelength about 121 nm falling on a hydrogen atom. The electron would be promoted to a higher energy state and the hydrogen atom is described as **excited**.

The technique of absorption spectroscopy is very common in chemistry labs. It is used particularly for molecular identification, because molecules have distinctive and complex absorption lines in the IR region of the spectrum. A sample cell containing a material, whose composition might be unknown, is illuminated by a continuous spectrum source and the absorption lines looked for. Molecules, as opposed to atoms, tend to have energy level differences in the IR, due to their vibrations and their rotations.

#### *Case study of spectroscopic absorption imaging*

Four pictures are shown on the lecture slides, courtesy of the US Geological Service. The case study illustrates the creation of a very detailed mineral map of an exposed rock region of Nevada, 10.5 km × 17 km wide. The image pixel size is 17 m. The same technique is already in use on Mars missions.



- 1) From images taken at 3 visible wavebands only, a full colour picture can be created of the area. The technique is discussed in the section on colour, coming up soon.
- 2) Data from 6 wavebands, including 3 in the IR, are combined to create a false colour picture. This picture highlights a range of minerals but 6 wavelengths are inadequate to determine precisely what the minerals are.
- 3) 224 channels, at wavelengths ranging from the deep violet to well into the infrared, allow the collection of a detailed spectrum of the reflected light for each pixel. The spectra can then be matched to reflectance spectra of known minerals. With 224 channels, even fine difference between related minerals can be distinguished.
- 4) There is enough detail in the spectra to distinguish, for example, the difference between low aluminium, medium aluminium and high aluminium muscovites.

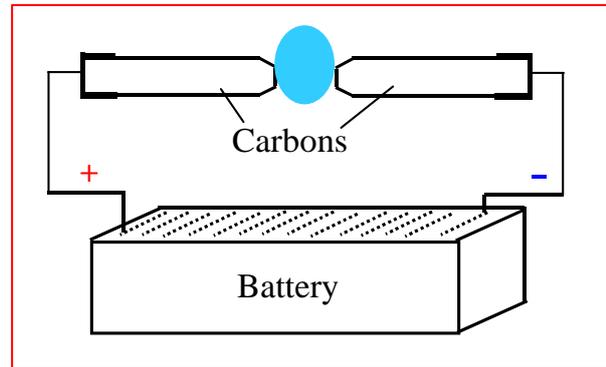
#### *Making light*

When you think of light, you think of the Sun. As far as light is concerned, nobody does it better. Sunlight is free and available to all, in varying amounts, without lifting a finger. It is true that there is an exploitation cost, such as providing windows in a house to let in the light, making solar cells to charge your batteries, and so on. Nothing in life is really free. The Sun generates energy from a sustained nuclear fusion reaction. Nothing is 'burnt', in the chemical sense of combining carbon, or hydrogen or sulphur with oxygen. The spectrum, as we've seen, is basically continuous but interrupted by a myriad of Fraunhofer lines.

Making light by burning things has been the traditional method here on Earth. For millennia, the burning was simply the campfire. Oil lamps with wicks appeared in Palaeolithic times, the oil being animal oil. A late 18<sup>th</sup> century improvement introduced circular wicks and glass chimneys, turning oil lamps into quite respectable sources of light. Candles go back to Roman times and, in terms of light output, modern candles are probably not much different from those of old, excepting that today's waxes are purer than the old wax or tallow candles. Gas lighting was an invention of a Scotsman, William Murdoch, in the very early 19<sup>th</sup> century. Gas lamps flourished in the 19<sup>th</sup> century, becoming standard fittings in town buildings. I'll say a bit more about gas lighting later.

*The arc lamp*

The arc lamp was the first electric light. It was created by Humphry Davy in 1808, when he was Professor at the Royal Institution in London, after a generous benefactor had given him an enormous battery. The arc is generated by touching together two carbon rods that are connected to a source of electricity capable of providing a current of many amps. Once the rods have been touched, then they are separated a little. The arc



keeps going, glowing at some 3300°C, brighter than several hundred candles, hotter than the hottest filament bulb we have today. Davy's discovery didn't catch on. No-one else had a battery that big and the cost of it far exceeded the value of the light, dazzling though it was.

