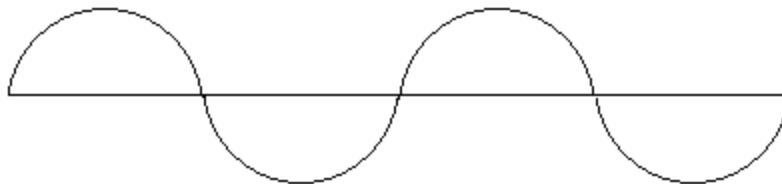
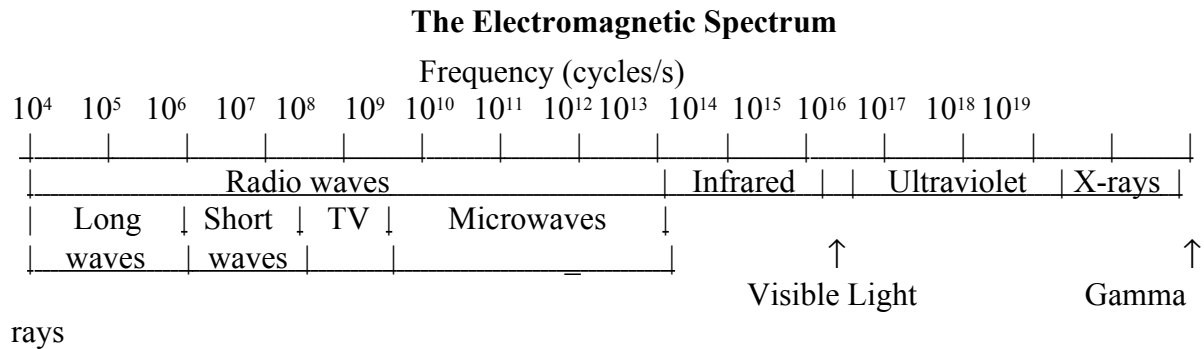


THE ELECTROMAGNETIC SPECTRUM

Electromagnetic radiation comes in a large range of frequencies and wavelengths. The range is referred to as "the electromagnetic spectrum". To give a few examples: in the range of 10^4 to 10^{12} Hz the electromagnetic spectrum has a portion called radio waves. From 10^{12} to 10^{14} is the range called infrared. From 10^{17} to 10^{19} we have what are called the X-rays. Microwaves are part of the radio wave portion in the shorter wavelength portion of the radio waves.

Infrared radiation consists of the range of frequencies that can make molecules of most substances vibrate internally. An increase in internal vibration is measured by an increase in temperature.



Of all these radiations, your eyes are able to sense only a very narrow band ranging from about 400 nm (violet) to 700 nm (red). This corresponds to 7.5×10^{14} to 4.3×10^{14} Hz. This narrow band is called the visible spectrum and consists of all the colours that can be seen from red through orange, yellow, green, blue, indigo and violet. White light is composed of all these colours in equal amounts. Red light borders on the left with infrared (under red) and violet borders on the right with ultraviolet (above violet).

Following the rules in the last worksheet the radio waves with a frequency of 10^4 Hz would have the longest wavelength of 30,000 metres. As we go to the right the wavelengths get smaller and smaller. Microwaves with a frequency of 10^{10} Hz have a wavelength of 3 mm. X-rays with a frequency of 10^{17} Hz have wavelengths of 3 nm. At the very end of the scale we have gamma and cosmic rays which have a frequency of 10^{19} Hz which means a wavelength of only 0.03 nm.

The Energy in a Light Wave

Why is it people will pay money to see reflections of light at a laser light show, or even spend time in a tanning booth yet they get paranoid when out on a hot sunny day because they fear skin cancer? The answer lies in what they've been told. Visible light is harmless, tanning booth light is safe if used in moderation but intense direct sunlight can be harmful. Why is that?

In 1900, Max Planck (1858-1947), a German physicist, coined the term "*photons*". It seems that light can be thought of in two ways. Either as waves with troughs and peaks or as bundles of energy. Anyone standing on a beach can attest to this. Have you even been knocked over by a wave? Max Planck proposed that electromagnetic radiation is emitted in tiny packets of energy which he called "photons". Each photon pulses with a frequency, ν , and travels at the speed of light.

Planck also proposed that photons with very high frequencies carried more energy than ones with lower frequencies. What he actually said was "*the energy of a radiation is proportional to its frequency.*" Albert Einstein later confirmed this and coined a new term, the *quantum* of energy.

$$\text{Energy of a photon}(E) = h\nu$$

where E = energy ν = frequency h = Planck's constant = 6.63×10^{-34} J·s

Lets take a look at a particular event in everyday life, photosynthesis in plants. Photosynthesis is started when a plant absorbs light. It is the frequency of light that is important. Therefore different kinds of light will have different effects on the efficiency of how plants do photosynthesis. We know that plants grow best in blue light (it just happens to have one of the higher frequencies) and not as well in red light. Plants do not grow well in green light which is a middle frequency because plant pigments do not absorb green light. Plants in fact reflect green light, which is why plants appear green to us humans!

