Lesson 3:
How Sunscreens Block: The Absorption of UV Light

Student Materials

Contents

Absorption of Light by Matter: Student Reading

Absorption is one of the ways in which light can interact with matter. In absorption, the energy of the light shining on an object (or a gas or liquid) is captured by the substance's molecules and used to move them from a ground (low energy) state to higher (excited) energy states.

Absorption by a Single Atom

Absorption can only occur when the energy packet carried by one photon of light is equal to the energy required to bring about transition between states. Since the energy of a photon is directly related to its frequency (by the formula E=hf where E is the photon's energy, f is its frequency, and h is Plancks' constant: 6.26 x 10^{-34} J s), then a given transition can only be caused by one specific frequency of light.

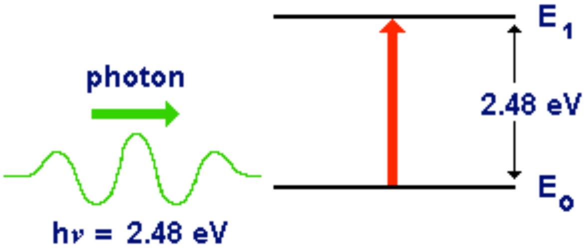


Figure 1: The energy of a photon must be exactly the same as the energy difference between levels for it to be absorbed [1]

The frequency of a photon is inversely related to its wavelength (by the formula λ=c/f where λ is the wavelength, f is the frequency and c is the speed of light in a vacuum). This means that the energy of the energy transition determines what frequency and hence what wavelength of light can be absorbed. The greater the energy of transition, the smaller the wavelength of light that can be absorbed.

What exactly are these different energy states that the atom can be in? Well, in atoms they correspond to the types of electron configurations in orbitals that are allowed by the rules of quantum mechanics. You may be familiar with transitions between electronic states if you've studied light emission by gasses. The main difference between the emission and the absorption of light is that in one case the light energy is being given off and in the other it is being captured, but in both cases the light energy of the photon absorbed or emitted corresponds to the energy difference between the two electronic states.

For example, Figure 2 shows several different energy levels and one possible transition for a hydrogen atom (in a gas). Each of the horizontal lines is an electronic state that the atom can be in and each the vertical line shows a possible "jump" or transition that could occur. Each possible transition (like the one shown) has a characteristic energy, and thus a specific wavelength of light that must be absorbed for it to occur. If we shine a flashlight at a sample of hydrogen atoms in a gas and measure the wavelengths of light that get through at the end, we would see an absorption spectrum like the one shown in the bottom of Figure 3. What you see is the full color spectrum, except for the “absorption lines”, the specific wavelengths of light whose energy was absorbed because it corresponds to the energy difference between states for helium atoms. Similarly, if we stimulate a hydrogen atom to emit light, it will do so only at these same characteristic wavelengths whose energy (which is proportional to its frequency given by E=hf) corresponds to the energy between possible electron states for the hydrogen atom.

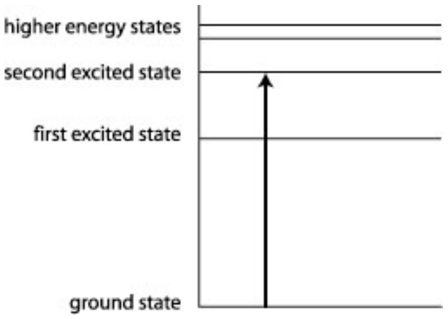
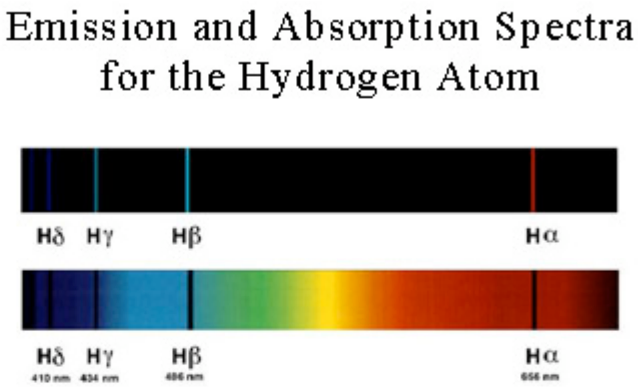


