ACTIVITY 5– THE CHEMICAL BEHAVIOR OF ATOMS

Background Information

The Nature of Light and the Electromagnetic Spectrum

What is light? It cannot be felt (although the heat from it may), it has no apparent resting mass, and it does not take up space. Even though light seems to be insubstantial, early scientists still considered it important enough to be one of the four earthly elements.

What our senses detect as light is actually radiation that is part of a continuum that makes up the **electromagnetic spectrum**. The electromagnetic spectrum is a continuum (the divisions between types of radiation are not highly defined, and there is some overlap in regions) of energies ranging from the high-energy gamma and x-rays to low energy radio- and microwaves. We should note that **radiation** is not limited to nuclear decay, but includes any energy emitted in all directions from a single source. The Sun is a common example of a source of radiation. Five million tons of the Sun's matter is converted into energy each second, and this energy is radiated into space. We see different colors because the radiation that reaches us is not all identical. Raindrops act as prisms to break the sunlight into a spectrum of colors, one differing from another due to their different wavelengths. The word "spectrum" comes from the Latin word for "ghost" and was first used by Newton when describing the rainbow band of colors.

What makes up light troubled early scientists. It could not be made up of tiny particles because when two beams of different wavelengths crossed each other, one remained unaffected by the other. If both beams were made of particles, they would collide and the beams would scatter. It was suggested that light might consist of very small waves, behaving much like the waves produced in water, advancing in straight lines. When **white light** (made of all the colors) passes through a prism, we see the colors of the rainbow—red, orange, yellow, green, blue, indigo and violet. It was proposed that the waves of light with the longest wavelength (red) would be refracted to a lesser degree than the shorter waves (indigo and violet).

The **visible range** of the electromagnetic spectrum lies from about 400 nm to 700 nm (1 nm = 10^{-9} m). The **wavelength (λ)** is the distance between successive peaks of a wave. The time it takes each full wave to pass a fixed point is the **frequency** (***f***). All light moves through a vacuum at the same rate of speed, about 3×10^8 m/s (the **speed of light, *c***). Frequency and wavelength are inversely related to each other, as the product of the two equals the speed of light: $c=\lambda f$.

Line Spectra

If energy is put into the hydrogen atom, the electron can be excited to different levels. When the electron falls back to a lower level it emits light. With the use of a spectroscope or diffraction grating lens, we detect a set of different colored lines in the visible region. These series of lines is called the **Balmer series**, named after Johann Balmer, a Swiss teacher. Balmer developed an equation that could be used to calculate wavelengths of the visible spectrum. This equation was later expanded to be able to calculate the wavelength of all of the lines of the hydrogen spectrum. This equation was called the Rydberg equation: $1/\lambda = R_H(1/n_1^2 - 1/n_2^2)$. Where the symbol λ stands for the wavelength, R_H is a constant (109678 cm^{-1}), and n_1 and n_2 are variables whose values are whole numbers that range from 1 to infinity (∞). The restriction is that $n_2 > n_1$ in order to get positive values for the wavelength.

Using the Rydberg equation, we find that when the electron falls from the 3rd level to the 2nd level it gives off light in the red part of the spectrum with a $\lambda = 656$ nm. Then if we wanted to find the amount of energy that is being emitted we could use Planck's equation: $\Delta E = hf$, where ΔE is the change in energy between the two levels, h is Planck's constant (6.626×10^{-34} J · s) and f is frequency of light (c/λ, where c is the speed of light, 3.00×10^8 m/s). Then solving the equation we find that light with a λ of 656 nm would emit 3.03×10^{-19}J or if you were talking in terms of moles your value would be 1.82×10^5J/mol or 182 kJ/mol or 43.6 kcal/mol. The SI unit of energy is the **joule** (J), although we may see the use of **calorie** (cal).

1 cal = 4.184 J

Expanding the line spectrum beyond the visible series, the **Lyman series** lies in the ultraviolet region and the **Paschen**, Brackett, and Pfund series lie in the infrared region. If an electron falls from the 2nd level to the 1st level, it will emit 984 kJ/mol or 235.2 kcal/mol of energy (ultraviolet). All of the energy levels of the hydrogen atom are satisfied with Rydberg's equation.

The Bohr Model

Balmer's work allowed the calculation of wavelengths for the lines seen in the line spectra, but did not explain *why* they exist. However, the regularity of the lines suggested that they had something to do with atomic structure.

One view of the atom was that electrons travel in circular orbits around the nucleus. This led to the obvious question: if electrons are negatively charged, and the nucleus is positively charged, and if opposite charges attract, what keeps the electrons from spiraling into the nucleus? It was known at the time that a negatively charged particle gives off electromagnetic radiation as it travels in a circular path around a positive center, losing energy in the process. But atoms seemed indefinitely stable.

In 1913 Niels Bohr, a young Danish physicist working in Rutherford's laboratory, combined ideas from classical physics and the newly emerging quantum theory to explain the structure of the hydrogen atom and the existence of line spectra. Bohr proposed the notion that the single electron of the hydrogen atom could occupy only certain energy levels. He called these energy levels orbits, and represented the energy difference between any two adjacent orbits as a single **quantum** of energy. Bohr described the energy of the electron in an atom as being "quantized." Whenever the electron was in a particular orbit, whatever its particular energy level, it did not either gain or lose energy. For the electron to move from one orbit to another, it must gain or lose a level of energy that corresponds to the energy difference between the two orbits. Bohr concluded that the line spectrum exhibits discrete lines because an atom's energy has only certain discrete levels, or states. When an electron is in the lowest energy state, it is said to be in its **ground state**. When an electron is in a higher energy state it is said to be in an excited state. Bohr worked out calculations that showed the permitted orbits that would result in the absorption or emission of discrete quanta of energy that perfectly matched the wavelengths of the lines in the hydrogen spectrum. Bohr was able to align the unseen interior structure of the atom with the observable lines in the hydrogen spectrum, and won the Nobel Prize in 1922 for this work.