What made carbon arc lights possible on a commercial scale, and indeed all of the electric light industry viable, was the development of dynamos that could supply hundreds of amps. Faraday, the discoverer of the phenomenon of electromagnetic induction, famously left it to others to develop the phenomenon for practical applications. Efforts to generate electricity from mechanical motion began on a desk-top scale later in the 1830s and continued more or less on that scale for another 25 years. It took until the 1860s before the first machines driven by steam engines of several horse-power (effectively kW) were tried. They were very inefficient by modern standards but led to the development of electric light in high value situations. For example, the Paris Universal Exhibition of 1867 demonstrated to the public a replica of an experimental electric lighthouse erected on the Brittany coast.

The break-through needed to progress electric lighting was the design of a more efficient dynamo. Gramme's machine of around 1870 was a key development, followed by that of Edison and related efforts. Just one industrial scale dynamo could then supply the electrical needs of many people. At first each building or institution had its own dynamo but in the late 1880s and early 1890s municipalities began installing generators to produce public electricity. Aberdeen's first power station was opened in 1893. It was the first power station in the country built by a municipality.

A serious problem with the carbon arc is that the carbons burn away. A host of different mechanical devices were invented to move the rods together at the same rate as they burnt. One widespread system was the 'Jablochkoff candle', in which the two carbon rods were run parallel to each other, separated by a packing of china clay that disintegrated at the same rate as the carbons burnt away. A pair of rods lasted 2½ to 3 hours. For lamps that had to last 8 hours, four pairs of rods were put in the same globe housing and the current manually switched between them by a commutator as each carbon neared the end of its life. The light was good, but you had to work hard for it. A typical carbon arc consumes about 500 W and produces light of several hundred candle power. Most of the light does not come from the arc itself but from glowing carbon. In an open environment, the carbons burn quickly. In an enclosure, the oxygen is quickly consumed and the carbons last many tens of hours after that.

### *The incandescent lamp*

The incandescent lamp that we're now familiar with isn't a development of the carbon arc but something different, based on a different principle. Thomas Alva Edison (1847 - 1931) was one of the great technological inventors. He did more than anyone to bring the incandescent lamp to the public. Both he and Joseph Swan in England had developed a viable lamp using a carbon filament. Neither of them had invented the concept of a filament light, which was over 30 years old, but they had both overcome the problem that the filaments used to burn out very quickly. The solution was to evacuate the glass bulb with one of the new high vacuum Sprengel mercury vapour pumps. Edison and Swan's carbon filaments were made in different ways and the bulbs fitted into the electrical circuit by different fittings. In spite of this, with some compulsion they joined forces to commercialise their ideas under the *Ediswan* banner. It was Edison in his huge development factory who did most to advance the carbon filament bulb. Ediswan had the monopoly on the supply of 'glow lamps' for 14 years from 1879.

### *Metal filament light-bulbs*

Carbon filament bulbs weren't very efficient, meaning that the fraction of electrical power consumed that appears as light was small, and in spite of the Sprengel pump they also didn't have the lifetime we now expect from a light bulb. The filaments lasted a reasonable time but the carbon evaporated and deposited a black film on the inside of the bulb. The bulbs were manufactured in clear glass and their narrow filaments could cast sharp, hard shadows. Such lighting is not flattering to personal appearance and electric lighting was not welcomed in households as enthusiastically by women as it was by men.

Around 1890 a few people tried to develop metal filament bulbs. The metal had to have a very high melting point. There were several candidates. First osmium was used, then tantalum and, later, tungsten. These 'heavy metal transition elements' all have exceptionally high binding energies and this leads to their high melting points and the comparatively high strength of thin filaments. Unfortunately, none of these metals can be drawn out nicely as thin wire. Making thin metal filaments was therefore a significant technological challenge.

The motivation was to develop a filament that runs hotter than a carbon filament and is therefore more efficient at producing light. The physics behind this is that the peak of the blackbody spectral curve is in the infra-red for all light bulbs. The nearer it is to the visible, the more efficient the bulb and Wien's law tells us that the only way to get it nearer the visible is to have a hotter filament. Nowadays, a typical metal filament bulb operates at about 2500°C. Wien's law shows that its wavelength of maximum emission is just over 1  $\mu\text{m}$ , which is still in the IR. Hence even modern metal filament lamps produce more IR than light but because they run at a higher temperature than carbon filament lamps, modern lamps are more efficient. One added characteristic of metal filament lamps is that you need a much longer filament to achieve the same resistance and hence pass the same current. In one way this is an advantage. A short but manageable length of filament can be made just as hot from a low voltage. Hence, torch bulbs, car lamps, miners' lamps and lamps of all sorts operating at 2 – 12 V become possible. This was a huge improvement on earlier technology. If you look at early car headlamps and bicycle lamps in a transport museum, you'll see that they produced their light by burning gas. The gas was acetylene, made in the lamp itself by dropping water onto calcium carbide.