Figure 2: Energy states and one transition for the hydrogen atom



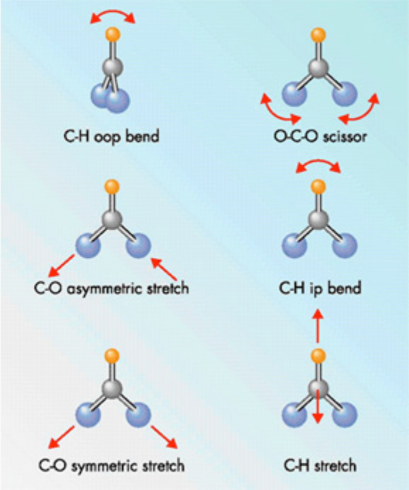
http://dosxx.colorado.edu/~bagenal/1010/SESSIONS/13.Light.html

Figure 3: Emission (top) and absorption (bottom) spectra for the hydrogen atom [2] Each line corresponds to a single transition

Absorption by Molecules

When we start dealing with molecules instead of atoms, the situation gets more complicated. When we talk about atoms being bonded together, it is not as if they are solidly glued to one another. The “glue” that hold the atoms together in a molecule is the attractive forces between the electrons and the different nuclei (intramolecular forces.) As electrons are constantly in motion, the strength of these attractive forces fluctuates over time allowing the nuclei to move back and forth and giving rise to molecular vibrations. As with electronic states, there are only certain vibrational modes which are possible. One simple example of molecular vibration can be seen in formate ($COOH^-$), a molecular anion. The six possible vibration modes for formate are shown in Figure 4.

Within each vibrational state, there are also multiple ways that the atoms in the molecule can rotate. These are called rotational states and again there are a limited number of possible rotational states that the molecule can be in. When we consider vibrational and rotational states, we realize that even without exciting any electrons (ground state), there are a bunch of different energy states that a molecule can be in. Not surprisingly, this is also true for the excited state. So now we have a situation in which the molecule can transfer from any of the ground states to any of the excited states and instead of a single energy of transition, we have a large group of energy gaps able to absorb light energy.



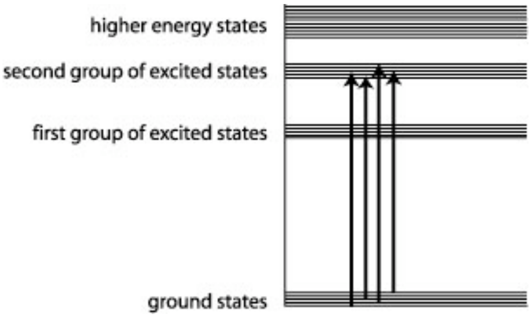


Figure 4: 6 Vibrational modes for formate ion ($COOH^{-}$) [3]

Figure 5: Multiple Transition Energies for Electrons in Molecules

The multiple energy transitions possible correspond to a range of photon energies and thus frequencies of light that can be absorbed. Comparing Figure 5 with Figure 2 we can see that where there was a single transition for an atom, there are multiple ones for a molecule. Each arrow corresponds to one transition. Since we now have clusters of energy levels (due to the different variations in rotational and vibrational energies possible), we have a cluster of multiple transitions with similar energies.

Thus light absorbing molecules produces multiple, closely spaced absorption lines which combine to form an absorption curve. Absorption curves can be described both by their absorption range and peak absorption wavelength as shown in Figure 6. Not that absorption is not uniform across the range; it is greatest at the peak and drops off rapidly. So absorption at the edges of the range is not very good.

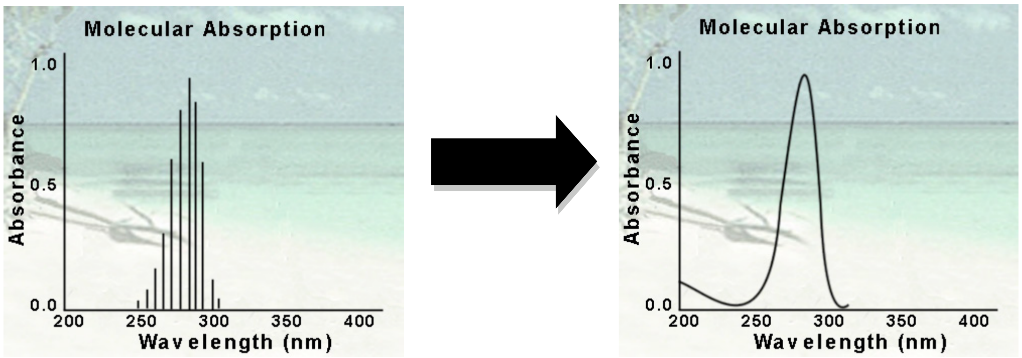


Figure 6: Absorption lines and curve spectrum for an organic molecule (Range = 245 – 305 nm) (Peak = 290 nm)

Absorption by Ionic Compounds

In inorganic compounds, there is no discrete atom or molecule to talk about. Instead the electrons belong to the positive nuclei as a collective group. Because so many atoms are close together and involved, there are a very large number of possible energy levels for electrons in both the ground and excited states as shown in Figure 7. Because there are so many possible energy states packed so closely together, we assume that electrons can have virtually any energy within each state and change between them very easily. Thus we call the group of possible energies for the ground states the "valence band" and the group of possible energies for the excited state the "conduction band".