To sum up: the higher the frequency the more energy a photon has. Since ultraviolet light has a higher frequency than infrared light, ultraviolet light photons each carry more energy than do infrared photons. Infrared photons warm you, ultraviolet photons burn you, cosmic rays are so tiny and so energetic they pass through you before any damage can be done.

Atomic Line Spectra

The spectrum of electromagnetic radiation that we looked at in the last set of notes is called the "continuous spectrum" because it contains the light of all colours. This spectrum is formed when the light from the sun, or any object is heated to a very high temperature. (You have of course heard of metals being heated until they were white hot). This light can then be spread out by passing it through a prism onto a screen. A rainbow is a continuous spectrum of visible light that has been spread out by tiny water droplets suspended in the air.

If we look at a pure gas like hydrogen or neon or anything else pure we do not get a continuous spectrum. When an electric current discharge passes through the gas the electric current *excites*, or energizes the atoms of the gas. The gas then releases this energy in the form of visible light as the atoms return to a lower energy state. When a beam of this light is passed through a prism

or a spectrometer we do not see a continuous spectrum. Instead, only a few colours are observed and these are in a series of individual lines. This series of lines is called the element's atomic spectrum.

Different elements produce different spectra. These different spectra are called the atomic spectra and are unique enough to be considered as characteristic as a fingerprint.

The equation $E=h\nu$ showed the simple relationship between the frequency of light and its energy. Atomic spectra show us that an atom produces only certain characteristic frequencies and this means that there are only certain characteristic energy changes taking place inside the atom. For example, in the atomic spectrum of hydrogen, there is a red line. That red line has a wavelength of 656 nm. If you do the math you'll see that the frequency is then 4.57×10^{14} Hz. Using the Planck's constant equation it can be determined that each photon of this light carries 3.03×10^{-19} J of energy. What is important here is that when hydrogen produces a red line in its atomic spectrum, its frequency is always 4.57×10^{14} Hz and the energy in each photon is always 3.03×10^{-19} J. It is always the same. This tells us that when an atom is excited and then loses energy, not just any arbitrary amount is lost. Only certain specific energy changes can occur, which means only certain specific frequencies of light are emitted.

In order to explain this we must use the following model. In an atom, an electron can have only certain definite amounts of energy and no others. The electron is restricted to certain energy levels and must use only these levels. We also say that the energy of the electron is *quantized*, meaning once again that the electron's energy in a particular atom can have only certain values and no others.

quantized - to have a certain specific quantity.

The energy of an electron in an atom can be compared to the potential energy of a ball on a staircase. The ball can only come to rest on a step, and on each step it will have some specific amount of potential energy. If the ball is raised to a higher step, then its potential energy will be increased as well. When the ball drops to a lower step, its potential energy decreases. But the ball cannot stop between steps. The ball can only rest at the specific energy levels specified by the steps. So it is with the electrons in an atom. The electron can only have energies corresponding to the set of electron energy levels in the atom. When an atom is supplied with energy, as in a gas discharge tube, an electron is raised from a low-energy level to a higher one. When the electron drops back, energy equal to the difference between the two levels is released and this energy gets emitted as a photon. Because only certain energy jumps can occur, only certain frequencies can appear in the spectrum.

The Bohr Model of the Atom

The problem with finding out that electrons were capable of existing only at certain energy levels was coming up with a model to explain these levels. In 1913 Niels Bohr (1885-1962), a Danish physicist, proposed a theoretical model for the hydrogen atom. He chose hydrogen because its atoms are the simplest, having only one electron about the nucleus, and because it produces the simplest spectrum with the fewest lines. In his model, Bohr imagined the electron to move around the nucleus following fixed paths, or orbits, much as a planet moves around the sun. His model also restricted the sizes of the orbits and the energy that the electron could have

in a given orbit. The equation Bohr derived for the energy of the electron included a number of physical constants such as the mass of the electron, its charge, and Planck's constant. It also contained an integer, n , that Bohr called a *quantum number*. Each of the orbits could be specified by its value of n .

Bohr's Atomic Model

Bohr found that the electron had the least energy when $n = 1$, which corresponds to the first Bohr orbit. This lowest energy state is called the ground state. This orbit also brings the electron closest to the nucleus.

When the hydrogen atom absorbed energy, as it does in a gas discharge tube, the electron is raised from the orbit $n = 1$ to a higher orbit such as $n = 2$ or $n = 3$ or even higher. Then when the electron drops back to a lower orbit, energy is emitted in the form of light. Since the energy of the electron in a given orbit is fixed, a drop from one particular orbit to another, say from $n=2$ to $n=1$, always releases the same amount of energy, and the frequency of light emitted because of this change in energy is always precisely the same.

Bohr's model of the atom was both a success and a failure. It successfully predicted the frequencies of the lines in the hydrogen spectrum, so it seemed to be valid. Nevertheless the model was a total failure when it tried to predict energy levels for atoms with more than one electron. Still the theory held some validity and is still used to introduce students to the concept of orbital shells and the first quantum number " n ".