To cut a long story short, the modern tungsten filament, gas filled, light bulb was invented by Irving Langmuir in 1913, in so much as one person can be given credit for the contributions

of many. The gas is a combination of nitrogen and argon, both natural constituents of our atmosphere. It is inserted after the bulb has been highly evacuated. The gas convects the evaporated tungsten up into the base of an inverted bulb, reducing the deposition on the sides. It also cools hot spots on the filament, prolonging its life. At the same time, Langmuir developed the spiralling of filaments, to increase their mechanical strength and reduce the effect of tungsten evaporation. Evaporation is less because any small element of the filament has some tungsten nearby onto which evaporated atoms can fall. For a long, thin filament this isn't the case. Langmuir's work with filaments was a significant contribution to his being awarded the Nobel Prize in Chemistry in 1932 "*for his discoveries and investigations in surface chemistry*". The spiralled filament had the added advantage of broadening the light source, hence reducing the glare and the sharpness of shadows. Concerns about these were finally removed by the introduction of 'pearl' light bulbs, whose insides were frosted using a hydrofluoric acid etch.

### *Resurgence of gas lighting*

Just before the end of the 19<sup>th</sup> century, Carl Auer von Welsbach developed his gas mantle that transformed the effectiveness of gas lighting. Previously, gas lighting had simply used a naked flame. Try reading a newspaper in a dark kitchen by the light of a gas cooker flame. On reflection, perhaps don't. You'll probably move the paper so near the flame as to risk setting it on fire. It's not quite a fair test of a gas light flame, which is made more useful by the presence of incandescent un-burnt carbon particles. Gas lamp burners were designed to produce a yellow flame. The same light enhancement process occurs in every candle flame. Because of the conspicuous defects of gas light flames, others beside Welsbach also tried to develop a gas mantle. It was Welsbach's version, though, that became adopted all over the world. His mantle was essentially a small cotton sock impregnated with rare earth oxides. When the mantle was first used, the cotton burnt away leaving a thin shell net of hardened oxides that glowed intensely white in the gas flame. Welsbach found that just 1% of cerium oxide added to a base of thorium oxide gave a particularly bright source. The mantles used in Calor gas lights today are, as far as I know, close to the original idea. The fabric is now synthetic and the oxides are coated to increase their hardness.

There are three reasons why mantles are so much better than a simple flame. First, the oxides are opaque, unlike the flame. Secondly, they become hotter than a naked flame because the burner is designed for complete combustion of the gas. Thirdly, the oxides emit more light than you would expect of a body at the temperature they are at, because of the phenomenon of luminescence. Much earlier in the 19th century, limelight used some of the same ideas. A piece of lime ( $\text{Ca}_2\text{O}$ ) was heated in an oxyhydrogen flame. It glowed intensely, creating the floodlight of the times. The invention of the limelight was made by the now forgotten Lt. Drummond and applied by him during the Ordnance Survey of Ireland and Scotland in 1826. Others quickly realised its value and not for nothing did the stars of the theatre want to be 'in the limelight'.

The gas mantle prolonged the use of mains gas lights by some 50 years in places. There were still many gas street lamps in Britain by the middle of the twentieth century. When electric lighting was first produced it had to compete only with gas flame illumination. Not only were electric lights brighter and steadier but they did not consume valuable oxygen in the air and did not create unwanted water vapour, carbon dioxide and even poisonous carbon monoxide in the rooms. Proponents of electric lighting, at least, claimed that being in a room for a good while with gas lighting gave people headaches and a feeling of lassitude. Gas was a fire hazard and after a while in use, gas lights damaged books, pictures and room decorations.

The gas industry should have been reeling from this attack by electric light on its flawed product but in the event town gas production continued to rise almost uninterrupted. Its saviour was not really the gas mantle but the rise of gas usage for domestic cooking and, to some extent, heating.

### *Gas discharge tubes*

Gas discharge tubes are sealed glass tubes containing an electrode at either end and a low-pressure gas within. They were investigated extensively in the second-half of the nineteenth century. They were investigated just to find out what on earth was going on when the gas within the tubes took on a multitude of different appearances under different circumstances. Some discharges looked a bit like the aurora, yet the strongest auroral spectral lines were not found. Names associated with this development were Gassiot, Spottiswode, De la Rue, Müller and Geissler. Line spectra produced by the tubes were investigated in detail to try to find the laws that related their wavelengths. There was little success here until Balmer at the very end of the century found a relationship that linked the wavelengths of the hydrogen lines, a relationship that now carries his name. No-one knew why this relationship should hold. Neither did Balmer. The secret was safe until the Bohr atom came along. Crookes in particular studied what happened with gas discharges at varying pressures and similar work led pretty directly to the discovery of ‘cathode rays’, that J.J. Thomson identified as electrons, and led Röntgen to the discovery of X-rays. Gas discharges produced a fairly feeble light and it wasn’t until the beginning of the 20th century that people began seriously to develop the phenomenon into a useful light source.

### *The fluorescent tube*

The ‘fluorescent tube’ is a more efficient light producer than a filament bulb. In other words, a greater fraction of the electrical power consumed appears as light. The light we see comes from the internal tube coating, not directly from the glowing gas within. Because a significant amount of its light comes from spectral lines, the colour of the tube can be varied comparatively easily by suitably choosing the light emitting material. No filters are needed to produce coloured lights.

The tube works by setting up a low pressure mercury discharge that emits a bright line spectrum, mainly in the UV. Inside the tube when it is running the vapour is in the form of a plasma, a conducting medium of positive ions and electrons. (Plasmas are a common state of matter in the universe, in stars, and can also be found in our upper atmosphere). The electric field within the tube accelerates the mercury ions until they have enough energy to become electronically excited upon collision. The excited mercury atoms rapidly emit their excess energy as light, mainly in the UV. Coated around the inside of the tube is a ‘phosphor’ that gives out spectral emission in the visible when struck with higher energy UV photons. This process is also known as fluorescence. With a suitable choice of phosphor, you can give the fluorescent tube any colour temperature you choose. Values from 2500K to 7500K are quite readily available. You can find out more about fluorescent tubes on your own.

### *Plasma screens*

Plasma screens are used in modern, flat-panel colour displays. They are typically bright, larger than most computer monitors, distortion free, flicker free and clearly visible from a wide range of angles. Essentially they use fluorescent tube technology for each pixel. The

plasma that emits the UV is not generated in mercury vapour but in a mixture of neon and xenon. You are invited to investigate their working via the World Wide Web.

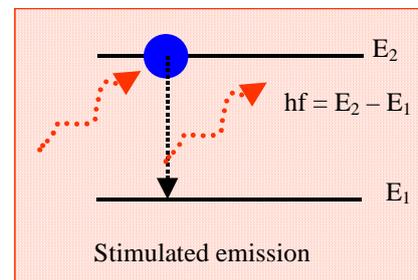
### Laser light

Laser: **L**ight **A**mplification by **S**timulated **E**mission of **R**adiation. Although lasers are common sources of light nowadays, they are really quite complicated devices, particularly when compared with a filament bulb. They do, though, have some very desirable characteristics. Laser light is very intense, it is highly directional and it is **coherent**. We'll see the significance of coherence later in our course. You can certainly tell that laser light is different, by the speckle pattern it shows when shone onto a white wall.

Why is laser light so dangerous? It is not the total power at all, which for many lasers is very small. The lasers in our labs produce just less than 1 mW light output. A standard, inefficient, 100 W bulb gives out a few watts of light. It is several thousand times more powerful. Why does 1 mW merit safety notices all around? It's all to do with the directionality. A 100 W bulb radiates into a solid angle of  $4\pi$  steradians. The laser may have a divergence of  $1/3^{\text{rd}}$  of a degree in two dimensions, which represents a solid angle of  $3.4 \times 10^{-5}$  steradians. The irradiance in  $\text{W sr}^{-1}$  is therefore over 100 times greater for the laser than for a 100 W bulb. You certainly wouldn't want to look at a 10,000 W bulb.

Moreover, if you look directly into a laser beam and your eyesight is good, the entire 1 mW is focused onto an area that may be 0.01 mm across. This gives a power density of  $(1 \times 10^{-3}) / (1 \times 10^{-5})^2 \text{ W m}^{-2}$  or  $1 \times 10^7 \text{ W m}^{-2}$ , which is much more than the retina can stand. The moral of these figures is that there is good reason to be very careful with lasers. The safest lasers are totally enclosed. For lasers more powerful than those in our labs, you need special goggles that allow you to see the scene around you but cut out almost all the laser light.

In spite of these important practical concerns, have you always wanted to know how a laser works? Now you will.



Stimulated emission is the reverse process to absorption of light. You'll remember that with absorption, an atom in a lower energy state is excited to a higher energy state by absorbing radiation whose photon energy is exactly equal to the difference in energy between the two states.

$$hf = E_2 - E_1 ,$$

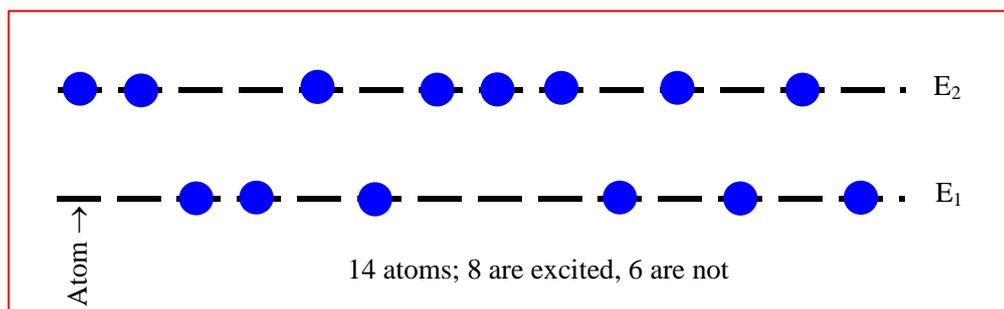
in our earlier notation. Stimulated emission of radiation happens when an excited atom at energy  $E_2$  is stimulated by radiation already present at just the right frequency to emit light and drop its energy level to  $E_1$ . The very same equation applies. What is also true is that the emitted light is in the same direction as the stimulating light and is *in phase* with the light that provided the stimulation.

Stimulated emission looks like a 'good thing' for making light. You start off with some light in the background and you end up with more light. However, with atoms at room temperature in the presence of the amount of blackbody radiation that is produced at room temperature,

then the chance of getting stimulated light emission is effectively zero. It was, in fact, Albert Einstein who worked out the relative probabilities of spontaneous emission, stimulated emission and absorption. The results are called the Einstein radiation coefficients, but I'm not going to dwell on the details. Lasers make stimulated emission work. But how? You need another idea first, that of **population inversion**.

### *Population inversion*

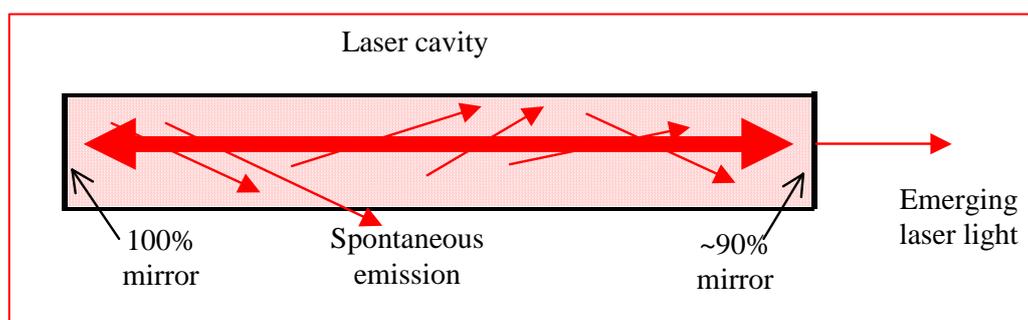
The next diagram shows a collection of similar atoms. Under normal circumstances, most atoms would be in the lower energy state. In the sketch, more atoms are in the upper level, or excited state. This is known as 'population inversion'.



If you try to create a population inversion in an energy level picked at random, using some method of exciting atoms to a higher energy level, then you are likely to fail. Most excited states decay very quickly, in less than 10 ps (a pico second  $\equiv 1.0 \times 10^{-12}$  s), by spontaneous emission. To create a population inversion you need to find an excited state that does not decay by the normal radiation process. Such states do occur and are called **metastable states**. With a metastable state, you can hope to create atoms in such a state faster than they eventually decay by some process or other. The process of interest in lasers is stimulated emission.