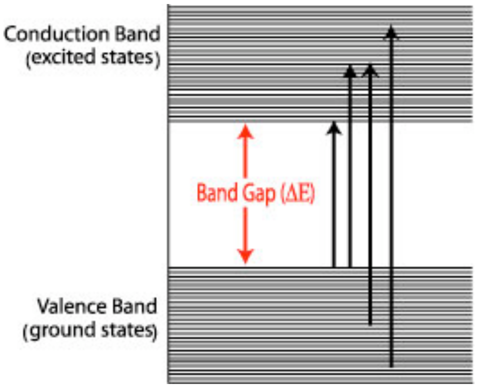


Figure 7: The large number of tightly packed energy levels possible for ground and excited electron states and possible transitions

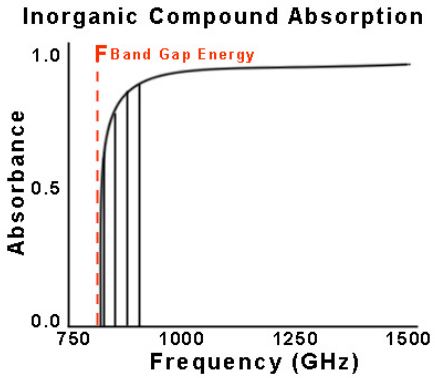


Figure 8: Absorption versus frequency graph for an inorganic compound. The smallest f of light that can be absorbed corresponds has an energy equal to the band gap energy

Electrons can transition from any energy value in the valence band to any energy value in the conduction band. The energy spacing between the two bands is called the "band gap" and is the minimum amount of energy that the substance can absorb. This corresponds to the minimum frequency (and maximum wavelength) of light that the substance can absorb as shown in Figure 8. Now instead of seeing an absorption peak, we see almost complete absorption up to a cut-off wavelength (which corresponds to the energy of the band gap). For example, zinc oxide (ZnO) has a particular band gap (minimum energy that can be absorbed). Using $E=hf$ and $c=\lambda f$ (where h is Planck's constant and c is the speed of light) we can calculate that this corresponds to light with a wavelength of 380 nm. Thus light with higher frequencies and energies is almost completely absorbed and light with lower frequencies and energies is not absorbed. If we think about absorption in relation to the wavelength of light (instead of the frequency) our graph gets reversed (remember that wavelength and frequency are inversely related by $c=\lambda f$). This leads to an absorption spectra with the characteristic sharp drop shown in Figure 9. Thus the *minimum frequency* corresponds to a *maximum wavelength* up to which the inorganic sunscreen ingredients can absorb.

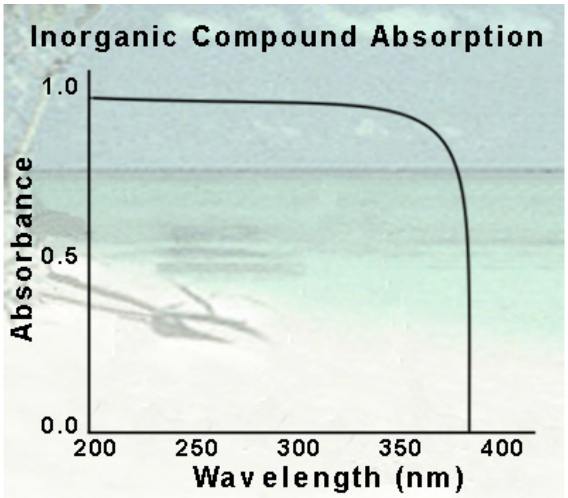


Figure 9: Absorption Spectrum for an Inorganic Compound (ZnO). Light with wavelengths less than 380 nm is almost completely absorbed. Light with wavelengths greater than 380 nm is not absorbed at all.