The Quantum Numbers

The theory of quantum mechanics tells us that in an atom, the electrons are found in orbitals, and each orbital has a characteristic energy. Orbital means "small orbit". We are interested in two properties of orbitals - their energies and their shapes. Their energies are important because we normally find atoms in their most stable states, which we call their *ground states*, in which electrons are at their lowest possible energies.

The Principal Quantum Number , n gives size of an atom or energy levels

The quantum number n is called the principle quantum number. You already know this as shell. The shell "K" has been given the value $n = 1$, the "L" shell has been given the value $n = 2$.

n	1	2	3	4	...
shell	K	L	M	N	...

The principle quantum number serves to determine the size of the orbital, or how far the electron extends from the nucleus. The higher the value of n the further from the nucleus we can expect to find it. As n increases so does the energy required as well because the further out from the nucleus you go the more energy the electron must have to stay in orbit. Bohr's work took into account only this first principle quantum number. His theory worked for hydrogen because hydrogen just happens to be the one element in which all orbitals having the same value of n also have the same energy. Bohr's theory failed for atoms other than hydrogen, however, because orbitals with the same value of n can have different energies when the atom has more than one electron.

Azimuthal or Angular Momentum Quantum Number, l gives shape of the orbitals

The secondary quantum number, l , divides the shells up into smaller groups of subshells called orbitals. The value of n determines the possible values for l . For any given shell the number of subshells can be found by $l = n - 1$. This means that for $n = 1$, the first shell, there is only $l = 1 - 1 = 0$ subshells. i.e. the shell and subshell are identical. When $n = 2$ there are two sets of subshells; $l = 1$ and $l = 0$. A number could be used to identify the subshell however to avoid confusion between the numerical values of n and those of l the l values are given a letter code.

value of l	0	1	2	3	4
letter designation	s	p	d	f	g

To designate a particular subshell we write the number of the shell itself followed by the subshell designator.

n	l	<i>This illustrates the relationship between "n" and "l".</i>
1	s	the first shell has one orbital type associated with it.
2	s p	the second shell has two orbital types associated with it.
3	s p d	etc
4	s p d f	
5	s p d f g	

The principle quantum number describes size and energy, but the second quantum number describes shape. The subshells in any given orbital differ slightly in energy, with the energy in the subshell increasing with increasing l . This means that within a given shell, the s subshell is lowest in energy, p is the next lowest, followed by d, then f, and so on. For example:

$$4s < 4p < 4d < 4f \text{ ---> increasing energy}$$

The Magnetic Moment Quantum Number, m_l gives orientation number of orbitals and directions

The third quantum number, m_l , is known as the magnetic quantum number. It splits the subshells into individual orbitals. This orbital describes how an orbital is orientated in space relative to other orbitals. i.e. It gives 3D information. The first "s" subshell has a magnetic number of "1". The "p" subshell has a magnetic number of "3". A simple numeric progression gives us:

s	p	d	f	<---name of subshell
1	3	5	7	<--- number of orbitals in that subshell
2	6	10	14	<--- number of electrons that will fit into that subshell.*

*An orbital can hold two electrons total. It may hold none, one or two but never more than two.

The Spin Magnetic Number, m_s

The fourth and final quantum number is used to indicate the orientation of the two electrons in each orbital. The values for m_s are $+1/2$ and $-1/2$. An atom is the most stable when its electrons

have the lowest possible energy. Electrons get the lowest possible energy when they occupy the lowest possible energy orbitals available. But what determines how the electrons "fill" the orbitals? Two electrons can fill each orbital. How can two electrons both with a negative charge, and therefore mutually repulsive stay together in the same orbital?

The concept of electron spin is based on the fact that electrons behave like tiny magnets. An electron spins about its axis much like a toy top. The revolving electrical charge generates a magnetic field. (The same effect makes electric motors and generators work.) The electron can spin in two directions, either clockwise or counter-clockwise. Using the Left Hand rule for magnetic fields the clockwise spinning electrons generate a north pole on top and a south pole on the bottom. The counter-clockwise spinning electrons generate a north pole on the bottom and a south pole on the top.

In 1925 an Austrian physicist, Wolfgang Pauli (1900-1958), expressed the importance of electron spin in determining electronic configurations. The Pauli exclusion principle states that no two electrons in the same atom may have identical values for all four quantum numbers. This means that the two electrons that fill any particular orbital must have opposite spins. What happens if an orbital contains only 1 electron? Then its magnetic field is not cancelled out and it can be attracted to other outside magnetic fields. Atoms having at least one an *unpaired* electron are paramagnetic and can be attracted to magnetic fields. Atoms with no *unpaired* electrons are said to be diamagnetic and are not seen to be magnetic.

In general, the number of electrons in a shell is $2n^2$.

shell	number of subshells	maximum number of electrons
1	1s	2
2	2s 2p	8
3	3s 3p 3d	18
4	4s 4p 4d 4f	32