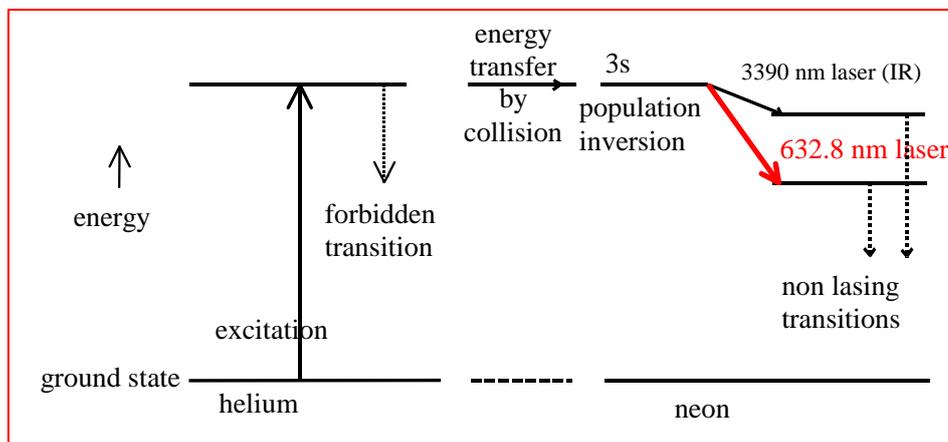
I said above that absorption and stimulated emission are opposing processes. Which one wins is quite simple. Under normal circumstances, absorption wins, because the number of atoms in the lower energy state is greater. There are more atoms able to absorb radiation than emit it. As we've already seen, absorption lines occur commonly and are not hard to create in the lab. Stimulated emission only wins if for some reason the population of excited states is inverted. To make a laser, an inverted population must be created, and maintained.

### *A simple laser cavity*



A simple laser cavity, such as you find in some HeNe gas lasers, consists of a tube with accurately parallel mirrors at either end. One mirror reflects completely, the other lets through a small fraction of light by virtue of having a very thin reflective coating. The Ne atoms are excited. I'll say how soon. Spontaneous emission occurs in all directions and most of the light goes off to the side of the tube and is lost. A small amount is reflected back and forward between the mirrors. As the radiation builds up in this direction, the likelihood of stimulated emission builds up and further stimulated emission is generated *just in the direction perpendicular to the mirrors*. Even more usefully, the stimulated light has both the same wavelength as the stimulating light and is in phase with it. The stimulated emission adds to the existing radiation in this one direction and increases the chances of yet more stimulated emission. This is the amplification process that is part of the laser's name. The intensity of light increases very quickly to such a level that the rate of creating population inversion can just keep up with the supply of light leaking out of one end. And there you have it: emerging light that is intense, highly directional and coherent. The HeNe laser produces red light at 632.8 nm.

### *The working of a HeNe laser*



Perhaps you noticed that one link in the chain of explanation was missing. How do we generate the population inversion of the Ne atoms? You can't just heat them up to excite them, because heating always creates more populated low energy states than higher energy states. This is a consequence of the Boltzmann distribution, mentioned much earlier. The method used in the HeNe laser is subtle.

An electrical discharge excites He atoms also present in the tube. It does so more easily than it can excite Ne atoms. One particular energy state of He at about 20.7 eV above the ground state (the energy of unexcited helium) is metastable, meaning that the He atoms can't radiate energy from this state in the normal way and drop to a lower energy. The metastable He atoms accumulate quickly so that there are large numbers of them in the tube. Now it just so happens that there is a metastable excited state of neon at more or less exactly the same energy of 20.7 eV above the ground state of neon. When an excited He atom collides with an unexcited Ne atom, it may well give up all its excitation energy to the neon. Because there are lots of metastable He atoms and lots of Ne atoms this process happens so often as to create a population inversion of excited, metastable, Ne atoms. Bingo, you have all the conditions for lasing in place. The scheme works, very well. Lasing can be maintained continuously. The final state of the Ne atoms once the laser light photon has been emitted by stimulated emission is itself well above the ground state and this excited state has a very short lifetime.

As a result, electrons in this state quickly decay to the ground state, helping to maintain the population inversion.

The HeNe system can also lase at two wavelengths in the IR. One of them is shown in the simplified diagram above.

**End of Light Sources & Spectra**

*JSR*