References

(Internet resources accessed December 2005)

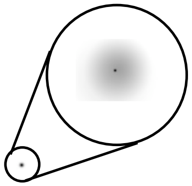
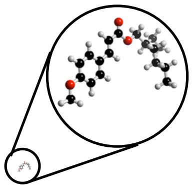
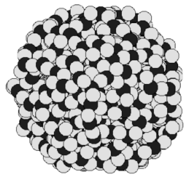
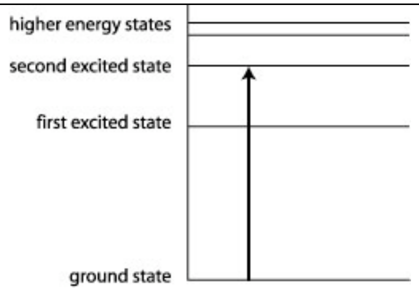
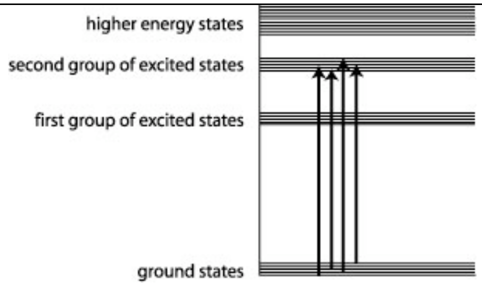
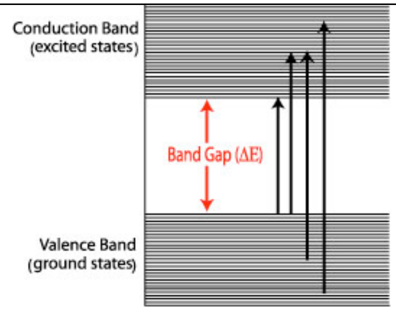
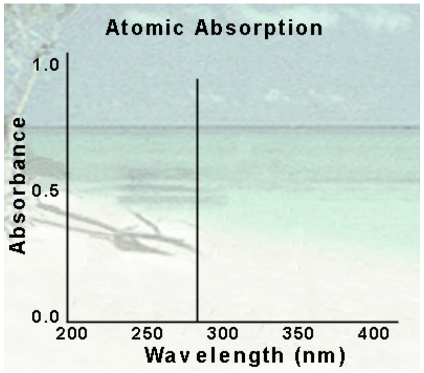
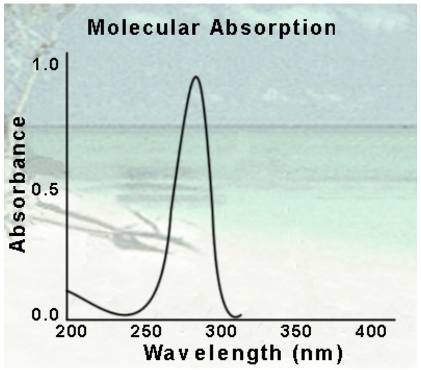
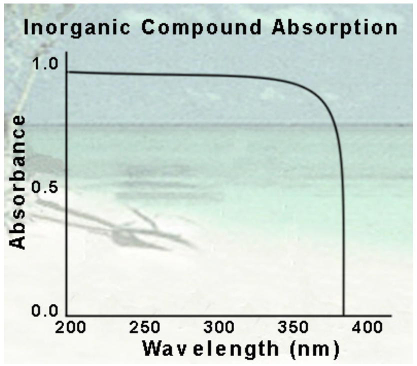
[1] http://members.aol.com/WSRNet/tut/absorbu.htm

[2] http://staff.imsa.edu/science/astro/astrometry/spectra/sld010.htm

[3] http://www.isis.rl.ac.uk/ISIS97/feature14.htm

[4] http://www.nptel.iitm.ernet.in/courses/IITMadras/CY101/lecture16new/lecture16.htm

Absorption Summary: Student Handout

	Atoms	Organic Molecules	Inorganic Compounds
Structure (not drawn to scale)	Nucleus with electron cloud 	Individual molecule 	Cluster of ions 
Energy Levels			
Absorption Spectrum	Absorbs specific λ 	Absorbs specific λ range 	Absorbs all UV < critical λ 
UV Protection	Minimal	Parts of UVA or UVB spectrum	Broad spectrum UVA and UVB

Name__________________________________ Date______________________ Period ______

Reflecting on the Guiding Questions: Student Worksheet

Think about the activity you just completed. What did you learn that will help you answer the guiding questions? Jot down notes in the spaces below.

1. What are the most important factors to consider in choosing a sunscreen?

What I learned in this activity:

What I still want to know:

2. How do you know if a sunscreen has “nano” ingredients?

What I learned in this activity:

What I still want to know:

3. How do “nano” sunscreen ingredients differ from most other ingredients currently used in sunscreens?

What I learned in this activity:

What I still want to know: